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VCE CHEMISTRY

Units 1 & 2

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Units 1 & 2

Jason Wallace, Angus Goatcher, Jamie Elliott, Jamie Garnham

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Edrolo

At Edrolo, we're transforming the way the students learn and teachers teach.

Our mission is simple: to improve education.

PUBLISHED IN AUSTRALIA BY Edrolo
321 Exhibition Street Melbourne VIC 3000, Australia

Ref: 2.1.2

First published 2022. Reprinted 2023

National Library of Australia Cataloguing-in-Publication data

TITLE: Edrolo VCE Chemistry Units 1 & 2
CREATOR: Edrolo et al.
ISBN: 978-1-922901-23-1
TARGET AUDIENCE: For secondary school age.
SUBJECTS: Chemistry--Study and teaching (Secondary)--Victoria
Chemistry--Victoria--Textbooks.
Chemistry--Theory, exercises, etc.
OTHER CREATORS/CONTRIBUTORS: Louise Lennard, Sophie Renshaw, Jack Schwenk, Chris Sparks, Christy Thomas

LAYOUT DESIGN: Christopher Dell and Edrolo
TYPESET BY: Daniel Douch, Dean Dragonetti, Yasmin Ramos, Eli Minuz, Maxwell Keesing
COVER DESIGN BY: Tasha Hassapis, Yev_1234/Shutterstock.com

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FEATURES OF THIS BOOK

Edrolo's VCE Chemistry Units 1 & 2 textbook has the following features.

Theory

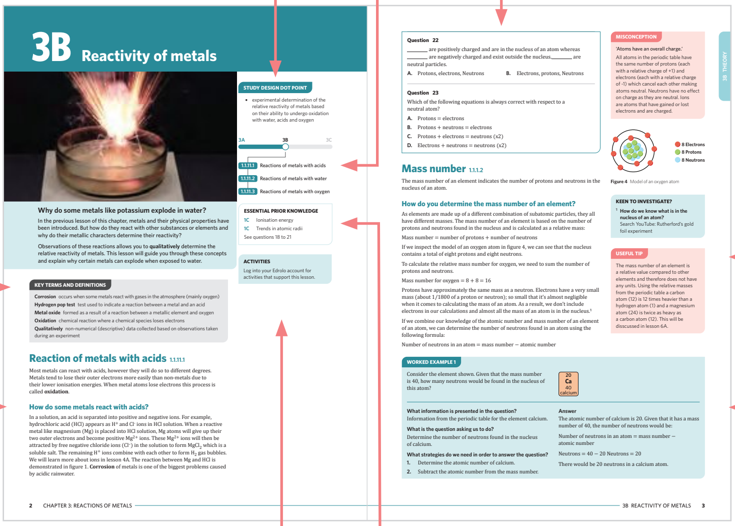
Study design dot points provide explicit links between the content covered in each lesson and the VCAA study design.

Key knowledge units break down the theory into smaller chunks that help navigate the corresponding theory lesson videos on the Edrolo online platform.

Progress questions test if students can remember bite-sized chunks of theory and overcome common misconceptions early in the learning journey. They are stepping stones between the content and exam-style questions.

Key terms are defined and highlighted to support learning

Guiding questions focus and simplify the key knowledge required by the VCAA study design.



Explore boxes include misconceptions, useful tips, strategy, and keen to investigate boxes.

Worked examples apply content to exam-style questions with strategies, full working and explanations of each step in the process.

Activities are provided on the Edrolo online platform and allow engagement and further understanding of the content.

Essential prior knowledge questions are designed to break down any misconceptions and highlight prerequisite knowledge from previous lessons that is 'essential' to understanding the lesson's content.

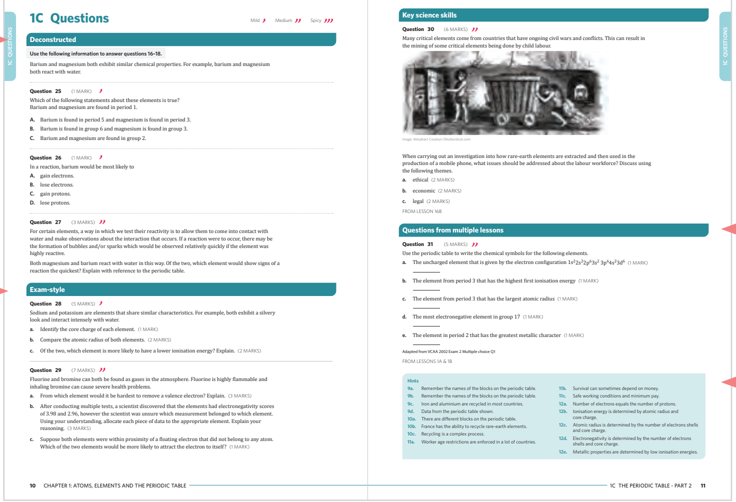
Questions

Spiciness rating of questions measures the difficulty and increases within each type of question type.

Key science skills questions develop inquiry skills students need for SACs and other assessments.

Deconstructed questions provide prompts for students to piece together the steps needed to answer an exam-style question.

Exam-style questions reflect the style of questions presented in end-of-year exams.



Questions from multiple lessons combine key knowledge from previous lessons to develop a deeper understanding of concepts.

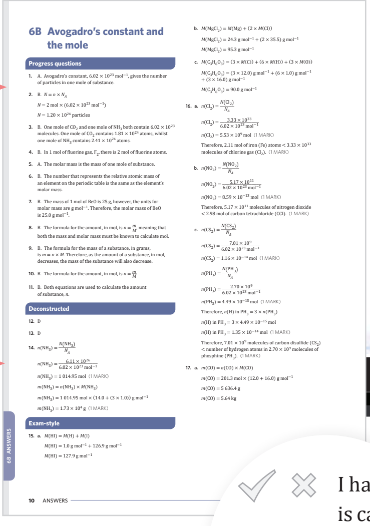
Hints are provided for questions to help give students a 'starting point' when they are unsure.

Answers

Worded and visual responses include exemplar answers to all questions to show what a full mark response looks like.

Online video solutions provide immediate feedback and extra guidance on how to answer questions.

Calculation responses show full working along with mark allocations to help students form good habits and to provide more detailed guidance for students.



6B Avogadro's constant and the mole

Progress questions

1. A. Avogadro's constant, 6.02×10^{23} mol^{-1}, gives the number of particles in one mole of substance.
2. B. $N = n \times N_A$

   $N = 2 \text{ mol} \times (6.02 \times 10^{23} \text{ mol}^{-1})$

   $N = 1.20 \times 10^{24}$ particles
3. B. One mole of CO_2 and one mole of NH_3 both contain 6.02×10^{23} molecules. One mole of CO_2 contains 1.81×10^{24} atoms, whilst one mole of NH_3 contains 2.41×10^{24} atoms.
4. B. In 1 mol of fluorine gas, F_2, there is 2 mol of fluorine atoms.
5. A. The molar mass is the mass of one mole of substance.
6. B. The number that represents the relative atomic mass of an element on the periodic table is the same as the element's molar mass.
7. B. The mass of 1 mol of BeO is 25 g, however, the units for molar mass are g mol^{-1}. Therefore, the molar mass of BeO is 25.0 g mol^{-1}.
8. B. The formula for the amount, in mol, is $n = \frac{m}{M}$, meaning that both the mass and molar mass must be known to calculate mol.
9. B. The formula for the mass of a substance, in grams, is $m = n \times M$. Therefore, as the amount of a substance, in mol, decreases, the mass of the substance will also decrease.
10. B. The formula for the amount, in mol, is $n = \frac{m}{M}$.
11. B. Both equations are used to calculate the amount of substance, n.

Deconstructed

12. D
13. D
14. $n(NH_3) = \frac{N(NH_3)}{N_A}$

    $n(NH_3) = \frac{6.11 \times 10^{26}}{6.02 \times 10^{23} \text{ mol}^{-1}}$

    $n(NH_3) = 1\,014.95$ mol (1 MARK)

    $m(NH_3) = n(NH_3) \times M(NH_3)$

    $m(NH_3) = 1\,014.95 \text{ mol} \times (14.0 + (3 \times 1.0)) \text{ g mol}^{-1}$

    $m(NH_3) = 1.73 \times 10^4$ g (1 MARK)

Exam-style

15. a. $M(HI) = M(H) + M(I)$

    $M(HI) = 1.0 \text{ g mol}^{-1} + 126.9 \text{ g mol}^{-1}$

    $M(HI) = 127.9 \text{ g mol}^{-1}$

    b. $M(MgCl_2) = M(Mg) + (2 \times M(Cl))$

    $M(MgCl_2) = 24.3 \text{ g mol}^{-1} + (2 \times 35.5) \text{ g mol}^{-1}$

    $M(MgCl_2) = 95.3 \text{ g mol}^{-1}$

    c. $M(C_3H_6O_3) = (3 \times M(C)) + (6 \times M(H)) + (3 \times M(O))$

    $M(C_3H_6O_3) = (3 \times 12.0) \text{ g mol}^{-1} + (6 \times 1.0) \text{ g mol}^{-1} + (3 \times 16.0) \text{ g mol}^{-1}$

    $M(C_3H_6O_3) = 90.0 \text{ g mol}^{-1}$
16. a. $n(Cl_2) = \frac{N(Cl_2)}{N_A}$

    $n(Cl_2) = \frac{3.33 \times 10^{33}}{6.02 \times 10^{23} \text{ mol}^{-1}}$

    $n(Cl_2) = 5.53 \times 10^9$ mol (1 MARK)

    Therefore, 2.11 mol of iron (Fe) atoms $< 3.33 \times 10^{33}$ molecules of chlorine gas (Cl_2). (1 MARK)

    b. $n(NO_2) = \frac{N(NO_2)}{N_A}$

    $n(NO_2) = \frac{5.17 \times 10^{11}}{6.02 \times 10^{23} \text{ mol}^{-1}}$

    $n(NO_2) = 8.59 \times 10^{-13}$ mol (1 MARK)

    Therefore, 5.17×10^{11} molecules of nitrogen dioxide $<$ 2.98 mol of carbon tetrachloride (CCl). (1 MARK)

    c. $n(CS_2) = \frac{N(CS_2)}{N_A}$

    $n(CS_2) = \frac{7.01 \times 10^9}{6.02 \times 10^{23} \text{ mol}^{-1}}$

    $n(CS_2) = 1.16 \times 10^{-14}$ mol (1 MARK)

    $n(PH_3) = \frac{N(PH_3)}{N_A}$

    $n(PH_3) = \frac{2.70 \times 10^9}{6.02 \times 10^{23} \text{ mol}^{-1}}$

    $n(PH_3) = 4.49 \times 10^{-15}$ mol (1 MARK)

    Therefore, $n(H)$ in $PH_3 = 3 \times n(PH_3)$

    $n(H)$ in $PH_3 = 3 \times 4.49 \times 10^{-15}$ mol

    $n(H)$ in $PH_3 = 1.35 \times 10^{-14}$ mol (1 MARK)

    Therefore, 7.01×10^9 molecules of carbon disulfide (CS_2) $<$ number of hydrogen atoms in 2.70×10^9 molecules of phosphine (PH_3). (1 MARK)
17. a. $m(CO) = n(CO) \times M(CO)$

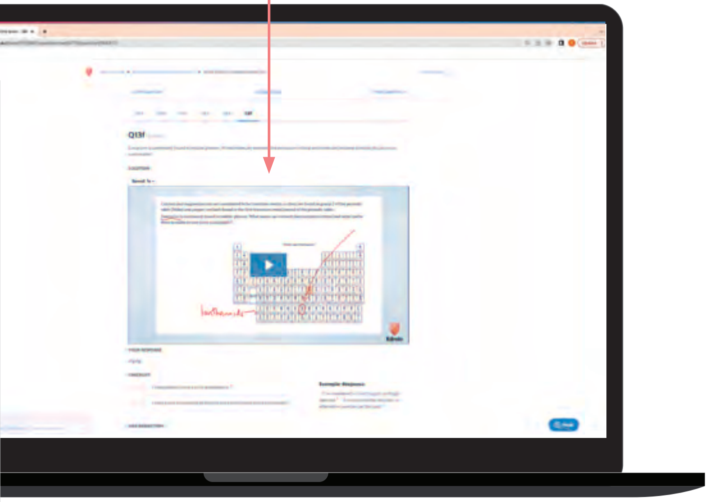
    $m(CO) = 201.3 \text{ mol} \times (12.0 + 16.0) \text{ g mol}^{-1}$

    $m(CO) = 5\,636.4$ g

    $m(CO) = 5.64$ kg

6B ANSWERS

10 ANSWERS



✓ ✗ I have des
is calculate

✓ ✗ I have ident

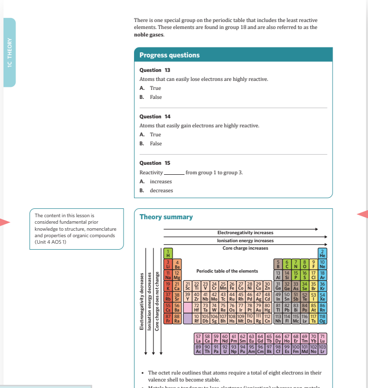
✓ ✗ I have ide

Checklists are provided for all answers (worth more than 1 mark) and they explain the purpose of each part of a full-mark answer to help students develop their scientific communication skills.

Reviews

The content in this lesson identifies the parts of the Unit 3 or 4 course the theory presented in the lesson is relevant to.

Theory summary is an outline of the knowledge from the entire lesson.



1C THEORY

There is one special group on the periodic table that includes the least reactive elements. These elements are found in group 18 and are also referred to as the **noble gases**.

Progress questions

Question 13

Atoms that can easily lose electrons are highly reactive.

A. True

B. False

Question 14

Atoms that easily gain electrons are highly reactive.

A. True

B. False

Question 15

Reactivity ______ from group 1 to group 3.

A. increases

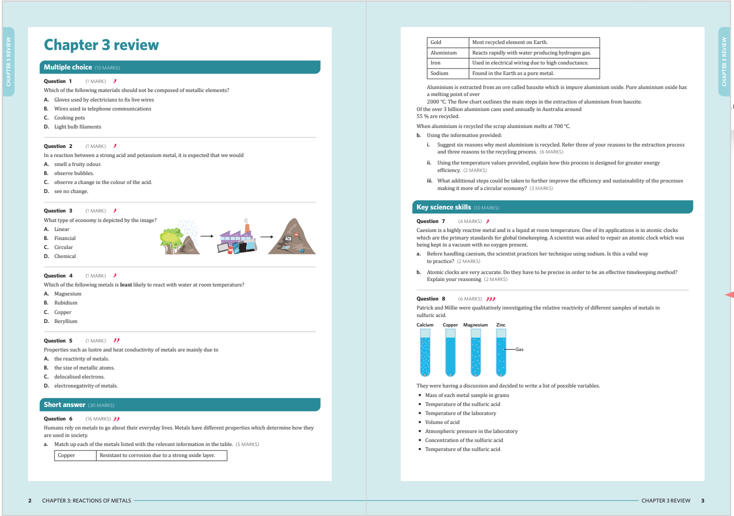
B. decreases

The content in this lesson is considered fundamental prior knowledge to structure, nomenclature and properties of organic compounds (Unit 4 AOS 1)

Theory summary

- The octet rule outlines that atoms require a total of eight electrons in their valence shell to become stable.
- Metals have a tendency to lose electrons (ionisation) whereas non-metals have a tendency to gain electrons (electronegativity).
- Metallic character decreases as we move across a period.
- Metalloids have characteristics of both metals and non-metals.
- Reactivity reflects the ease at which an atom will lose or gain electrons to satisfy the octet rule.

ELEMENTS AND THE PERIODIC TABLE



CHAPTER 3 REVIEW

Chapter 3 review

Multiple choice (10 MARKS)

Question 1 (1 MARK)

Which of the following materials should not be composed of metallic elements?

A. Gloves used by electricians to fix live wires

B. Wires used in telephone communications

C. Cooking pots

D. Light bulb filaments

Question 2 (1 MARK)

In a reaction between a strong acid and potassium metal, it is expected that we would

A. smell a fruity odour.

B. observe bubbles.

C. observe a change in the colour of the acid.

D. see no change.

Question 3 (1 MARK)

What type of economy is depicted by the image?

A. Linear

B. Financial

C. Circular

D. Chemical

Question 4 (1 MARK)

Which of the following metals is **least** likely to react with water at room temperature?

A. Magnesium

B. Rubidium

C. Copper

D. Beryllium

Question 5 (1 MARK)

Properties such as lustre and heat conductivity of metals are mainly due to

A. the reactivity of metals.

B. the size of metallic atoms.

C. delocalised electrons.

D. electronegativity of metals.

Short answer (30 MARKS)

Question 6 (16 MARKS)

Humans rely on metals to go about their everyday lives. Metals have different properties which determine how they are used in society.

a. Match up each of the metals listed with the relevant information in the table. (5 MARKS)

Copper	Resistant to corrosion due to a strong oxide layer.
Gold	Most recycled element on Earth.
Aluminium	Reacts rapidly with water producing hydrogen gas.
Iron	Used in electrical wiring due to high conductance.
Sodium	Found in the Earth as a pure metal.

2 CHAPTER 3: REACTIONS OF METALS

Aluminium is extracted from an ore called bauxite which is impure aluminium oxide. Pure aluminium oxide has a melting point of over 2000 °C. The flow chart outlines the main steps in the extraction of aluminium from bauxite.

Of the over 3 billion aluminium cans used annually in Australia around 55 % are recycled.

When aluminium is recycled the scrap aluminium melts at 700 °C.

b. Using the information provided:

   i. Suggest six reasons why most aluminium is recycled. Refer three of your reasons to the extraction process and three reasons to the recycling process. (6 MARKS)

   ii. Using the temperature values provided, explain how this process is designed for greater energy efficiency. (2 MARKS)

   iii. What additional steps could be taken to further improve the efficiency and sustainability of the processes making it more of a circular economy? (3 MARKS)

Key science skills (10 MARKS)

Question 7 (4 MARKS)

Caesium is a highly reactive metal and is a liquid at room temperature. One of its applications is in atomic clocks which are the primary standards for global timekeeping. A scientist was asked to repair an atomic clock which was being kept in a vacuum with no oxygen present.

a. Before handling caesium, the scientist practices her technique using sodium. Is this a valid way to practice? (2 MARKS)

b. Atomic clocks are very accurate. Do they have to be precise in order to be an effective timekeeping method? Explain your reasoning (2 MARKS)

Question 8 (6 MARKS)

Patrick and Millie were qualitatively investigating the relative reactivity of different samples of metals in sulfuric acid.

They were having a discussion and decided to write a list of possible variables.

- Mass of each metal sample in grams
- Temperature of the sulfuric acid
- Temperature of the laboratory
- Volume of acid
- Atmospheric pressure in the laboratory
- Concentration of the sulfuric acid
- Temperature of the sulfuric acid

CHAPTER 3 REVIEW 3

Chapter reviews consists of end-of-year exam-style questions that cover content from that chapter or area of study. They are structured into three sections: multiple choice, short answer and key science skills questions.

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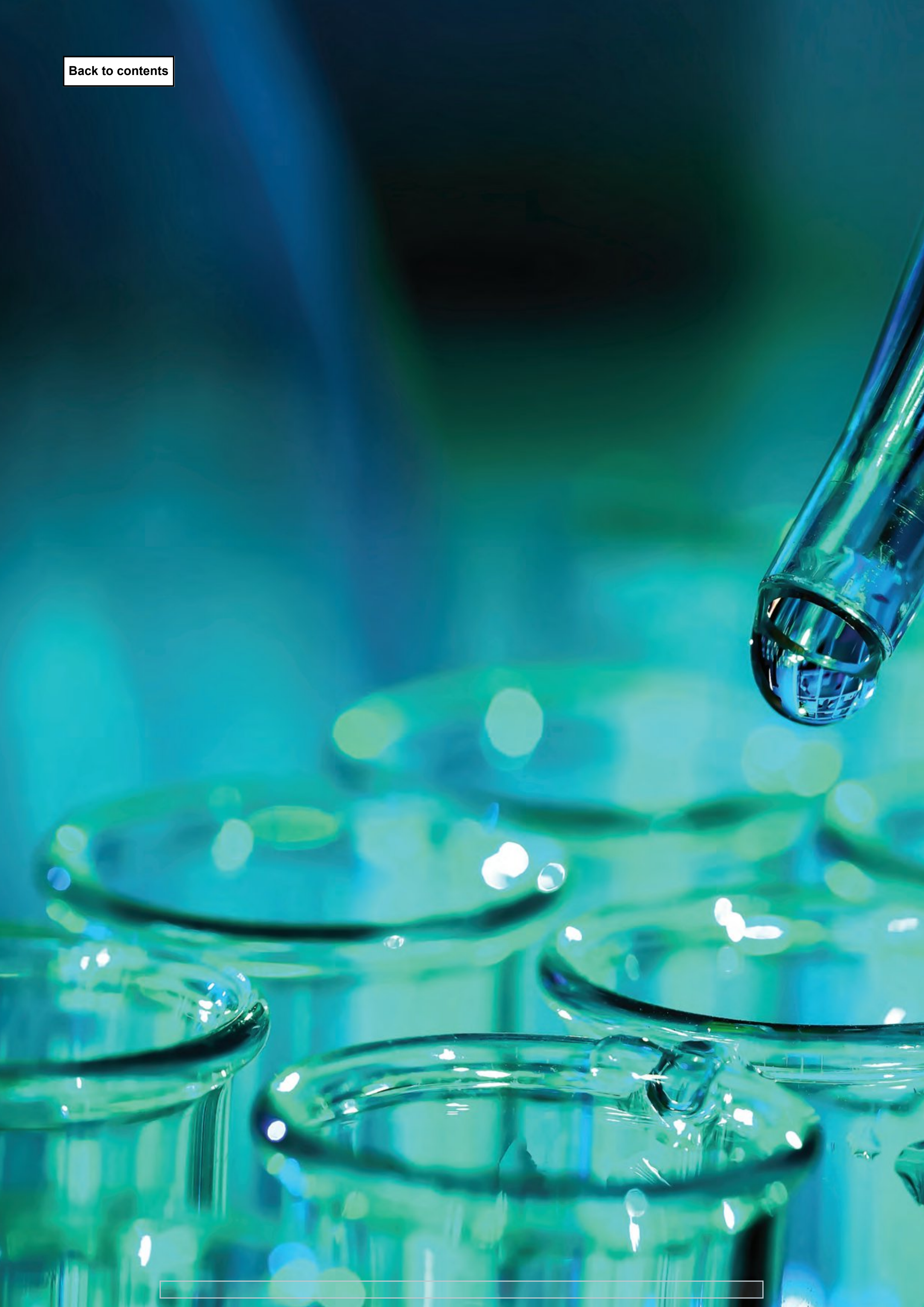

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1

UNIT 1

How can the diversity of materials be explained?

The development and use of materials for specific purposes is an important human endeavour. In this unit students investigate the chemical structures and properties of a range of materials, including covalent compounds, metals, ionic compounds and polymers. They are introduced to ways that chemical quantities are measured. They consider how manufacturing innovations lead to more sustainable products being produced for society through the use of renewable raw materials and a transition from a linear economy towards a circular economy.

Students conduct practical investigations involving the reactivity series of metals, separation of mixtures by chromatography, use of precipitation reactions to identify ionic compounds, determination of empirical formulas, and synthesis of polymers.

Throughout this unit, students use chemistry terminology including symbols, formulas, chemical nomenclature and equations to represent and explain observations and data from their own investigations and to evaluate the chemistry-based claims of others.

A student-directed research investigation into the sustainable production or use of a selected material is to be undertaken in Area of Study 3. The investigation explores how sustainability factors such as green chemistry principles and the transition to a circular economy are considered in the production of materials to ensure minimum toxicity and impacts on human health and the environment. The investigation draws on key knowledge and key science skills from Area of Study 1 and/or Area of Study 2.

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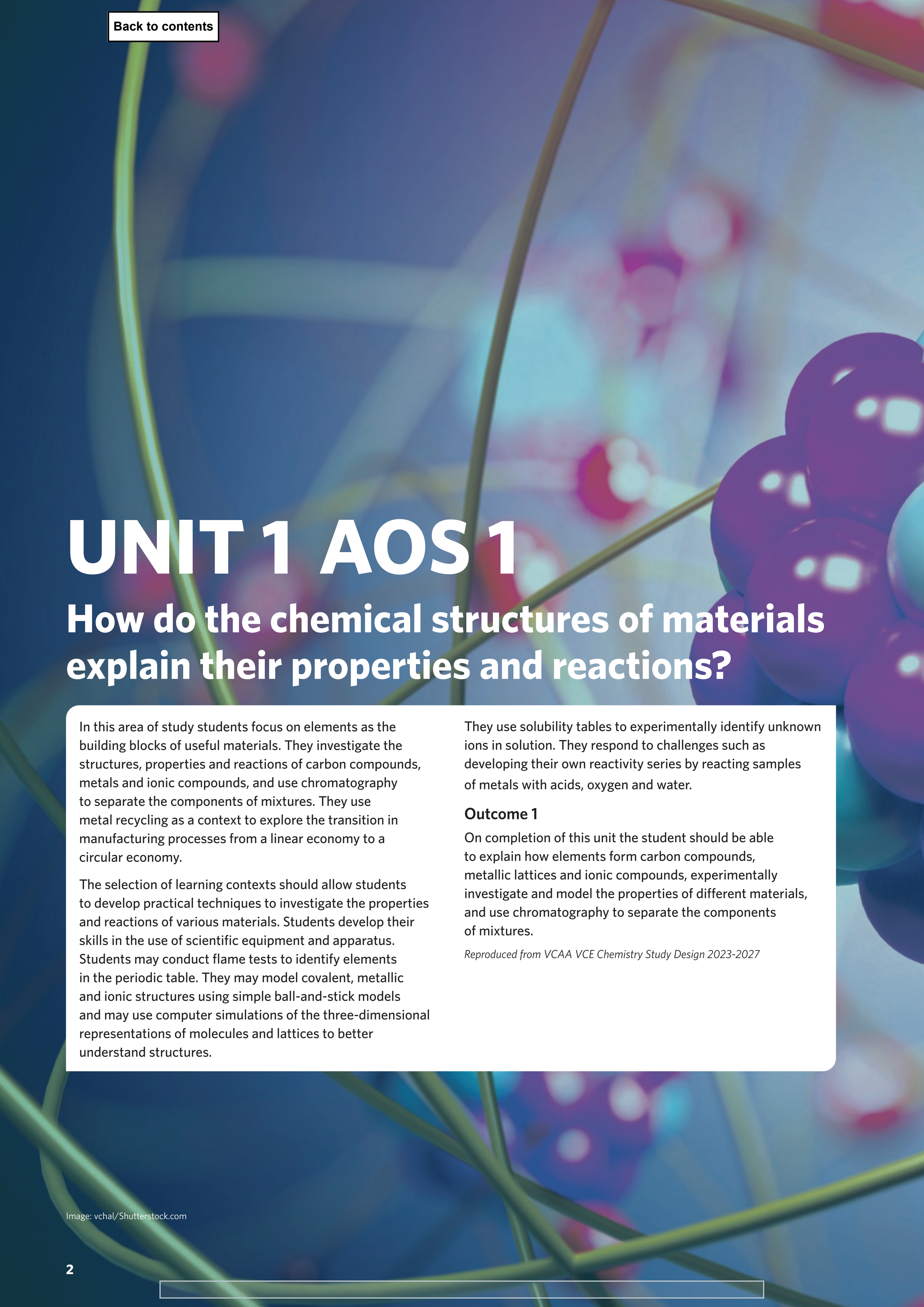

UNIT 1 AOS 1

How do the chemical structures of materials explain their properties and reactions?

In this area of study students focus on elements as the building blocks of useful materials. They investigate the structures, properties and reactions of carbon compounds, metals and ionic compounds, and use chromatography to separate the components of mixtures. They use metal recycling as a context to explore the transition in manufacturing processes from a linear economy to a circular economy.

The selection of learning contexts should allow students to develop practical techniques to investigate the properties and reactions of various materials. Students develop their skills in the use of scientific equipment and apparatus. Students may conduct flame tests to identify elements in the periodic table. They may model covalent, metallic and ionic structures using simple ball-and-stick models and may use computer simulations of the three-dimensional representations of molecules and lattices to better understand structures. They use solubility tables to experimentally identify unknown ions in solution. They respond to challenges such as developing their own reactivity series by reacting samples of metals with acids, oxygen and water.

Outcome 1

On completion of this unit the student should be able to explain how elements form carbon compounds, metallic lattices and ionic compounds, experimentally investigate and model the properties of different materials, and use chromatography to separate the components of mixtures.

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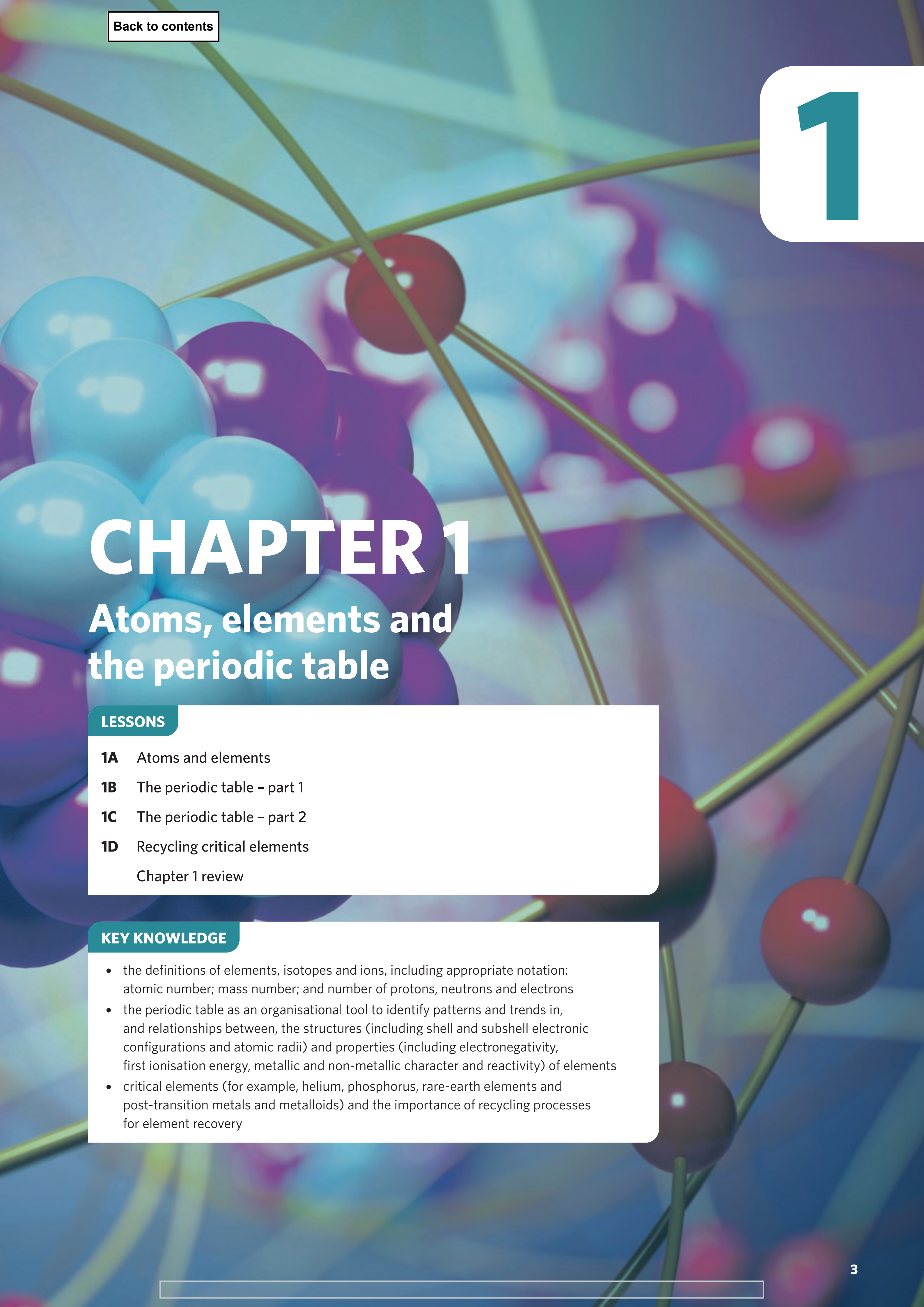

1

CHAPTER 1

Atoms, elements and the periodic table

LESSONS

KEY KNOWLEDGE

- the definitions of elements, isotopes and ions, including appropriate notation: atomic number; mass number; and number of protons, neutrons and electrons
- the periodic table as an organisational tool to identify patterns and trends in, and relationships between, the structures (including shell and subshell electronic configurations and atomic radii) and properties (including electronegativity, first ionisation energy, metallic and non-metallic character and reactivity) of elements
- critical elements (for example, helium, phosphorus, rare-earth elements and post-transition metals and metalloids) and the importance of recycling processes for element recovery

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1A Atoms and elements

STUDY DESIGN DOT POINT

- the definitions of elements, isotopes and ions, including appropriate notation: atomic number; mass number; and number of protons, neutrons and electrons

1A 1B 1C 1D

1.1.1.1 Atomic number

1.1.1.2 Mass number

1.1.1.3 Isotopes

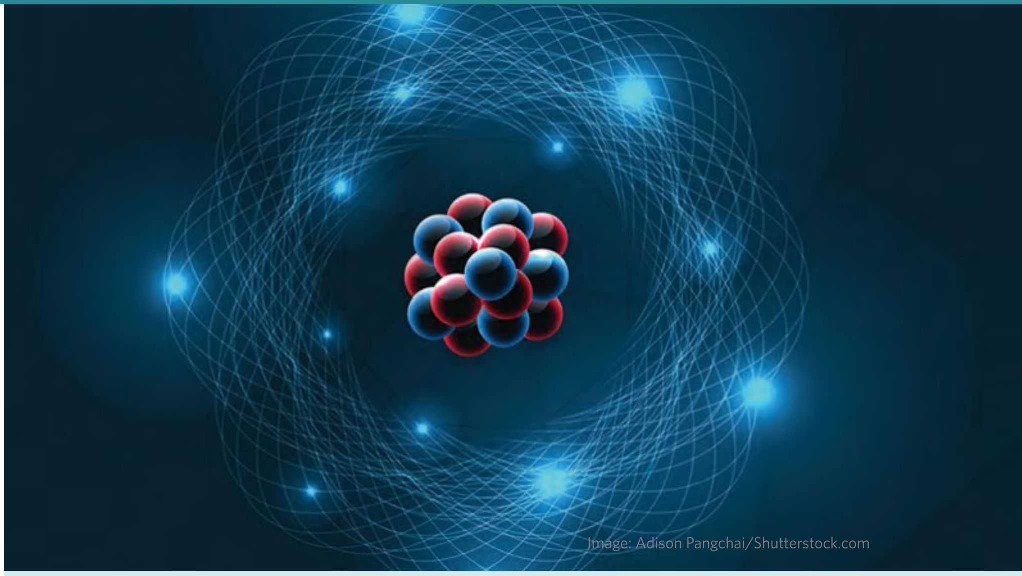
ESSENTIAL PRIOR KNOWLEDGE

- There are many different types of atoms that exist
- Subatomic particles

See questions 1–2.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.



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How much volume of an atom do subatomic particles occupy?

In chemistry **atoms** consist of three fundamental different components – **protons** and **neutrons** which form the **nucleus** and **electrons** which surround the nucleus. In this lesson we'll discuss how to accurately describe the structure of **elements** and learn about the numerical characteristics that define elements.

KEY TERMS AND DEFINITIONS

Atom	smallest unit of matter used in VCE chemistry
Atomic number	number of protons in the nucleus of an atom
Chemical symbol	abbreviation used to represent a chemical element
Compound	two or more atoms of different elements bound together
Electron	negatively charged particle that exists outside the nucleus of an atom
Element	pure substance containing only a single type of atom
Ion	atom that has lost or gained electrons to become a charged particle
Isotope	variants of an element which have the same atomic number but a different number of neutrons in their nuclei
Mass number	mass of an atom, equal to the sum of protons and neutrons
Molecule	when two or more non-metal atoms bond by sharing electrons
Neutron	neutral particle inside the nucleus of an atom
Nucleus	region at the centre of an atom that contains protons and neutrons
Periodic table	table of chemical elements in which elements are arranged in order of increasing atomic number
Proton	positively charged particle inside the nucleus of an atom
Subatomic particle	particle that exists inside the atom

Atomic number 1.1.1.1

All elements have different **atomic numbers**.

How is the atomic number of an element determined?

Atoms consist of **subatomic particles** (protons, neutrons and electrons figure 1) and to date, scientists have discovered 118 different types of atoms, referred to as elements.

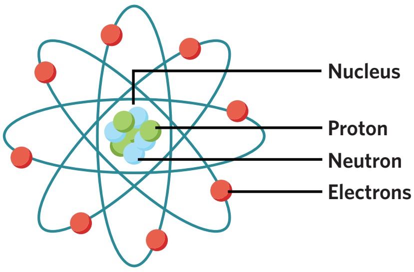

Figure 1 Protons (+) are attracted to the electrons (-) in an atom

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With so many different elements to keep track of, Russian chemist Dimitri Mendeleev developed a way of organising these into a table we now refer to as the periodic table shown in figure 2. We will learn about the **periodic table** in great detail in the next few lessons.

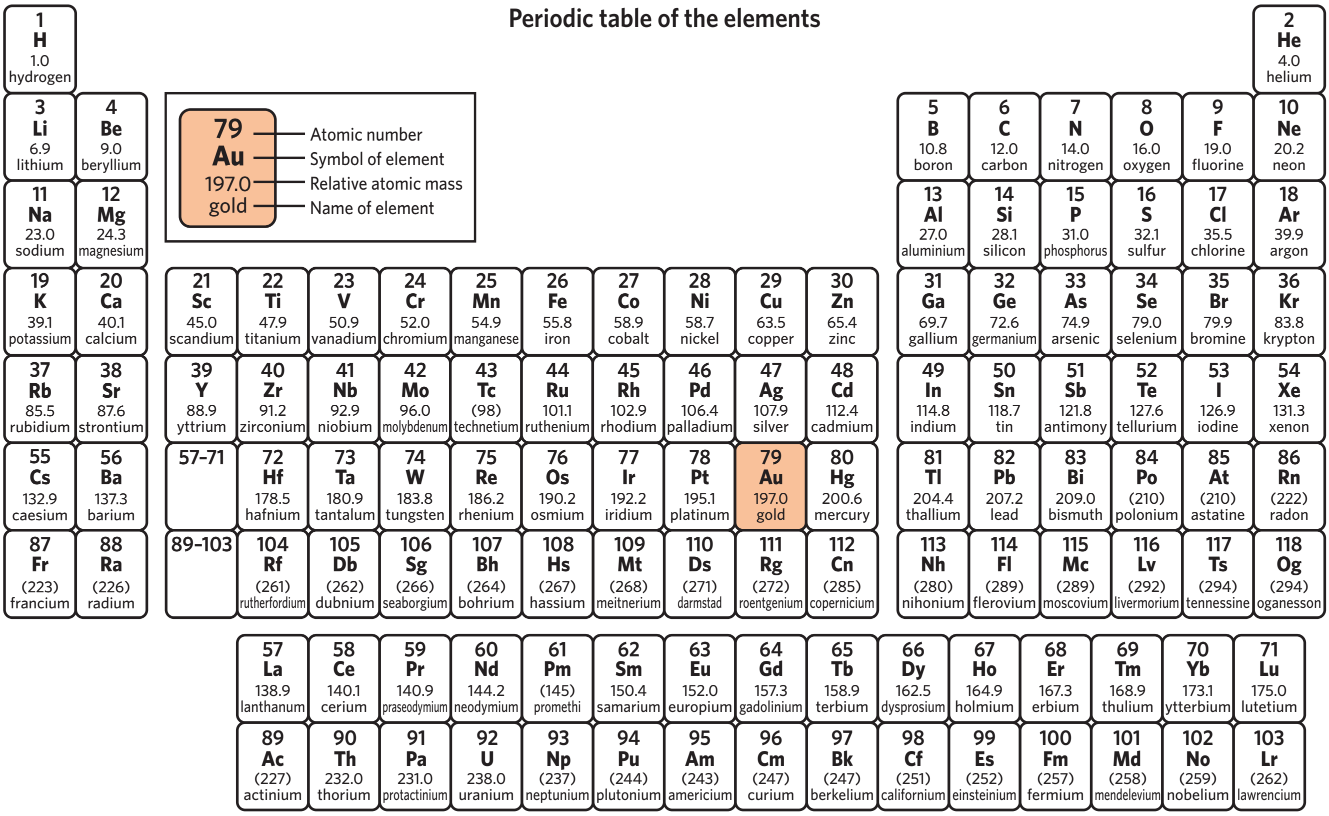

Figure 2 The periodic table

Although we have names for each of the different elements, chemists generally represent each element using its **chemical symbol**. Each box in the table represents a new element, as seen by the different names and symbols. When two or more atoms bond by sharing electrons, they form **molecules** such as O_2 (oxygen). When these atoms come from different elements, such as H_2O (water), they form **compounds**. We will now explore the information displayed for each element on the periodic table in further detail. As shown in figure 3, the number located at the top of each element indicates how many protons are in the nucleus of each atom, and is referred to as the atomic number.

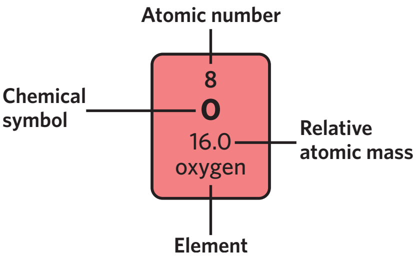

Figure 3 Notation of elements in the periodic table

The atomic number is unique to every element and gives information about the number of protons found in the nucleus of the atom. Looking at the example of oxygen in figure 3, oxygen has an atomic number of 8, and by no coincidence, is also the 8th element on the periodic table. This means that the identity of an element is based on the number of protons it has in its nucleus, and does not depend on the number of neutrons or electrons. When an atom gains or loses electrons it becomes an **ion**. As this change only results in the atom becoming charged, it does not affect the number of either protons or neutrons.

MISCONCEPTION

'Protons, electrons and neutrons are the only subatomic particles.'

Subatomic particles like quarks and muons have been discovered by physicists and at the last count there are a total of 36 elementary and composite subatomic particles.

Progress questions

Question 1

The atomic number describes the number of

A. neutrons in an atom.

B. protons in an atom.

Continues →

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MISCONCEPTION

'Atoms have an overall charge.'

All atoms in the periodic table have the same number of protons (each with a relative charge of +1) and electrons (each with a relative charge of -1) which cancel each other making atoms neutral. Neutrons have no effect on charge as they are neutral. Ions are atoms that have gained or lost electrons and are charged.

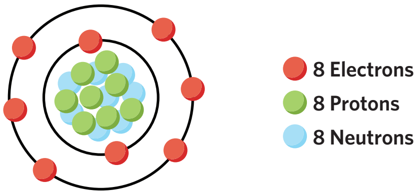

Figure 4 Model of an oxygen atom

USEFUL TIP

The mass number of an element is a relative value compared to other elements and therefore does not have any units. Using the relative masses from the periodic table a carbon atom (12) is 12 times heavier than a hydrogen atom (1) and a magnesium atom (24) is twice as heavy as a carbon atom (12). This will be disscussed in lesson 6A.

KEEN TO INVESTIGATE?

[1] **How do we know what is in the nucleus of an atom?**
Search YouTube: Rutherford's gold foil experiment

Question 2

________ are positively charged and are in the nucleus of an atom whereas ________ are negatively charged and exist outside the nucleus. ________ are neutral particles.

A. Protons, electrons, Neutrons

B. Electrons, protons, Neutrons

Question 3

Which of the following equations is always correct with respect to a neutral atom?

A. Protons = electrons

B. Protons + neutrons = electrons

C. Protons + electrons = neutrons (x2)

D. Electrons + neutrons = neutrons (x2)

Mass number 1.1.1.2

The **mass number** of an element indicates the number of protons and neutrons in the nucleus of an atom.

How do you determine the mass number of an element?

As elements are made up of a different combination of subatomic particles, they all have different masses. The mass number of an element is based on the number of protons and neutrons found in the nucleus and is calculated as a relative mass:

Mass number = number of protons + number of neutrons

If we inspect the model of an oxygen atom in figure 4, we can see that the nucleus contains a total of eight protons and eight neutrons.

To calculate the relative mass number for oxygen, we need to sum the number of protons and neutrons.

Mass number for oxygen = 8 + 8 = 16

Protons have approximately the same mass as a neutron. Electrons have a very small mass (about 1/1800 of a proton or neutron); so small that it's almost negligible when it comes to calculating the mass of an atom. As a result, we don't include electrons in our calculations and almost all the mass of an atom is in the nucleus.[1]

If we combine our knowledge of the atomic number and mass number of an element of an atom, we can determine the number of neutrons found in an atom using the following formula:

Number of neutrons in an atom = mass number − atomic number

WORKED EXAMPLE 1

Consider the element shown. Given that the mass number is 40, how many neutrons would be found in the nucleus of this atom?

20
Ca
40
calcium

What information is presented in the question?

Information from the periodic table for the element calcium.

What is the question asking us to do?

Determine the number of neutrons found in the nucleus of calcium.

What strategies do we need in order to answer the question?

1. Determine the atomic number of calcium.
2. Subtract the atomic number from the mass number.

Answer

The atomic number of calcium is 20. Given that it has a mass number of 40, the number of neutrons would be:

Number of neutrons in an atom = mass number − atomic number

Neutrons = 40 − 20 Neutrons = 20

There would be 20 neutrons in a calcium atom.

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Progress questions

Question 4

The relative mass number is determined by

A. multiplying the mass of a proton by the number of protons.

B. adding the total number of neutrons and protons in the nucleus of an atom.

Question 5

A neutral atom with 10 electrons would have

A. 10 neutrons.

B. 10 protons.

Question 6

A neutral atom with 10 protons would always have

A. 10 neutrons.

B. an atomic number of 10.

Question 7

Elements on the periodic table are numbered in increasing order of the number of

A. protons.

B. electrons.

Question 8

The mass of a proton is approximately the same amount as

A. a neutron.

B. an electron.

Isotopes 1.1.1.3

The same element can have different mass numbers.

How many elements have multiple isotopes?

Individual elements can exist in different forms known as **isotopes**. Isotopes of an element contain the same number of protons (and therefore have the same atomic number), however, they vary in the number of neutrons and therefore have a different mass (see figure 5).

Considering the mass number is the total number of protons and neutrons in the atom, the mass numbers of protium, deuterium and tritium would be 1, 2 and 3 respectively. The mass number is shown as a superscript to the upper left hand side of the chemical symbol. This value can be used to determine the identity of the isotope as shown in figure 6.

Of the first 82 elements in the periodic table, 71 of them are considered to have at least one stable isotope (not radioactive[2]) in nature. Tin is the winner with 10 stable isotopes ranging in mass from 112 to 124.

STRATEGY

The convention used to represent isotopes is different to what appears in the VCE Data Book. Mass and atomic number are swapped around.

MISCONCEPTION

'The atomic masses that appear on the periodic table are actual masses of the atoms.'

The relative atomic masses shown on the periodic table are the relative average mass of all of the isotopes of that element. This will be covered in more detail in lesson 6A.

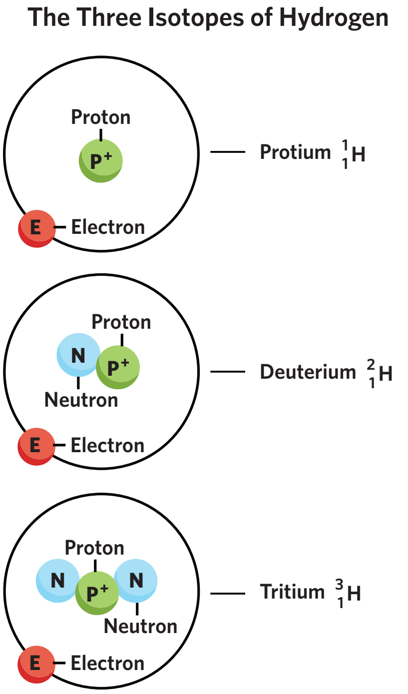

Figure 5 Different isotopes of hydrogen

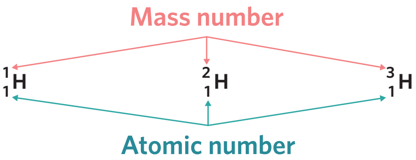

Figure 6 Conventions used to represent isotopes

KEEN TO INVESTIGATE?

[2] **Why are some isotopes radioactive?** Search YouTube: What are radioactive isotopes?

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Progress questions

Question 9

The number of neutrons in an atom with an atomic number of 11 and a mass number of 23 is

A. 12.

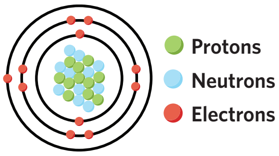
B. 34.

Question 10

The mass number of the following atom is

A. 22.

B. 33.



Question 11

Isotopes have

A. atoms with a different number of protons but the same number of neutrons.

B. atoms with a different number of neutrons but the same number of protons.

Question 12

Isotopes of an element all have the same physical properties.

A. True

B. False

Theory summary

- The atomic number reflects the number of protons in the nucleus of an atom.
- Different elements have different atomic numbers.
- The mass number is the sum of neutrons and protons in the nucleus of an atom.
- Isotopes have the same number of protons but a different number of neutrons (and therefore a different mass).

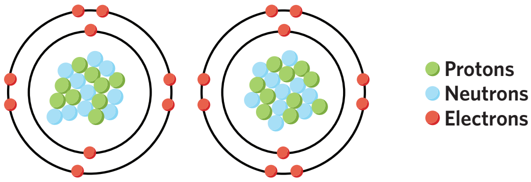
1A Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 13–15.

Atoms are the most basic building blocks of living things. In order to be able to produce different kinds of substances, we require a whole range of different types of atoms. Take for example the following two atoms:



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Question 13 (1 MARK)

These atoms

A. are isotopes of each other.

B. have the same number of neutrons.

C. have the same number of protons.

D. are radioactive.

Question 14 (1 MARK)

The atomic number of each atom is/are

A. 19 and 20 respectively.

B. 10.

C. 9 and 10 respectively.

D. 7 and 8 respectively.

Question 15 (3 MARKS)

Describe the difference in mass number between the two atoms.

Exam-style

Question 16 (1 MARK)

Dorothy is a nuclear physicist who devotes much of her time to her love of isotopes. One afternoon she decides that she wishes to explore the potential of artificially synthesising some zinc isotopes, and carries out some inquiries. Identify the key characteristic of zinc that she would change in order to develop different isotopes.

Question 17 (3 MARKS)

When an atom becomes radioactive, it can undergo a process called neutron release. During neutron release, a neutron is effectively 'ejected' from the nucleus.

Consider an atom with the nucleus shown.

Assuming all of the protons and neutrons are visible, describe what would happen to the atom if it were to undergo neutron release.

Question 18 (13 MARKS)

Alicia conducted an experiment to detect subatomic particles in newly discovered atoms. Her observations are given in the table below:

Fictitious elemental symbol	Protons	Neutrons	Electrons	Mass number
R	17	17	17	
I	18	18	18	
T	18	19	18	
A	19	20	19	

a. Determine the mass number of each element. (4 MARKS)

b. Which of the two elements are isotopes of each other? (1 MARK)

c. How should the table be updated to reflect your answer in part b? (2 MARKS)

d. What is the complete notation including the symbol, atomic number and mass number for element A? (3 MARKS)

e. If elements R and A came together to create RA, would RA be classified as a molecule or a compound? (3 MARKS)

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Key science skills

Question 19 (5 MARKS) 🌶🌶🌶

Matthew, a year 11 chemistry student, wanted to compare two unknown elements. He believed that by developing an experiment which could measure and compare the mass number of two elements, he would be able to distinguish the two elements from one another. In preparation for his experiment, Matthew developed a set of scales that could detect extremely small masses. Prior to conducting the experiments, he made sure to calibrate the scales.

a. Why is it important for Matthew to calibrate the scales prior to an experiment? (1 MARK)

b. Consider the following statement from Matthew:

"Every element is different, and therefore will have a different mass number to other elements."

Comment on the accuracy of his statement. (4 MARKS)

FROM LESSON 16D

Hints

16.	Isotopes have a varying number of neutrons.
17.	Isotopes have a varying number of neutrons.
18a.	Calculating mass number involves only the protons and neutrons.
18b.	Isotopes are derived from the same element.
18c.	Isotopes have the same atomic number.
18d.	The representation involves only symbols and numbers.
18e.	Compounds involve two or more different elements.
19a.	Errors can affect experimental results.
19b.	Isotopes have different masses.

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1B The periodic table – part 1

Image: Pavel Sapozhnikov/Shutterstock.com

Who created the periodic table and why did they leave gaps?

In this lesson, we will learn how an element's **electron configuration** is written, how elements are organised in the periodic table, and the periodic trend in the size of atoms.

STUDY DESIGN DOT POINT

- the periodic table as an organisational tool to identify patterns and trends in, and relationships between, the structures (including shell and subshell electronic configurations and atomic radii) and properties (including electronegativity, first ionisation energy, metallic and non-metallic character and reactivity) of elements

1A 1B 1C 1D

1.1.2.1 Electronic configurations

1.1.2.2 & 1.1.2.4 The basic organisation of the periodic table & atomic radii

1.1.2.3 Blocks on the periodic table

ESSENTIAL PRIOR KNOWLEDGE

1A Subatomic particles

1A An element

See questions 3–4.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

KEY TERMS AND DEFINITIONS

Aufbau principle rule that states subshells are filled by electrons from the lowest to the highest energy level

Blocks assortment of elements into discrete categories based on which subshell their valence electrons are found in

Electron configuration arrangement of electrons in shells and/or subshells

Energy shells or energy levels orbits containing different levels of energy, around the nucleus of an atom where electrons are found according to the Bohr and Schrodinger models

Groups columns in the periodic table

Orbitals regions with the highest probability of finding electrons

Pauli exclusion principle rule that states an orbital can't hold more than 2 electrons

Periods rows in the periodic table

Periodicity characteristics of elements in a period

Valence shell outermost energy shell

Valence electrons electrons in the outermost shell of an atom

Electronic configurations 1.1.2.1

We can use the Bohr and Schrodinger models of the atom to write the electron configurations of the shells and subshells of different atoms.

How are electron shells configured around the nucleus of an atom?

The electron configurations show information about the number of electrons in each atom and where they are located in terms of electron shells (table 1).

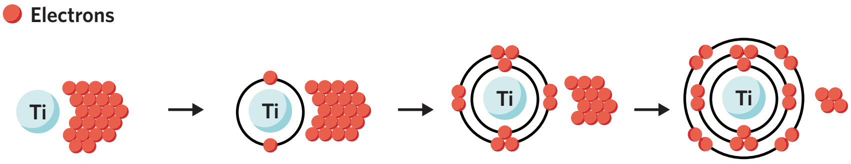
In the Bohr model, when electrons occupy shells in an atom, there are three general rules that need to be followed:

1. Each electron shell holds a different number of electrons.
2. Electrons will fill shells closest to the nucleus first.
3. Electrons fill and empty shells in a particular order.

Table 1 Difference between shells in the Bohr model

Distance from nucleus	Shell number	Maximum number of electrons in the shell
Closest	1	2
↓	2	8
↓	3	18
Farthest	4	32

Figure 1 shows how electrons fill in an atom of titanium, starting with the inner shell.



Electrons found in titanium.

Two electrons fill the first shell.

Eight electrons fill the second shell.

Electrons begin to fill the third shell as the first two shells are full.

Figure 1 Order of filling of electrons in an atom

WORKED EXAMPLE 1

Scientists discovered an atom with a total of 28 electrons. What is the electron shell configuration for this atom?

What information is presented in the question?

The atom contains 28 electrons.

What is the question asking us to do?

Write the electron configuration for the atom.

What strategies do we need in order to answer the question?

1. Recall the maximum number of electrons that can fill each shell.
2. Fill each shell with electrons until it is full.
3. Write the electron configuration for the atom.

Answer

We know that the capacity of each shell is:

Shell number	Max number of electrons it can hold
1	2
2	8
3	18
4	32

However, once the third shell reaches eight electrons, it will fill the fourth shell with two electrons first before electrons will start to fill shell number three again.

According to the question, there are 28 electrons in total. Therefore the order of filling would be:

Order of filling	Number of electrons filling the shell	Total electrons placed in a shell
First shell	2	2
Second shell	8	10
Third shell	8	18
Fourth shell	2	20
Third shell	8	28

We can see that the third shell is filled twice, with a total of 16 electrons. Therefore the electronic configuration for this atom is 2, 8, 16, 2.

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How are electrons configured within shells and subshells?

The Schrödinger model of the atom is a more accurate and complex model than the Bohr shell model. Instead of electrons occupying fixed orbits, electrons occupy regions around the nucleus. Schrödinger's model of the atom takes this principle into account and defines **orbitals** as regions of space where electrons are likely to be found.

One of the key rules about orbitals is the **Pauli exclusion principle** which states that the maximum number of electrons in an orbital is 2. Therefore, it could hold 0, 1 or 2 electrons. However, similar to the Bohr model, electrons are said to be found in shells which vary in **energy level**. These shells are numbered starting at 1, and also reflect the order of energy levels. Shells contain smaller subshells, however, not all shells contain the same type of subshells, as seen in table 2 and figure 2.

Table 2 The different shells, subshells and orbitals in an atom

Number of shell	Subshell(s) in shell	Number of orbitals in subshell	Maximum number of electrons
1	1s	1	2
2	2s	1	2
	2p	3	6
3	3s	1	2
	3p	3	6
	3d	5	10
4	4s	1	2
	4p	3	6
	4d	5	10
	4f	7	14

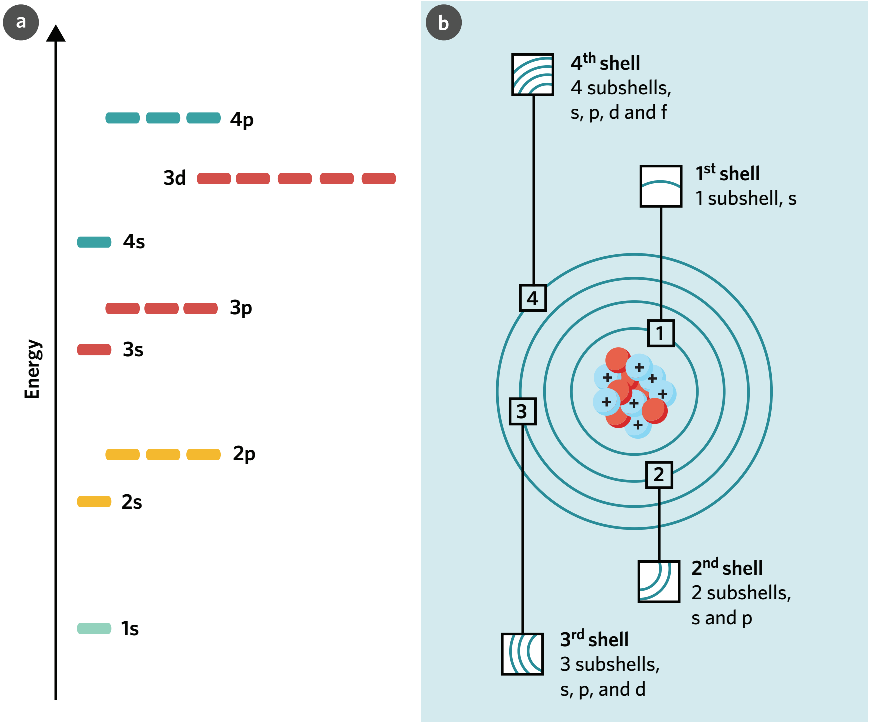

Figure 2 Subshells and shells around the nucleus (a) and by increasing energy (b)

USEFUL TIP

When electrons are filling either the p, d or f subshells they generally spread out when they can so that only 1 electron is in each orbital due to the fact that electrons repel each other. When electrons have to pair up in a subshell they do so by spinning in opposite directions, causing a slight attraction, thus making an electron pair.

How do we write electron configurations using shells and subshells?

We can also use the Schrodinger model of the atom to write the electron configurations of different elements by counting how many electrons are in each subshell. In order to do this, we need to understand the **Aufbau principle**, which states that electrons fill the lowest **energy shells or energy levels** first. Based on experiments, it's been shown that the energy levels found in atoms can be represented in figure 3.

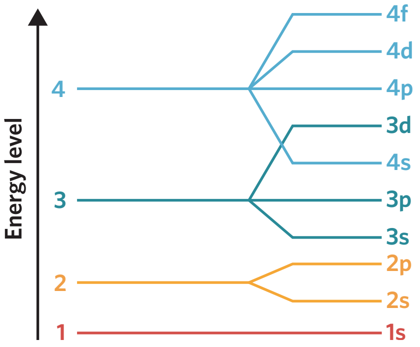

Figure 3 Different energy levels found in atoms

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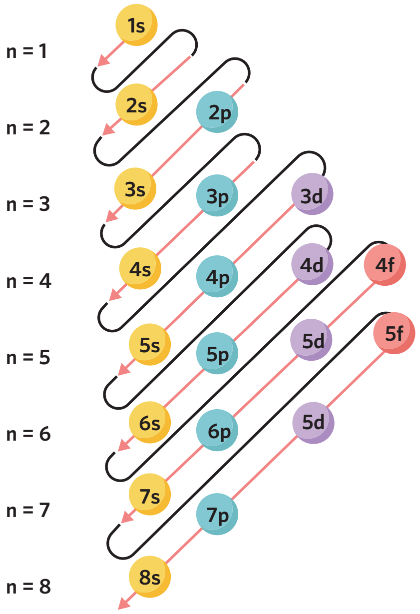

Figure 4 The electron filling pattern of subshell energy levels

While each energy level represents an energy shell, we can see that the order does not necessarily show that electrons fill shells completely before moving onto the next. Electrons prefer to fill subshells with lower energy first, as they are more stable. As such, we can see from figure 4 that electrons will begin to fill subshell 1s first, followed by 2s, 2p and so on. However, when electrons have filled subshell 3p, they then start to fill 4s before 3d. This is due to the fact that shell 4s has lower energy than 3d. This whole process can be represented by the geometric design in figure 4.

USEFUL TIP

Sometimes the subshell energy levels are not in sequential order. For example, the 4s subshell is lower in energy than the 3d subshell. It takes less energy to fill the one orbital in the 4s subshell than the five orbitals in the 3d subshell.

The standard notation when writing electron configurations is to write the shell number first, the type of subshell second and then in superscript above the subshell, the number of electrons. Consider an atom of sulfur, which contains a total of 16 electrons. We know that electrons fill subshells as outlined in figure 5. We also need to remember that the s subshell is able to hold two electrons, the p subshell is able to hold six electrons and the d subshell is able to hold 10 electrons. Every subshell needs to be filled with electrons before moving onto the next subshell. Using this information, we can determine the electron configuration for sulfur to be:

$$1s^2 2s^2 2p^6 3s^2 3p^4$$

Shell Number of electrons Subshell

Based on the electron configuration, we can see that sulfur contains a total of three shells.

WORKED EXAMPLE 2

Write out the complete electron configuration for an atom of oxygen which has 8 electrons.

What information is presented in the question?

The atom in question is oxygen which has 8 electrons.

What is the question asking us to do?

Find the complete electron configuration of oxygen by using the Schrödinger model.

What strategies do we need in order to answer the question?

1. Write out the order of the energy levels of shells and subshells.
2. Fill the subshells with the maximum number of electrons they can hold in this order, remembering the s subshell can hold two electrons, the p subshell can hold six electrons and the d subshell can hold 10 electrons.

Answer

There are eight electrons in an oxygen atom.

The order of energy subshells is:

$1s<2s<2p<3s\ldots$
The s–subshell can hold two electrons and the p–subshell can hold six electrons.

Therefore, there will be two electrons in the 1s subshell, two electrons in the 2s subshell and the remaining four electrons in the p–subshell. Therefore, the electron configuration of oxygen is: $1s^2 2s^2 2p^4$

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Progress questions

Question 1

An atom that contains 11 electrons would contain

A. 3 energy shells.

B. 11 energy shells.

Question 2

Orbitals are

A. mathematical visualisations of regions where electrons are likely to be found.

B. real and measurable bubbles around the nucleus.

Question 3

What shell and subshell will the 28th electron of an atom fill?

A. 4p

B. 3d

Question 4

Which of the following correctly shows the electronic configuration of a sodium atom?

A. $1s^2 2s^2 3p^6 4s^1$

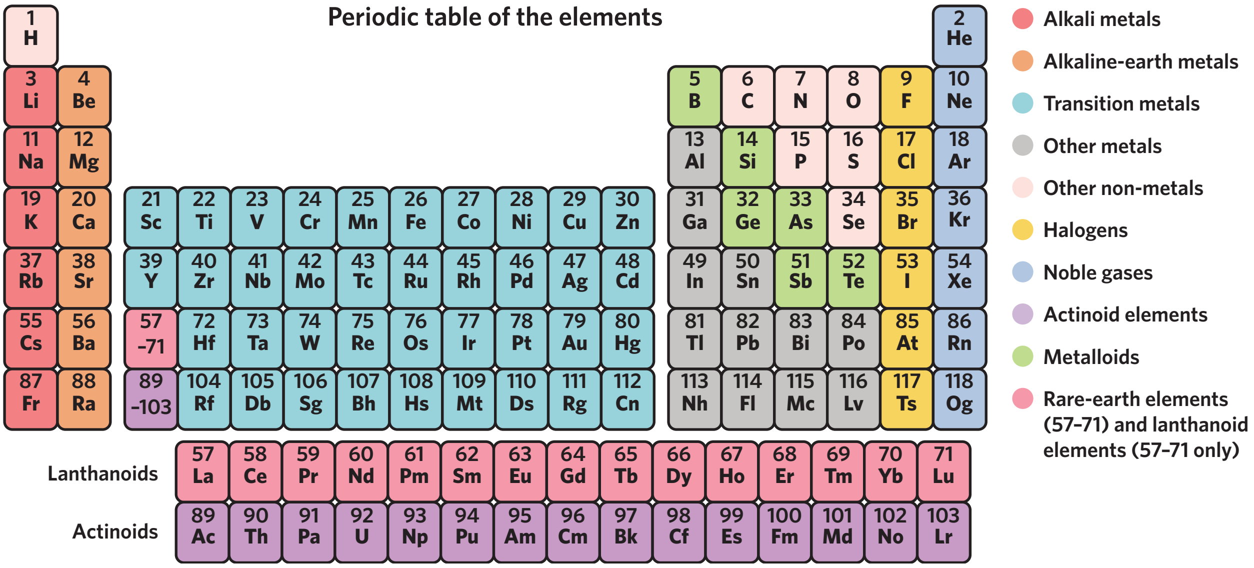
B. $1s^2 2s^2 2p^6 3s^1$

The basic organisation of the periodic table & atomic radii 1.1.2.2 & 1.2.2.4

The periodic table is arranged in a systematic manner.

How is the periodic table organised?

As we explored in the previous lesson, elements behave differently, and the periodic table helps scientists represent all the different elements in a way that demonstrates their properties and trends. A modern version of the periodic table is shown in figure 5.



Periodic table of the elements

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 -71	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 -103	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og
Lanthanoids			57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinoids			89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

- ● Alkali metals
- ● Alkaline-earth metals
- ● Transition metals
- ● Other metals
- ● Other non-metals
- ● Halogens
- ● Noble gases
- ● Actinoid elements
- ● Metalloids
- ● Rare-earth elements (57–71) and lanthanoid elements (57–71 only)

Figure 5 A modern periodic table

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From a broad perspective, we can see that the colours in the periodic table highlight the many different types of elements that exist, ranging from metals to non-metals. Within these two categories, the elements can be further classified into sub–categories.

How can a period explain trends in the periodic table?

At first glance, we can notice that elements are arranged by their atomic number, which increases from left to right across rows. As we learned in the previous lesson, the atomic number reflects the number of protons in the nucleus of an atom, therefore we can appreciate that as we progress from left to right on the periodic table, the number of protons increases.

Rows in the periodic table are referred to as **periods**. Elements in the same period show **periodicity**;[1] a repeating trend in properties, including:

- same number of occupied electron shells, and
- decreasing atomic radius.

A special characteristic of periods on the periodic table is that elements in the same period contain the same number of occupied electron shells as seen in figure 6.

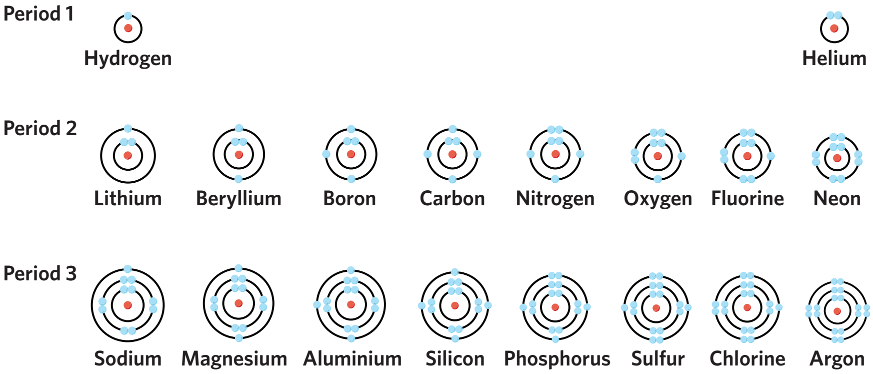

Figure 6 The electron arrangement in different elements

If we look carefully at figure 6, we can also see that although elements in the same period contain the same number of electron shells, the radius of the atom actually becomes smaller.

The atomic radius of an atom is estimated (due to the fact that an atom does not have a defined boundary) from the nucleus to the outer shell of an atom as shown in figure 7. The outermost shell of an atom is known as the **valence shell**. Electrons that are found in the valence shell are known as **valence electrons**. Valence electrons are an important contributor to the behaviour of an atom; a concept that we will investigate further.

The number of protons in an atom increases with atomic number across a period. Given there will be more protons in the nucleus, there would be a greater positive charge in the nucleus. As each period contains the same number of occupied shells, atoms towards the end of the period would have a greater ability to attract the valence electrons towards the nucleus compared to elements at the start of the period, due to the increase in number of protons (and therefore positive charge). This would result in the valence shell being pulled closer to the nucleus, reducing the atomic radius. This is summarised in figure 8.

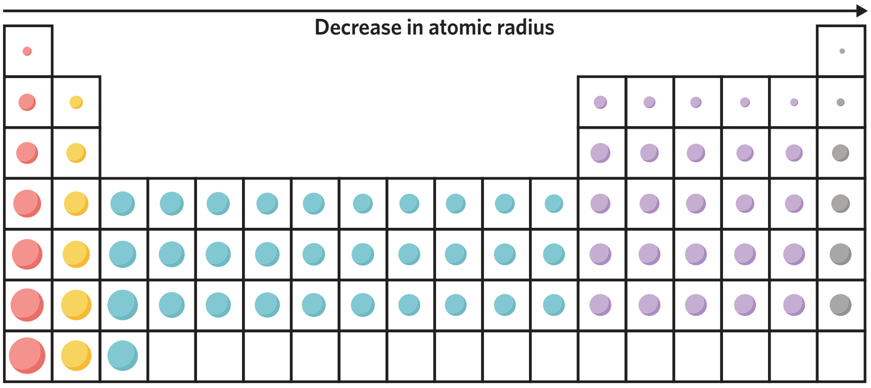

Figure 8 Atomic radius decreases across a period

KEEN TO INVESTIGATE?

[1] **What other blocks of elements are there?**
Search: interactive periodic table rs

MISCONCEPTION

'Atoms get bigger as you add more electrons.'

Moving from left to right across the period, there is an increase in the number of protons and therefore also electrons. However, if we investigate further, we will realise that the increase in protons is exactly the reason why the size of the atom decreases as we move along a period because the attraction of the protons to the valence electrons is stronger.

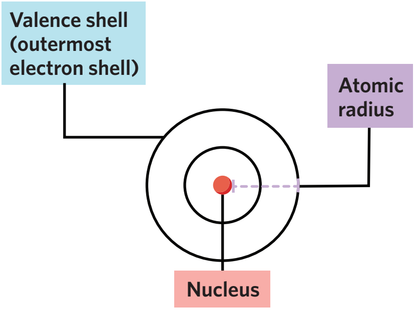

Figure 7 The method of measuring atomic radius

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However, we need to remember that in–between the valence shell and the nucleus are inner shells of occupied electrons.. As we know, electrons are negatively charged particles, and therefore would repel each other when they come into close contact, as shown in figure 9.

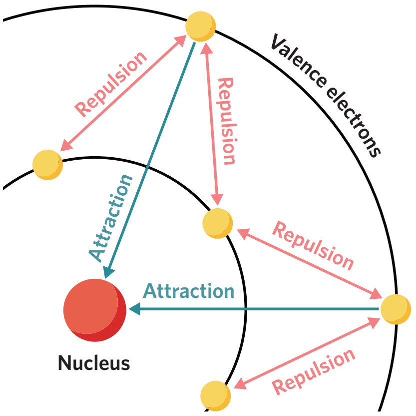

Figure 9 Repulsion of electrons in an atom

So far, we have been investigating the periodic table by looking at the properties of elements in the same period (row), however, we can also look at the table from a different perspective: investigating elements in the same column.

How can a group explain trends in the periodic table?

The columns of a periodic table are known as groups. **Groups** are numbered from left to right from 1 to 18, where each different column represents a different group. Consider figure 10, which shows the first few elements in a variety of different groups.

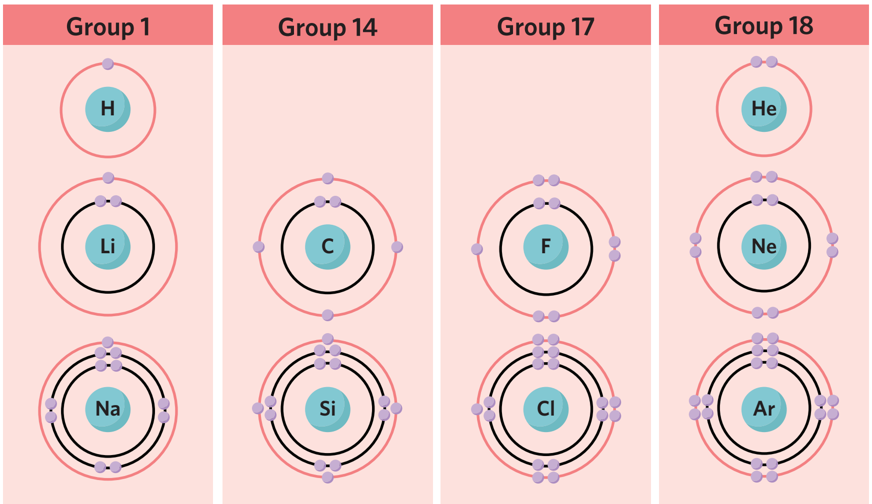

Figure 10 Electron arrangement of elements in the same group

If we look closely, we can notice that the valence shell in each element (in red) of each group contains the same number of valence electrons. Hydrogen, lithium and sodium, all found in group 1, all contain one valence electron. The elements in group 14 all contain four valence electrons. This pattern is evident throughout all groups except helium in group 18. Due to the fact that elements in the same group contain the same number of valence electrons, they all behave similarly.

Although elements in the same group have the same number of valence electrons, they differ in the total number of occupied electron shells. As we move down a group, we can see that there is an additional electron shell. This makes sense considering when we move down a group, we are essentially entering a new period. As such, we experience the addition of another occupied electron shell.

The increase in the number of electron shells as we progress down a group results in an increase in atomic radius as seen in figure 11. As we know, this means that there is an increase in the distance between the nucleus and the valence shell. Consequently, the valence electrons experience a weaker force of attraction towards the nucleus. Again, the electrons found in between the valence shell and nucleus also have a shielding effect which contributes to the atomic radius increasing down a group.

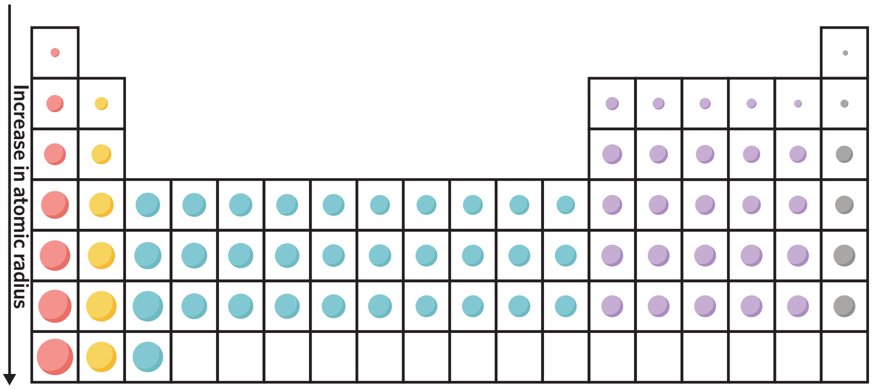

Figure 11 Atomic radius increases down a group

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Progress questions

Question 5

Elements with the same number of occupied energy shells are found in the same

A. group.

B. period.

Question 6

Elements with three valence electrons can be found in the same

A. group.

B. period.

Question 7

Potassium can be found in

A. group 1, period 4.

B. group 4, period 1.

Question 8

The number of protons ________ as we move across a period.

A. decreases

B. increases

Question 9

The greater the number of electrons the larger the atom.

A. True

B. False

Question 10

Atomic radius increases across a period.

A. True

B. False

Question 11

Atomic radius increases as the number of protons increases.

A. True

B. False

Blocks on the periodic table 1.1.2.3

Elements can be grouped based on their highest occupied energy subshell.

How can the periodic table be divided into blocks?

As we learned earlier in this lesson, the electron configuration of atoms involves a system whereby electron shells are divided up into discrete energy levels – called subshells – from the lowest energy level to the highest energy level. Combined with our understanding of the information reflected by the periods of the periodic table, we can now appreciate the organisation of elements into **blocks** as shown in figure 12.

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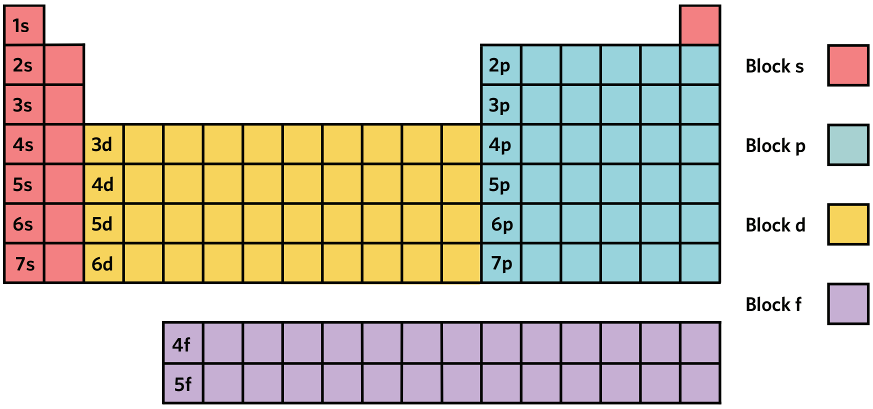


Figure 12 Elemental blocks in the periodic table

The periodic table is also organised into different categories called blocks. Specifically, elements are placed into blocks according to which subshell their highest energy, valence electrons fill. As we can see in figure 12, there are four different blocks, each corresponding to a subshell: s, p, d or f. Elements that fall within the s–block will have the s subshell as its outermost occupied subshell. For example, lithium, which is found in period two, falls within the s–block and has a 2s subshell as its highest energy subshell. This pattern is the same for all of the other blocks.

USEFUL TIP

The best way to identify an atom based on its electron configuration is to simply add up all of the electrons, because this is equal to the atomic (proton) number of the element.

WORKED EXAMPLE 3

In which block would carbon be found on the periodic table? Include the electron configuration of carbon to support your answer.

What information is presented in the question?

The element in question is carbon.

What is the question asking us to do?

Find the block in which carbon is found in the periodic table.

What strategies do we need in order to answer the question?

1. Write the electron configuration for carbon.
2. Identify the last subshell in carbon as shown in the electron configuration.

Answer

As a neutral atom, carbon contains a total of six electrons. Therefore the electron configuration for carbon is $1s^2 2s^2 2p^2$. As a result, carbon would be found in the p-block as its outermost electrons are in a p-subshell.

Progress questions

Question 12

Elements in the s block

A. have the same number of valence electrons.

B. have the same outer subshell type.

Question 13

Elements in the same period belong in the same block.

A. True

B. False

Question 14

An element found in the p block

A. will have the same number of energy shells.

B. will have a maximum of six electrons occupying their outer subshell.

Continues →

Question 15

Oxygen can be found in block

A. s.

B. p.

Theory summary

- Schrödinger's subshell model of the atom replaced Bohr's shell model and describes electrons inhabiting orbitals as well as the shells and subshells of atoms.
- Groups (columns) of elements in the periodic table all have the same number of outer valence electrons.
- Periods (rows) of elements in the periodic table all have the same number of occupied energy shells.
- Atomic radius decreases across a period and increases down a group.
- There are four blocks on the periodic table: s, p, d and f.
- Blocks on the periodic table reflect the highest energy subshell found in the atoms of an element.

1B Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 16-18.

Selenium is an element found in trace amounts in everyday foods. A sample of selenium is examined to determine its chemical properties. A second element is studied, sulfur, which exists abundantly on Earth as a yellow crystalline solid.

Images (left to right): Bjoern Wylezich, Sebastian Janicki/Shutterstock.com

Question 16 (1 MARK) 🌶

Sulfur is found in group

A. 13.

B. 14.

C. 15.

D. 16.

Question 17 (1 MARK) 🌶

In comparison to sulfur, selenium has

A. the same number of occupied electron shells.

B. a smaller number of protons in the nucleus.

C. less valence electrons.

D. a larger atomic radius.

Question 18 (2 MARKS)

What is the subshell electron configuration of sulfur?

Exam-style

Question 19 (1 MARK)

The maximum number of electrons that can occupy a 4p subshell is

A. 2

B. 6

C. 10

D. 14

VCAA 2002 Exam 2 Multiple choice Q2

Question 20 (3 MARKS)

The periodic table acts like a list that shows all of the elements that currently exist.

Consider the following elements:

Ca V P Si Be

a. Which elements would be found in the same period? (1 MARK)

b. Which elements would be found in the same block? (1 MARK)

c. Which element would most likely have the greatest atomic radius? (1 MARK)

Question 21 (5 MARKS)

Two elements were selected at random, both belonging to period 4. Element A contains 32 electrons and element B contains 20 electrons.

a. If both atoms were neutral, identify the number of protons in each element. (1 MARK)

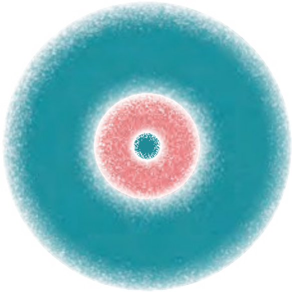
b. Write the subshell electron configuration for each element. (1 MARK)

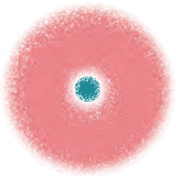
c. Of the two elements, which would have a smaller atomic radius? Explain. (3 MARKS)

Key science skills

Question 22 (5 MARKS)

Scientists have developed a high-powered electron microscope that is able to visualise atoms, however not the way we're used to seeing it in textbooks.





To test the machine, a scientist obtained two elements, lithium and sodium, to be viewed under the microscope. The scientist wanted to see whether the clarity of the image would increase if the atom was left under the electron microscope for a longer period of time.

a. Identify the dependent variable. (1 MARK)

b. The results from the test can be seen below:

Test 1

	1 min	**5 mins**	**10 mins**
Clarity	Not very clear	Clear	Semi–clear

Test 2

	1 min	**5 mins**	**10 mins**
Clarity	Clear	Clear	Not very clear

i. What kind of data was collected in the experiment? (1 MARK)

ii. Comment on the reproducibility of the results. (3 MARKS)

FROM LESSONS 16A, 16B & 16D

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Questions from multiple lessons

Question 23 (4 MARKS)

The discovery of element 113 was claimed by teams of Russian and American scientists in February 2004. It has since been given the name nihonium and the symbol Nh. Nh undergoes rapid radioactive decay but atoms of Nh have been identified with a mass number of 283 and also with a mass number of 284.

a. State the number of subatomic particles in an uncharged Nh atom of mass number 284.
protons ___________ electrons ___________ neutrons ___________ (1 MARK)

b. In what group and period is Nh located in the periodic table?
group ___________ period ___________ (1 MARK)

c. Give the symbol of the element that is expected to be most similar to Nh in chemical properties. (1 MARK)

d. In terms of atomic structure, explain why the atomic radius of Nh is predicted to be smaller than that of Francium. (1 MARK)

Adapted from VCAA 2007 Exam 2 Short answer Q2b

FROM LESSON 1A

Hints

19. The maximum number of electrons is fixed in each subshell.
20a. Periods are viewed as rows in the periodic table.
20b. Blocks are viewed as sections in the periodic table according to valence electrons.
20c. The atomic radius increases down a group.
21a. Neutral atoms have no overall charge.
21b. Electron configurations show the electron arrangement of elements.
21c. Atomic size decreases across a period.
22a. The independent variable is changed by the experimenter.
22bi. Data can be categorised based on the information collected.
22bii. Reproducibility relates to the consistency of experimental data between experiments.
23a. mass number = protons + neutrons
23b. Valence electrons determine the group and number of shells determines the period.
23c. Elements are grouped together for a specific reason.
23d. Atomic radius changes across and down the periodic table.

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1C The periodic table – part 2

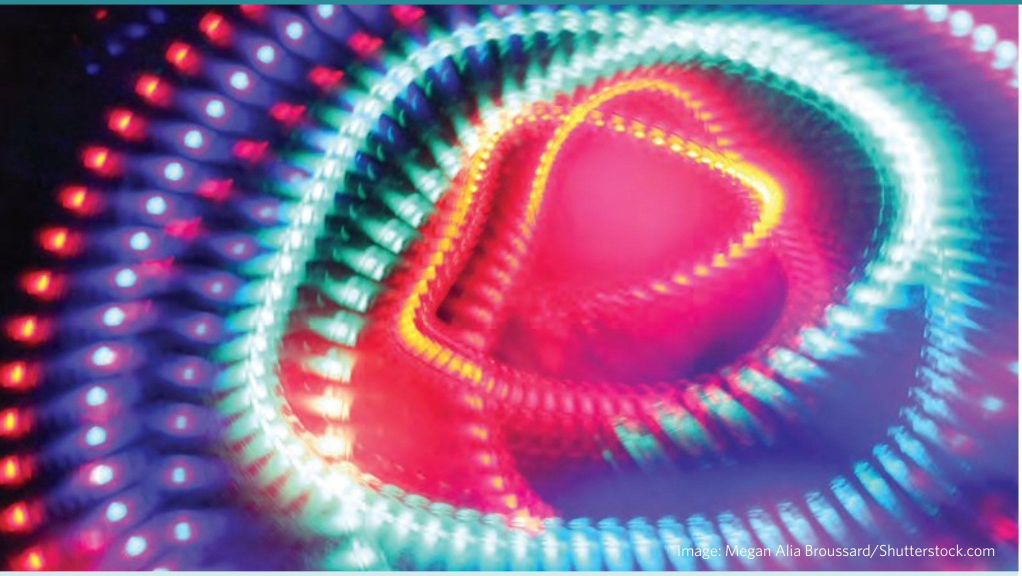
Image: Megan Alia Broussard/Shutterstock.com

How are some trends of elements on the periodic table related to the movement of electrons?

In this lesson, we will learn about more trends in the periodic table. More specifically, with regards to their **electronegativity**, ionisation energy, **metallic character** and **reactivity.**

KEY TERMS AND DEFINITIONS

Core charge	attractive force between the positive nucleus and valence electrons, measured as the difference in the number of protons and inner electrons
Electronegativity	how strongly an atom attracts electrons towards itself
First ionisation energy	energy required to remove the first valence electron from an atom
Ion	atom that has lost or gained electrons to become a charged particle
Metallic character	chemical properties associated with metals
Metalloids	elements which possess properties that are in-between those of metals and non-metals
Noble gases	unreactive gases found in group 18 of the periodic table
Octet rule	general principle which states that atoms like to have 8 electrons in their valence shell for stability
Reactivity	tendency of atoms to undergo a chemical reaction

Electronegativity 1.1.2.5

Elements attract valence electrons differently.

Why do some elements attract electrons more strongly than others?

Different atoms have the ability to attract electrons differently depending on the number of protons in their nucleus and their atomic radius. The ability of an atom to attract electrons to itself is known as electronegativity. The greater the electronegativity, the greater the ability of the atom to attract electrons towards its nucleus. This will be an important concept when we explore chemical bonding later in the book.

STUDY DESIGN DOT POINT

- the periodic table as an organisational tool to identify patterns and trends in, and relationships between, the structures (including shell and subshell electronic configurations and atomic radii) and properties (including electronegativity, first ionisation energy, metallic and non-metallic character and reactivity) of elements

1A 1B **1C** 1D

1.1.2.5 Electronegativity
1.1.2.6 First ionisation energy
1.1.2.7 Metallic character
1.1.2.8 Reactivity

ESSENTIAL PRIOR KNOWLEDGE

1B Valence electrons
1B Atomic radius
See questions 5–6.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

MISCONCEPTION

'Helium is the most electronegative element as it is in the top right hand corner of the periodic table.'

Group 18 elements – known as the noble gases – already have a full valence shell and therefore, according to the **octet rule**, do not want to attract any more electrons. Hence, they are not considered to be electronegative.

The quantitative measurement of electronegativity is called the Pauling Scale, where the greater the number, the more electronegative the element. As we can see in figure 1, fluorine has the highest number on the Pauling Scale (4.0), and therefore is the most electronegative element on the periodic table. Caesium, on the other hand, has the lowest Pauling value (0.7), and is therefore the least electronegative element on the periodic table. For this study design, we are not required to know the Pauling Scale in any detail.

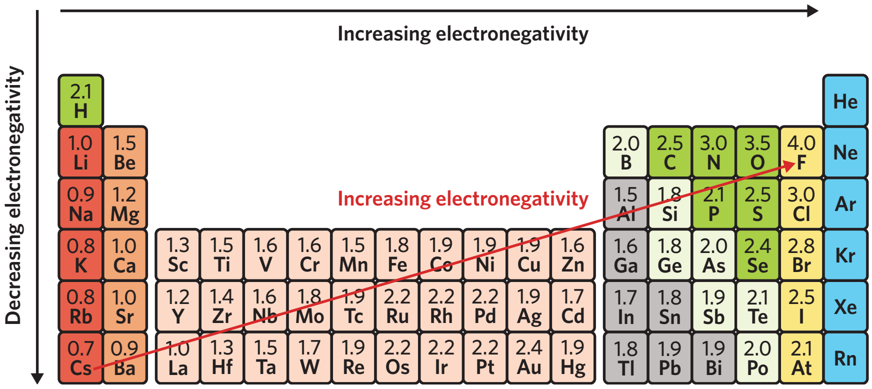

Figure 1 Trends in electronegativity based on the Pauling Scale

The trend shown by the red arrow in figure 1 aligns with our understanding of the number of protons in the nucleus and atomic radius of elements in the periodic table. Elements with a greater number of protons in the nucleus and smaller atomic radius can be found towards the end of each period and further up in a group. As such, the trend in electronegativity follows the same pattern, where the most electronegative elements are found towards the upper right hand corner of the periodic table.

Progress questions

Question 1

Highly electronegative elements will

A. have a large atomic radius.

B. strongly attract electrons.

Question 2

The electronegativity varies across a period due to the

A. change in size of atomic radii.

B. addition of more electron shells.

Question 3

How does electronegativity vary down a group?

A. Increases

B. Decreases

First ionisation energy 1.1.2.6

The properties of atoms affect the amount of energy required to remove electrons from an atom.

Why do some elements lose their electrons more easily than others?

Elements are represented on the periodic table as neutral atoms. This means that the number of electrons and protons is the same, and therefore the negative and positive charges cancel each other out. However, this is not always how elements exist.

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Given enough energy, we are able to remove electrons from an atom. More specifically, atoms are able to lose their valence electrons.[1] When this happens, atoms become **ions**; charged atoms. The amount of energy required to remove the first valence electron from an atom in its gaseous state is known as the **first ionisation energy**. This depends on the attraction of the valence electrons to the positively charged nucleus.

As we have already learned, the atomic radius of an atom decreases as we move across a period and increases as we move down a group, which affects the attraction between the nucleus and the valence electrons. This concept - the attraction between the nucleus and valence electrons - is related to the **core charge** of an atom, which describes how strongly the nucleus is able to attract valence electrons. The core charge takes into consideration the electrons between the valence shell and the nucleus, also known as the inner-shell electrons, and can be expressed as: core charge = atomic number (number of protons) - number of inner-shell electrons.

The greater the core charge, the greater the attraction between the valence shell and the nucleus, and therefore, the more energy required to remove valence electrons from the atom. We also need to consider the impact of the inner shell electrons on the ability of the nucleus to attract the valence electrons. The inner electrons can shield the valence electron from the full attractive force of the nucleus as seen in figure 2, and this is why we need to consider these electrons when calculating the core charge.

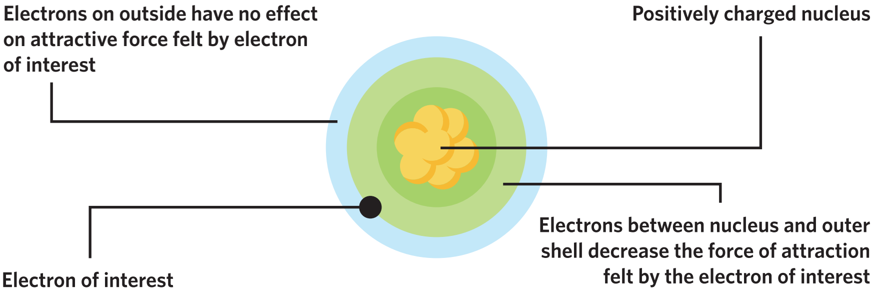
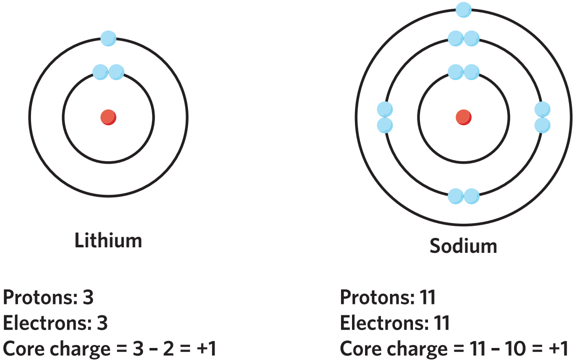


Figure 2 Illustration of the shielding effect of inner shell electrons

Consider a lithium and sodium atom as shown in figure 3, both of which occur in group 1.



Protons: 3
Electrons: 3
Core charge = 3 – 2 = +1

Protons: 11
Electrons: 11
Core charge = 11 – 10 = +1

Figure 3 The deconstruction of a lithium and sodium atom

We can see that both lithium and sodium have a core charge of one. Therefore we would expect the valence electrons from both atoms to experience the same level of attraction towards the nucleus. However, sodium contains more inner-shell electrons and has a larger atomic radius. Therefore, not only would there be more shielding from inner-shell electrons, but the valence electron is much further from the nucleus compared to lithium. As a result, sodium is not able to hold onto its valence electron as strongly as lithium, and thus would require less energy to remove the valence electron. Hence, the first ionisation energy for sodium would be less than that of lithium.

We can also compare lithium with fluorine as shown in figure 4, both of which are found in period 2.

KEEN TO INVESTIGATE?

[1] **How can the movement of electrons create light?**
Search YouTube: electron energy and light spectra

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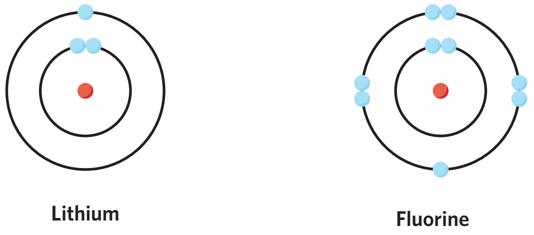


Protons: 3
Electrons: 3
Core charge = 3 - 2 = +1

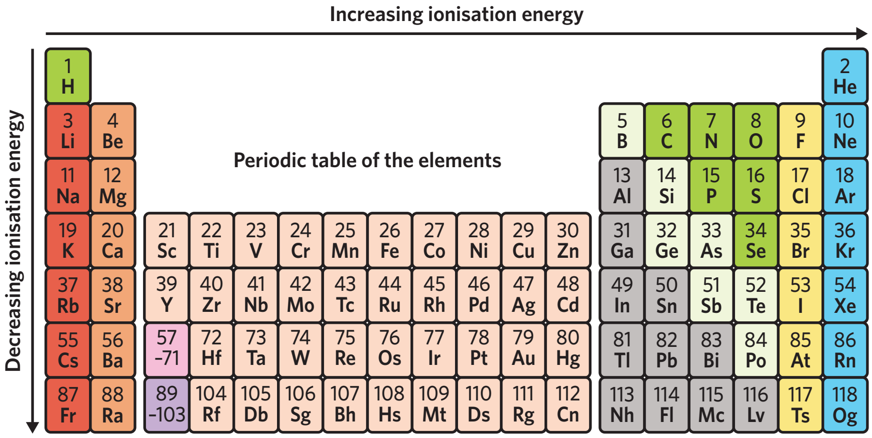
Protons: 9
Electrons: 9
Core charge = 9 - 2 = +7

Figure 4 The deconstruction of a lithium and fluorine atom

As we can see, fluorine has a greater core charger and smaller atomic radius than lithium, both of which allow fluorine to better attract the valence electrons. As a result, fluorine would have a greater first ionisation energy than lithium. The overall trend of ionisation energy can be seen in figure 5.

USEFUL TIP

Ionisation (loss of electrons) is the opposite of electronegativity (gain of electrons).



1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 -71	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 -103	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

Figure 5 Trend in ionisation energy

We can now appreciate that both core charge and atomic radius affect the first ionisation energy of an atom.

WORKED EXAMPLE 1

As you move down the elements of group 1 of the periodic table, what is the trend in the first ionisation energy and electronegativity?

What information is presented in the question?

A group is equivalent to a column in the periodic table.

What is the question asking us to do?

Identify the trends in ionisation energy and electronegativity down a group.

What strategies do we need in order to answer the question?

1. Recall the trend in ionisation energies down a group.

2. Recall the trend in electronegativity down a group.

Answer

Ionisation energy and electronegativity both decrease as you go down a group on the periodic table.

Adapted from VCAA 2002 Exam 2 Multiple choice Q3

Progress questions

Question 4

The core charge ________ as we move across a period.

A. increases

B. decreases

Question 5

The larger the atom, the greater the core charge.

A. True

B. False

Question 6

Atomic radius increases across a period.

A. True

B. False

Question 7

Atomic radius increases as core charge increases.

A. True

B. False

Question 8

Ionisation energy increases as we move

A. up a group.

B. down a group.

Question 9

Ionisation energy decreases as atomic size increases.

A. True

B. False

Metallic character 1.1.2.7

Some metallic elements can easily lose valence electrons.

Why do some metals lose their valence electrons more easily than others?

Elements in group 1 and 2, known as the alkali and alkaline earth metals respectively, exhibit similar physical properties. For example, both types of metals are shiny, have a silvery appearance and are good conductors of electricity. Scientists define a metal in the periodic table as an element that readily loses electrons to form a positively charged ion. Elements that have a similar physical appearance and chemical properties to the elements in group 1 and 2 possess high metallic character**.** Metals will be further investigated in the next chapter.

To identify the trend in metallic character, the elements can first be organised by how easily electrons are lost. As we've learned earlier, it is easier to remove electrons from an atom as we move down a group, than when we move across a period from left to right. As a result, metallic character increases as we move down a group and decreases as we move across a period from left to right. This is summarised in figure 6.

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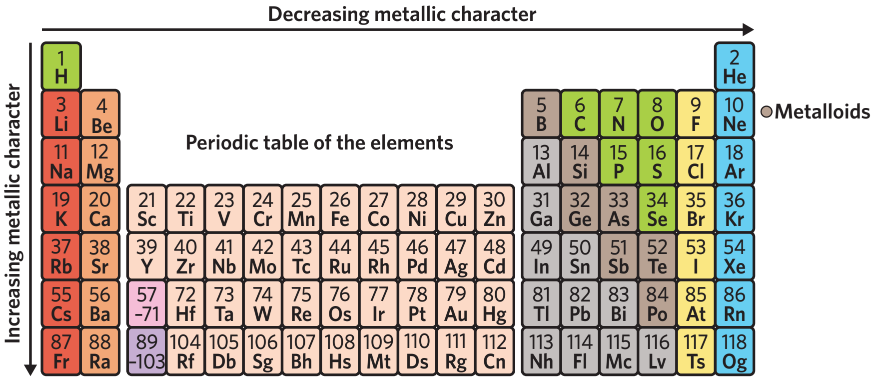

Figure 6 Trend in the metallic character

Metalloids are an interesting group of elements that display a mixture of both metallic and non-metallic properties. Silicon, considered a metalloid, is able to conduct electricity however is not shiny or silvery in appearance. Some other metalloids include boron, arsenic and tellurium as shown in figure 6.

WORKED EXAMPLE 2

The elements in group 4 of the periodic table become more metallic in character as we go down the group. Provide an explanation for this trend.

What information is presented in the question?

Group 4 is equivalent to a column in the periodic table.

What is the question asking us to do?

Explain the trends in metallic behaviour down a group.

What strategies do we need in order to answer the question?

1. Recall the trend in metallic behaviour down a group.
2. Explain why the metallic behaviour changes.

Answer

The size of the atom increases down a group which means the valence electron is more easily lost, a defining feature of metallic behaviour.

Adapted from VCAA 2004 Exam 2 Multiple choice Q12

Progress questions

Question 10

Metals

A. contain five or more valence electrons.

B. are able to easily lose electrons.

Question 11

Metallic character increases as we move across a period.

A. True

B. False

Question 12

Metalloids are the most metallic elements on the periodic table.

A. True

B. False

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Reactivity 1.1.2.8

The reactivity of an element depends on various factors.

Why are some elements more reactive than others?

In the previous lesson, we learned about the features affecting an atom's ability to attract or lose electrons, including the number of electron shells present, the number of valence electrons in the outer shell, the size of the atomic radius and the core charge of the atom. These concepts are important as they help us to understand the reactivity of different elements. Reactivity describes how easy it is for an atom of an element to gain or lose electrons. The easier it is for this to occur, the more reactive the element.

Whether atoms gain or lose electrons depends on the number of valence electrons already present in the atom. For example, metallic elements found in group 1 and 2, which contain one and two valence electrons respectively, would rather lose their valence electrons to have a full valence shell as shown in figure 7. In contrast, non-metallic elements in groups 16 and 17 which contain six and seven valence electrons respectively, would prefer to gain one or two more electrons to satisfy the octet rule as shown in figure 7.

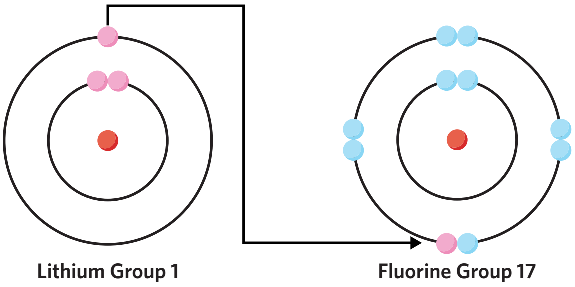

Figure 7 Behaviours of atoms in different groups of the periodic table

Based on the previous lesson we know that atomic radius, electron shielding and core charge can all impact how easy it is for atoms to lose or gain electrons. Reactivity across a period is a multi-faceted concept, as the criterion changes as we move across different types of elements. While we define reactivity of metals as how easily metals lose electrons, we define reactivity of non-metals by how easily non-metals gain electrons. As such, groups 16 and 17 elements can also be described to be highly reactive, as they have a great tendency to gain electrons based on their high electronegativity and core charge. A summary of the factors affecting reactivity can be seen in table 1.

Table 1 Summary of the impact of different properties on reactivity

Property	Effect on reactivity	Description
Increase in atomic radius	↑ for elements that tend to lose electrons ↓ for elements that tend to gain electrons	A large atomic radius allows valence electrons to be easily lost from the valence shell, however it reduces the ability of atoms to attract electrons due to the shielding electrons.
Increase in core charge	↓ for elements that tend to lose electrons ↑ for elements that tend to gain electrons	An increase in core charge allows atoms to better hold onto their valence electrons and gain electrons more easily.
Increase in electronegativity	↓ for elements that tend to lose electrons ↑ for elements that tend to gain electrons	Electronegativity relates to an atom's ability to attract electrons. Atoms with high electronegativity are more likely to gain electrons.

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There is one special group on the periodic table that includes the least reactive elements. These elements are found in group 18 and are also referred to as the **noble gases**.

Progress questions

Question 13

Atoms that can easily lose electrons are highly reactive.

A. True

B. False

Question 14

Atoms that easily gain electrons are highly reactive.

A. True

B. False

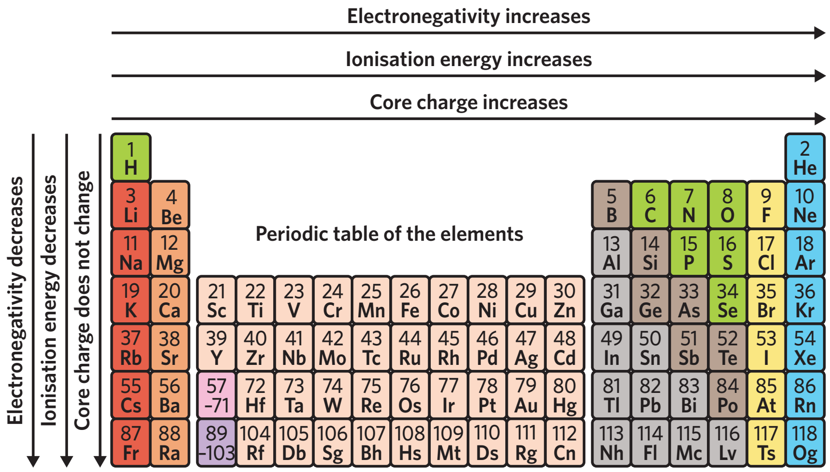
Question 15

Reactivity ________ from group 1 to group 3.

A. increases

B. decreases

Theory summary



- The octet rule outlines that atoms require a total of eight electrons in their valence shell to become stable.
- Metals have a tendency to lose electrons (ionisation) whereas non-metals have a tendency to gain electrons (electronegativity).
- Metallic character decreases as we move across a period.
- Metalloids have characteristics of both metals and non-metals.
- Reactivity reflects the ease at which an atom will lose or gain electrons to satisfy the octet rule.

The content in this lesson is considered fundamental prior knowledge to structure, nomenclature and properties of organic compounds (Unit 4 AOS 1).

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1C Questions

Mild Medium Spicy

Deconstructed

Use the following information to answer questions 16–18.

Barium and magnesium both exhibit similar chemical properties. For example, barium and magnesium both react with water.

Question 16 (1 MARK)

Which of the following statements about these elements is true?

A. Barium and magnesium are found in period 1.

B. Barium is found in period 5 and magnesium is found in period 3.

C. Barium is found in group 6 and magnesium is found in group 3.

D. Barium and magnesium are found in group 2.

Question 17 (1 MARK)

In a reaction, barium would be most likely to

A. gain electrons.

B. lose electrons.

C. gain protons.

D. lose protons.

Question 18 (3 MARKS)

For certain elements, a way in which we test their reactivity is to allow them to come into contact with water and make observations about the interaction that occurs. If a reaction were to occur, there may be the formation of bubbles and/or sparks which would be observed relatively quickly if the element was highly reactive.

Both magnesium and barium react with water in this way. Of the two, which element would show signs of a reaction the quickest? Explain with reference to the periodic table.

Exam-style

Question 19 (5 MARKS)

Sodium and potassium are elements that share similar characteristics. For example, both exhibit a silvery look and interact intensely with water.

a. Identify the core charge of each element. (1 MARK)

b. Compare the atomic radius of both elements. (2 MARKS)

c. Of the two, which element is more likely to have a lower ionisation energy? Explain. (2 MARKS)

Question 20 (7 MARKS)

Fluorine and bromine can both be found as gases in the atmosphere. Fluorine is highly flammable and inhaling bromine can cause severe health problems.

a. From which element would it be hardest to remove a valence electron? Explain. (3 MARKS)

b. After conducting multiple tests, a scientist discovered that the elements had electronegativity scores of 3.98 and 2.96, however the scientist was unsure which measurement belonged to which element. Using your understanding, allocate each piece of data to the appropriate element. Explain your reasoning. (3 MARKS)

c. Suppose both elements were within proximity of a floating electron that did not belong to any atom. Which of the two elements would be more likely to attract the electron to itself? (1 MARK)

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Question 21 (8 MARKS)

Two students are preparing for a presentation to their chemistry class. The notes the students use for the presentation are composed of a series of statements and justifications in order to tackle the misconceptions of the class. A segment of their notes is given.

Complete the table by identifying whether each statement is correct or incorrect and give a justification for your choice.

Statement	Correct/Incorrect	Justification
Atoms with high reactivity are expected to have high electronegativity.		
Metals are more reactive than non-metals.		
Metals can be found on the left side of the periodic table.		
First ionisation energy is highest for the noble gases.		

Key science skills

Question 22 (9 MARKS)

A scientist discovers a beaker hidden in a cupboard containing a solution of aqueous sodium chloride (NaCl). In order to further explore the properties of the sodium chloride, the scientist designs a practical investigation.

a. The first step of this investigation is researching the substance.

  i. The scientist must classify the substance as belonging to one of the following categories: atom, element, compound or mixture. Which category is most appropriate? (1 MARK)

  ii. The research conducted requires the period and group to be identified for each component of the substance. Identify both the period and group of Na and Cl. (2 MARKS)

Now the scientist must begin executing an experiment. A range of elements in the same period as sodium and chlorine are analysed and the following results are obtained:

Sample	A	B	C	D	E
First ionisation energy (kJ mol^{-1})	494	577	785	997	1250

b. Explain the trend shown in the table, assuming the samples are listed in order of ascending mass number, with reference to core charge. (3 MARKS)

c. Identify the dependent variable in this experiment. (1 MARK)

d. A fellow colleague of the scientist notes that the scales used to weigh the sodium chloride prior to conducting the first ionisation energy analysis were not calibrated correctly. What type of error would this induce? Explain your answer. (2 MARKS)

FROM LESSONS 16C, 16A & 16D

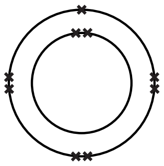
Questions from multiple lessons

Question 23 (7 MARKS)

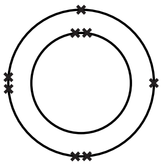
Scientists use experimental results to determine properties and characteristics of different elements and molecules.

Consider the following uncharged atoms.

a



b



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A scientist conducting an experiment recorded the data shown.

	Element A	Element B
Ionisation energy (kJ mol^{-1})	1681	1313.9
Atomic radius (pm)	42	48
Period		

a. Complete the missing numbers in the table provided. (2 MARKS)

b. Using all the information presented in the table, determine the name of elements A and B. Explain your choices. (5 MARKS)

FROM LESSON 1B

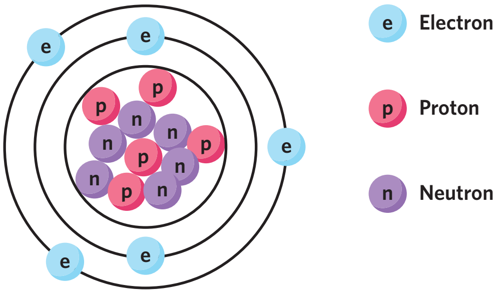
Question 24 (2 MARKS)

Consider the following element.

a. Write the electron configuration for this element. (1 MARK)

b. In which block would we find this element on the periodic table? (1 MARK)

FROM LESSON 1B



Hints

19a. Core charge is calculated based on the number of protons and inner electrons.

19b. Atomic radius is impacted by the number of protons and energy shells.

19c. Ionisation energy indicates how difficult it is for a valence electron to be removed from an atom.

20a. The attraction between the positive nucleus and valence electrons is what keeps the valence electrons in the atom.

20b. Elements with a strong ability to attract electrons are more electronegative.

20c. Elements with a higher electronegativity are able to more strongly attract electrons.

21. The atomic structure of an element impacts the properties of the element.

22ai. Substances are classified based on their composition.

22aii. Periods and groups are like coordinates for an element.

22b. The ability to remove electrons depends on the strength of attraction to the nucleus.

22c. The dependent variable is the variable affected by the change in the independent variable.

22d. Errors are categorised differently depending on their effect on the experiment.

23a. Periods indicate the number of energy shells.

23b. Elements with higher core charges have a greater tendency to attract valence electrons.

24a. Electrons fill an atom's electron shells in a particular order.

24b. Blocks on the periodic table are based on the outermost subshell.

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1D Recycling critical elements

STUDY DESIGN DOT POINT

- critical elements (for example, helium, phosphorus, rare-earth elements and post-transition metals and metalloids) and the importance of recycling processes for element recovery

1A 1B 1C **1D**

1.1.3.1 Critical elements

1.1.3.2 Recycling elements

ESSENTIAL PRIOR KNOWLEDGE

1B	The periodic table
1B	Blocks on the periodic table

See questions 7–8.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: Andrey Suslov/Shutterstock.com

How many critical elements does it take to make a mobile phone?

Humans continue to utilise more and more elements in their daily lives to the extent that a common mobile phone can contain more than 70 different elements. A consequence of this escalating demand has led to some vital elements becoming listed as ‘**endangered**’. In this lesson we will learn about these **critical elements** and how they can be **recycled** for future generations.

KEY TERMS AND DEFINITIONS

Critical elements	elements that are vital and considered in short supply unless they are recycled as sources could become depleted in 50–100 years
Endangered	elements that have a limited supply left on Earth
Lanthanoids	15 naturally occurring metallic chemical elements, whose atomic numbers fall between 57 (Lanthanum) and 71 (Lutetium)
Recycled	converting a waste product into a reusable product
Sustainable	can be produced at a rate that is greater than consumption without compromising future generations
Transition metals	elements that occupy groups 3 to 12 on the periodic table

Critical elements 1.1.3.1

Elements that are vital for industry and technology play an important role in the development of the human race.

Why are some elements considered to be critical?

There are many elements that are in limited supply. Figure 1 highlights the estimated timescales remaining for the supply of many of these vital elements. The supply, demand and discovery of new reserves of these elements are constantly changing. Table 1 outlines the uses and sources of some critical elements.

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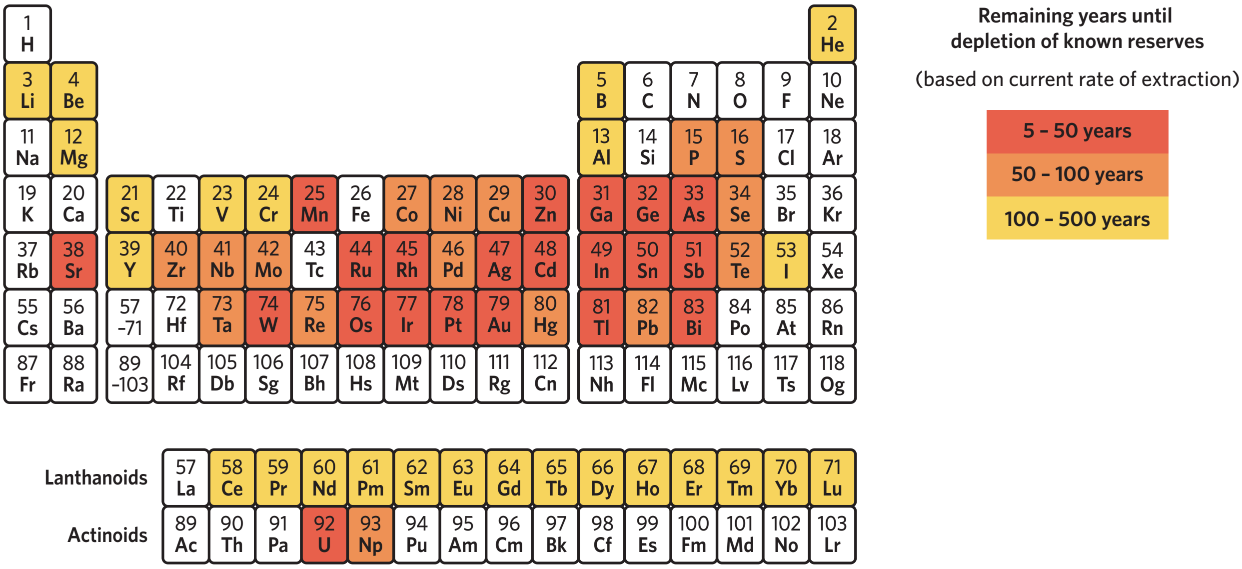

Figure 1 Estimated depletion timescales of critical elements

Table 1 Critical elements are all around us

Critical element(s)	Uses and sources
Helium	Used in party balloons, as a coolant and as a mixture in deep-sea diving. Can be produced from either radioactive sources or natural gas at a high cost.
Phosphorus	Used in fertilisers, present in meat and as a key element in living organisms. Only limited sources are found on Earth in the form of phosphate salts.
Transition metals	Used heavily in electronics. We only have small reserves left on Earth. Examples include iridium, platinum, osmium and palladium.
Indium as an example of a post transition metal/metalloid	Used in touch screens, automotive glass and solar panels. It is a metalloid (see figure 2) and is extracted as a byproduct of zinc mining and sources are dwindling.
Rare-earth	Have many uses (see figure 3) and are mostly from the **lanthanoids** on the periodic table (see figure 4). They are difficult to extract from rocks.

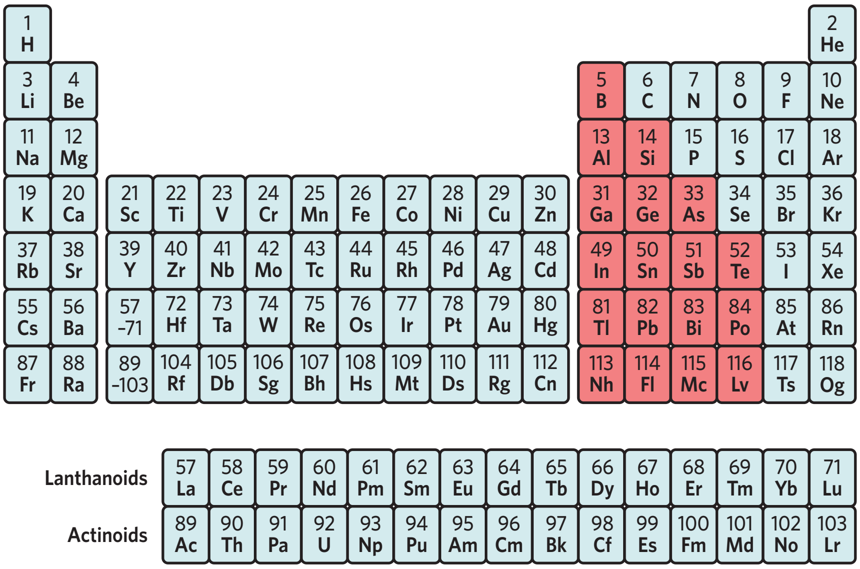

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Figure 2 Location of the post-transition metals and metalloids (shown in red)

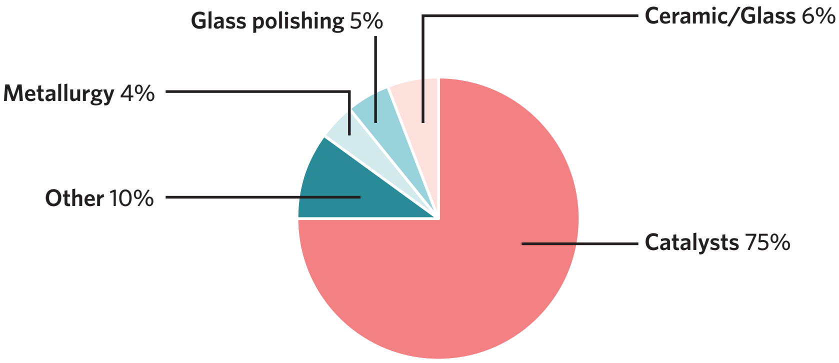

Figure 3 Uses of rare-earth elements

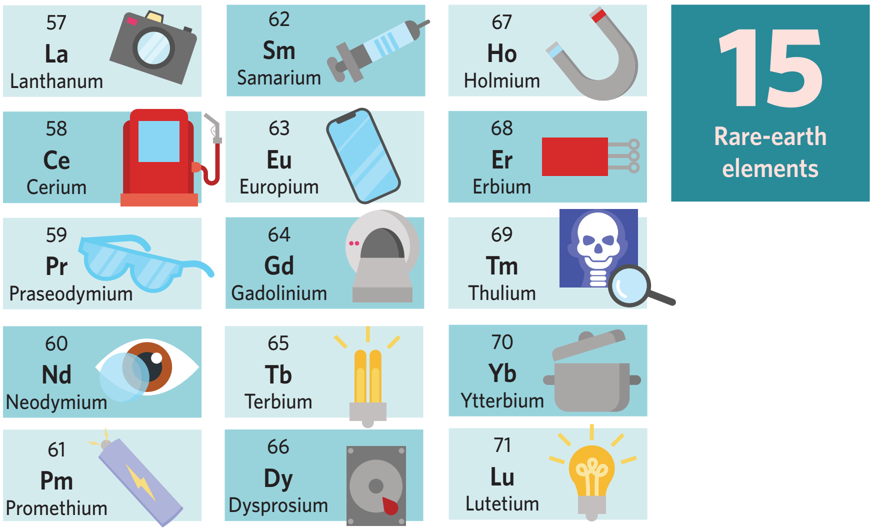

Figure 4 Rare-earth elements

MISCONCEPTION

'Elements are used up.'

Only helium (as it escapes into space) and radioactive elements (as they break down into other elements) are actually used up, meaning the supply on Earth is reduced. Most elements have the potential to be recycled.

Progress questions

Question 1

With regards to a chemical element, critical means that it is important to a particular product or material.

A. True

B. False

Question 2

Which of the following elements could be classified as metalloids?

A. Sc, Ti, and V

B. Ga, In, and Bi

Question 3

What name is given to the f-block that contains most of the rare-earth elements?

A. Actinoids

B. Lanthanoids

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Recycling elements 1.1.3.2

The demand for these critical elements is increasing at a rate that is not **sustainable**.

How can these critical elements be used more sustainably?

As both the population of the world and dependency on materials continues to increase, the rate at which Earth's reserves are being used is not sustainable. Chemical processes need to meet the demands of humans and must also take into account the requirements for future generations. The solution is to improve and innovate methods of recycling, as many critical elements are currently being buried in landfill sites (see table 2). The recycling of metals will be discussed in greater detail in lesson 3C.

Table 2 Using critical elements more sustainably

Critical element(s)	Recycling and more sustainable options
Helium	• Collect and reuse helium liquid that has been used as a coolant in MRI scanners • Ban the use of helium in party balloons
Phosphorus	• Reduce the use of artificial fertilisers • Increase the use of natural fertilisers • Remove phosphorus from sewage and reuse • Use algae to remove the phosphate from polluted waterways
Transition metals	• Recycle batteries[1] which contain copper, cobalt, manganese and nickel • Repurpose them instead of throwing them into landfill • Improve recycling to the same level as iron and aluminium
Indium as an example of a post transition metal/metalloid	• Improve collection methods • Increase the duration that a device is used for before getting a new one • Use bacteria to remove indium • Find non-metal alternatives to indium
Rare-earth	• Improve manufacturing process to use less • Increase the amount that is recycled • Make products more recyclable

MISCONCEPTION

'Rare-earth elements are rare.'

These elements are not rare but they are difficult, expensive and sometimes dangerous to produce.

KEEN TO INVESTIGATE?

[1] **What is the rate of recycling of batteries in Australia?**
Search YouTube: Recycled Batteries from envirostream

WORKED EXAMPLE 1

Give an example and a common use of a transition metal that is considered to be a critical element.

What information is presented in the question?

The critical element must be a transition metal.

What is the question asking us to do?

Identify an element and give an example of how it is used.

What strategy do we need in order to answer the question?

Locate the transition metals in the periodic table.

Answer

Nickel is found in many rechargeable batteries.

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Progress questions

Question 4

Given that Ni, Co, and Fe are all magnetic they could be more easily separated from other metals.

A. True

B. False

Question 5

Which of the following is an element recovery method?

A. Removing phosphorus from human and animal waste

B. Using an alternative to helium in party balloons

The content in this lesson is considered fundamental prior knowledge to the concept of sustainability which is embedded throughout Units 3&4.

Theory summary

- Many elements are both endangered and vital to humans and are therefore classified as critical.
- Many critical elements on the periodic table are found in blocks called transition metals, metalloids and lanthanoids.
- Recycling these elements and finding alternatives is crucial in creating more sustainable chemical processes.

1D Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 6–8.

In 1669, Hennig Brand evaporated and condensed over 5000 litres of urine to produce and discover phosphorus.

Question 6 (1 MARK) 🌶

Phosphorus is essential to all life because it is

A. an unreactive non-metal.

B. the building block of atoms.

C. exists in either white or red.

D. involved in many cellular reactions.

Question 7 (1 MARK) 🌶

Phosphorus is commonly found in waterways as

A. phosphate.

B. phosphine.

C. red phosphorus.

D. white phosphorus.

Question 8 (4 MARKS) 🌶🌶🌶

Explain how humans come to have phosphorus in their urine and how this phosphorus could be recycled.

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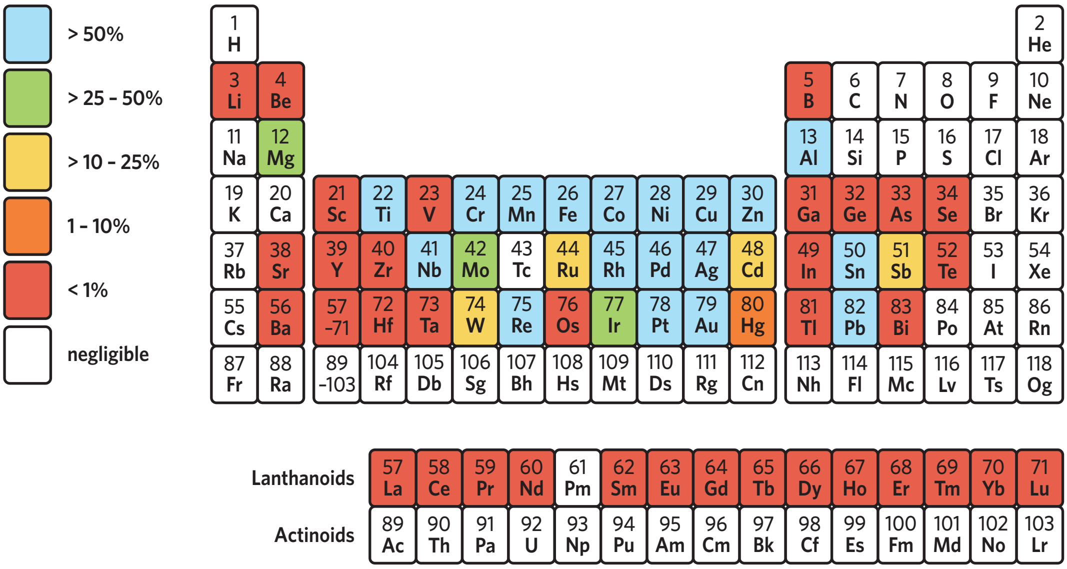
Exam-style

Question 9 (6 MARKS) 🌶🌶

The periodic table shows the rates at which elements are recycled. Using the information provided in the table and your knowledge answer the following questions.



🟦	> 50%
🟩	> 25 - 50%
🟨	> 10 - 25%
🟧	1 - 10%
🟥	< 1%
⬜	negligible

1 H ⬜																	2 He ⬜
3 Li 🟥	4 Be 🟥											5 B 🟥	6 C ⬜	7 N ⬜	8 O ⬜	9 F ⬜	10 Ne ⬜
11 Na ⬜	12 Mg 🟩											13 Al 🟦	14 Si ⬜	15 P ⬜	16 S ⬜	17 Cl ⬜	18 Ar ⬜
19 K ⬜	20 Ca ⬜	21 Sc 🟥	22 Ti 🟦	23 V 🟥	24 Cr 🟦	25 Mn 🟦	26 Fe 🟦	27 Co 🟦	28 Ni 🟦	29 Cu 🟦	30 Zn 🟦	31 Ga 🟥	32 Ge 🟥	33 As 🟥	34 Se 🟥	35 Br ⬜	36 Kr ⬜
37 Rb ⬜	38 Sr 🟥	39 Y 🟥	40 Zr 🟥	41 Nb 🟦	42 Mo 🟩	43 Tc ⬜	44 Ru 🟨	45 Rh 🟦	46 Pd 🟦	47 Ag 🟦	48 Cd 🟨	49 In 🟥	50 Sn 🟦	51 Sb 🟨	52 Te 🟥	53 I ⬜	54 Xe ⬜
55 Cs ⬜	56 Ba 🟥	57 -71 🟥	72 Hf 🟥	73 Ta 🟥	74 W 🟨	75 Re 🟦	76 Os 🟥	77 Ir 🟩	78 Pt 🟦	79 Au 🟦	80 Hg 🟧	81 Tl 🟥	82 Pb 🟦	83 Bi 🟥	84 Po ⬜	85 At ⬜	86 Rn ⬜
87 Fr ⬜	88 Ra ⬜	89 -103 ⬜	104 Rf ⬜	105 Db ⬜	106 Sg ⬜	107 Bh ⬜	108 Hs ⬜	109 Mt ⬜	110 Ds ⬜	111 Rg ⬜	112 Cn ⬜	113 Nh ⬜	114 Fl ⬜	115 Mc ⬜	116 Lv ⬜	117 Ts ⬜	118 Og ⬜

Lanthanoids	57 La 🟥	58 Ce 🟥	59 Pr 🟥	60 Nd 🟥	61 Pm ⬜	62 Sm 🟥	63 Eu 🟥	64 Gd 🟥	65 Tb 🟥	66 Dy 🟥	67 Ho 🟥	68 Er 🟥	69 Tm 🟥	70 Yb 🟥	71 Lu 🟥
Actinoids	89 Ac ⬜	90 Th ⬜	91 Pa ⬜	92 U ⬜	93 Np ⬜	94 Pu ⬜	95 Am ⬜	96 Cm ⬜	97 Bk ⬜	98 Cf ⬜	99 Es ⬜	100 Fm ⬜	101 Md ⬜	102 No ⬜	103 Lr ⬜

a. What is the name of the block of elements that are the most recycled? (1 MARK)

b. What is the name of the block of elements that are the least recycled? (1 MARK)

c. Suggest a reason why this difference in the rates of recycling in these blocks exists. (2 MARKS)

d. Comment on the sustainability of using elements like lithium in batteries. (2 MARKS)

Question 10 (5 MARKS) 🌶🌶

Electric vehicles utilise an incredible amount of rare-earth elements. Some countries like France have dedicated recycling plants used to isolate rare-earth elements. Most of the world's rare-earth metals come from China and to a lesser extent from Australia.

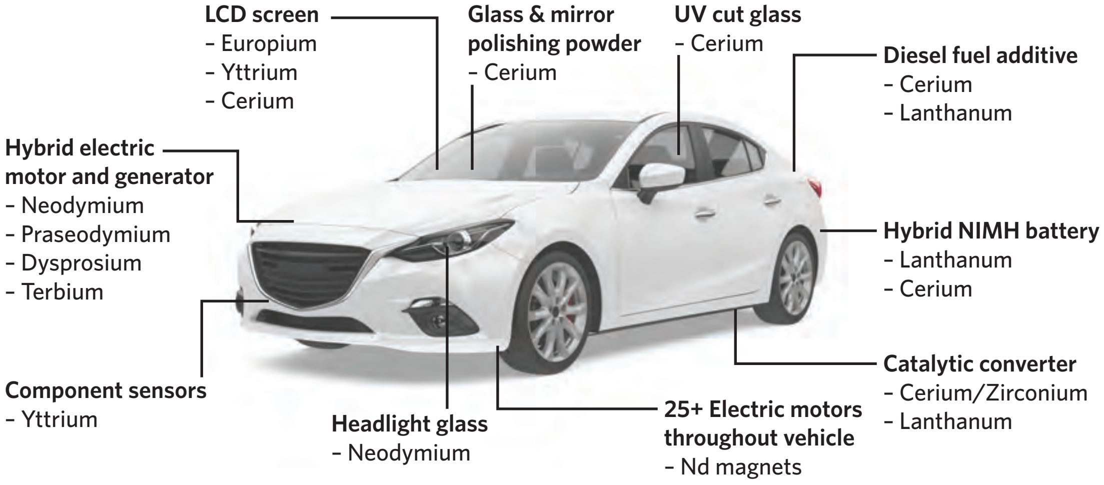


Image: Nerthuz/Shutterstock.com

a. Name the two elements used in the electric car shown that are not classified as rare-earth elements. (1 MARK)

b. Give an advantage France has over other countries with regards to the sourcing of rare-earth elements. (2 MARKS)

c. Why might other countries not have dedicated rare-earth element recycling plants? (2 MARKS)

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Key science skills

Question 11 (6 MARKS)

Many critical elements come from countries that have ongoing civil wars and conflicts. This can result in the mining of some critical elements being done by child labour.

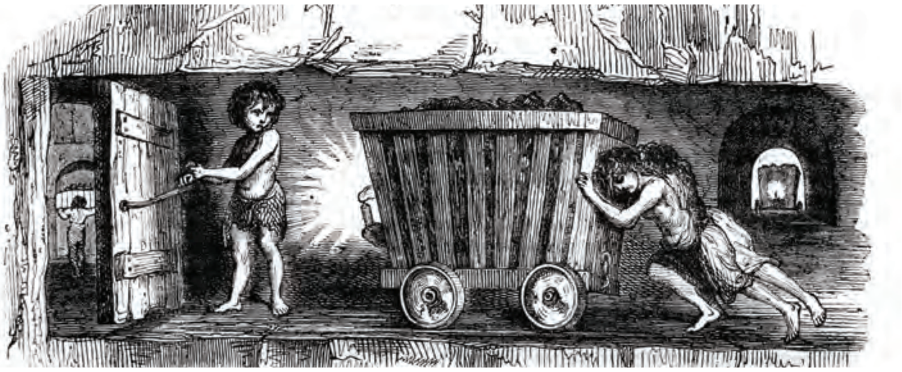

Image: Morphart Creation/Shutterstock.com

When carrying out an investigation into how rare-earth elements are extracted and then used in the production of a mobile phone, what issues should be addressed about the labour workforce? Discuss using the following themes.

a. Ethical (2 MARKS)

b. Economic (2 MARKS)

c. Legal (2 MARKS)

FROM LESSON 16B

Questions from multiple lessons

Question 12 (5 MARKS)

Use the periodic table to write the chemical symbols for the following elements.

a. The uncharged element that is given by the electron configuration $1s^22s^22p^63s^2\ p^64s^23d^6$ (1 MARK)

b. The element from period 3 that has the highest first ionisation energy (1 MARK)

c. The element from period 3 that has the largest atomic radius (1 MARK)

d. The most electronegative element in group 17 (1 MARK)

e. The element in period 2 that has the greatest metallic character (1 MARK)

Adapted from VCAA 2002 Exam 2 Multiple choice Q1

FROM LESSONS 1A & 1B

Hints

9a. Remember the names of the blocks on the periodic table.

9b. Remember the names of the blocks on the periodic table.

9c. Iron and aluminium are recycled in most countries.

9d. Data from the periodic table shown.

10a. There are different blocks on the periodic table.

10b. France has the ability to recycle rare-earth elements.

10c. Recycling is a complex process.

11a. Worker age restrictions are enforced in a lot of countries.

11b. Survival can sometimes depend on money.

11c. Safe working conditions and minimum pay.

12a. Number of electrons equals the number of protons.

12b. Ionisation energy is determined by atomic radius and core charge.

12c. Atomic radius is determined by the number of electrons shells and core charge.

12d. Electronegativity is determined by the number of electrons shells and core charge.

12e. Metallic properties are determined by low ionisation energies.

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Chapter 1 review

Multiple choice (10 MARKS)

Question 1 (1 MARK) 🌶

According to the subshell model, what is the electronic configuration of the atom in period 3, group 2 of the periodic table?

A. $1s^22s^22p^63s^2$

B. $1s^22s^22p^6$

C. 2, 8, 2

D. 2, 3

Question 2 (1 MARK) 🌶

Which of the following statements regarding atomic number is **incorrect**?

A. Atomic numbers are unique to each element.

B. The atomic number represents the number of electrons in a neutral atom.

C. The atomic number is the number of neutral particles in the nucleus.

D. The periodic table is organised with respect to atomic number.

Question 3 (1 MARK) 🌶

Which of the following is not considered to be a critical element?

A. La

B. Ce

C. Ca

D. Lu

Question 4 (1 MARK) 🌶

Which of the following options correctly matches the properties of elements to their trend in the periodic table?

	Property	**Across a period (left to right)**	**Down a group**
A.	Electronegativity	Increases	Increases
B.	Core charge	Increases	Decreases
C.	First ionisation energy	Decreases	Increases
D.	Metallic character	Decreases	Increases

Question 5 (1 MARK) 🌶🌶

The noble gases (helium to radon) have an outer shell electron configuration of

A. s^2.

B. s^2p^6.

C. either s^2 or s^2p^6.

D. either s^2p^6 or $s^2p^6d^{10}$.

VCAA 2005 Exam 2 Multiple choice Q17

Question 6 (1 MARK)

Which one of the following alternatives lists the atoms of aluminium, calcium, sulfur and chlorine in order of increasing electronegativity?

A. (lowest) Al, S, Cl, Ca (highest)

B. (lowest) Ca, Al, S, Cl (highest)

C. (lowest) Cl, S, Al, Ca (highest)

D. (lowest) S, Ca, Al, Cl (highest)

VCAA 2006 Exam 2 Multiple choice Q16

Question 7 (1 MARK)

Which one of the following alternatives lists the atoms of chlorine, fluorine, magnesium and potassium in order of increasing atomic radius?

A. (smallest) K, Mg, Cl, F (largest)

B. (smallest) F, Mg, Cl, K (largest)

C. (smallest) K, F, Mg, Cl (largest)

D. (smallest) F, Cl, Mg, K (largest)

VCAA 2006 Exam 2 Multiple choice Q17

Question 8 (1 MARK)

Which one of the following alternatives lists the atoms of chlorine, magnesium, neon and phosphorus in order of increasing first ionisation energy?

A. (smallest) Mg, P, Cl, Ne (largest)

B. (smallest) Ne, Cl, P, Mg (largest)

C. (smallest) Cl, Mg, Ne, P (largest)

D. (smallest) P, Mg, Cl, Ne (largest)

VCAA 2006 Exam 2 Multiple choice Q18

Question 9 (1 MARK)

When Dimitri Mendeleev developed the periodic table he left gaps for as yet undiscovered elements. On the basis of the position of these gaps and in relation to these undiscovered elements, Mendeleev was able to predict

A. their electron configurations.

B. the occurrence of their isotopes.

C. many of their physical properties.

D. their atomic numbers and mass numbers.

VCAA 2006 Exam 2 Multiple choice Q19

Question 10 (1 MARK)

Which of the following equations is always true where A = atomic number, P = number of protons, N = number of neutrons, and M = mass number?

A. $M = 2 \times P$

B. $M = A + P$

C. $M = 2 \times N$

D. $M = A + N$

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Short answer (30 MARKS)

Question 11 (11 MARKS)

Through experimentation, it is determined that a neutral atom of an element has an electron configuration of $1s^22s^22p^63s^23p^63d^24s^2$.

a. What is the name, chemical symbol, atomic number and relative atomic mass of this element? (3 MARKS)

b. What block of the periodic table is this element found in? Justify your answer. (2 MARKS)

c. Compare the electronegativity of this element with copper. (2 MARKS)

d. Compare the metallic character of this element with zirconium. (2 MARKS)

e. Is this element more likely to react by gaining or losing electrons? Explain your answer (2 MARKS)

Question 12 (7 MARKS)

a. Predict the trend in the following properties of the second period elements, moving from lithium to fluorine, giving an explanation in support of each prediction. (4 MARKS)

Property	Increases or decreases?	Explanation for predicted trend
Atomic size		
Electronegativity		

VCAA 2004 Exam 2 Short answer Q6b

b. There is also a general trend in first ionisation energy both across periods and down groups of the periodic table.

What is meant by the term ionisation energy? (1 MARK)

c. Given that the atomic size of potassium is greater than that of sodium, explain why the first ionisation energy of sodium is greater than that of potassium. (2 MARKS)

Adapted from VCAA 2004 Exam 2 Short answer Q6c

Question 13 (12 MARKS)

Calcium and magnesium are not considered to be transition metals as they are found in group 2 of the periodic table. Nickel and copper are both found in the first transition metal period of the periodic table.

a. Give two reasons why calcium and magnesium can be placed in the same vertical group. (2 MARKS)

b. Explain why the first series of transition metals are placed in the same period. (1 MARK)

c. The electronegativity of magnesium (1.31) is greater than that of calcium (1.00). Give a brief explanation for this difference. (2 MARKS)

d. Write the electron configuration, in terms of shells and subshells, for the calcium and nickel atoms. (2 MARKS)

e. Explain why there are exactly 10 elements in each of the periods of transition metals. (3 MARKS)

f. Europium is commonly found in mobile phones. What makes an element like europium critical and what can be done to make its use more sustainable? (2 MARKS)

Adapted from VCAA 2005 Exam 2 Short answer Q2c

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Key science skills (9 MARKS)

Question 14 (9 MARKS)

Antoine Lavoisier was a French chemist who conducted many experiments to prove that when reactants completely react together, no mass is lost in the formation of the products; he called this the law of conservation of mass. One of these experiments involved placing mercury in a sealed glass jar and then heating it to see its reaction with oxygen.

a. Write a possible aim for this experiment. (1 MARK)

b. Write a possible hypothesis for this experiment. (2 MARKS)

c. Explain what data, and its type, would need to be collected in order to prove the law of conservation of mass. (3 MARKS)

d. In one experiment, the jar wasn't sealed properly, and so some of the product escaped. What effect would this have on the experiment and what type of error is it? (2 MARKS)

e. Safety standards were not very good in the 18th century. Suggest one precaution that should be taken when handling toxic mercury. (1 MARK)

FROM LESSONS 16A, 16B & 16D

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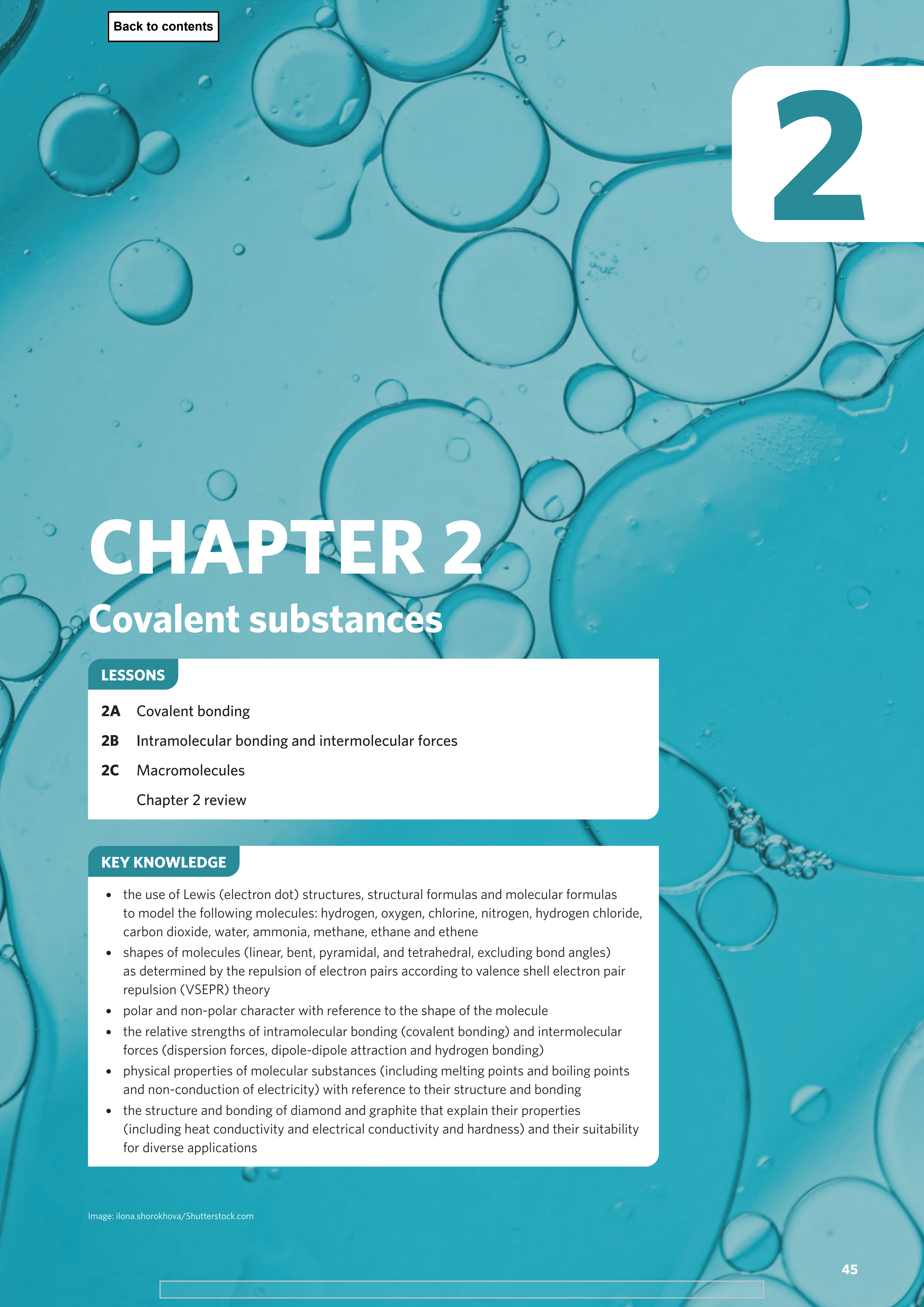

CHAPTER 2

Covalent substances

LESSONS

KEY KNOWLEDGE

- the use of Lewis (electron dot) structures, structural formulas and molecular formulas to model the following molecules: hydrogen, oxygen, chlorine, nitrogen, hydrogen chloride, carbon dioxide, water, ammonia, methane, ethane and ethene
- shapes of molecules (linear, bent, pyramidal, and tetrahedral, excluding bond angles) as determined by the repulsion of electron pairs according to valence shell electron pair repulsion (VSEPR) theory
- polar and non-polar character with reference to the shape of the molecule
- the relative strengths of intramolecular bonding (covalent bonding) and intermolecular forces (dispersion forces, dipole-dipole attraction and hydrogen bonding)
- physical properties of molecular substances (including melting points and boiling points and non-conduction of electricity) with reference to their structure and bonding
- the structure and bonding of diamond and graphite that explain their properties (including heat conductivity and electrical conductivity and hardness) and their suitability for diverse applications

Image: ilona.shorokhova/Shutterstock.com

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2A Covalent bonding

STUDY DESIGN DOT POINTS

- the use of Lewis (electron dot) structures, structural formulas and molecular formulas to model the following molecules: hydrogen, oxygen, chlorine, nitrogen, hydrogen chloride, carbon dioxide, water, ammonia, methane, ethane and ethene
- shapes of molecules (linear, bent, pyramidal, and tetrahedral, excluding bond angles) as determined by the repulsion of electron pairs according to valence shell electron pair repulsion (VSEPR) theory
- polar and non-polar character with reference to the shape of the molecule

2A 2B 2C

1.1.4.1 Representing molecules
1.1.5.1 Shapes of molecules
1.1.6.1 Properties determining polarity

ESSENTIAL PRIOR KNOWLEDGE

1B Valence shell
1C Octet rule
1C Electronegativity
See questions 9–11.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

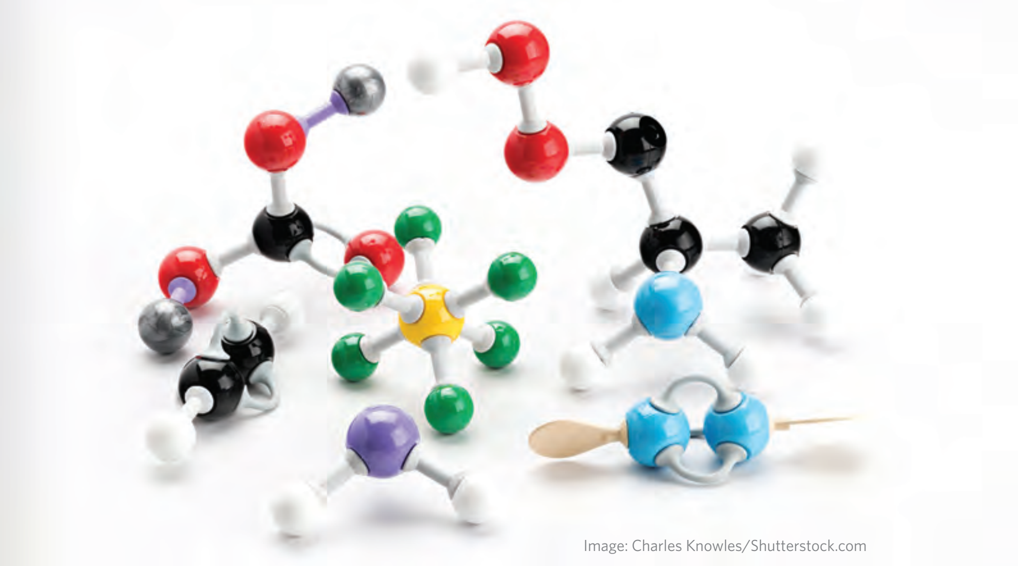

Image: Charles Knowles/Shutterstock.com

How accurate are covalent bonding models?

In this lesson, we'll investigate the bonds which hold **molecules** together, and how these molecules are represented. The shape and types of covalent bonds within covalent substances have important effects on a molecule's **polarity**.

KEY TERMS AND DEFINITIONS

Bond pair pair of valence electrons that is shared between atoms in a covalent bond

Covalent bonding bonding resulting from the sharing of electrons between non-metal atoms

Diatomic molecule covalently bonded pair of atoms of the same or a different non-metal element

Electrostatic attraction attraction between opposite charges

Intramolecular bonds bonding between atoms within a molecule

Lewis structure representation of a molecule where valence electrons are shown as dots

Lone pair electron pairs of valence electrons not shared with another atom in a covalent bond

Molecule when two or more non-metal atoms are covalently bonded together

Molecular formula actual number of atoms in a molecule

Non-polar covalent bond covalent bond between two atoms of similar electronegativity

Permanent dipole permanent area of partial positive or partial negative charge on a molecule

Polar covalent bond covalent bond between atoms of differing electronegativities

Polar molecule asymmetrical molecule with a partially positively charged end and a partially negatively charged end

Polarity property of having a partial positive and negative charge

Structural formula shows all the atoms and bonds in a molecule or compound

Unpaired electron electron in an orbital by itself

Valence shell electron pair repulsion theory (VSEPR) model used to predict the different shapes of molecules based on the repulsion between electron pairs

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Representing molecules 1.1.4.1

Covalent bonding is the sharing of a pair or pairs of electrons between non-metallic atoms. Covalent bonding can be modelled in multiple ways.

What is a covalent bond?

A molecule is a group of non-metal atoms chemically bonded together. Bonds between atoms of the same molecule are known as **intramolecular bonds**, an example of which is shown in figure 1.

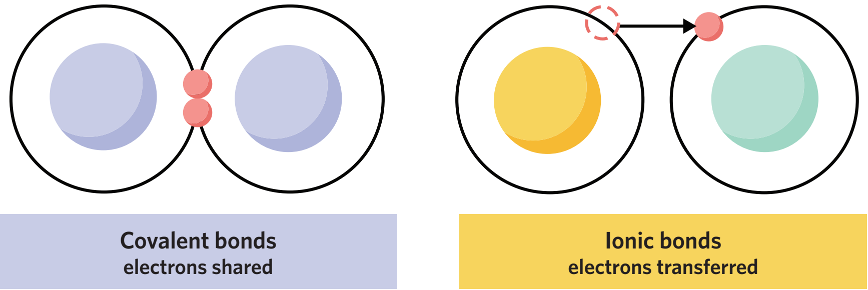


Figure 1 The difference between ionic and covalent bonding

A covalent bond is formed when non-metallic atoms share electrons in order to fill their valence shell. This sharing occurs in accordance with the octet rule, which states that atoms are most stable with eight electrons in their outer shell. Two atoms might share one, two, three or more pairs of electrons to satisfy this rule (figure 2). Electrons shared between atoms are electrostatically attracted to the positive nucleus of both atoms. This force of **electrostatic attraction** holds the atoms together. The more electron pairs shared between two atoms, the stronger the covalent bond will be. There are some important exceptions to the octet rule, such as hydrogen, which is most stable with two electrons in its outer shell (figure 3).

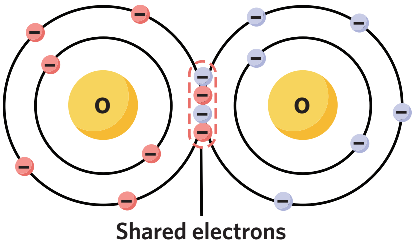


Figure 2 Double covalent bond between two oxygen atoms

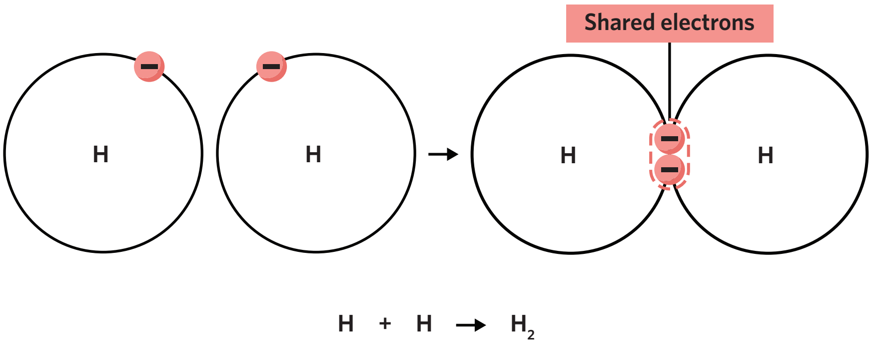


Figure 3 A **diatomic molecule** of hydrogen

How can we represent covalently bonded molecules?

Lewis structures (also known as electron dot formulas, electron dot diagrams or Lewis diagrams) are a useful way of representing covalent bonds. These diagrams illustrate the number of valence electrons that an atom has, with dots representing electrons, as shown in figure 4. They do not include inner shell electrons because these do not contribute to the bonding or shape of molecules.

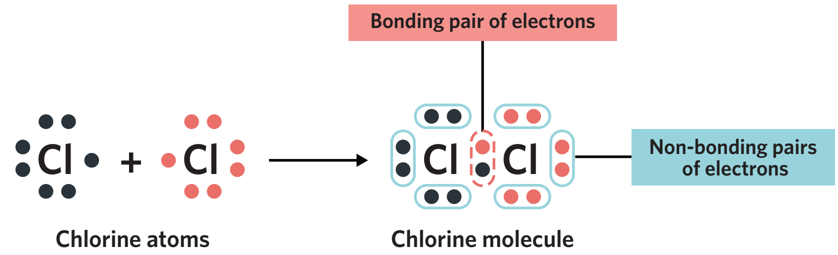


Figure 4 Lewis diagram illustrating the covalent bonding in a diatomic molecule of chlorine

WORKED EXAMPLE 1

Draw an electron dot diagram for the molecule PF_3.

What information is presented in the question?

The molecule is PF_3.

What is the question asking us to do?

Draw an electron dot diagram.

What strategies do we need in order to answer the question?

1. Use the periodic table to determine the number of electrons in each atom of P (phosphorus) and F (fluorine).
2. Determine the number of valence electrons in each atom by finding the electron configuration.
3. Place four of the valence electrons of each atom around the atomic symbol, and then start pairing the electrons up.
4. Draw an electron dot diagram taking into account that a single bond involves two shared electrons.

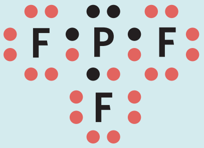
Answer

Phosphorus has an atomic number of 15, which means in a neutral atom of phosphorus, there are 15 electrons. Using our rules of electron configurations, there will be two electrons in the first shell, eight electrons in the second shell and the remaining five electrons in the third shell. This gives an electron configuration of $1s^22s^22p^63s^23p^3$.

This means that the number of valence electrons is five.

Fluorine has an atomic number of 9 which means in a neutral atom there are nine electrons. Using our knowledge of electronic configuration, there will be two electrons in the first shell and seven electrons in the second shell. This gives an electron configuration of $1s^22s^22p^5$. Therefore, fluorine has seven valence electrons.

In order to fill their valence shells, the fluorine atoms each form a single bond with the phosphorus atom. Therefore, the electron dot diagram is:



WORKED EXAMPLE 2

Draw an electron dot diagram for a molecule of carbon dioxide (CO_2).

What information is presented in the question?

The molecule is CO_2.

What is the question asking us to do?

Draw an electron dot diagram.

What strategies do we need in order to answer the question?

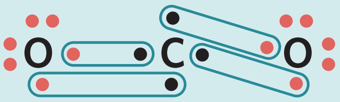
1. Use the periodic table to determine the number of electrons in each atom of C (carbon) and O (oxygen).
2. Determine the number of valence electrons in each atom by finding the electron configuration.
3. Determine how many electron pairs must form in order to fill the outer shells of the atoms.
4. Draw an electron dot diagram taking into account that a double bond involves four shared electrons.

Answer

Carbon has an atomic number of 6 which means in a neutral atom of carbon there are six electrons. Using our rules for electron configurations, there will be two electrons in the first shell and four electrons in the second shell. This gives an electron configuration of $1s^22s^22p^2$. This means that the number of valence electrons is four. Since there are only four valence electrons, they will be equally spaced around the atomic symbol. Furthermore, carbon is likely to be the central atom in the electron dot diagram because it has the greatest number of unpaired valence electrons.

Oxygen has an atomic number of 8 which means in a neutral atom there are eight electrons. Using our knowledge of electronic configuration, there will be two electrons in the first shell and six electrons in the second shell. This gives an electron configuration of $1s^22s^22p^4$. Therefore, oxygen has six valence electrons.

Since the carbon atom needs four electrons (has four **unpaired electrons**) to satisfy the octet rule, and each oxygen atom needs two electrons (has two unpaired electrons) to satisfy the octet rule, each oxygen can form a double bond with the central carbon atom as shown below.



After arranging this into an electron dot diagram, the final answer is:

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Structural formulas offer a clearer approach to representing molecules but do not provide information about valence electrons, only the bonds, as shown in figure 5. One line represents a shared pair of electrons or a single covalent bond.

The **molecular formula** depicts the actual number of each type of element present in a compound (figure 6). Molecular formulas do not provide any information about structure or bonding.

Table 1 contains the Lewis structures, structural formulas, and molecular formulas for a number of common molecules.

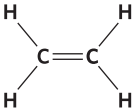
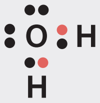
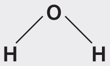
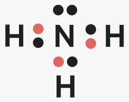
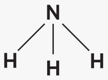
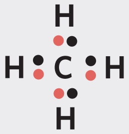
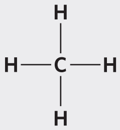
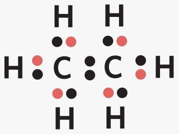
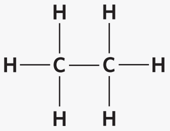
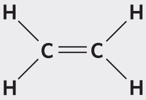


Figure 5 The structural formula of ethene

C_2H_6

Figure 6 The molecular formula of ethane

Table 1 Modelling molecules

Element or compound	Lewis structure	Structural formula	Molecular formula
Hydrogen gas **single bond**		H—H	H_2
Oxygen gas **double bond**		O=O	O_2
Chlorine gas **single bond**		Cl—Cl	Cl_2
Nitrogen gas **triple bond**		N≡N	N_2
Hydrogen chloride **single bond**		H—Cl	HCl
Carbon dioxide **2 double bonds**		O=C=O	CO_2
Water **2 single bonds**			H_2O
Ammonia **3 single bonds**			NH_3
Methane **4 single bonds**			CH_4
Ethane **7 single bonds**			C_2H_6
Ethene **double bond** **4 single bonds**			C_2H_4

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Progress questions

Question 1

A molecule consists of a

A. single atom.

B. group of atoms bonded together.

Question 2

Electrons are transferred in a covalent bond.

A. True

B. False

Question 3

All atoms obey the octet rule.

A. True

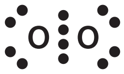
B. False

Question 4

Which of the following depicts a double covalent bond?

A.

B.



Shapes of molecules 1.1.5.1

The shapes of molecules are dictated by the repulsion between electron pairs.

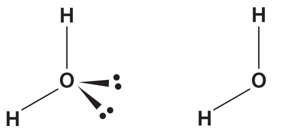
How do electron pairs determine the shape of molecules?

Valence shell electron pair repulsion theory (VSEPR) provides a way of predicting the shapes of molecules according to the electron pairs in the valence shell. Since like charges repel each other, negatively charged pairs of electrons are most stable when they are as far apart as possible. Think about trying to force the negatively charged poles of two magnets together.

The key to learning molecular shapes is to examine the number of **lone pairs** and **bond pairs** of electrons around the central atom (figure 7). Lone pairs of electrons are electrons not involved in covalent bonding, while bond pairs are those that are involved in bonding. Lone pairs of electrons repel bond pairs to a greater extent that a bond pair repels another bond pair. This extra repulsion changes the shape of a tetrahedral structure to pyramidal (1 lone pair and 3 bond pairs) and angular (2 lone pairs and 2 bond pairs).

USEFUL TIP

In three-dimensional molecule structures dotted lines are used to represent bonds going 'into the page', while thick, triangle-shaped lines are used to represent bonds coming 'out of the page'.



What's there **What we see**

Figure 7 The effect of lone pairs on molecular shape

USEFUL TIP

Double (O=O) and triple covalent bonds (N≡N) should be treated as single bonds when considering molecular shape.

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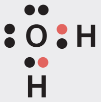
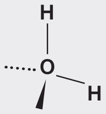
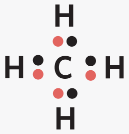
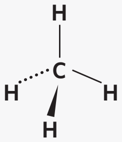
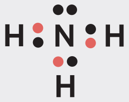
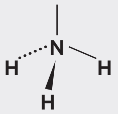
The different shapes that covalent molecules adopt are summarised in table 2.

Table 2 Shapes of molecules

Compound name and formula	Lewis Structure	Number of lone pairs around central atom	Number of bond pairs around central atom	Shape diagram	Shape
Carbon dioxide, CO_2		0	2	O=C=O	Linear
Hydrogen chloride, HCl		3	1	H—Cl	Linear
Water, H_2O		2	2		Angular, V-shaped or Bent
Methane, CH_4		0	4		Tetrahedral
Ammonia, NH_3		1	3		Pyramidal

WORKED EXAMPLE 3

According to VSEPR theory, what is the shape of NI_3?

What information is presented in the question?

The shape of NI_3 needs to be determined using valence shell electron pair repulsion theory.

What is the question asking us to do?

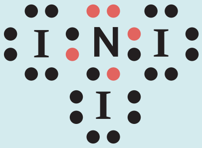
Describe the shape of NI_3.

What strategies do we need in order to answer the question?

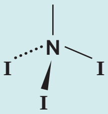
1. Locate the central atom of the molecule.
2. Consider the valence shell of the central atom, how many bond pairs and how many lone pairs are there?
3. Use this information to determine the shape of the molecule.

Answer

Nitrogen[1] is the central atom. It has three bonded pairs and one lone pair in its valence shell.



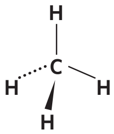
Thus, the shape of the molecule is pyramidal.



Progress questions

Question 5

The shape pictured is



A. tetrahedral.

B. pyramidal.

Question 6

When it comes to determining the shape of a molecule, a double bond is the same as a single bond.

A. True

B. False

KEEN TO INVESTIGATE?

[1] **How does lightning help plants by breaking a triple bond?**
Search YouTube: why is lightning good for agriculture?

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Properties determining polarity 1.1.6.1

By examining the symmetry of any given molecule, we can determine its polarity.

What does the shape of a molecule tell us about its polarity?

A **polar covalent bond** is formed between two covalently bonded atoms with differing electronegativities, as shown in figure 8, while a **non-polar covalent bond** forms between two atoms with similar electronegativity. However, it is not just these bonds that influence a molecule's polarity (figure 9).

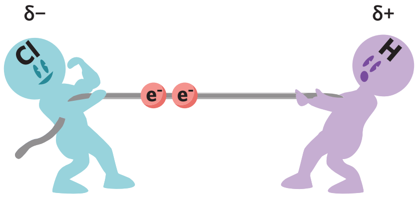

Figure 8 Electronegativity modelled as a tug of war

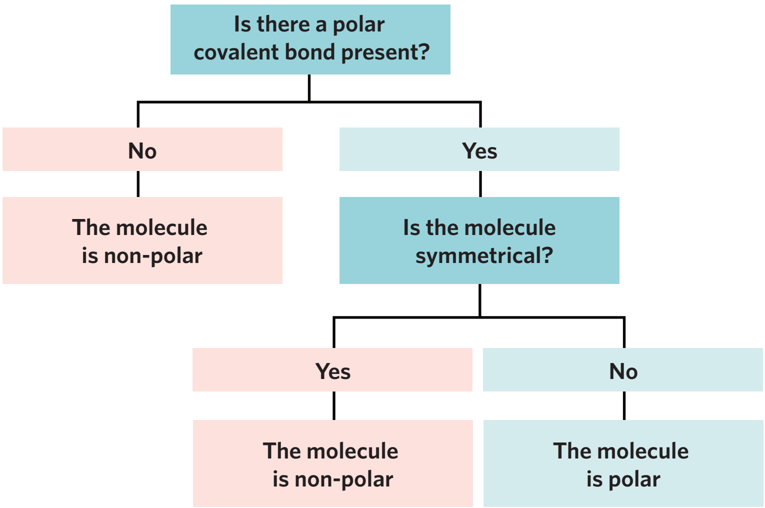

Figure 9 Flowchart for determining polarity

MISCONCEPTION

'A molecule containing a polar covalent bond must be a polar molecule.'

The presence of a polar covalent bond does not necessarily mean that the molecule is polar overall. The positioning of a molecule's **permanent dipoles** must be asymmetrical for it to be classed as polar.

In the case of hydrogen chloride, the electron pair orbits closer to the chlorine atom, due to its greater electronegativity. As a result, the chlorine atom has a greater electron density than the hydrogen atom, causing chlorine to possess a partial negative charge and hydrogen to possess a partial positive charge. These regions are called permanent dipoles. Figure 10 is a common visualisation of this idea with the blue area representing the partially negatively charged, electron-rich region and the red region representing the partially positively charged, electron-deficient region.

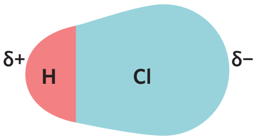

Figure 10 Electron distribution in a polar covalent bond

It is easy to work out if a molecule such as hydrogen chloride is symmetrical and hence whether it is polar or non-polar. Polyatomic molecules can be a bit trickier at first but by examining the permanent dipoles, it can become quite simple. If the permanent dipoles of a molecule cancel each other out, then that molecule is said to be symmetrical and hence, non-polar. However, if the permanent dipoles do not cancel, then the molecule is polar. A summary of this information is provided in table 3.

USEFUL TIP

The lowercase Greek letter delta (δ) is used to represent the partial nature of a charge, followed by the '+' or '−' sign to indicate whether the charge is positive or negative.

MISCONCEPTION

'δ+' has the same charge as a positive ion and 'δ-' has the same charge as a negative ion.'

The delta (δ) symbol represents only a partial charge while an ion has a complete charge.

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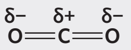
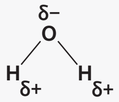
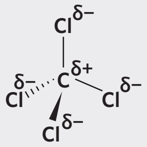
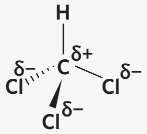
Table 3 Determining polarity using permanent dipoles

Molecule	Permanent dipoles	Symmetry	Polarity
Carbon dioxide		Symmetrical. The permanent dipoles cancel each other out.	Non-polar molecule that has two polar bonds.
Water		Asymmetrical. The permanent dipoles do not cancel.	**Polar molecule** that has two polar bonds.
Tetrachloromethane		Symmetrical. The permanent dipoles cancel each other out.	Non-polar molecule that has four polar bonds.
Trichloromethane		Asymmetrical. The permanent dipoles do not cancel.	Polar molecule that has three polar bonds.

WORKED EXAMPLE 4

HBr and CBr_4 are both compounds which have a highly electronegative bromine atom. One molecule is polar and the other is non-polar. Determine which is polar and which is non-polar and justify your answer.

What information is presented in the question?

The molecules in question are HBr and CBr_4.

What is the question asking us to do?

Identify the polar and the non-polar molecule.

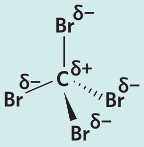
What strategies do we need in order to answer the question?

1. Determine whether there are polar covalent bonds present in either molecule.
2. Determine whether the molecules are symmetrical overall.
3. Establish which molecule is polar and which is non-polar.

Answer

Bromine is a highly electronegative element which means that it forms a polar covalent bond with either hydrogen or carbon. HBr has two permanent dipoles which do not cancel. The molecule is hence polar.

CBr_4 has a tetrahedral shape and even though it consists of four polar covalent bonds, it is non-polar overall due to its symmetry.



Progress questions

Question 7

Non-polar covalent bonds always occur

A. between atoms of the same species.

B. between atoms of different species.

Question 8

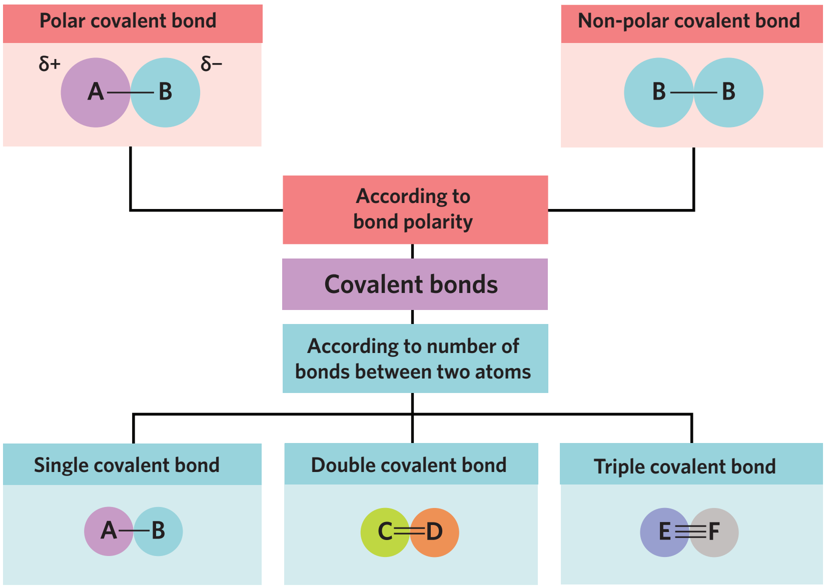
__________ polar molecules have polar covalent bonds but __________ molecules with polar covalent bonds are polar.

A. All, not all

B. Not all, all

Theory summary

- Covalent bonding involves the sharing of electrons between atoms.
- Atoms share electrons to fill their outer shell, satisfying the octet rule.
- Lone and bond pairs of electrons are the key to a molecule's shape.
- Lone pairs of electrons repel bond pairs to a greater extent than a bond pair repels another bond pair.
- Examining the symmetry (or lack thereof) of a molecule's permanent dipoles helps determine if a molecule is polar or non-polar.



2A Questions

Mild Medium Spicy

Deconstructed

Use the following information to answer questions 9-11.

Tom would like to draw the electron dot diagram of a molecule called Tilden's reagent, NOCl (nitrosyl chloride) which is used to classify different alcohols. He is unaware of the number of bonds and whether there are double or even triple bonds present.

Question 9 (1 MARK)

How many valence electrons are there in nitrogen, oxygen and chlorine respectively?

A. 5, 6, 7

B. 6, 7, 5

C. 7, 6, 7

D. 6, 5, 7

Question 10 (1 MARK)

Some atoms have more valence electrons than others. In order to obey the octet rule, which atom needs to have the most electrons shared with it?

A. Nitrogen

B. Oxygen

C. Chlorine

D. All of the above

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Question 11 (2 MARKS)

Based on the above information, draw the electron dot diagram of nitrosyl chloride, NOCl.

Exam-style

Question 12 (7 MARKS)

Hydrogen peroxide (H_2O_2) and oxygen gas (O_2) both have an oxygen atom covalently bonded to another oxygen atom.

a. Identify the type of oxygen-oxygen covalent bond present in O_2. Explain your answer. (2 MARKS)

b. Identify the type of oxygen-oxygen covalent bond present in H_2O_2. Explain your answer (2 MARKS).

c. Hydrogen peroxide undergoes explosive chemical reactions whilst oxygen gas is a stable molecule. With reference to the type of covalent bonds present in both molecules, offer a possible explanation for this dramatic difference in properties. (3 MARKS)

Question 13 (9 MARKS)

For each of the following molecules, draw the electron dot diagram, describe the shape, and state whether the molecule is polar or non-polar.

a. CH_4 (3 MARKS)

b. CO_2 (3 MARKS)

c. NH_3 (3 MARKS)

Question 14 (5 MARKS)

Oxygen difluoride, OF_2 was first noticed in 1929 and reacts with many metals, forming oxides.

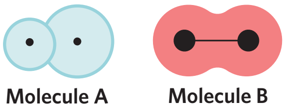
a. Draw the electron dot diagram of oxygen difluoride. (1 MARK)

b. Circle the lone pairs of electrons. What effect do they have on the shape of the molecule? (2 MARKS)

c. Is OF_2 polar or non-polar? If it is polar, annotate the diagram using appropriate conventions to show the permanent dipoles. (2 MARKS)

Question 15 (4 MARKS)

The figures shown represent the electron clouds surrounding atoms of two different diatomic molecules.



Which of the two molecules would be considered polar? Explain.

Key science skills

Question 16 (5 MARKS)

Peter and Paul were investigating the type of bonding present in unknown chemical compounds by using a technique called boiling point determination. In the experiment, they used a machine which boils a small sample of a chemical compound and gives an accurate boiling point value.

The results collected after investigating substance A and B using this machine are shown below.

	Substance A	Substance B
Boiling point (°C)	−183	756

a. Identify the independent variable. (1 MARK)

b. Write an appropriate aim for the experiment. (1 MARK)

c. What type of data is collected in the table? (1 MARK)

d. Which substance, A or B, is more likely to be a diatomic molecule? Explain your answer. (2 MARKS)

FROM LESSONS 16A & 16B

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Questions from multiple lessons

Question 17 (5 MARKS)

The periodic table acts like a list that shows all of the elements that currently exist. Consider the following elements:

Ca Vanadium Phosphorus Si Be

a. Which elements would be found in the same period? (1 MARK)

b. Which elements would be found in the same group? (1 MARK)

c. Which element would most likely have the greatest atomic radius? (1 MARK)

d. Which element would have the lowest first ionisation energy? (1 MARK)

e. Which element would have the most metallic character? (1 MARK)

FROM LESSONS 1B & 1C

Hints

12a. The type of covalent bond depends on the number of electrons shared in the bond.

12b. The type of covalent bond depends on the number of electrons shared in the bond.

12c. Double covalent bonds are stronger than single covalent bonds.

13a. The periodic table can help to determine the number of valence electrons in each atom and the central atom is the atom with the greatest number of unpaired valence electrons.

13b. The periodic table can help to determine the number of valence electrons in each atom and the central atom is the atom with the greatest number of unpaired valence electrons.

13c. The periodic table can help to determine the number of valence electrons in each atom and the central atom is the atom with the greatest number of unpaired valence electrons.

14a. The periodic table can help to determine the number of valence electrons in each atom.

14b. Lone pairs of electrons are those that are not involved in bonding.

14c. Polar molecules are asymmetrical in shape.

15. Polarity depends on the distribution of electron density.

16a. The independent variable is deliberately manipulated (or changed) by the experimenter.

16b. The aim is the purpose/objective of an experiment.

16c. Qualitative data relates to observations while quantitative data is numerical.

16d. Diatomic molecules are gases at room temperature.

17a. Periods are viewed as rows in the periodic table.

17b. Groups are viewed as columns in the periodic table.

17c. The atomic radius increases down a group.

17d. The first ionisation energy decreases down the group.

17e. Metallic character increases down a group.

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2B Intramolecular bonding and intermolecular forces

Image: Light and Vision/Shutterstock.com

Why does the water form a mask over this swimmer's face?

In the previous lesson, we learned about the covalent bonds that hold atoms together to form molecules. In this lesson, we will explore the forces that hold adjacent molecules together.

KEY TERMS AND DEFINITIONS

Boiling point	temperature at which a liquid transitions into a gas
Hydrogen bonding	a special type of permanent dipole-dipole attraction between molecules (not a covalent bond) that results from the attractive force between a hydrogen atom covalently bonded to a very electronegative atom such as an N, O, or F atom
Instantaneous dipole	partial positive and negative charge that appears in an atom/molecule due to the random movement of electrons
Intermolecular forces	forces that attract adjacent molecules to each other
Kinetic energy	the level of movement of a particle, molecule or molecular substance
Melting point	temperature at which a substance changes state from solid to liquid
Non-conductivity	the inability of a molecule to conduct electricity
Molecular substance	substance with two or more non-metal atoms joined by covalent bonds
Permanent dipole-dipole attraction	electrostatic force of attraction between the permanent dipoles of adjacent polar molecules
Temporary dipole-dipole attraction (dispersion force)	electrostatic force of attraction between two instantaneous dipoles of adjacent molecules

STUDY DESIGN DOT POINTS

- the relative strengths of intramolecular bonding (covalent bonding) and intermolecular forces (dispersion forces, dipole-dipole attraction and hydrogen bonding)
- physical properties of molecular substances (including melting points and boiling points and non-conduction of electricity) with reference to their structure

2A 2B 2C

1.1.7.1	Covalent bonding
1.1.7.2	Temporary dipole-dipole attractions (dispersion forces)
1.1.7.3	Permanent dipole-dipole attractions
1.1.7.4	Hydrogen bonding

ESSENTIAL PRIOR KNOWLEDGE

2A	Covalent bonding
2A	Polar bonds

See questions 12–13.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Covalent bonding 1.1.7.1

Intramolecular bonds, such as covalent bonds, are the bonds between the atoms of a molecule, while **intermolecular forces** are the forces of attraction between molecules.

What are the relative strengths of intramolecular (covalent) bonds?

There are two factors we are going to examine that affect the strength of a covalent bond – bond length, and the number of electrons shared. In table 1, we can see that going down group 17 (the halides), hydrogen forms progressively weaker covalent bonds.

MISCONCEPTION

'Intermolecular forces can also be referred to as intermolecular bonds.'

The only types of bond are covalent, ionic and metallic. The interactions between molecules are called forces, as they are much weaker than the aforementioned bonds.

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This is because the bonding electrons are one shell further out as we move down the group. This makes the bond length longer and the electrostatic forces of attraction holding the bond together weaker.

Table 1 Bond strength in the hydrogen halides

Halide	Formula	Energy to break bond (kJ mol^{-1})	Bond length (picometres) (pm)
Hydrogen fluoride	HF	568	91
Hydrogen chloride	HCl	432	127
Hydrogen bromide	HBr	366	141
Hydrogen iodide	HI	298	161

Table 2 highlights the relationship between number of electrons shared and bond strength. With more bonding electrons, the forces pulling the two atoms together are going to be stronger. This increases the energy needed to break the bond and decreases the bond length.

Table 2 Number of bonding electrons influencing bond strength

Bond type	Energy to break bond (kJ mol^{-1})	Bond length (pm)
C-C	348	154
C=C	614	134
C≡C	839	120

Progress questions

Question 1

Intramolecular bonds are

A. the bonds between the atoms of a molecule.

B. the forces of attraction between molecules.

Question 2

The longer the covalent bond, the __________ the bond. The more electrons shared in a covalent bond, the __________ the bond.

A. stronger, weaker

B. weaker, stronger

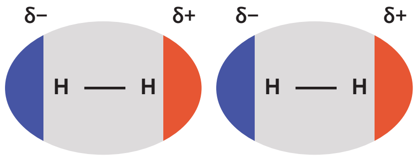

Figure 1 Temporary dipole-dipole attraction between two molecules

MISCONCEPTION

'Not all covalent molecular substances have dispersion forces.'

All covalent molecules have dispersion forces. Even if other types of intermolecular forces are operating, they can all exist simultaneously.

Temporary dipole-dipole attractions (dispersion forces) 1.1.7.2

Due to their constant movement, there are moments when electrons in a molecule will concentrate in one region, creating an area of slightly negative charge and one with a slightly positive charge. This is called an **instantaneous dipole**.

What are the relative strengths of dispersion forces?

An instantaneous dipole in one molecule will induce an instantaneous dipole of opposite charge in an adjacent molecule, creating a force of electrostatic attraction between the two molecules. This is called a **temporary dipole-dipole attraction** or **dispersion force**, as shown in figure 1.

The more electrons a molecule has, the greater the possibility that an instantaneous dipole will form and induce another dipole in a nearby molecule. This means that there are stronger dispersion forces between larger molecules (greater surface area with more electrons) than smaller molecules (smaller surface area with fewer electrons) as shown in figure 2.

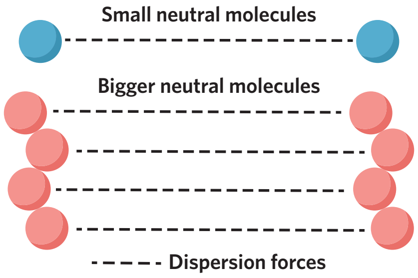

Figure 2 Dispersion forces between molecules of different size

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WORKED EXAMPLE 1

Methane and ethane are both hydrocarbons. Methane has the molecular formula CH_4 and ethane has the molecular formula C_2H_6. Which molecule would have the strongest dispersion forces and why?

What information is presented in the question?

The molecular formulas of methane and ethane.

What strategies do we need in order to answer the question?

1. Use the periodic table to determine the number of electrons in each molecule.
2. Use the information from step 1 to determine which molecule has the strongest dispersion forces.
3. Answer the question with a justifying statement.

Answer

According to the periodic table, carbon has 6 electrons (atomic number 6) and hydrogen has 1 electron (atomic number 1). Therefore, methane (CH_4) has $1 \times 6 + 4 \times 1 = 10e^-$ and ethane (C_2H_6) has $2 \times 6 + 6 \times 1 = 18e^-$. Since ethane has more electrons than methane and a greater surface area, ethane will have much stronger dispersion forces between molecules than methane.

Progress questions

Question 3

When determining the relative strength of dispersion forces, it can be useful to determine the

A. number of electrons.

B. number of neutrons.

Question 4

Dispersion forces arise from

A. permanent dipole moments caused by polar covalent bonds.

B. instantaneous dipole moments caused by randomly moving electrons.

Permanent dipole-dipole attractions 1.1.7.3

Permanent dipole-dipole attractions are another type of intermolecular force that are formed from the electrostatic attraction between polar molecules (figure 3).

What are the relative strengths of permanent dipole-dipole forces?

Permanent dipole-dipole attractions only occur between polar molecules and not between symmetrical non-polar molecules with polar bonds. Recall from 2A that non-polar molecules do not have permanent dipoles, only instantaneous dipoles.

Permanent dipole-dipole attractions are considerably harder to disrupt than dispersion forces, as the attractive forces between molecules are greater. The strength of a permanent dipole-dipole attraction depends on the difference in electronegativity of the atoms involved in the bond.

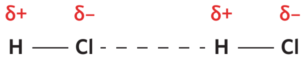

Figure 3 Permanent dipole-dipole forces between HCl molecules

WORKED EXAMPLE 2

HBr and CBr_4 both have a highly electronegative bromine atom. With respect to their intermolecular forces of attraction, which molecule would have the strongest permanent dipole-dipole attractive forces?

What information is presented in the question?

The molecules in question are HBr and CBr_4.

What is the question asking us to do?

Compare the strength of each molecule's dipole-dipole forces.

What strategies do we need in order to answer the question?

1. Compare the electronegativities of the atoms covalently bonded together to determine whether there are polar covalent bonds present in either molecule.
2. Determine whether the molecules are polar overall using symmetry.
3. Answer the question with a justifying statement.

Answer

Bromine is a highly electronegative element which forms polar covalent bonds with hydrogen or carbon.

HBr has a linear shape which means overall it is a polar molecule.

CBr_4 has a tetrahedral shape and even though it consists of four polar covalent bonds, it is non-polar overall due to symmetry.

Therefore, since HBr is a polar molecule, it can form permanent dipole-dipole forces, whereas CBr_4 is a non-polar molecule with polar bonds, meaning it can only form dispersion forces (instantaneous dipole-dipole forces). Hence, HBr has the strongest permanent dipole-dipole interactions.

MISCONCEPTION

'Hydrogen bonds are intramolecular bonds and as strong as ionic, covalent, and metallic bonds.'

Hydrogen bonding is a type of intermolecular force, not an intramolecular bond. Therefore, hydrogen bonds are not really bonds, despite the name.

MISCONCEPTION

'All molecules containing oxygen and hydrogen form hydrogen bonds.'

In order for hydrogen bonds to form between molecules, a hydrogen must be directly bonded to a nitrogen, oxygen or fluorine atom.
Remember the 'N.O.F gang'!

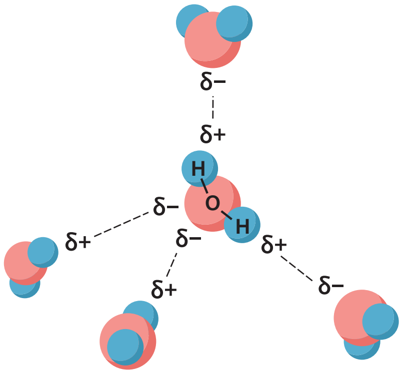

Figure 4 Hydrogen bonding

Progress questions

Question 5

Permanent dipole-dipole forces are ___________ than dispersion forces and occur between ___________ molecules.

A. stronger, polar

B. weaker, non-polar

Question 6

A permanent dipole-dipole attraction can occur between two non-polar molecules.

A. True

B. False

Hydrogen bonding 1.1.7.4

Hydrogen bonding is a type of permanent dipole-dipole attraction that only occurs between molecules in which hydrogen is bonded directly to one of nitrogen, oxygen or fluorine.

What are the relative strengths of hydrogen bonding forces?

Hydrogen bonding occurs between water molecules as shown in figure 4. In all instances of hydrogen bonding, the hydrogen atom has a partial positive charge and is attracted to a nitrogen, oxygen or fluorine atom with a partial negative charge on an adjacent molecule.

The considerable difference in electronegativity between hydrogen and these elements creates a highly polar bond. In turn, this means that the electrostatic attraction between the permanent dipoles is stronger than those observed in typical dipole-dipole attractions. In fact, this difference in electronegativity is so pronounced that the positive nucleus of hydrogen is exposed when bonded to one of nitrogen, oxygen or fluorine. This is what makes the dipole-dipole interactions in hydrogen bonding the strongest of all intermolecular forces.

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Progress questions

Question 7

Hydrogen bonding is the strongest form of permanent dipole-dipole forces.

A. True

B. False

Question 8

Given that fluorine is more electronegative than chlorine, would hydrogen fluoride or hydrogen chloride have the strongest intermolecular forces?

A. Hydrogen fluoride

B. Hydrogen chloride

Properties of covalent compounds 1.1.8.1

The intermolecular forces between molecules are responsible for a number of different physical properties including **melting point**, **boiling point** and **non-conductivity**. These forces are also responsible for many real world phenomena, including the Water Strider's ability to walk on water.[1]

How do intramolecular bonding and intermolecular forces affect the physical properties of covalent compounds?

The strength of intermolecular forces reflects the amount of energy required to break the forces between molecules. Molecules that exist in the solid state have low **kinetic energy** and therefore cannot overcome the intermolecular forces that are holding them together. When there is enough energy to overcome these forces, the substance is able to transition into a liquid, which has weaker intermolecular forces. The temperature at which this occurs is known as the melting point. Finally, with enough kinetic energy, molecules can break all intermolecular forces to exist as a gas. When substances change from a liquid to a gaseous state, this is known as the boiling point which is shown in figure 5. This idea is further explored in table 3.

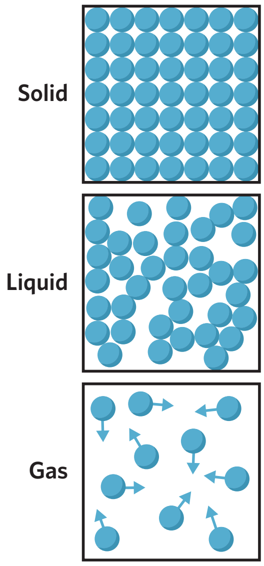
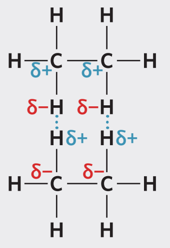
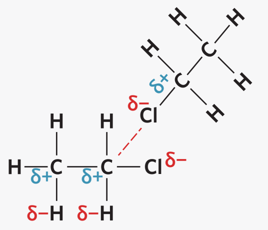
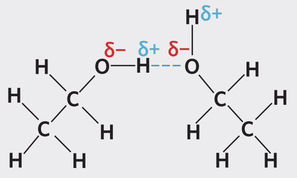


Figure 5 Arrangement of molecules in different states of matter

KEEN TO INVESTIGATE?

[1] **How can the Water Strider walk on water?**
Search: How Water Striders Walk on Water

Table 3 The influence of intermolecular forces on the physical properties of molecules

Molecule	Diagram	Intermolecular forces	State at room temperature (25°C)	Boiling point (°C)
Ethane, C_2H_6		Dispersion forces	Gas	−89.0
Chloroethane, C_2H_5Cl		Dispersion forces, permanent dipole-dipole attraction	Gas	12.3
Ethanol, C_2H_5OH		Dispersion forces, hydrogen bonding	Liquid	78.4

WORKED EXAMPLE 3

Explain why methanol (CH_3OH) has a higher boiling point than methane (CH_4) with reference to intermolecular forces.

What information is presented in the question?

The molecules in question are CH_3OH and CH_4.

What is the question asking us to do?

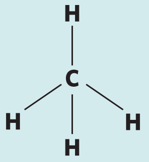
Compare the difference in boiling points with respect to intermolecular forces.

What strategies do we need in order to answer the question?

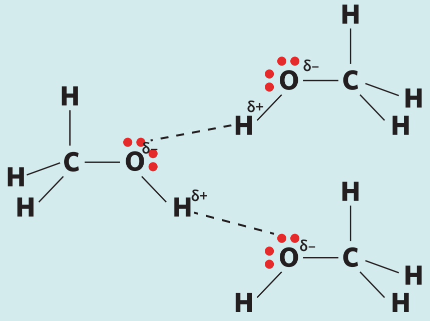
1. Identify the intermolecular forces present in each molecule.
2. Relate the strength of intermolecular forces to the boiling point.
3. Answer the question with a justifying statement.

Answer

Methane is a non-polar molecule as it is symmetrical.



Methanol is a polar molecule because it is asymmetrical and has a polar covalent bond that occurs between the oxygen and hydrogen in the OH group. This means that the highly electronegative oxygen atom is pulling electrons away from the exposed nucleus of the hydrogen atom. As a result, there is a permanent dipole-dipole interaction. Furthermore, the partially positive hydrogen atom on one molecule is able to form a hydrogen bond with a partially negative oxygen atom on a different molecule.



Although both molecules have dispersion forces of roughly equal strength, methanol has hydrogen bonding as well. As a result, methanol will have stronger intermolecular forces and, therefore, a higher boiling point (by 226.3 °C) as more energy is required to break the hydrogen bonds between the methanol molecules. Methanol has a boiling point of 64.7 °C, whereas methane's boiling point is –161.6 °C.

The strength of intermolecular forces is approximately 100 times less than that of intramolecular bonding like covalent bonds or ionic bonds. Table 4 uses water to illustrate this point.

Table 4 Intramolecular bonds vs. intermolecular forces strength comparison

Molecule	Description of bond/force	Temperature at which bonds/forces are broken
Water, H_2O	Intramolecular covalent bonds between hydrogen and oxygen atoms	3000 °C
	Intermolecular hydrogen bonding between hydrogen (δ+) and oxygen (δ−) atoms	100 °C

Generally, covalent molecular substances do not conduct electricity. This is because there are no free or mobile charged particles. However, there are exceptions which will be discussed in lesson 2C.

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Progress questions

Question 9

Water has remarkable properties including a high boiling point for its size because

A. it has no dispersion forces.

B. it has hydrogen bonding.

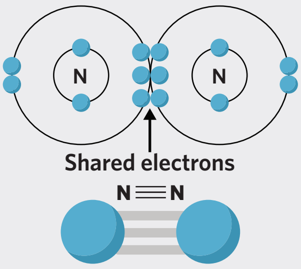
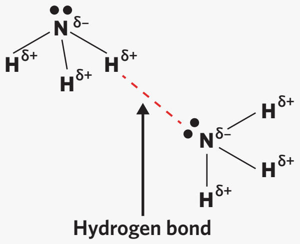
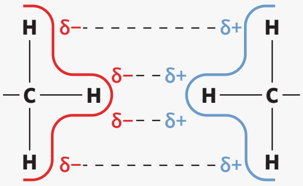
Question 10

Molecules with a __________ surface area have __________ interactions with other molecules.

A. small, stronger

B. large, stronger

Theory summary

Bond type (strongest to weakest)	Visual	Physical properties
Covalent (intramolecular)		N/A
Hydrogen bonding (intermolecular)		Highest melting and boiling points
Permanent dipole-dipole attraction (intermolecular)		Intermediate melting and boiling points
Dispersion forces (intermolecular)		Lowest melting and boiling points

The content in this lesson is considered fundamental prior knowledge to structure, nomenclature and properties of organic compounds (Unit 4 AOS 1).

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2B Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

The following information is to be used for questions 11–13.

Depending on environmental conditions, molecules can exist in many different forms. For example, water on the planet of 55 Cancri e, has an average temperature of 1400 °C and therefore exists predominantly as vapour. Water, however, exists on earth predominantly in liquid form.

Consider the following statement:

At room temperature, F_2 and Cl_2 are gases, Br_2 is a liquid, and I_2 is a solid.

Question 11 (1 MARK) 🌶

The intermolecular forces acting between the different molecules are

A. dipole-dipole interactions.
B. hydrogen bonds.
C. dispersion forces.
D. ionic bonds.

Question 12 (1 MARK) 🌶

The main difference between these molecules is

A. the type of intramolecular bonds holding the molecule together.
B. the type of intermolecular forces holding the molecules together.
C. the size of the atoms.
D. the strength of the intramolecular bonds holding the molecule together.

Question 13 (3 MARKS) 🌶🌶

Explain why the molecules above exist in different physical states at room temperature.

Exam-style

Question 14 (3 MARKS) 🌶

Consider the following compounds: CF_4, CCl_4, CBr_4.

a. Which has the strongest intramolecular bonds? (1 MARK)

b. Which has the strongest dispersion forces between molecules? (1 MARK)

c. Which has the highest boiling point? (1 MARK)

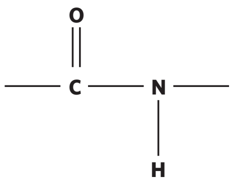
Question 15 (6 MARKS) 🌶🌶

Proteins are large organic molecules that are important in biology and their function is heavily linked to their structure. One of the important features of proteins is their ability to form hydrogen bonds to stabilise their structure.

a. Why might hydrogen bonds, as opposed to dispersion forces, be important in stabilising protein structure? (2 MARKS)

b. The bond that connects the components of a protein together is called the peptide bond and its structure is shown. It consists of a carbon-oxygen double bond and a nitrogen-hydrogen single bond. Draw a possible hydrogen bonding interaction between two different peptide bonds and label the hydrogen bond. (2 MARKS)

c. Suggest one factor that could disrupt the stability of the interactions between two peptide bonds. How do you think this would affect the function of the protein? (2 MARKS)



Peptide (also called amide) bond

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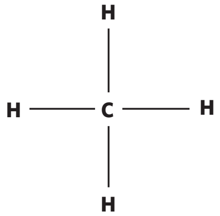
Question 16 (6 MARKS)

Water is one of the most important substances in the world. Not only is it required to maintain bodily functions, it is also involved in many agricultural and material production processes.

a. Draw and label hydrogen bonding occurring between four water molecules (in the liquid state). (2 MARKS)



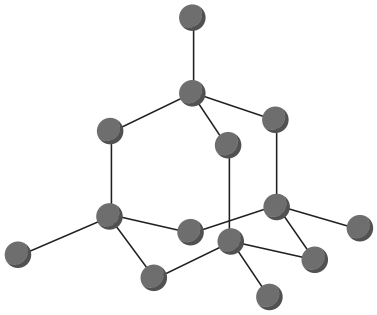
b. Consider the molecule shown. It has a similar molecular mass to water, however the melting point of this molecule is −182 °C whereas the melting point of water is 0 °C. Explain this difference in melting points with respect to the intermolecular forces present. (2 MARKS)

c. A student predicted that tetradecane ($C_{14}H_{30}$) which only has dispersion forces would also have a lower melting point than water. This, however, is not the case. Explain how this is possible. (2 MARKS)

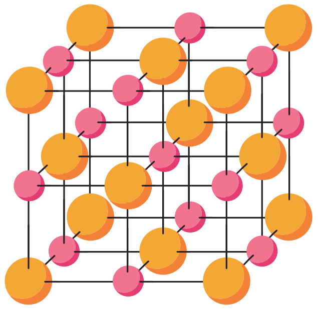
Key science skills

Question 17 (5 MARKS)

Kath and Kim wanted to conduct an experiment to see the differences in boiling points of two different substances containing the two structures shown.



Substance A



Substance B

It is known that substance A is a covalent compound, whereas substance B is an ionic compound. When conducting the experiment, the students used 5 g of substance A and 4 g of substance B.

The results of the experiment are shown:

	Substance A	Substance B
Boiling point (°C)	145	120

a. Identify the independent variable. (1 MARK)

b. Write an appropriate aim for the experiment. (1 MARK)

c. Comment on the validity of the results of the experiment. (3 MARKS)

FROM LESSONS 16A & 16D

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Questions from multiple lessons

Question 18 (9 MARKS)

Trichloromethane, $CHCl_3$, also known as a chloroform, was a common sedative used up until the 1960s.

a. Draw the three-dimensional structure of trichloromethane, labeling any partial positive and negative charges. (2 MARKS)

b. List all the intermolecular forces that are present in liquid trichloromethane and explain how these forces arise. (4 MARKS)

c. Chloroform is a dense, yet volatile liquid. With respect to the strength of the intermolecular forces present, justify why it is a liquid at room temperature. (3 MARKS)

FROM LESSON 2A

Hints

14a. The longer the covalent bond, the less energy it takes to break it.

14b. The molecule with the most electrons will have the strongest dispersion forces.

14c. The compound with the strongest intermolecular forces will have the highest boiling point.

15a. Hydrogen bonds and dispersion forces differ greatly in strength.

15b. A hydrogen bond occurs between a hydrogen atom bonded to one of N, O, F and the lone pair of electrons on another highly electronegative atom.

15c. Intermolecular forces can be disrupted with temperature.

16a. Water is a highly polar compound.

16b. Water molecules in ice adopt a crystalline structure.

16c. Melting point is affected by the type(s) of intermolecular forces present.

17a. The independent variable is the variable that is manipulated.

17b. The aim indicates the goal of an experiment.

17c. The validity of an experiment is affected by the variables changed in the experiment.

18a. There are four atoms around the central carbon meaning it will adopt a tetrahedral shape.

18b. Dispersion forces are present in every molecule.

18c. State of matter and intermolecular forces are interrelated.

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2C Macromolecules

Image: Napolenka/Shutterstock.com

How can these materials both be made of pure carbon?

Many covalent compounds exist as liquids and gases at SLC. In this lesson, we will be introducing carbon **allotropes** and explaining why these are solid and the differences in their structures and bonding.

KEY TERMS AND DEFINITIONS

Allotrope one of the different physical forms an element can take due to different structural arrangements of atoms

Covalent layer lattice atoms covalently bonded together in a two-dimensional arrangement to form layers, held together by weak dispersion forces

Covalent network lattice regular three-dimensional arrangement of atoms covalently bonded together extending outwards

Lattice a series of atoms arranged in a regular pattern

Macromolecule molecule containing a very large number of atoms such as a giant atomic covalent lattice of diamond or graphite

Tetravalent describes an atom with four electrons in its valence shell

Carbon 1.1.9.1

Carbon is an extremely versatile element, found in abundance across the universe.

Why is carbon such a versatile element?

One of the key characteristics of carbon is that it is **tetravalent**. Carbon atoms can form up to 4 covalent bonds (single, double and triple) with different elements. Often, carbon atoms form covalently bonded chains, which act as the backbone of molecular substances. There are also many complex three-dimensional structures that carbon can take on.

There are a number of pure forms in which carbon can exist, including graphite and diamond. These structures are giant atomic **lattices**, called **macromolecules**. They contain only carbon and differ only in their spatial orientations. The different structures an element can take on are called allotropes.[1]

Allotropes are not exclusive to carbon based structures – figure 1 shows various allotropes of sulfur.

STUDY DESIGN DOT POINT

- the structure and bonding of diamond and graphite that explain their properties (including heat conductivity and electrical conductivity and hardness) and their suitability for diverse applications

2A 2B 2C

1.1.9.1 Carbon

1.1.9.2 Carbon lattice structures

ESSENTIAL PRIOR KNOWLEDGE

2A	Covalent bonding
2B	Melting point

See questions 14–15.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

KEEN TO INVESTIGATE?

[1] **What is a buckyball of carbon?**
Search YouTube: Buckyball: Tiny Carbon Soccer Balls

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MISCONCEPTION

'Macromolecules are strictly molecules related to biological processes, such as nucleic acids and proteins.'

The term 'macromolecule' is not exclusive to biology. It essentially means 'giant molecule.' In this case, macromolecules describe giant atomic lattices.

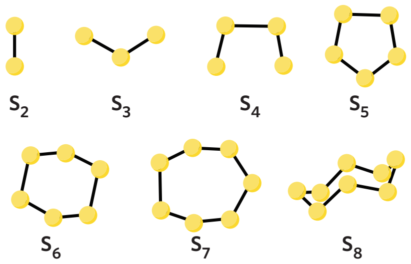


Figure 1 Allotropes of sulfur

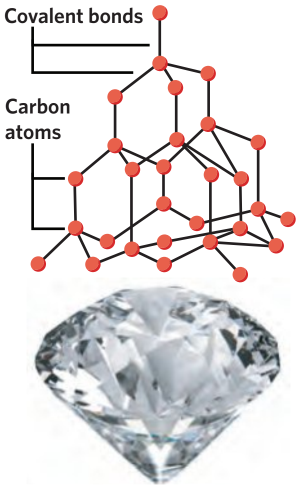


Image: magicoverlay/Shutterstock.com

Figure 2 Different representations of diamond

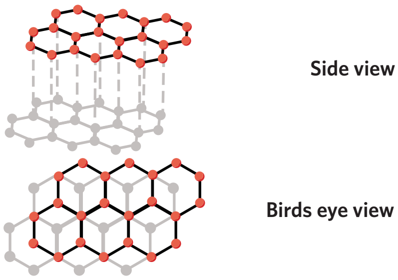


Figure 3 Structure of graphite

MISCONCEPTION

'Only metals conduct electricity.'

Various non-metals that have free mobile electrons are able to conduct electricity, e.g. graphite.

Progress questions

Question 1

Tetravalent atoms contain

A. four valence electrons.

B. six valence electrons.

Question 2

Carbon allotropes are

A. molecules with different amounts of carbon present.

B. molecules consisting of carbon atoms arranged differently.

Carbon lattice structures 1.1.9.2

Carbon atoms can form giant atomic lattices. We will explore some of these carbon allotropes throughout this lesson.

How are the structures of diamond and graphite made entirely of carbon?

Perhaps one of the most well known examples of a pure carbon compound is diamond. Diamond is made up of repeating units of carbon organised into a three-dimensional tetrahedral lattice structure, as shown in figure 2. Each carbon atom in the lattice is bonded to four other carbon atoms (single covalent bonds). The overall structure of diamond is called a **covalent network lattice**.

Graphite, like diamond, is an allotrope of carbon, however its structure is very different to diamond's. Graphite's structure consists of layers of hexagonal, two-dimensional carbon lattices held together by weak dispersion forces, as seen in figure 3. Within each layer, carbon atoms are covalently bonded to other carbon atoms, much like diamond. Therefore, graphite is referred to as a **covalent layer lattice**.

It is also important to understand that there is a substantial distance between the layers, due to the relatively weak dispersion forces. These weak dispersion forces between layers allow the layers to easily slide over each other which makes graphite a useful lubricant.

What are the properties and real world applications of diamond and graphite?

Table 1 Properties of diamond

Property	Explanation
High melting point	A lot of energy is required to disrupt the covalent bonds of the lattice.
Hard	The rigid 3-D tetrahedral lattice structure of diamond makes diamond hard as when pressure is exerted on the structure it is reduced by spreading it throughout the tetrahedral structure.
Brittle (substance that does not deform before breaking apart)	The rigid lattice structure does not allow diamond to be bent. It will just break.
No electrical conductivity (flow of electricity through the movement of charged particles)	Each carbon takes part in four single covalent bonds. This means that there are no free electrons available to conduct electricity.
High thermal conductivity (ability to transfer heat)	The strong covalent bonds can absorb a lot of heat.
Insoluble (unable to be dissolved)	The strength of the covalent bonds in the structure of diamond cannot be overcome by the intermolecular forces of a solvent like water.

Table 2 Applications of diamond

Application	Property that makes it suitable for this use
Cutting tools	The hardness of diamond resists wear, and enhances durability (figure 4).
Thermal conductor in electrical components	Diamond has high thermal conductivity.
Optical components	Diamond lasers can be produced due to diamond's ability to transmit heat and light very effectively.
Abrasive	Diamond is able to induce friction on other objects without experiencing wear itself due to its hardness.

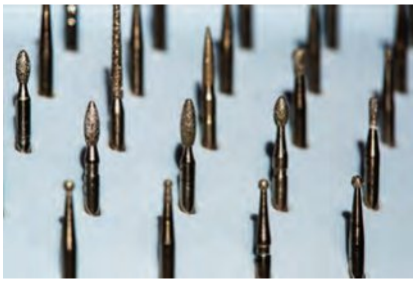


Image: ANDRIY B/Shutterstock.com

Figure 4 Diamond-coated dental tools

The structural differences between graphite and diamond cause these allotropes to have vastly different properties. Notably, graphite is able to conduct electricity. In graphite, each carbon atom is bonded to three other carbon atoms, leaving one unpaired electron. These electrons become delocalised and are able to move throughout the layers of carbon lattice allowing graphite to conduct electricity (figure 5). The other properties of graphite are summarised in table 3.

WORKED EXAMPLE 1

Joe is a mechanical engineer, and he suggests that a circular saw with a diamond blade should be used for cutting concrete. What property of diamond would lead Joe to make this suggestion? With reference to its structure, explain why diamond has this property.

What information is presented in the question?

It has been suggested that a diamond blade should be used for cutting concrete.

What is the question asking us to do?

Determine why a diamond blade would be optimal for this task.

What strategies do we need in order to answer the question?

1. Recall the properties of diamond.
2. Apply the properties of diamond to this application.
3. Recall the structure of diamond, and use it to explain why diamond possesses these properties.

Answer

Diamond is a very hard and durable substance, which would make it ideal for use on a circular saw.

Diamond possesses these properties due to its structure. Diamond is a covalent network lattice with carbon atoms covalently bonded to four other carbon atoms in a network lattice arrangement that extends outwards. Therefore, this lattice structure means that carbon atoms are held together by strong covalent bonds, making diamond hard and rigid.

Table 3 Properties of graphite

Property	Description
High melting point	A lot of energy is required to break the covalent bonds of the lattice.
High thermal conductivity	When heat is applied the hexagonal lattice of covalently bonded atoms are able to vibrate about a very strong fixed position passing on heat energy without having their covalent bonds broken.
High electrical conductivity	Graphite has delocalised electrons that can move through the structure and carry an electrical charge.
High thermal conductivity	Layers of graphite readily slide over each other.
Soft, slippery feeling	Layers of graphite held together by dispersion forces are not as tightly packed as covalently bonded carbon atoms.
Insoluble	The strength of the covalent bonds in the structure of graphite cannot be overcome by the intermolecular forces of a solvent like water.

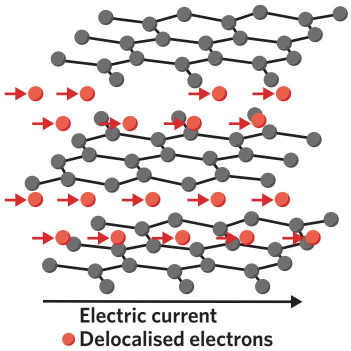


Figure 5 Movement of delocalised electrons found in graphite

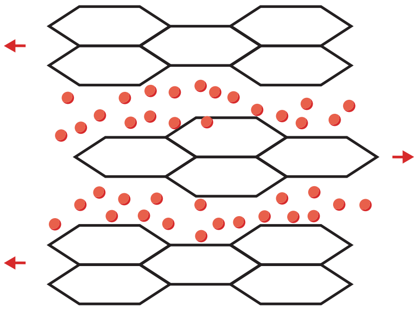


Figure 6 Movement of carbon layers within graphite

One of the most common uses for graphite is in pencils. Layers of graphite can slide past each other with minimal force, easily depositing on paper (figure 6). Other applications of graphite are detailed in table 4.

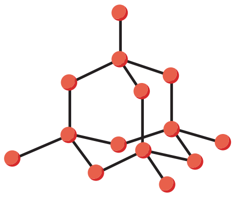
Table 4 Applications of graphite

Application	Property that makes it suitable for this use
Carbon brushes in electrical motors	Graphite is able to conduct electricity, and therefore can transfer current from a stationary wire to moving parts in electrical motors.
Electrode in batteries	Graphite is an inert (unreactive) and electrically conductive material.
Industrial lubricant	Layers of graphite are able to slide over each other, which, when used as a lubricant, reduces friction in machinery.

Progress questions

Question 3

The following structure is an example of



A. graphite.

B. diamond.

Question 4

Graphite is a

A. two-dimensional covalent layer lattice.

B. three-dimensional covalent network lattice.

Question 5

The density of graphite is

A. higher than diamond's, as the lack of covalent bonds decreases the distances between layers.

B. lower than diamond's, as the lack of covalent bonds increases the distances between layers.

Question 6

Both diamond and graphite are good conductors of electricity.

A. True

B. False

Question 7

Both diamond and graphite are insoluble in water.

A. True

B. False

Question 8

The high melting points of diamond and graphite can be explained by

A. the strength of their intramolecular bonds.

B. the presence of delocalised electrons.

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Theory summary

- Carbon molecules can be structured in different ways, known as allotropes, each with their own unique properties.
- Carbon allotropes include diamond and graphite.
- The structure of diamond is a rigid three dimensional network of covalently bonded carbon atoms.
- Diamond is a hard, thermally conductive substance that does not conduct electricity.
- The structure of graphite consists of layers of hexagonal carbon lattices held together by dispersion forces.
- Graphite's weak intermolecular forces enable layers to be rubbed off, and delocalised electrons enable the conduction of electricity.

2C Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 9–11.

A scientist wishes to explore the different properties of diamond and graphite. To do this, he sets up an experiment where a sharpened piece of each of these substances is rubbed against a piece of wood.

Question 9 (1 MARK) 🌶

What would be expected to happen in the case of graphite?

A. A pencil mark is left on the wood.

B. The graphite crumbles due to friction.

C. A cut is made into the wood.

D. A thin layer of wood remains on the graphite.

Question 10 (1 MARK) 🌶

Which would be more likely to happen in the case of diamond?

A. A pencil mark is left on the wood.

B. The diamond crumbles due to friction.

C. A cut is made into the wood.

D. A thin layer of wood remains on the diamond.

Question 11 (4 MARKS) 🌶

Explain the expected effect on the wood when diamond and graphite are rubbed against it. Make reference to the structures of graphite and diamond, and how these contribute to their different properties.

Exam-style

Question 12 (7 MARKS) 🌶🌶

Electric motors use carbon brushes in order to conduct electrical current from a stationary wire to a moving component in the motor. Graphite is often used to make these carbon brushes.

a. Describe the property of graphite that enables it to be used for this purpose. (3 MARKS)

b. Diamond is another material being considered for use as a carbon brush, however this is quickly ruled out. Provide two reasons why diamond would be ineffective for use as a carbon brush. (2 MARKS)

c. Extremely high temperatures are able to be generated in electric motors. Would this compromise the functionality of the graphite carbon brushes? Explain your answer with reference to graphite's structure. (2 MARKS)

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Question 13 (8 MARKS)

In dentistry, tools tipped with diamonds are used to drill and polish cavities. This is because the presence of diamonds on the tip enhances the abrasive property of these tools, or makes the tool more effective at shaping or polishing cavities through friction.

a. Draw the atomic structure of diamond, showing bonding between atoms. (2 MARKS)

b. With reference to the structure of diamond, explain why diamond is an effective abrasive. (2 MARKS)

c. Evaluate the effectiveness of graphite for use in dental tools. Compare its effectiveness to that of diamond. (4 MARKS)

Key science skills

Question 14 (3 MARKS)

A team of laboratory clinicians decide to carry out an experiment comparing the properties between four carbon allotropes – diamond, graphite, graphene and fullerenes. The team set up a series of experiments to assess the electrical and thermal conductivity of each of the allotropes.

The team of laboratory clinicians knew before conducting the experiment that most of the carbon allotropes would be good electrical conductors as well as good conductors of thermal energy. Upon analysis of the results, the results on thermal conductivity were as expected, however the results showed that none of the allotropes conduct electricity. The clinicians concluded that an error must have been made.

a. What type of error has likely taken place? (1 MARK)

b. Suggest a way of avoiding such errors. (1 MARK)

c. The clinicians repeated the thermal conductivity tests. Again, the results centered around the expected temperature value, however, there was significant deviation between the trials.

What can be concluded about the results? (1 MARK)

FROM LESSON 16D

Questions from multiple lessons

Question 15 (4 MARKS)

Diamonds are formed under extreme pressure and heat in sedimentary rock. A chemistry student proposes that, because these conditions are difficult to replicate, diamond should be considered a critical element.

a. Evaluate the student's proposal. (3 MARKS)

b. Phosphorus is considered a critical element. Identify one way in which it can be used more sustainably. (1 MARK)

FROM LESSON 1D

Hints

12a. Electric currents involve the movement of electrons.
12b. The structure of diamond is different to that of graphite.
12c. The structure of a compound impacts the ability of the substance to withstand heat.
13a. Diamond has a strong carbon lattice.
13b. Diamond is a strong material.
13c. Graphite and diamond are structurally different substances.
14a. Different types of errors affect results differently.
14b. Changes need to be made to ensure the accuracy of the experiment.
14c. Consider accuracy and precision.
15a. Consider what a critical element is.
15b. Think about ways in which phosphorus could be recycled.

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Chapter 2 review

Multiple choice (10 MARKS)

Question 1 (1 MARK)

Which of the following statements regarding covalent bonding is true?

A. Covalent bonding results from the electrostatic attraction between cations and anions.

B. Electrons are transferred between atoms in covalent bonds.

C. Electrons are evenly distributed between atoms in all covalent bonds.

D. The more electrons shared in a covalent bond, the stronger the bond.

Question 2 (1 MARK)

Which of the following statements regarding molecular representations is correct?

A. Electron dot diagrams represent the bonds between atoms with lines.

B. Structural formulas indicate the valence electrons of each atom.

C. Molecular formulas depict the actual number of atoms in a molecule.

D. Lewis diagrams represent bonded electrons and lone electrons differently.

Question 3 (1 MARK)

Identify the option which incorrectly matches a molecule to its shape according to VSEPR.

	Molecule	Shape
A.	NF_3	Trigonal pyramidal
B.	CO_2	Linear
C.	C_2H_2	Bent
D.	NH_4^+	Tetrahedral

Question 4 (1 MARK)

The following molecules all contain polar covalent bonds. Which of the following have a permanent dipole?

I. CCl_4

II. H_2O

III. NH_3

IV. BeF_2

A. I, II, II, IV

B. I, IV

C. I, II, III

D. II, III

Question 5 (1 MARK)

Which molecule has the strongest dispersion forces?

A. Hydrogen

B. Carbon dioxide

C. Ethane

D. Water

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Question 6 (1 MARK)

Which of the following statements is not correct?

A. Permanent dipole-dipole forces vary in strength depending on the difference in electronegativity of atoms present.

B. Hydrogen bonds are a type of permanent dipole-dipole force.

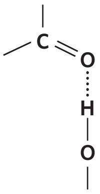
C. Permanent dipole-dipole forces occur whenever a polar covalent bond is present.

D. In both dispersion forces and permanent dipole-dipole forces, there is a partial electrostatic attraction between molecules.

Question 7 (1 MARK)

Which of the following involves a hydrogen bond?

1.



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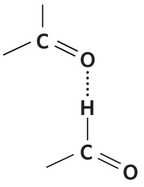

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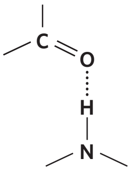
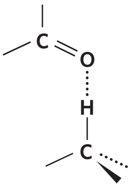


4.



A. 1, 2, 3 and 4

B. 1 and 2

C. 1, 2, and 3

D. 1 and 3

Question 8 (1 MARK)

Which of the following lists the chemical species in order of decreasing boiling point?

A. H_2O, CH_4, HCl, H_2

B. H_2O, HCl, CH_4, H_2

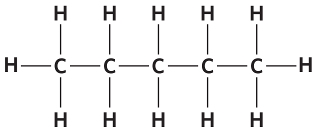
C. H_2, HCl, CH_4, H_2O

D. H_2, CH_4, HCl, H_2O

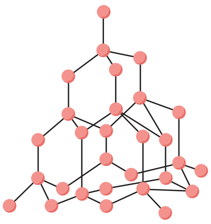
Question 9 (1 MARK)

Which of the following shows the structure of diamond?

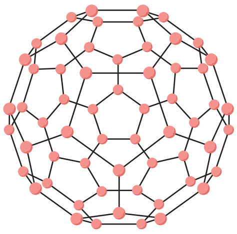
A.



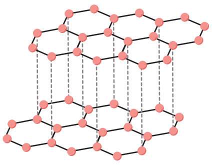
B.



C.



D.



Question 10 (1 MARK)

Which of the following statements regarding diamond is incorrect?

A. Diamond has a relatively high melting point.

B. Diamond is too brittle to be used in cutting tools.

C. There are no delocalised electrons in diamond's structure.

D. Each carbon atom in diamond's structure is bonded to 4 other carbon atoms.

Short answer (30 MARKS)

Question 11 (11 MARKS)

Hydrazine (N_2H_4) is an inorganic compound that is used in the space industry as a rocket propellant and is produced by the following reaction:

$N_2(g) + 2H_2(g) \rightarrow N_2H_4(l)$

a. Draw the electron dot diagrams of nitrogen and hydrogen. (2 MARKS)

b. Draw the electron dot diagram of $N_2H_4(l)$. (1 MARK)

c. Based on the Lewis diagram of $N_2H_4(l)$ and with reference to VSEPR, explain the shape of the molecule. It may help to use a shape diagram. (3 MARKS)

d. Is this molecule polar? Justify your answer. (2 MARKS)

e. A student wanted to compare the boiling points of C_2H_4 and N_2H_4. Which molecule would have the greatest boiling point? Explain your answer. (3 MARKS)

Question 12 (8 MARKS)

Hydrogen gas (H_2), oxygen gas (O_2) and nitrogen gas (N_2) are all diatomic molecules.

a. Draw the Lewis structures of H_2, O_2 and N_2. (3 MARKS)

b. Are these molecules polar or nonpolar? (1 MARK)

c. With reference to the relative strengths of different types of covalent bonds, order the molecules in increasing order in terms of energy required to break the intramolecular bonds. (2 MARKS)

d. In the atmosphere, hydrogen gas, oxygen gas and nitrogen gas exist in fixed ratios. What type of intermolecular forces exists in a mixture of these gases? (2 MARKS)

Question 13 (11 MARKS)

Diamond and graphite are two allotropes of carbon with many real-world applications. The basis for these applications lies within their structure.

a. Show, through diagrams, the lattice structures of these allotropes. (4 MARKS)

b. With reference to your diagram, explain why graphite is used:

  i. as an industrial lubricant. (2 MARKS)

  ii. in a battery-powered electric toothbrush. (2 MARKS)

c. Identify the structural properties of diamond that make it inappropriate for the previously listed applications. (3 MARKS)

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Key science skills (10 MARKS)

Question 14 (10 MARKS)

Samantha is currently undertaking a summer internship at a chemical engineering firm. She is investigating the differences between different forms of intermolecular forces.

In order to determine the relative strengths of different intermolecular forces, she proposes an experiment where she heats samples of $C_3H_8(l)$, $C_3H_7OH(l)$ and $C_3H_7Cl(l)$ until they boil.

a. Write a simple method for this experiment, remembering to specify equipment. (3 MARKS)

b. Identify the independent variable and dependent variable, as well as one controlled variable. (3 MARKS)

c. Explain how the data collected from the selected measurement device could be used to determine the strength of intermolecular forces between the different compounds. (3 MARKS)

d. As part of her internship, the company has a duty of care to Samantha and so, they conduct a thorough occupational health and safety review of her experiment. Identify a possible safety hazard in her experiment and how it could be minimised. (1 MARK)

FROM LESSONS 16B, 16D & 16E

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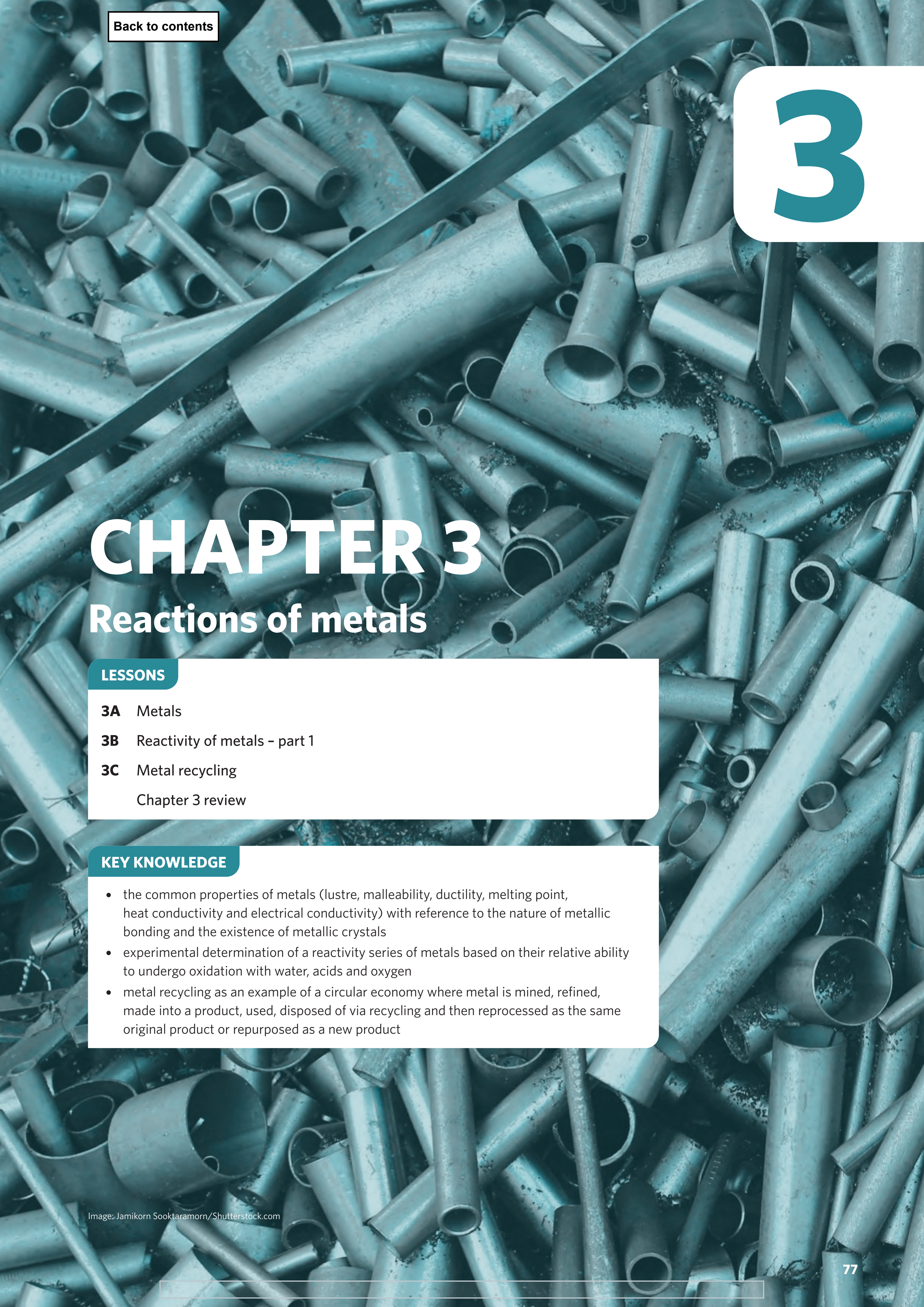

3

CHAPTER 3

Reactions of metals

LESSONS

KEY KNOWLEDGE

- the common properties of metals (lustre, malleability, ductility, melting point, heat conductivity and electrical conductivity) with reference to the nature of metallic bonding and the existence of metallic crystals
- experimental determination of a reactivity series of metals based on their relative ability to undergo oxidation with water, acids and oxygen
- metal recycling as an example of a circular economy where metal is mined, refined, made into a product, used, disposed of via recycling and then reprocessed as the same original product or repurposed as a new product

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3A Metals

STUDY DESIGN DOT POINT

- the common properties of metals (lustre, malleability, ductility, melting point, heat conductivity and electrical conductivity) with reference to the nature of metallic bonding and the existence of metallic crystals

3A 3B 3C

1.1.10.1 Metallic bonding

1.1.10.2 Properties of metals

ESSENTIAL PRIOR KNOWLEDGE

1C Ionisation energy

See question 16.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

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How do the properties of tungsten allow it to melt at 3410 °C?

Metals are the most abundant type of elements found on Earth and are used in our daily lives. In this lesson, we will learn about how metal particles are arranged in a solid **crystal lattice** and their special properties.

KEY TERMS AND DEFINITIONS

Cation positive ion formed when an atom loses its valence electron(s)

Crystal lattice (metallic) atoms of one type of metal element that are metallically bonded and organised in a pattern

Ductility ability to be hammered or stretched into a thin shape without breaking

Electrical conductivity ability to allow an electric current to flow through

Electrostatic of attraction attraction between opposite charges

Heat conductivity ability to allow heat to pass through

Lustre shiny and glossy appearance

Malleability ability to deform under force without breaking

Metallic bonding the electrostatic force of attraction between delocalised electrons and cations in a metallic lattice structure

Ore deposit in Earth's outermost layer containing metals and other minerals

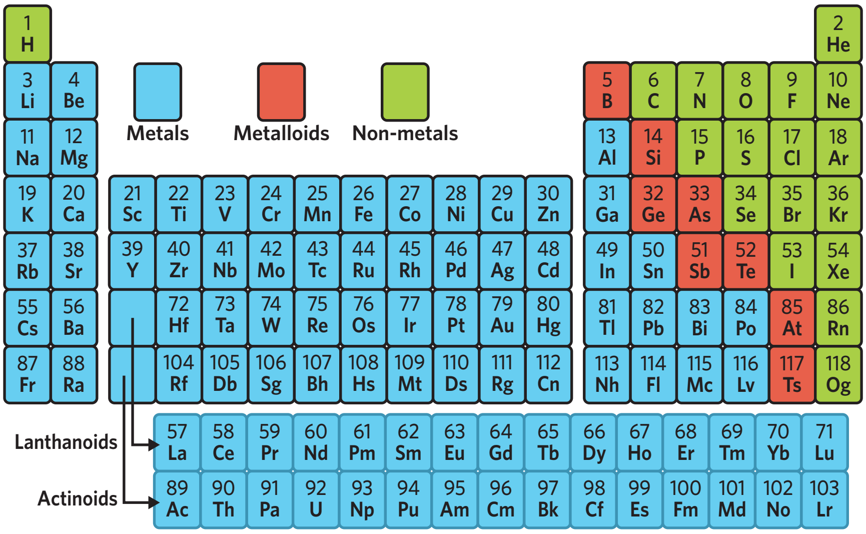
Sea of delocalised electrons electrons that freely move between metal cations in the metallic bonding model

Metallic bonding 1.1.10.1

The **metallic bonding** model explains how metal atoms bond to each other when in solid form.

What does the metallic bonding crystal lattice model look like?

Metals are extracted from deposits of minerals formed over long periods of time. Pure metals that we use are extracted from **ores** in the Earth's crust; the outermost layer of the Earth. Metals are the most common elements found in the periodic table as shown in figure 1.



🟦 Metals 🟥 Metalloids 🟩 Non-metals

🟩 1 H																	🟩 2 He
🟦 3 Li	🟦 4 Be											🟥 5 B	🟩 6 C	🟩 7 N	🟩 8 O	🟩 9 F	🟩 10 Ne
🟦 11 Na	🟦 12 Mg											🟦 13 Al	🟥 14 Si	🟩 15 P	🟩 16 S	🟩 17 Cl	🟩 18 Ar
🟦 19 K	🟦 20 Ca	🟦 21 Sc	🟦 22 Ti	🟦 23 V	🟦 24 Cr	🟦 25 Mn	🟦 26 Fe	🟦 27 Co	🟦 28 Ni	🟦 29 Cu	🟦 30 Zn	🟦 31 Ga	🟥 32 Ge	🟥 33 As	🟩 34 Se	🟩 35 Br	🟩 36 Kr
🟦 37 Rb	🟦 38 Sr	🟦 39 Y	🟦 40 Zr	🟦 41 Nb	🟦 42 Mo	🟦 43 Tc	🟦 44 Ru	🟦 45 Rh	🟦 46 Pd	🟦 47 Ag	🟦 48 Cd	🟦 49 In	🟦 50 Sn	🟥 51 Sb	🟥 52 Te	🟩 53 I	🟩 54 Xe
🟦 55 Cs	🟦 56 Ba	🟦	🟦 72 Hf	🟦 73 Ta	🟦 74 W	🟦 75 Re	🟦 76 Os	🟦 77 Ir	🟦 78 Pt	🟦 79 Au	🟦 80 Hg	🟦 81 Tl	🟦 82 Pb	🟦 83 Bi	🟦 84 Po	🟥 85 At	🟩 86 Rn
🟦 87 Fr	🟦 88 Ra	🟦	🟦 104 Rf	🟦 105 Db	🟦 106 Sg	🟦 107 Bh	🟦 108 Hs	🟦 109 Mt	🟦 110 Ds	🟦 111 Rg	🟦 112 Cn	🟦 113 Nh	🟦 114 Fl	🟦 115 Mc	🟦 116 Lv	🟥 117 Ts	🟩 118 Og

Lanthanoids	🟦 57 La	🟦 58 Ce	🟦 59 Pr	🟦 60 Nd	🟦 61 Pm	🟦 62 Sm	🟦 63 Eu	🟦 64 Gd	🟦 65 Tb	🟦 66 Dy	🟦 67 Ho	🟦 68 Er	🟦 69 Tm	🟦 70 Yb	🟦 71 Lu
Actinoids	🟦 89 Ac	🟦 90 Th	🟦 91 Pa	🟦 92 U	🟦 93 Np	🟦 94 Pu	🟦 95 Am	🟦 96 Cm	🟦 97 Bk	🟦 98 Cf	🟦 99 Es	🟦 100 Fm	🟦 101 Md	🟦 102 No	🟦 103 Lr

Figure 1 Metals make up most of the periodic table.

As we learned in lesson 1C, metals have low ionisation energies and therefore require a relatively small amount of energy to remove outer valence electrons from their atoms. As a result, it is relatively easy for metal atoms to lose electrons from their outer valence shell and become positive ions called **cations**. Figure 2 shows how a metal atom loses its electrons.

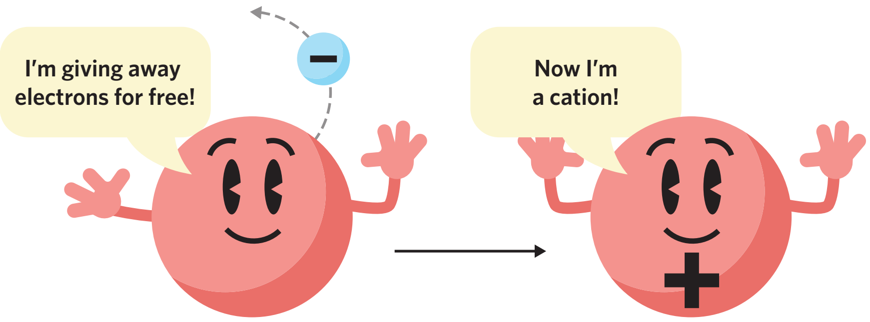


Figure 2 Drawing of a metal atom becoming a metal ion

MISCONCEPTION

'When you mix two different metal atoms like copper and zinc they form a metallic bond to each other.'

Metal atoms do not bond with other different metals atoms, they can only form mixtures of different metal atoms which are called alloys.

When in solid form, metals are organised in a crystal lattice structure as shown in figure 3. Electrons removed from the (outer) valence shells of metal atoms are able to move freely within the crystal lattice structure between the positively charged metal cations. These electrons are called a **sea of delocalised electrons**.

The **electrostatic attraction** between the negatively charged delocalised electrons and positively charged metal cations holds the metal cations closely to each other in fixed positions and maintains the structure of the metal. The strong electrostatic force of attraction between the delocalised electrons and cations is referred to as metallic bonding.

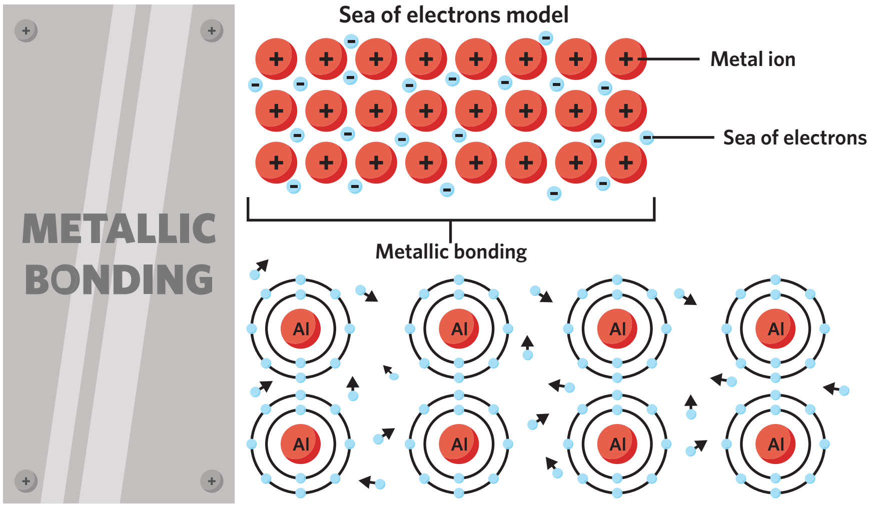


Figure 3 Crystal lattice model of bonding in aluminium

USEFUL TIP

Metallic bonding is non-directional as the electrostatic forces of attraction present between the cations and the delocalised electrons are in all directions (unlike ionic and covalent bonding which are directional forces).

Progress questions

Question 1

Mercury can be extracted from the red compound shown in the image called cinnabar.

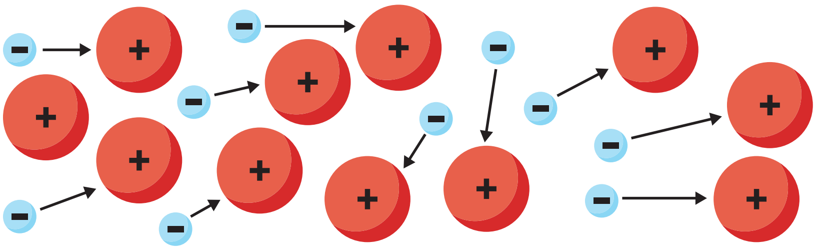
Source: Dan Olsen/Shutterstock.com

These sources of metals are called

A. ores.

B. crusts.

Use the following information to answer questions 2–3.



Question 2

The small freely moving particles illustrated in the image provided are referred to as

A. localised electrons.

B. delocalised electrons.

Question 3

In the metallic bonding model, metal cations are held in fixed positions and kept closely packed together by

A. the directional electrostatic force between delocalised electrons and metal cations.

B. the non-directional electrostatic force between delocalised electrons and metal cations.

Properties of metals 1.1.10.2

Metals exhibit different properties that allow them to be used for many different purposes.

Why do metals have unique properties?

Given that the structure of metals can be explained using the metallic bonding model, most metals share common properties as shown in table 1.

Table 1 Properties of metals

Property	Description
Malleability	Can be bent and hammered into different shapes without breaking
Ductility	Can be drawn into thin wire
Heat conductivity	Allows heat to pass through
Electrical conductivity	Allows electric current to flow through
High melting and boiling point	High melting and boiling temperatures
Lustre	Looks shiny and reflective

The ability of metals to be ductile and malleable is due to the non-directional electrostatic forces of attraction between the cations and delocalised electrons.

When a mechanical force is applied on metals to bend, hammer or to change their shapes, layers of cations are able to roll over each other whilst still being held in place by the electrostatic force of attraction with the delocalised electrons as shown in figure 4. As a result, the structure can take another shape without breaking.

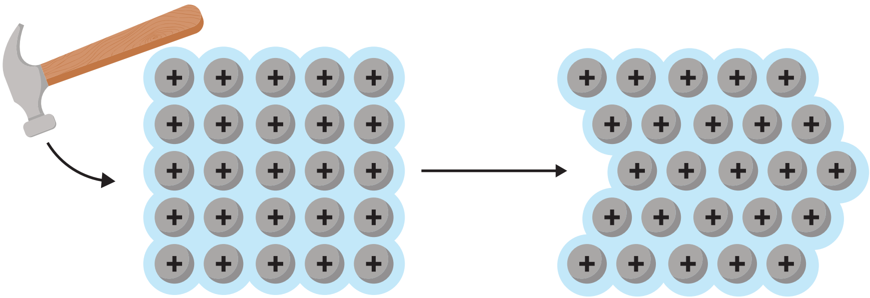


Figure 4 The effect of an applied force on the metallic bonding structure

Metals are considered to be good heat conductors, which can be explained through the metallic bonding model.

- When a piece of metal absorbs heat, the rise in temperature causes the increase in kinetic energy of delocalised electrons and metal ions.
- With this increase in energy, the metallic ions can vibrate more rapidly (about a fixed position) and delocalised electrons can move faster, transferring the kinetic energy to nearby ions and electrons.
- This creates a chain reaction that allows the energy to be passed through the metal as seen in figure 5.
- When the piece of metal is in contact with another conductive object, the kinetic energy carried by moving electrons will transfer to the object in the form of heat.

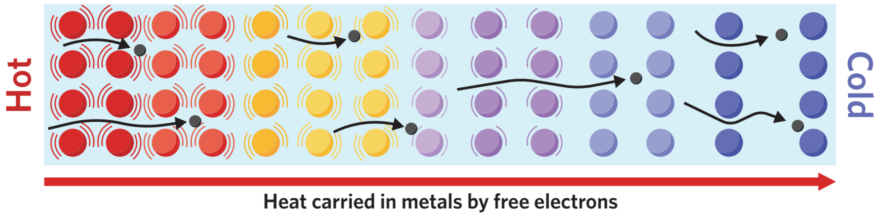


Figure 5 The conduction of heat through a metallic substance

USEFUL TIP

Boiling point gives a more useful indication of metallic bond strength than melting point within a metal as this is the measure of how much heat it actually takes to break the metallic bonds between the cations.

As we can appreciate, the conduction of heat requires particles to transfer the energy from a hot region to a cold region relatively quickly. Therefore, if we were to place a piece of metal into a beaker of iced water, the metal would be able to quickly transfer heat energy from the metal (hot region) to the water (cold region). As a result, the metal would feel cold relatively quickly.

The metallic bonding model can also be used to explain the electrical conductivity of metals, which refers to a metal's ability to conduct electricity. As shown in figure 6, in a basic circuit, electrons are continuously moving away from the negative electrode of a power source towards the positive electrode.

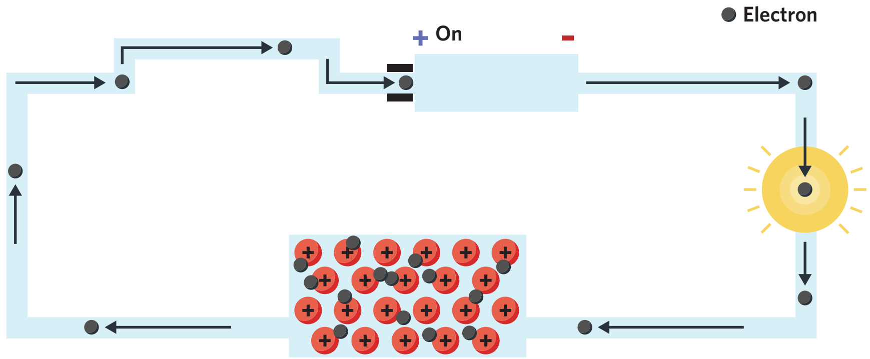


Figure 6 Using the metallic bonding model to explain the electrical conductivity of metals

Due to the fact that metals contain a sea of delocalised electrons when connected to the circuit, there is a movement of delocalised electrons away from the negative electrode and towards the positive electrode, which generates a current in the circuit. In the case of the circuit in figure 6, this results in the light being turned on. This property is why electrical wires are made from metallic substances.

Metals are generally hard (except for groups 1 and 2) and have relatively high melting and boiling points. This is due to the strong non-directional electrostatic force of attraction that holds metal cations and the sea of delocalised electrons together in all directions (figure 7). As a result, a large amount of heat is required in order to break the metallic bonds and deform the structure of metal crystal lattices.

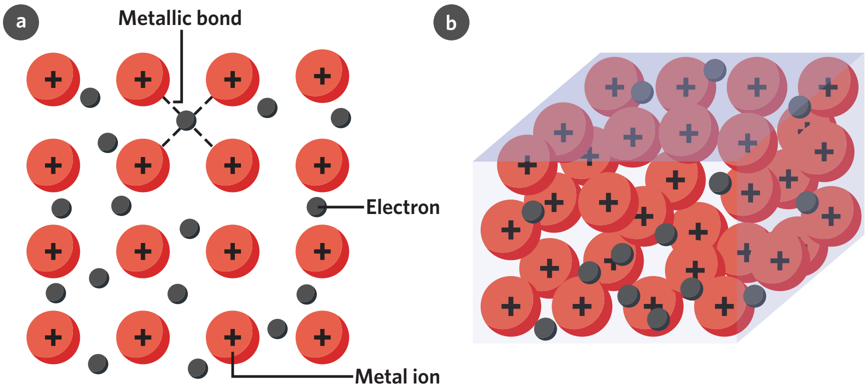


Figure 7 Crystal lattice metallic bonding models in 2-D **(a)** 3-D **(b)**

The shiny or lustrous nature of metals is due to the sea of delocalised electrons, which can reflect light of any wavelength.

WORKED EXAMPLE 1

Use the metallic bonding model to explain why copper is widely used to make wires that conduct electricity.

What information is presented in the question?

Use the metallic bonding model.

Copper is widely used to make wires that conduct electricity.

What is the question asking us to do?

Explain why copper wire is widely used to make wires that conduct electricity.

What strategies do we need in order to answer the question?

1. Describe the metallic bonding model of copper.
2. Use the metallic bonding model to describe what occurs when copper wire is connected to an electrical source.
3. Conclude the electrical conductivity of copper.

Answer

In the metallic bonding model of copper, the sea of delocalised electrons move freely in between copper cations. Therefore, when one end of the copper wire is connected to the positive electrode and the other end is connected to the negative electrode of an electrical source, delocalised electrons move away from the negative electrode and towards the positive electrode, which generates a current in the circuit that lights up the bulb. Hence, copper is widely used to make wires that conduct electricity.

What are the limitations of the metallic bonding model?

Even though the metallic bonding model can be used to explain some properties of metals, it isn't able to explain the variations in the properties shown by different metals. Particularly, it cannot be used to fully explain:

- the different melting and boiling points of different metals,
- the difference in electrical conductivity of different metals,
- and magnetic properties shown by iron, nickel and cobalt.[1]

As the explanation for these properties falls outside the study design, we will not be going into any further detail about the metallic bonding model.

KEEN TO INVESTIGATE?

[1] **Why are only these transition metals naturally magnetic?**
Search YouTube: Magnetic properties of the transition elements

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Progress questions

Question 4

Some properties of metals can be explained by the metallic bonding model.

A. True

B. False

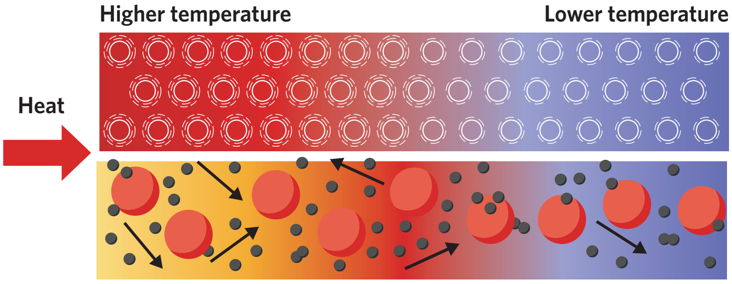
Question 5

Metals are malleable due to

A. the non-directional electrostatic force between delocalised electrons and metal cations that hold metal cations closely and tightly together in the crystal lattice.

B. the directional electrostatic force between localised electrons and metal anions that hold metal cations closely and tightly together in the crystal lattice.

Question 6

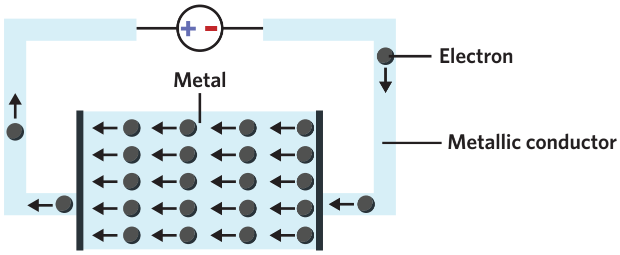


When a piece of metal is heated, the ________ electrons absorb the heat to increase their ________. Therefore, they move faster and transfer the kinetic energy to the ________ electrons nearby. As a result, heat is transferred along the piece of metal.

A. delocalised, kinetic energy, delocalised

B. localised, chemical energy, delocalised

Question 7



The electrical conductivity of metals can be explained by the movement of

A. delocalised electrons away from the positive electrode and towards the negative electrode.

B. delocalised electrons away from the negative electrode and towards the positive electrode.

Question 8

Many metals have a boiling point that is over 2500 °C. This suggests that

A. metallic bonds are relatively strong.

B. delocalised electrons are able to move quickly through the lattice.

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Theory summary

- In the metallic bonding model, the non-directional electrostatic forces of attraction between the sea of delocalised electrons and metal cations keep metal cations closely packed together.
- Metals exhibit lustre, malleability, ductility, heat and electrical conductivity which can be explained by the metallic bonding model.
- A pure sample of a metal has a crystalline lattice structure.
- The model of metallic bonding has limitations.

3A Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 9–11.

As shown in the periodic table, there are many different types of metals that exist on Earth. The table provided shows the melting points of different metals.

Metal	Melting point (°C)
Iron	1535
Aluminium	660
Copper	1083
Magnesium	650

Question 9 (1 MARK) 🌶

Based on the information provided by the table, which metal could be considered as being most resistant to melting?

A. Iron
B. Copper
C. Aluminium
D. Magnesium

Question 10 (1 MARK) 🌶

With regards to the metallic bonding, order each metal based on the strength of their bonds from strongest to weakest.

A. Iron, aluminium, copper, magnesium
B. Magnesium, copper, aluminium, iron
C. Iron, copper, aluminium, magnesium
D. Copper, iron, magnesium, aluminium

Question 11 (4 MARKS) 🌶🌶

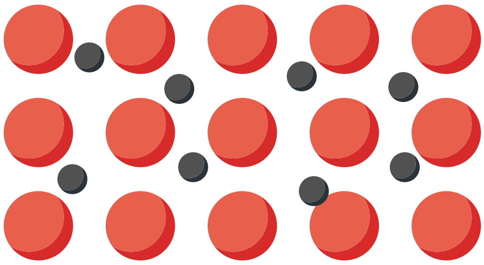
A building company wanted to build the basic frame structure for a four storey concrete apartment block. Using your understanding of bond strength explain which of the metals listed above would allow for the most sturdy structure?

Exam-style

Question 12 (5 MARKS)

The structure of metals allows them to exhibit a wide range of properties.

a. Label the structure shown assuming that it is a representation of the structure of a metal. (2 MARKS)



b. Gold is known to be very malleable. As a result, it is able to be made into many different types of products. Explain, using the model shown, how gold is able to be malleable. (3 MARKS)

Question 13 (3 MARKS)

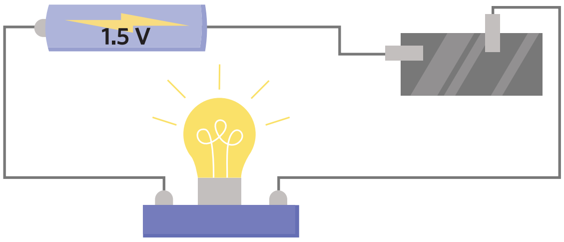
Nick conducted an experiment to investigate the properties of metals. He placed a metal spoon into a bowl of boiled water. After 30 seconds, he had a 'burning' feeling when he touched the spoon.

a. What property of metals does this phenomenon demonstrate? (1 MARK)

b. Explain the identified property of metals using the metallic bonding model. (2 MARKS)

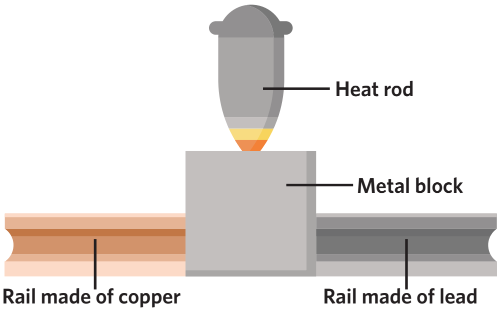
Question 14 (3 MARKS)

A student argued that metals conduct electricity because metal cations can move from one electrode to the other. Is the student's explanation correct? Justify your answer with the metallic bonding model.



Question 15 (7 MARKS)

With a rising population, the city of Dalton commissioned an engineering company to develop a railway using materials that were able to withstand high levels of heat whilst being low conductors of heat. As part of their investigation, they developed an experiment involving two metals as shown in the diagram. Two pieces of rail track developed using different metals were joined to a metal block. The engineers then applied heat to the metal block through an electric heat rod at increasing temperatures. Observations made during the experiment are shown in the table.



Temperature (°C)	Copper rail track	Lead rail track
200	Unchanged	Unchanged
400	Unchanged	The rail was beginning to lose its shape
600	It was glowing red but the structure remained unchanged.	No solid metal left, there was a lot of hot liquid in place of where the piece of rail used to be.

a. Which of the two metals used would be considered as having a higher melting point? Explain. (4 MARKS)

b. Is it possible to use the results provided to compare the heat conductivity of both metals? Explain. (3 MARKS)

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Key science skills

Question 16 (3 MARKS)

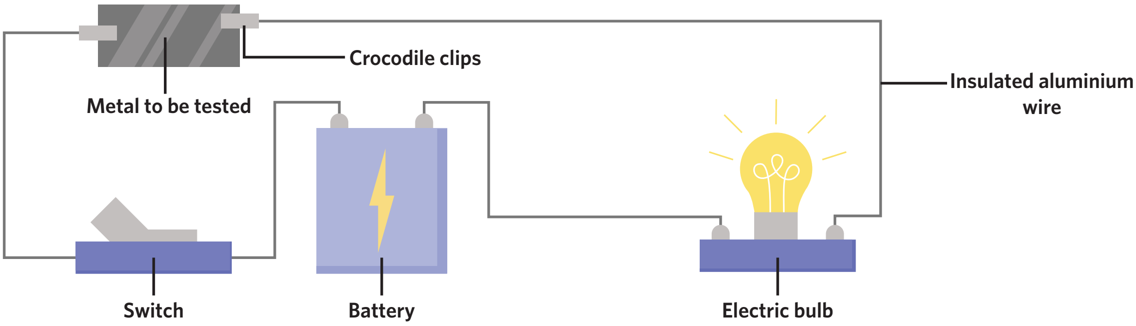
Gabriella conducted an experiment to confirm if her ring is entirely made of silver.

Materials

- 1 × electric bulb
- 1 × battery
- 1 × switch
- Insulated aluminium wires
- 2 × crocodile clips
- The ring
- 1 × piece of pure silver

Method

The diagram of the electrical circuit designed by Gabriella.



Results

Object	Observation
The ring	Dull light from the bulb
Piece of pure silver	Strong and dazzling light from the bulb

a. From the obtained results, what can Gabriella conclude about the material of her ring? Explain your answer. (2 MARKS)

b. Gabriella's friend told her that her conclusion might not be correct because the electricity might not be completely transferred through the wires. Suggest one change Gabriella can make to increase the validity of her experiment. Explain your suggestion. (1 MARK)

FROM LESSONS 16C & 16D

Questions from multiple lessons

Question 17 (4 MARKS)

A biomedical engineer wants to choose a material to design hip implants. Due to the wide range of movement of the human hips, the hip implants are required to withstand strong bending force without breaking. The engineer is considering choosing either titanium or a plastic (covalently bonded) material.

a. Should the engineer choose titanium or plastic to create hip implants? (1 MARK)

b. Justify your answer. (3 MARKS)

FROM LESSONS 1A & 2B

Hints

12a.	In general, metals have low ionisation energies.
12b.	The nature of bonds allows substances to exhibit particular properties.
13a.	'Burning' feeling when touching the spoon.
13b.	Increased temperature can increase kinetic energy of delocalised electrons and metals in the metallic bonding model.
14.	Delocalised electrons can move freely in the metallic bonding model.
15a.	Changes to the structure of a substance can indicate its properties.
15b.	Even with similar structures, metals exhibit slightly different properties.
16a.	Same observations should be obtained from the same pure metals.
16b.	Different metal wires have different electrical conductivities.
17a.	Consider the relative strength of the two materials.
17b.	Properties of metals and non-metals.

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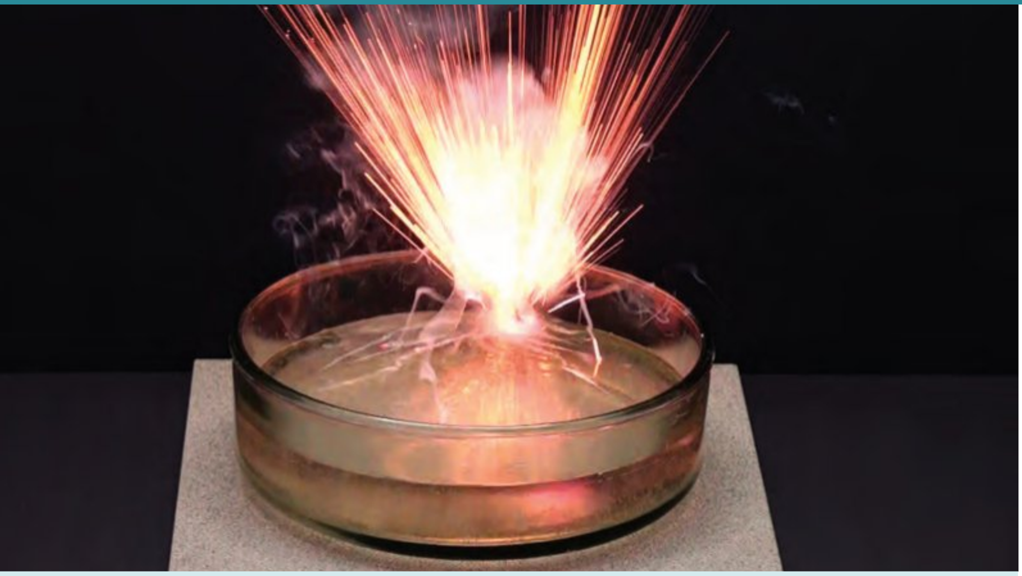
3B Reactivity of metals – part 1



Why do some metals like potassium explode in water?

In the previous lesson of this chapter, metals and their physical properties have been introduced. But how do they react with other substances or elements and why do their metallic characters determine their reactivity?

Observations of these reactions allows you to **qualitatively** determine the relative reactivity of metals. This lesson will guide you through these concepts and explain why certain metals can explode when exposed to water.

STUDY DESIGN DOT POINT

- experimental determination of the relative reactivity of metals based on their ability to undergo oxidation with water, acids and oxygen

3A 3B 3C

1.1.11.1 Reactions of metals with acids

1.1.11.2 Reactions of metals with water

1.1.11.3 Reactions of metals with oxygen

ESSENTIAL PRIOR KNOWLEDGE

1C Ionisation energy

1C Trends in atomic radii

See questions 17–18.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

KEY TERMS AND DEFINITIONS

Corrosion	occurs when some metals react with gases in the atmosphere (mainly oxygen)
Hydrogen pop test	test used to indicate a reaction between a metal and an acid
Metal oxide	formed as a result of a reaction between a metallic element and oxygen
Oxidation	chemical reaction where a chemical species loses electrons
Qualitative	describes non-numerical (descriptive) data collected based on observations taken during an experiment

Reactions of metals with acids 1.1.11.1

Most metals can react with acids, however they will do so to different degrees. Metals tend to lose their outer electrons more easily than non-metals due to their lower ionisation energies. When metal atoms lose electrons this process is called **oxidation**.

How do some metals react with acids?

In a solution, an acid is separated into positive and negative ions. For example, hydrochloric acid (HCl) appears as H^+ and Cl^- ions in HCl solution. When a reactive metal like magnesium (Mg) is placed into HCl solution, Mg atoms will give up their two outer electrons and become positive Mg^{2+} ions. These Mg^{2+} ions will then be attracted by free negative chloride ions (Cl^-) in the solution to form $MgCl_2$ which is a soluble salt. The remaining H^+ ions combine with each other to form H_2 gas bubbles. We will learn more about ions in lesson 4A. The reaction between Mg and HCl is demonstrated in figure 1. **Corrosion** of metals is one of the biggest problems caused by acidic rainwater.

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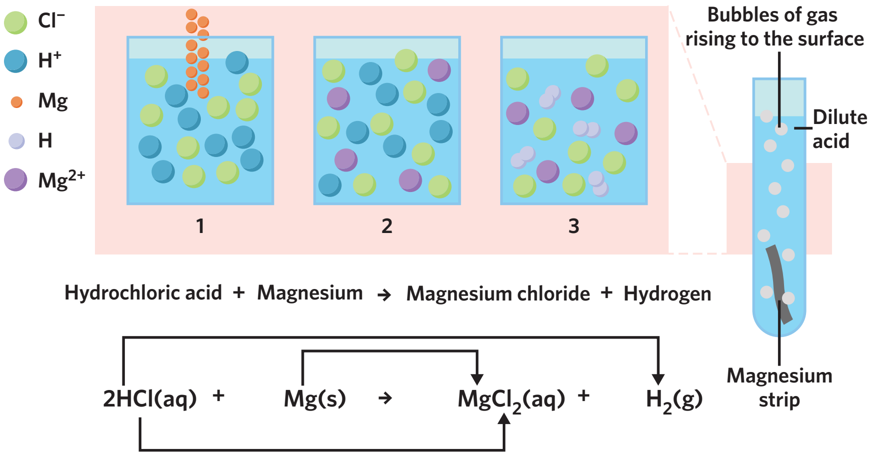


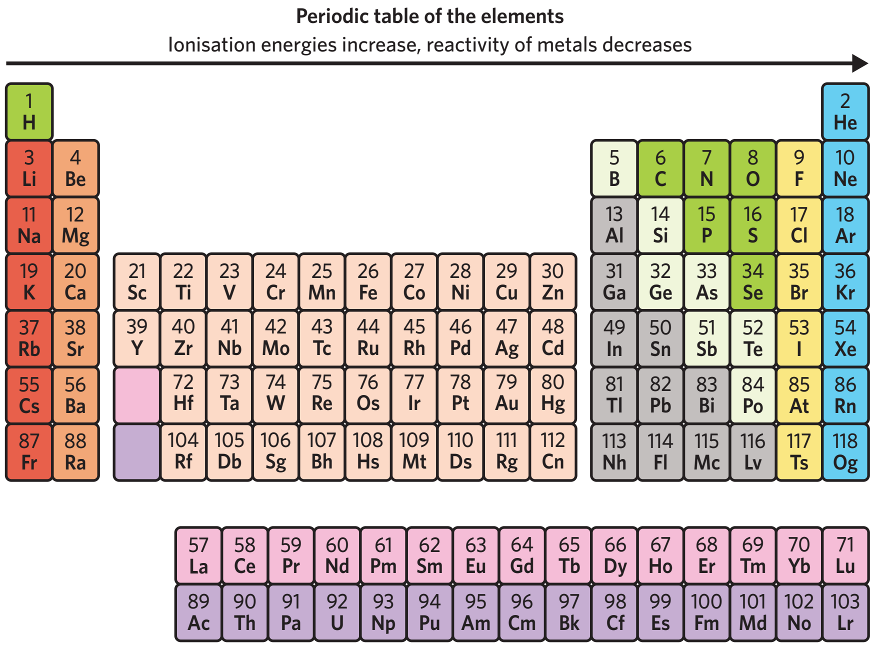
Figure 1 Demonstration of the reaction between Mg(s) and HCl(aq) – atoms and ions are not drawn to scale.

By looking at the reaction equation between Mg(s) and HCl(aq) in figure 1, a general equation can be developed to show the reactions between reactive metals and acids. The ionic salt formed may or may not dissolve in water depending on which acid is used - this will be covered in more detail in lesson 4C.

acid + reactive metal → ionic salt + hydrogen gas

Why are some metals more reactive in acids than others?

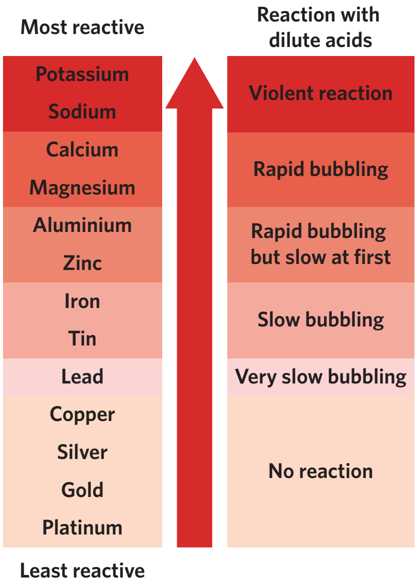
Since reactions between reactive metals and acids produce hydrogen gas, qualitative observations of the amount of gas bubbles produced can be used to indicate their relative reactivity. However, from lesson 1C, we know that some metals have higher ionisation energies than others, so they react differently with acids. For example, calcium (Ca) has a higher ionisation energy than potassium (K) due to its smaller atomic radius and greater number of protons in the nucleus. Consequently, more energy is required for calcium to lose electrons than for potassium to lose electrons. Therefore, the reaction of calcium with acids is slower. Figure 2 shows the trends in the reactivity of metals in the periodic table.



1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og
			57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
			89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Figure 2 Trends in the periodic table

In the periodic table, as we go across a period from the left to the right, ionisation energy increases. This means that the reactivity of metals decreases, and therefore the reactions with acids occur more slowly with less hydrogen gas bubbles produced. On the other hand, as we move down a group from the top to the bottom, ionisation energies decrease and the reactivity of metals increases. How rapidly different metals react with acid is shown in figure 3.



Most reactive	Reaction with dilute acids
Potassium Sodium	Violent reaction
Calcium Magnesium	Rapid bubbling
Aluminium Zinc	Rapid bubbling but slow at first
Iron Tin	Slow bubbling
Lead	Very slow bubbling
Copper Silver Gold Platinum	No reaction

Least reactive

Figure 3 The trend in reactivity of different metals with acids

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WORKED EXAMPLE 1

Does a reaction of tin with HCl or a reaction of sodium with HCl occur more quickly?

Use the difference in ionisation energies of the two metals to justify your answer.

What information is presented in the question?

Two different metals.

What is the question asking us to do?

Determine the reaction of which metal with HCl occurs more quickly.

What strategies do we need in order to answer the question?

1. Compare the ionisation energies of the metals.
2. Use the difference in ionisation energies to compare the reactivity of the metals.
3. Use the difference in reactivity of the metals to determine which metal reacts more quickly with acid.

Answer

Sodium has lower ionisation energy than tin, so it is easier for sodium to lose its outer electrons, meaning that sodium is more reactive than tin. Therefore, sodium reacts more quickly with HCl.

As reactions of some metals with acids occur slowly and do not produce enough bubbles of hydrogen gas to see, a **hydrogen pop test** can be used to indicate a successful metal-acid reaction. When a burning splint is placed inside a tube, a "pop" sound will be heard if there is hydrogen gas produced from a metal-acid reaction. This is because hydrogen gas is highly flammable and therefore will readily react with oxygen in the air in the presence of a flame. The use of a hydrogen pop test can be seen in figure 4.

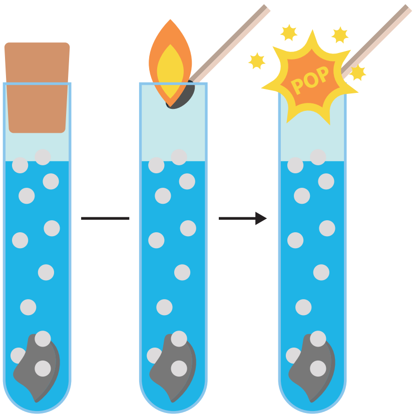

Figure 4 Stages of the hydrogen 'pop' test

Progress questions

Question 1

Some metals are able to react with acids.

A. True

B. False

Question 2

In a solution, metals tend to undergo oxidation becoming positive ions because

A. they tend to lose electrons.

B. they tend to gain electrons.

Question 3

Across the periods from the left to the right in the periodic table, the reactivity of metals with acid

A. increases.

B. decreases.

Question 4

One of the products of a reaction between a metal and acid is

A. oxygen gas.

B. hydrogen gas.

Continues →

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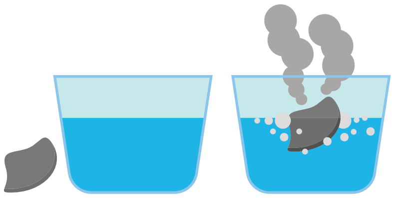
Question 5

All metallic elements on the periodic table react with water to the same extent.

A. True

B. False



Reactions of metals with water 1.1.11.2

Some metals can react with water while other metals cannot.

Why are some metals explosive in water?

While metals play an important role in industry, there are some certain rules in storing and disposing of metals. This is because metals in group 1 like lithium, sodium, potassium, rubidium and caesium (alkali metals) and metals in group 2 like calcium and magnesium can readily react with water[1] as shown in figure 5. For this reason, group 1 and 2 metals are often stored under oil to ensure they do not come into contact with water or oxygen.

KEEN TO INVESTIGATE?

[1] **Is this a safe way to dispose of sodium?**
Search YouTube: The disposal of surplus sodium 1947

Figure 5 Explosive reaction between sodium and water

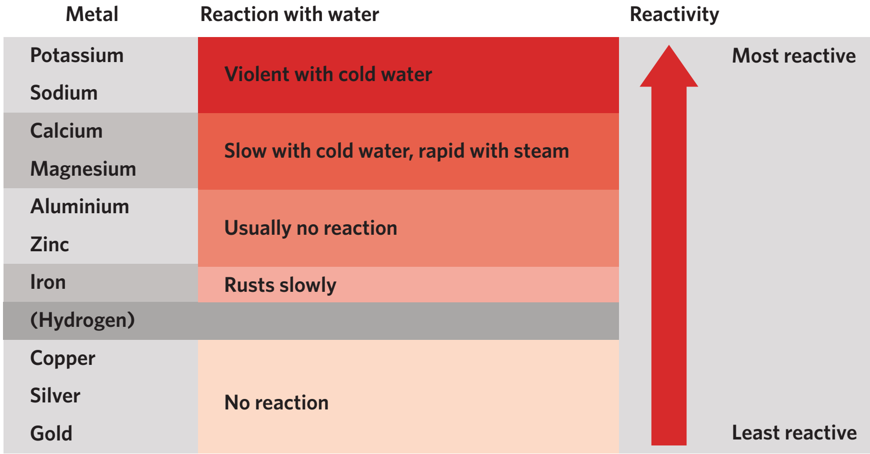
A general equation demonstrating a reaction of a reactive metal with water can be represented by:

metal + water → metal hydroxide + hydrogen gas

For example, a reaction between sodium and water will produce sodium hydroxide and hydrogen gas as shown (equations will be covered in more detail in lesson 4B):

sodium + water → sodium hydroxide + hydrogen gas

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$



Metal	Reaction with water	Reactivity
Potassium	Violent with cold water	Most reactive
Sodium		
Calcium	Slow with cold water, rapid with steam	
Magnesium		
Aluminium	Usually no reaction	
Zinc		
Iron	Rusts slowly	
(Hydrogen)		
Copper	No reaction	
Silver		
Gold		Least reactive

Figure 6 The trend in reactivity of common metals with water

Not all metals react with water in the same way. Across a period in the periodic table from the left to the right, ionisation energy of metals increases, meaning that it is more difficult for the metals to lose electrons and therefore the reactivity of the metals with water decreases. For example:

- group 4 metals react more slowly with water in comparison to group 1 metals,
- and ionisation energy of metals decreases down a group, meaning it is easier for them to lose electrons; they will therefore react more quickly down a group.

The trend of reactivity of common metals with water can be seen in figure 6.

USEFUL TIP

Although some metals like magnesium react very slowly with water at room temperature they react vigorously with water vapour (steam).

Progress questions

Question 6

A group 1 metal was placed into a beaker of water. When ignited, the gas produced ___________ makes a "pop" sound in a hydrogen test.

A. would

B. would not

Question 7

The product of a reaction between calcium and water is

A. calcium hydroxide.

B. calcium oxide.

Question 8

Which of the following gases can form an explosive mixture with oxygen?

A. Hydrogen

B. Carbon dioxide

MISCONCEPTION

'All metals rust.'

Rust is only formed when iron reacts with both oxygen and water to produce a red/brown solid called hydrated iron (III) oxide or more commonly rust.[2] Other metals do not rust – they corrode or oxidise.

KEEN TO INVESTIGATE?

[2] **Why is rusting like an electronic cell?** Search YouTube: Rusting of iron IPTV

Reactions of metals with oxygen 1.1.11.3

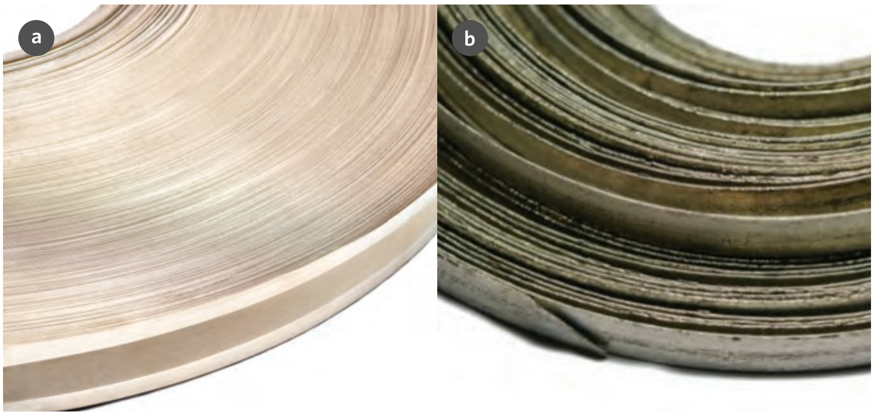
Corrosion occurs due to the reactions between metals and oxygen.

How does the corrosion of metals occur?

It is a common occurrence to come across some form of corrosion. When exposed to oxygen over a certain period of time, metals react with oxygen to form compounds called **metal oxides**. Most metals react with oxygen except for platinum (Pt) and gold (Au). The reactions between those that can react with oxygen have the general equation.

metal + oxygen gas → metal oxide

For example, if a piece of magnesium ribbon is left in the atmosphere, a reaction between magnesium and oxygen will take place as shown in figure 7.



Images (left to right): Stephen Barnes/andregric/shutterstock.com

Figure 7 Pure magnesium ribbon **(a)** is corroded (oxidised) at its surface when left exposed to oxygen gas **(b)**.

The chemical equation for this reaction is:

magnesium + oxygen gas → magnesium oxide

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

The trend of reactivity of metals with oxygen is similar to the trend of reactivity with water and acid. In the periodic table, the reactivity of metals with oxygen decreases as we go across a period from the left to the right, and increases as we move down a group. The trend of the reactivity of common metals with oxygen is shown in figure 8.

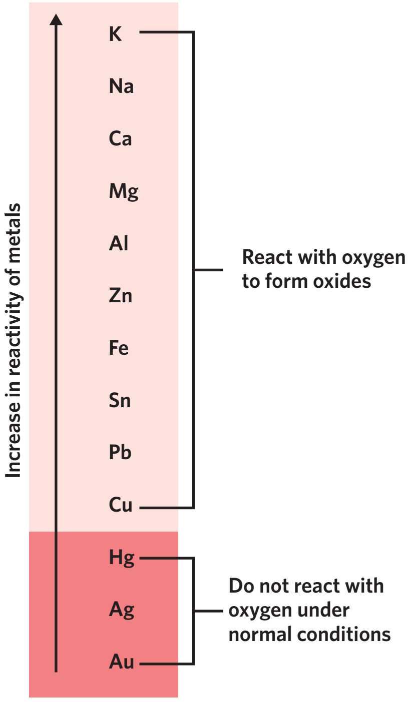


Figure 8 The trend in reactivity of common metals with oxygen

MISCONCEPTION

'Aluminium is unreactive.'

We often use aluminium foil which is unreactive. This is because the aluminium is covered in a protective oxide layer which forms a barrier (aluminium oxide) preventing it from further reacting with oxygen.

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USEFUL TIP

Although all oxidation reactions involve the loss of electrons, there are sometimes easier ways to spot if something has undergone oxidation. The gaining of oxygen or the loss of hydrogen by an element, compound, or species also indicates that oxidation has occurred.

Progress questions

Question 9

A toy truck has rust built up on it after being left exposed to the air and water for many years. This is the result of a reaction

A. between iron and carbon dioxide.

B. between iron, oxygen and water.

Image: Sergey Skleznev/shutterstock.com

Question 10

All metals can react with oxygen under normal conditions.

A. True

B. False

Question 11

__________ can react with oxygen under normal conditions, whereas __________ cannot.

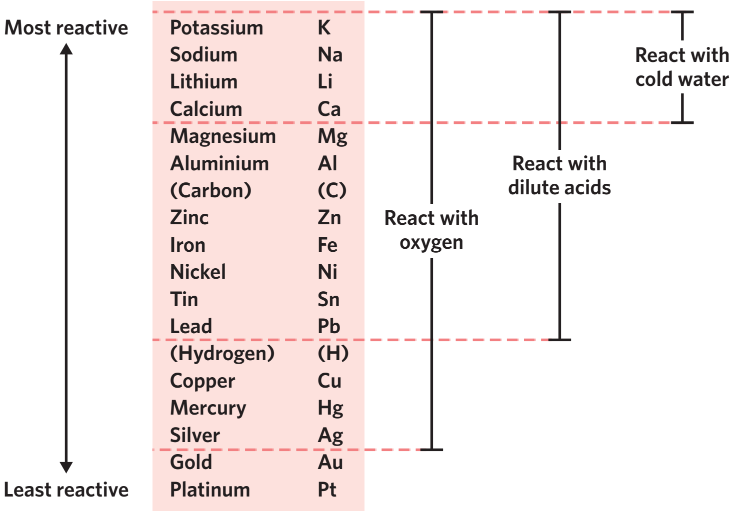
A. Zinc, silver

B. Silver, zinc

Theory summary

- A reaction between a metal and an acid will produce an ionic salt and hydrogen gas.
- A hydrogen pop test can be used to detect reactions between metals and acids.
- A reaction between a metal and water will produce a metal hydroxide and hydrogen gas.
- A reaction between a metal and oxygen will produce a metal oxide.
- Qualitative relative reactivity can be found by observing the amount of bubbles produced or the extent to which metals react (e.g. explosions with water).
- Metals react differently with acids, water and oxygen.

Reactivity series of metals



The content in this lesson is considered fundamental prior knowledge to primary galvanic cells and fuel cells as sources of energy (Unit 3 AOS 1).

3B Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 12–14.

Giao is a chemical engineer. While working in the laboratory on a humid day, she noticed that there was a small explosion with some gas produced on the shelf where she left some metals exposed to the air. No materials were burned by her colleagues during that time and the room temperature was maintained constant.

Question 12 (1 MARK) 🌶

What condition of the room can lead to the explosion?

A. Humidity

B. Temperature

C. Atmospheric pressure

D. All of the above

Question 13 (1 MARK) 🌶

What types of metals may have been affected in the explosion?

A. Group 3 metals with low ionisation energies

B. Group 1 metals with low ionisation energies

C. Group 2 metals with high ionisation energies

D. Transition metals

Question 14 (3 MARKS) 🌶🌶

Provide a possible cause of the small explosion and the presence of gas. Justify your answer including the type of metals that might have been affected in the explosion.

Exam-style

Question 15 (1 MARK) 🌶

Which of the following metals react with water at room temperature?

A. Fe, Mg, Na, K

B. Ni, Ag, Na, K

C. Na, K, Ca, Li

D. Hg, Cd, Na, K

Question 16 (1 MARK) 🌶

Which row in the table below correctly identifies the relative reactivity of the metals shown?

		Reacts with dilute acid	Reacts with oxygen gas	Reacts with water at room temp.	React with water vapour (steam)
A.	Copper	✓	✓	✓	✓
B.	Magnesium	✓	✓	✓	✓
C.	Calcium	✓	✗	✓	✗
D.	Sodium	✗	✓	✓	✓

Question 17 (2 MARKS) 🌶🌶

Two boxes A and B contain two different metal pieces, calcium and lithium. Thomson placed one piece of metal from each box into a tube containing dilute sulfuric acid and observed the results shown in the table.

Identify the metals contained in box A and B. Justify your answer with reference to the difference in ionisation energies of the two metals.

Box	Observation
A	Bubbles appear very slowly
B	More bubbles were produced at a faster rate

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Question 18 (2 MARKS)

An engineer wants to choose a metal to make agrimotors for farmers to use during the rainy season. Which metal, iron or barium, should the engineer choose? Explain your answer.

Question 19 (5 MARKS)

A chemistry student is given two jars of metals lithium and zinc which have lost their labels. The student suggests that the type of metals in each jar can be determined by placing 1 g of metal from each jar into a beaker of dilute hydrochloric acid, HCl.

a. Is the student's method correct? Explain. (2 MARKS)

b. Depending on the type of metals involved in a reaction, it is possible to distinguish between two metals based on their reactive nature. Based on what you learned in the lesson, suggest another method that can be used to determine the metal in each jar. Include balanced chemical equation(s) in your answer. (3 MARKS)

Key science skills

Question 20 (8 MARKS)

5 g of solid copper turnings was mixed with solid calcium and solid magnesium. A group of chemistry students tried to separate pure solid copper from the mixture by the following steps.

1. Place cold water into the mixture and wait for 5 minutes.
2. Separate the remaining solid mixture and the solution produced.
3. Place dilute sulfuric acid into the remaining solid mixture.
4. Collect the remaining solid which should be pure copper.

a. Identify the metals in the solid mixture in step 2. Justify your answer including the balanced chemical reaction(s). (3 MARKS)

b. One student in the group suggested that concentrated sulfuric acid should be placed into the mixture first instead of the cold water. Is the student's suggestion correct? Explain your answer. (2 MARKS)

c. After finishing all four steps, the students weighed the collected sample of copper and realised that they collected 5.3 g of copper. Explain why the mass could be greater than 5.0 g and suggest one change that the students can make to obtain the original 5.0 g of pure copper. (3 MARKS)

FROM LESSON 16B

Questions from multiple lessons

Question 21 (4 MARKS)

Use the list of symbols of elements to answer the following questions. Be, P, N, Mg, Na, Fe.

a. Which of the elements are not considered to be metals? (1 MARK)

b. Which metal is likely to be used for making cars? (1 MARK)

c. Which of the elements is considered a critical element? Suggest a way that we can preserve its supply? (2 MARKS)

FROM LESSONS 1D & 3A

Hints

15. Only group 1 and 2 metals react with water at room temperature.

16. All metals react with oxygen (except gold and platinum), and some react with steam but not water at room temperature.

17. Different metals have different ionisation energies.

18. Rain water reacts differently with the metals.

19a. Two different metals can be distinguished if different observations are obtained from their reactions with the same substance.

19b. Choose a substance that the metals react with differently.

20a. Only some metals react in cold water.

20b. Many metals can react with acids.

20c. Only some metals react completely in very dilute acids.

21a. Look at the periodic table locations.

21b. Reactive metals are not very useful.

21c. Most critical elements are non-metals.

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3C Metal recycling

Image: JasminkaM/Shutterstock.com

How are metals recycled?

Metals are valuable resources that can be **recycled** many times without degrading their properties. The **sustainable** process of recycling metals is crucial to preserving natural resources, the environment and reducing energy consumption.

STUDY DESIGN DOT POINT

- metal recycling as an example of a circular economy where metal is mined, refined, made into a product, used, disposed of via recycling and then reprocessed as the same original product or repurposed as a new product

3A 3B 3C

1.1.12.1 Circular economy

1.1.12.2 Metal recycling

ESSENTIAL PRIOR KNOWLEDGE

3A Metals do not form compounds with other metals

See question 19.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

KEY TERMS AND DEFINITIONS

Atom economy mass of desired products divided by mass of all reactants given as a percentage conversion

Circular economy a continuous cycle that focuses on the optimal use and re-use of resources from the extraction of raw materials through to production of new materials, followed by consumption and re-purposing of unused and waste materials

Green chemistry principles aimed at reducing the chemical-related impact on both humans and the environment through dedicated sustainability management programs

Life cycle the management of the production and processing, storage, transportation, use, and disposal of a chemical product

Linear economy operates on a ‘take-make-dispose’ model, making use of resources to produce products that will be discarded after use

Recycle convert a waste product into a reusable product

Renewable resources comes from sources that naturally renew themselves at a rate that allows them not to become depleted

Sustainable can be produced at a rate that is greater than consumption without compromising future generations

Circular economy 1.1.12.1

In the past chemical manufacturers have not factored in what happens to a chemical product after it has been used.

What is a circular economy?

Resources are at risk of becoming scarce due to a growing global population and rising prosperity levels. Available resources must therefore be used as efficiently as possible.

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KEEN TO INVESTIGATE?

[1] **What are the main differences between these economies?**
Search YouTube: Circular economy vs linear economy champions

A transition from a **linear economy** towards a **circular economy**[1] is increasingly being adopted by society and industry as a strategy to achieve more sustainable development. Figure 1 shows the differences in these two types of economy.

VCAA Study design 2023–2027 p21

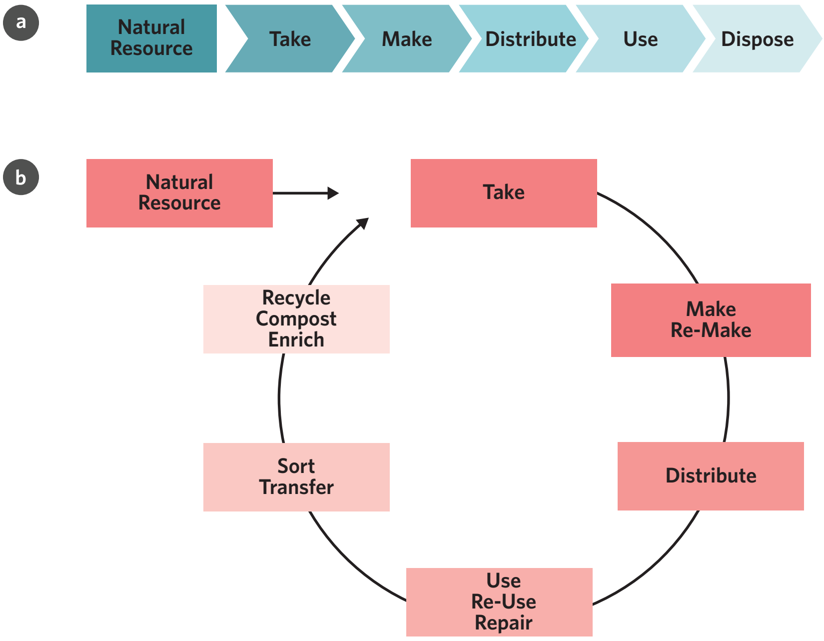

Figure 1 **(a)** Linear economy – open **life cycle** **(b)** Circular economy – closed life cycle

Put simply, a circular economy is when the materials and resources from a product are returned back into the current economy which is a closed product cycle. A linear economy on the other hand is when a product is essentially disposed of in landfill after it has been used.

In Australia, new materials that are useful for society tend to be produced through a linear economy, in which products are purchased, used and then thrown away. Increasingly, manufacturing companies are moving towards a circular economy, which seeks to reduce the environmental impacts of production and consumption while enabling economic growth through more productive use of natural resources and creation of less waste.

How can chemistry contribute to a circular economy?

Society relies on the chemical industry for a variety of products through various chemical processes and pathways. To improve the life cycle management of manufactured goods and their associated processes, a broad and global transformation towards a safe and more sustainable chemical future is required. The creation, use, re-use, disposal and elimination of chemicals should be achieved with the least possible adverse effects on human health and the environment while still providing economic and social benefits.

Green chemistry is the design of new chemical products and manufacturing processes that are safer and more sustainable than traditionally used products and processes. It is underpinned by a set of 12 principles that aim to minimise the impact of the product or process on the environment, Earth's resources, human health and the viability of other living organisms. These principles are based on reducing risk, minimising the production of unwanted by-products and wastes[2],and limiting the amount of energy used and raw materials (particularly non-renewable raw materials) consumed. Creativity and innovation are required to apply green chemistry principles to chemical manufacturing as new products are designed and produced. Of the 12 Principles of Green Chemistry developed in 1991by Paul T Anastas and John C Warner, the following four contribute directly to a circular economy.

MISCONCEPTION

'Reuse and repurpose are the same as recycle.'

Recycling is a process that involves the conversion of materials, whereas, reuse and repurpose does not involve conversion of materials.

KEEN TO INVESTIGATE?

[2] **How can gold be recycled?**
Search YouTube: Recycling Gold from Electronic Waste

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- **Atom economy**: Processes/pathways should be designed to maximise incorporation of all reactant materials used in the process into the final product.
- Use of **renewable resources:** feedstocks: Raw materials or feedstocks should be made from renewable (mainly plant-based) materials, rather than from fossil fuels, whenever practicable.
- Prevention of wastes: It is better to prevent waste than to treat or clean up waste after it has been produced.
- Design for energy efficiency: Processes/pathways should be designed for maximum energy efficiency and with minimal negative environmental and economic impacts.

VCAA Study design 2023–2027 p21

Progress questions

Question 1

What is the main difference between a linear and circular economy?

The product life cycle of

A. both a linear and circular economy are an open cycle.

B. a linear economy is open, but a circular economy is a closed cycle.

Question 2

Sustainable development is characterised by a

A. linear economy product life cycle.

B. circular economy product life cycle.

Question 3

Which of the following is an example of a sustainable product?

A. Glass milk bottle

B. Single-use plastic bag

Metal recycling 1.1.12.2

Most metals can be recycled repeatedly without degrading their properties.

Why should all steel items be recycled?

Steel, which is made mainly from iron (and small amounts of carbon and other metals), is the most recycled material on Earth. It is extremely sustainable as it can be infinitely recycled as it does not lose any of its inherent strength qualities. Another advantage of recycling iron is that due to its magnetic properties it can easily be separated from other metals. Table 1 shows some of the other most commonly recycled metals and their uses.

MISCONCEPTION

'Most metals are magnetic.'

Of the 80 or so metals on Earth only three are naturally magnetic – iron, nickel and cobalt.

Table 1 Commonly recycled metals and their uses

Metal	Uses
Iron	Food cans and construction
Aluminium	Drinking and food cans
Copper	Electrical appliances and electrical wires
Gold	Computers and jewellery
Lead	Roofs and construction

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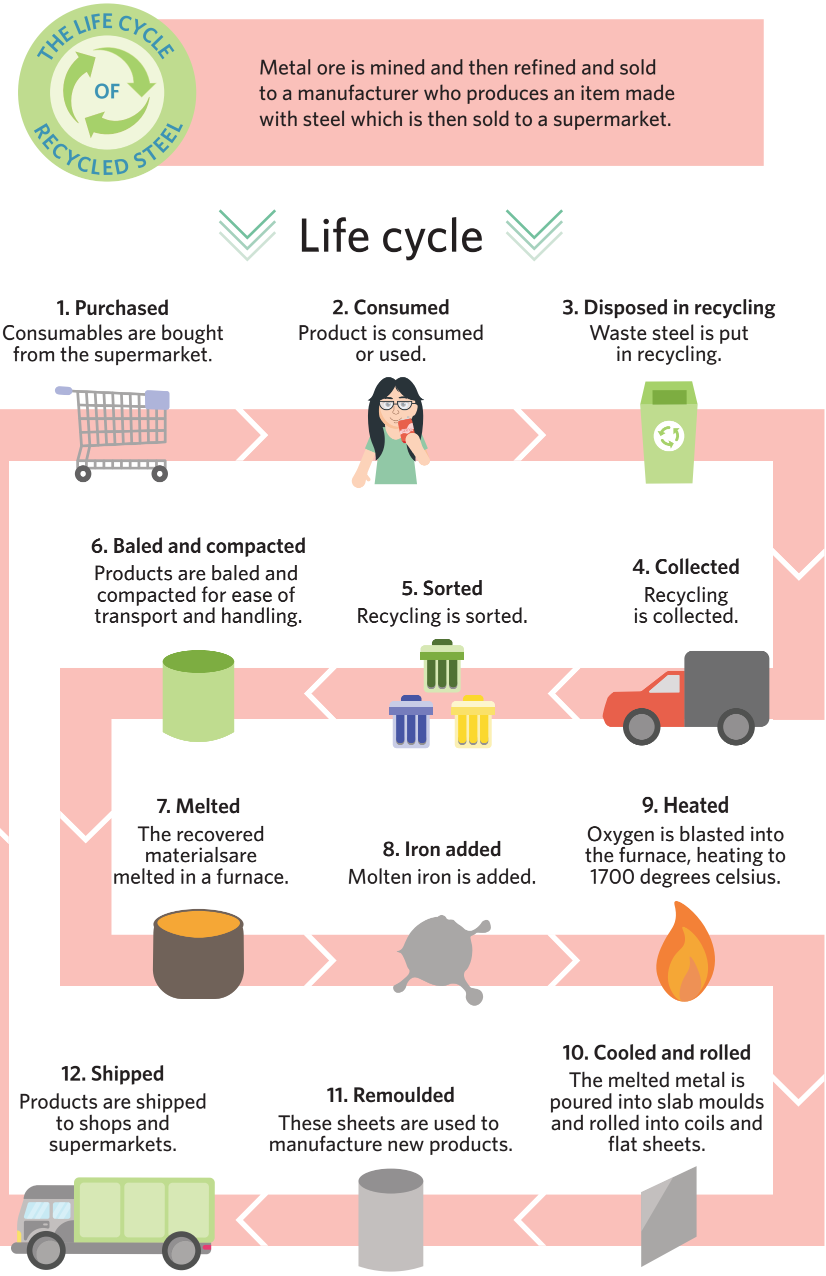
Figure 2 shows the complex life cycle of recycled steel.



Metal ore is mined and then refined and sold to a manufacturer who produces an item made with steel which is then sold to a supermarket.

1. Purchased Consumables are bought from the supermarket.

2. Consumed Product is consumed or used.

3. Disposed in recycling Waste steel is put in recycling.

4. Collected Recycling is collected.

5. Sorted Recycling is sorted.

6. Baled and compacted Products are baled and compacted for ease of transport and handling.

7. Melted The recovered materialsare melted in a furnace.

8. Iron added Molten iron is added.

9. Heated Oxygen is blasted into the furnace, heating to 1700 degrees celsius.

10. Cooled and rolled The melted metal is poured into slab moulds and rolled into coils and flat sheets.

11. Remoulded These sheets are used to manufacture new products.

12. Shipped Products are shipped to shops and supermarkets.

Figure 2 Typical life cycle of a steel can

It is estimated 200 billion steel cans are produced each year, around 25% of which contain recycled steel. Using recycled steel saves around 70% more energy, reduces air pollution by around 85% and reduces water use by 40%.

Source TATA steel

What are the advantages of recycling metals?

- Reduce energy consumption: the process of recycling waste metal is far more energy efficient; for example, producing a new aluminium can from recycled aluminium requires 95% less energy than a can produced from the ore.
- Reduce landfill: diverts waste from landfill sites as metals take a long time to break down.
- Reduce greenhouse gas emissions[3]: less fossil fuels are required in the production of materials made from recycled metals, thereby reducing the amount of greenhouse gases created which can help combat climate change.
- Conserve natural resources: mining is a disruptive process that depletes natural resources and can have adverse effects on wildlife, local communities and the environment.
- Endless: most metals can be recycled an unlimited number of times.

KEEN TO INVESTIGATE?

[3] **Why are some metals key to a more sustainable future?**
Search: six metals are key to a low-carbon future

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WORKED EXAMPLE 1

Using the following infographic, a student concludes that 'The recycling process has a net energy of zero'. Justify whether the student is correct.

What information is presented in the question?

The circular economy of aluminium cans.

What is the question asking us to do?

Justify with reasons why the student's statement is or is not correct.

What strategies do we need in order to answer the question?

1. Decide on whether the statement is correct or not.
2. Carefully look over the infographic highlighting important points to be used as justification.

Answer

Although recycling aluminium saves a lot of energy it still requires an input of energy to supply the recycling process. From the infographic it can be seen that the following processes require an input of energy:

- Transporting the cans to be recycled
- Machine to crush and clean the cans
- Melting the cans
- Rolling the cans
- Producing new cans
- Transporting the cans to the supermarket

Therefore the process of recycling is not net zero and the statement is incorrect.

Progress questions

Question 4

Nearly all metals are recyclable. Which of the following are not metals that can be recycled?

A. Tin, nickel, vanadium

B. Uranium, plutonium, sodium

Question 5

What is a good reason why two thirds of the food cans found on supermarket shelves are made from steel?

A. Iron requires less energy to be extracted than most metals.

B. Aluminium is lighter and stays shiny for longer.

Question 6

Advantages of recycling metals include reducing the

A. number of steel cans manufactured worldwide.

B. amount of natural resources mined.

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Theory summary

- A circular economy is much more sustainable than a linear economy.
- Metals can be recycled over and over again.
- Green chemistry principles contribute to the development of a circular economy.
- The life cycle of a circular economy is complex and has many advantages.

3C Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 7–9.

The following image shows a crane sorting out metal at a recycling centre. Dealing in scrap metal is big business as the value of scrap metal continues to rise.

Image: Silarock/Shutterstock.com

Question 7 (1 MARK) 🌶

What property allows certain metals containing iron, nickel or cobalt to be separated out in this way?

A. Electric

B. Density

C. Magnetic

D. Electrostatic

Question 8 (1 MARK) 🌶

Why does the cost of some scrap metals keep rising?

A. Due to an increased amount of recycling

B. Production of pure metal has decreased

C. Because mining the ore has become more expensive

D. Greater competition from other companies

Question 9 (3 MARKS) 🌶🌶

Outline the main energy inputs that would be involved in the recycling of scrap iron before selling the purified metal back to a manufacturer.

Exam-style

Question 10 (1 MARK) 🌶

Green chemistry principles apply across the ________ of a chemical product including its design, manufacture, use, and after its use.

A. properties **B.** efficiency **C.** sustainability **D.** life cycle

Question 11 (7 MARKS)

a. Define what is meant by the term 'sustainable chemistry'. (2 MARKS)

b. For recycling purposes, scrap metals are often classified into ferrous and non-ferrous. Explain the basis of this classification. (2 MARKS)

c. What property of metals allows them to be produced in a circular economy, and why is this property important? (3 MARKS)

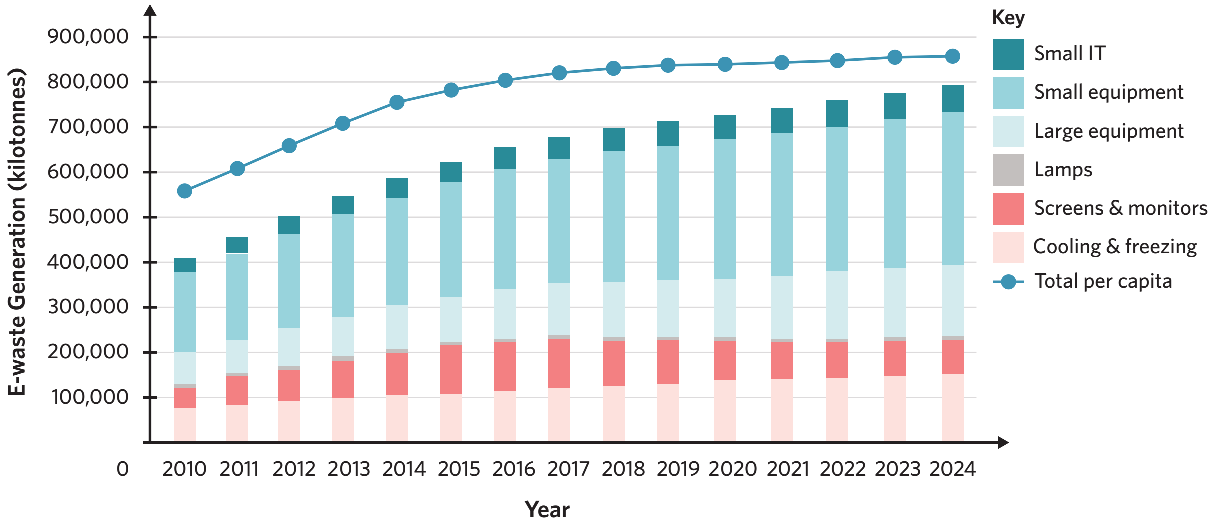
Question 12 (2 MARKS)

Steel requires around 55% less energy to produce when it is made from recycled materials. Give two specific reasons why using 55% less energy is beneficial to the environment.

Key science skills

Question 13 (6 MARKS)

E-waste is a term used when talking about the recovery of precious metals like gold, silver and copper metal from used electronic devices.



a. What is the overall trend shown in the chart? (1 MARK)

b. Which type of equipment is the largest contributor to e-waste? (1 MARK)

c. Which type of e-waste equipment has been decreasing since 2016? (1 MARK)

d. Why is it important that precious metals are recovered by specialised recycling companies? (3 MARKS)

FROM LESSONS 16B & 16C

Question 14 (5 MARKS)

Using only the information provided in the table, write an overall conclusion about the benefits of recycling metals.

Recycling metals	
Fewer raw materials (ores) are used and ores are a non-renewable resource.	The process of extracting a metal from its ore is more costly and requires lots of energy compared to recycled metals.
Recycled metals are generally impure.	Reduces waste that will be sent to landfill.
Recycling provides lots of jobs.	Mining and landfill negatively affect the environment.
Recycling some metals does not meet the demands.	Sorting metals costs lots of time and money.

FROM LESSON 16E

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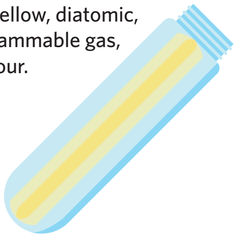
Questions from multiple lessons

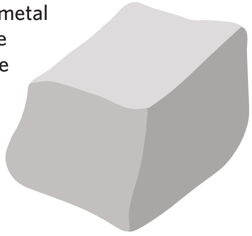
Question 15 (6 MARKS)

Fluorine gas and sodium metal are both very reactive elements that react vigorously with water.

Fluorine is a pale yellow, diatomic, highly corrosive, flammable gas, with a pungent odour.



Sodium is a silvery-white metal belonging to group 1 of the periodic table, which is the alkali metals group.



They react according to the following equations:

fluorine + water → hydrogen fluoride + oxygen gas
$2F_2(g) + 2H_2O(l) \rightarrow 4HF(aq) + O_2(g)$

sodium + water → sodium hydroxide + hydrogen gas
$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

a. What type of chemical reaction does sodium undergo when it reacts with water? (1 MARK)

b. How could you test for the presence of hydrogen gas in this reaction? (2 MARKS)

c. Why is fluorine gas referred to as a diatomic gas? (1 MARK)

d. Explain why both fluorine gas and sodium metal react very vigorously with water. (2 MARKS)

FROM LESSONS 1B, 2A & 3B

Hints

10. Consider what is described by the phrase 'design, manufacture, use, and after its use.'

11a. Check out the definitions at the start of the lesson.

11b. Ferrous is the latin name for an element which has the symbol Fe.

11c. Metals do not form compounds and alloys are mixtures which can be separated.

12. Focus on how less energy used will impact the environment.

13a. Use the blue trend line to see the pattern (per capita)

13b. Use the colour coded sections of the bars.

13c. Use the colour coded sections of the bars.

13d. Precious metals are in high demand which outstrips supply.

14. Address each statement and try to group any that are similar, do not just repeat the statements. Synthesise them into a concise sentence.

15a. The sodium lost an electron.

15b. A procedure for an explosive test.

15c. All of group 17 elements exist in pairs.

15d. The life and times of an electron can differ.

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Chapter 3 review

Multiple choice (10 MARKS)

Question 1 (1 MARK)

In which group on the periodic table would we find the most reactive metals?

A. Group 17

B. Group 4

C. Group 2

D. Group 1

Question 2 (1 MARK)

Sodium, rubidium and caesium are all alkali metals. Which of the following options orders the metals in increasing order of reactivity?

A. Rubidium, sodium, caesium

B. Sodium, caesium, rubidium

C. Caesium, rubidium, sodium

D. Sodium, rubidium, caesium

Question 3 (1 MARK)

In chemistry, rust is also referred to as a

A. metallic hydrolyte.

B. hydrated iron (III) oxide.

C. dihydrogen oxide.

D. metalloid.

Question 4 (1 MARK)

Gold jewellery is produced by heating solid gold to a temperature above 1064 °C and allowing it to melt so that it can be poured into different shaped moulds. This suggests that metallic substances

A. are highly combustible.

B. are expensive.

C. contain strong bonds.

D. are easily malleable.

Question 5 (1 MARK)

A company wanted to develop a new chemical product to be used for their newest car model. Before the company starts to produce this new chemical, they should undertake a study to determine the complete ___________ of the product.

A. open cycle

B. life cycle

C. recycle

D. reuse

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Question 6 (1 MARK)

Which of the following materials should not be composed of metallic elements?

A. Gloves used by electricians to fix live wires

B. Wires used in telephone communications

C. Cooking pots

D. Light bulb filaments

Question 7 (1 MARK)

In a reaction between a strong acid and potassium metal, it is expected that we would

A. smell a fruity odour.

B. observe bubbles.

C. observe a change in the colour of the acid.

D. see no change.

Question 8 (1 MARK)

What type of economy is depicted by the image?

A. Linear

B. Financial

C. Circular

D. Chemical

Question 9 (1 MARK)

Which of the following metals is **least** likely to react with water at room temperature?

A. Magnesium **B.** Rubidium **C.** Copper **D.** Beryllium

Question 10 (1 MARK)

Properties such as lustre and heat conductivity of metals are mainly due to

A. the reactivity of metals.

B. the size of metallic atoms.

C. delocalised electrons.

D. electronegativity of metals.

Short answer (30 MARKS)

Question 11 (16 MARKS)

Humans rely on metals to go about their everyday lives. Metals have different properties which determine how they are used in society.

a. Match up each of the metals listed with the relevant information in the table

Copper	Resistant to corrosion due to a strong oxide layer.
Gold	Most recycled element on Earth.
Aluminium	Reacts rapidly with water producing hydrogen gas.
Iron	Used in electrical wiring due to high conductance.
Sodium	Found in the Earth as a pure metal.

(5 MARKS)

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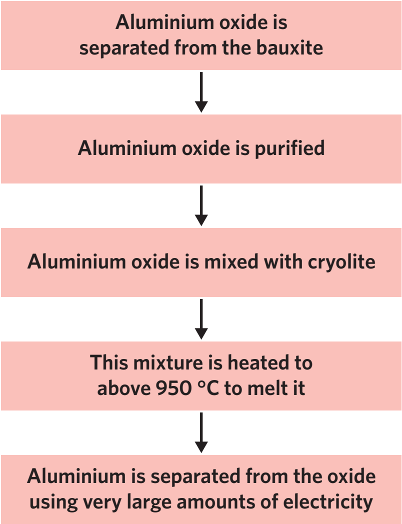


b. Aluminium is extracted from an ore called bauxite which is impure aluminium oxide. Pure aluminium oxide has a melting point of over 2000 °C. The flow chart outlines the main steps in the extraction of aluminium from bauxite.

Of the over 3 billion aluminium cans used annually in Australia around 55% are recycled.

When aluminium is recycled the scrap aluminium melts at 700 °C.

Using the information provided:

i. Suggest six reasons why most aluminium is recycled. Refer three of your reasons to the extraction process and three reasons to the recycling process. (6 MARKS)

ii. Using the temperature values provided, explain how this process is designed for greater energy efficiency. (2 MARKS)

iii. What additional steps could be taken to further improve the efficiency and sustainability of the processes making it more of a circular economy? (3 MARKS)

Question 12 (5 MARKS)

A materials scientist wants to create a new type of alloy to use as a wire in a heart pacemaker. Heart pacemakers send electrical pulses to the heart through wires inserted into the body's veins as shown in the image. Therefore, the scientist needs the alloy to exhibit the following properties:

- Electrical conductivity
- Malleability
- Ductility
- Durability

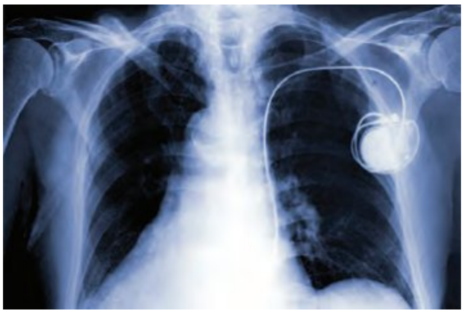


Image: khuruzero/Shutterstock.com

a. What is an alloy? (1 MARK)

b. With reference to metallic bonding, explain why metals are able to conduct electricity. Use a diagram to explain your answer. (3 MARKS)

c. Why would the wires in a heart pacemaker need to be malleable? (1 MARK)

Question 13 (9 MARKS)

Frank has been given three pieces of grey metal which look identical. He knows that one metal is magnesium, one metal is barium and one metal is beryllium. He wants to determine the identity of each metal based on their chemical reactions and properties.

a. Which metal would be expected to have the greatest reactivity? Justify your answer with reference to trends in the periodic table. (2 MARKS)

b. Write the equation of the chemical reaction that would occur between solid barium and water. If this reaction initially occurred in a sealed test tube, what would happen when a flame is placed inside the test tube after removing the seal? (2 MARKS)

c. Frank proposes that he could also determine the reactivity series by leaving the metals out in the open air. What would happen to the metal samples over time (include any relevant equations) and how could this be used to differentiate between them? (5 MARKS)

Key science skills (10 MARKS)

Question 14 (4 MARKS)

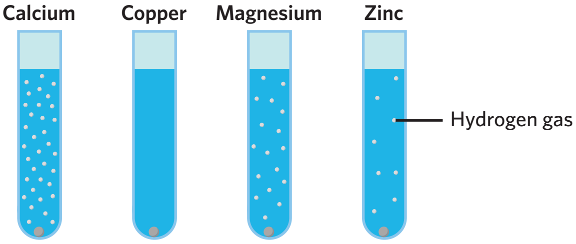
Caesium is a highly reactive metal and is a liquid at room temperature. One of its applications is in atomic clocks which are the primary standards for global timekeeping. A scientist was asked to repair an atomic clock which was being kept in a vacuum with no oxygen present.

a. Before handling caesium, the scientist practices her technique using sodium. Is this a valid way to practice? (2 MARKS)

b. Atomic clocks are very accurate. Do they have to be precise in order to be an effective timekeeping method? Explain your reasoning. (2 MARKS)

Question 15 (6 MARKS)

Patrick and Millie were qualitatively investigating the relative reactivity of different samples of metals in sulfuric acid.



They were having a discussion and decided to write a list of possible variables.

- Mass of each metal sample in grams
- Temperature of the sulfuric acid
- Temperature of the laboratory
- Volume of acid
- Atmospheric pressure in the laboratory
- Concentration of the sulfuric acid
- Temperature of the sulfuric acid

a. Which one of these variables cannot be controlled in this experiment? (1 MARK)

b. What are the independent and dependent variables in this experiment? (2 MARKS)

c. What is an example of a variable that is impossible to control even though it will have a significant impact on the amount of hydrogen gas bubbles produced? (2 MARKS)

d. How could the validity of the methodology used be improved? (1 MARK)

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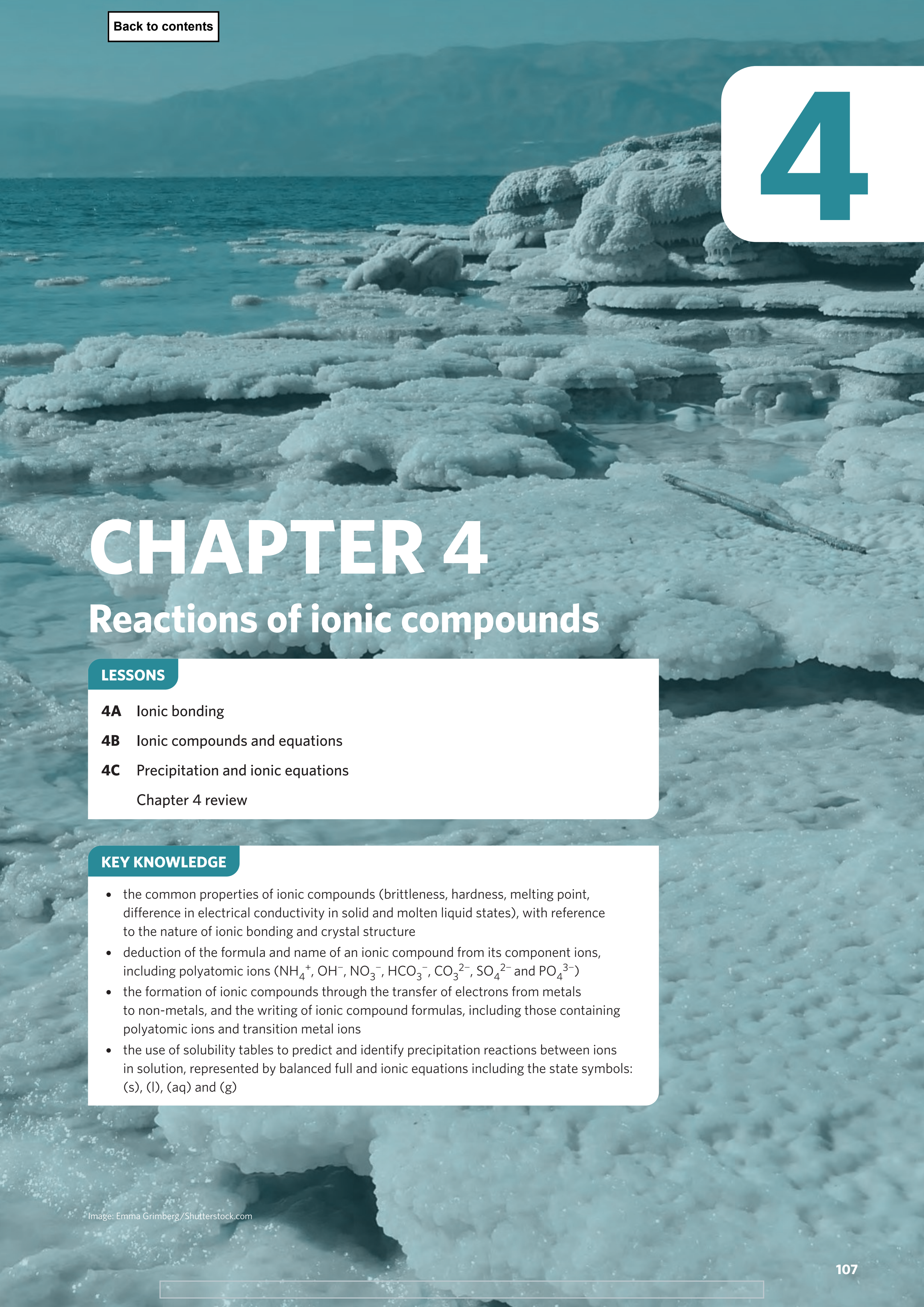

CHAPTER 4

Reactions of ionic compounds

LESSONS

4A Ionic bonding
4B Ionic compounds and equations
4C Precipitation and ionic equations
Chapter 4 review

KEY KNOWLEDGE

- the common properties of ionic compounds (brittleness, hardness, melting point, difference in electrical conductivity in solid and molten liquid states), with reference to the nature of ionic bonding and crystal structure
- deduction of the formula and name of an ionic compound from its component ions, including polyatomic ions (NH_4^+, OH^-, NO_3^-, HCO_3^-, CO_3^{2-}, SO_4^{2-} and PO_4^{3-})
- the formation of ionic compounds through the transfer of electrons from metals to non-metals, and the writing of ionic compound formulas, including those containing polyatomic ions and transition metal ions
- the use of solubility tables to predict and identify precipitation reactions between ions in solution, represented by balanced full and ionic equations including the state symbols: (s), (l), (aq) and (g)

Image: Emma Grimberg/Shutterstock.com

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4A Ionic bonding

STUDY DESIGN DOT POINT

- the common properties of ionic compounds (brittleness, hardness, melting point, difference in electrical conductivity in solid and molten liquid states), with reference to the nature of ionic bonding and crystal structure

4A 4B 4C

1.1.13.1 Ionic bonding model

1.1.13.2 Properties of ionic compounds

ESSENTIAL PRIOR KNOWLEDGE

1B The Bohr model

See question 20.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

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How do ionic bonds form crystals?

The unique properties of ionic compounds make them a versatile ingredient in our daily lives. In this lesson, we will explore the nature of ionic crystals, and the unique properties that they exhibit.

KEY TERMS AND DEFINITIONS

Anion	negative ion formed when an atom gains one or more valence electrons
Brittle	tendency of a material to fracture (break) when subjected to pressure
Coordination number	the number of atoms, molecules or ions bonded to a chemical species
Crystal lattice (ionic)	symmetrical, three-dimensional arrangement of ions
Hardness	resistance to deformation when subjected to pressure
Ion	atom that has lost or gained electrons to become a charged particle
Melting point	temperature at which a substance changes its state from solid to liquid
Molten	a pure substance that has been liquefied by heat

Ionic bonding model 1.1.13.1

The ionic bonding model gives insight into the nature of ionic bonding.

How can the ionic bonding model explain crystal lattice structures?

As we learned in lesson 1B, the Bohr models allow us to see the ratio of cations and **anions** in an ionic bond, as shown in figure 1. However, they do not show how **ions** are arranged in 3 dimensions.

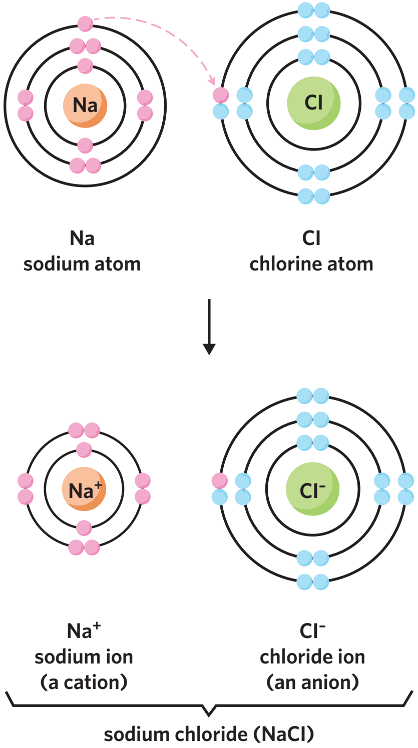

Figure 1 Ionic bond between a sodium and a chloride ion

An ionic compound is actually a three-dimensional structure consisting of multiple ions bonding to each other, rather than a flat, two-dimensional structure of a single cation and a single anion. This is because ionic bonds can be formed between a single cation and multiple anions if they are close enough to each other. Similarly, ionic bonds an be formed between a single anion and multiple cations. In an ionic compound, many cations and anions are bound to each other forming a three-dimensional structure called a **crystal lattice.**

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As shown in figure 2, the crystal lattice of sodium chloride contains multiple sodium and chloride ions bound to each other by electrostatic force. The more ions bound to one another, the greater the electrostatic force holding the ions in place, and thus, the more stable the lattice. Hence, to maximise the stability of the lattice, a sodium ion is bound to multiple chloride ions.

Since, in a crystal lattice, chloride ions are close to each other and sodium ions are close to each other, there is also a repulsive force due to like charges between chloride ions and between sodium ions. However, the electrostatic force between oppositely charged sodium ions and chloride ions is much stronger than the repulsive force, meaning that the structure of the crystal lattice remains relatively stable.

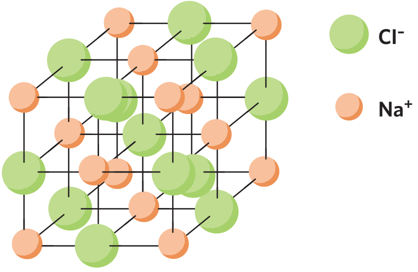

Figure 2 Crystal lattice formed by sodium and chloride ions

USEFUL TIP

The **coordination number** tells us how many other atoms, molecules, or ions are bound to a given ion. For example, in an ionic lattice of sodium chloride, the coordination number of sodium is 6, meaning that one sodium cation is bound to 6 chloride anions. Likewise, one chloride anion is bound to 6 sodium cations.

MISCONCEPTION

'There are covalent bonds in a crystal lattice.'

There are no covalent bonds in a crystal lattice as electrons are not shared between atoms – they are transferred from a metal atom to a non-metal atom.

Progress questions

Question 1

Bohr models can show ________ but cannot demonstrate ________.

A. how ionic bonds are formed, how ions are arranged in space

B. the ratio of cations and anions, how ionic bonds are formed

---

Question 2

In a sample of an ionic compound, each cation is only bound to the anion(s) it donates its electrons to.

A. True

B. False

Question 3

A crystal lattice of an ionic compound is ________ structure because ________.

A. a flat 2D, each cation is only bound to the non-metal ion(s) it donates its electrons to

B. a 3D, a single cation can be bound to multiple anions if they are close enough to each other

---

Question 4

In a sample of sodium chloride, the bond between a sodium ion and the chloride ion it donates its electron to is an ionic bond, while the bond between that sodium ion and the other chloride ions is another type of attractive force.

A. True

B. False

Properties of ionic compounds 1.1.13.2

The physical properties of ionic compounds are determined by the characteristics of ionic bonds.

How can the ionic model explain the physical properties of ionic compounds?

Hardness

Drawing on the ionic bonding model, we know that the ionic bonds between positive and negative ions in a crystal lattice are strong, meaning a large amount of force is required to disrupt these bonds. Therefore, ionic compounds have a high degree of **hardness**. However, when a strong enough force is applied to an ionic compound, ions with the same charge will be pushed closer to each other.

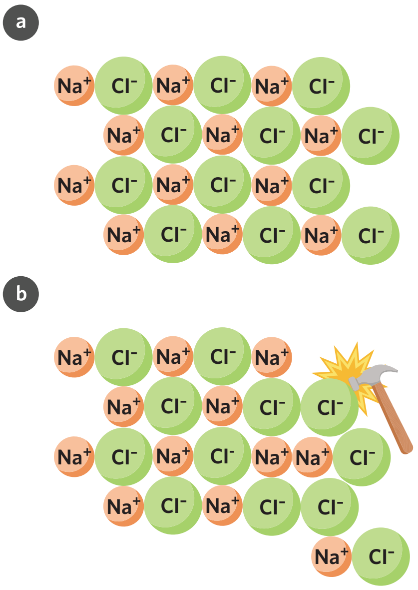

Figure 3 The sodium chloride lattice can shatter when subjected to a strong mechanical force.

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Brittleness

The repulsive force between ions of like charge is strong enough to shatter the lattice. As a result, ionic compounds are also described as being **brittle.**[1] As shown in figure 3, when the crystal lattice of sodium chloride is struck with a hammer, the lattice is disrupted in a way that allows negative chloride ions to be closer to each other. This creates a strong repulsive force between negative chloride ions, causing the lattice to split.

KEEN TO INVESTIGATE?

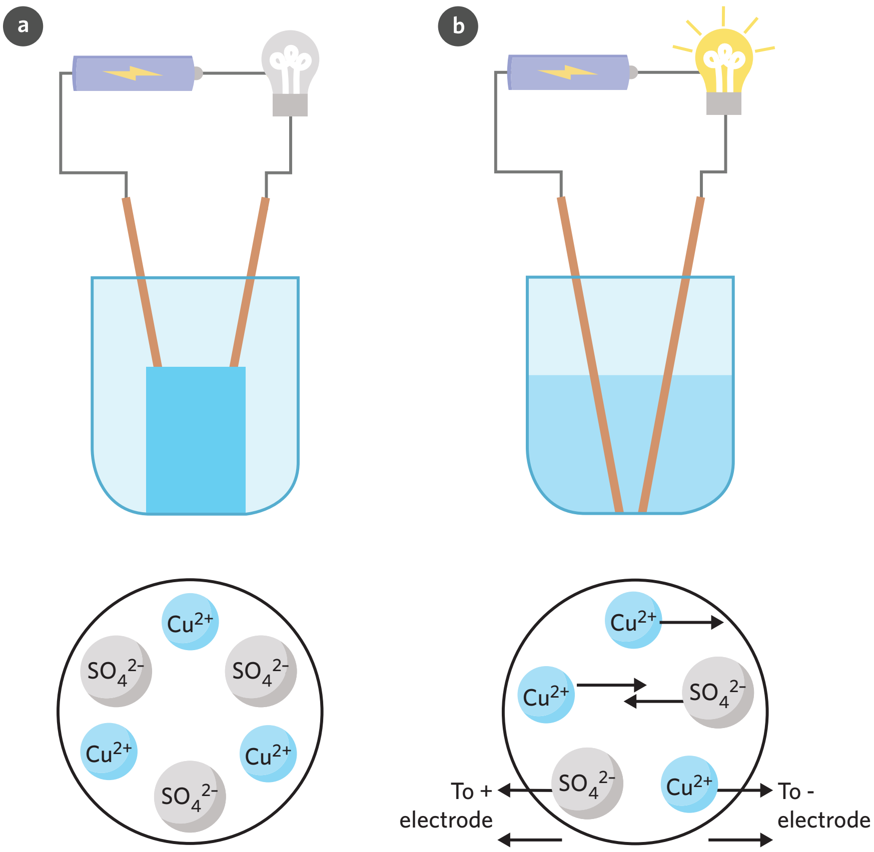
[1] **Why are ionic compounds brittle?**
Search YouTube: Brittleness of an Ionic Crystal

Melting point

Due to the fact that the ionic bonds between positive and negative ions within crystal lattices are strong, a lot of energy is required to break these bonds. Hence, ionic compounds have relatively high **melting points** compared to other substances. For example, the melting point of magnesium oxide, MgO, is approximately 2852 °C.

Conductivity

The ability of ionic compounds to conduct electricity depends on their physical states. For a substance or material to be electrically conductive, it must have mobile charged particles to carry an electric current. However, according to the ionic bonding model, cations and anions are fixed in a crystal lattice, meaning there is no movement of charged particles. As such, an ionic compound in its solid state is unable to conduct electricity, as shown in figure 4a.



A current cannot flow through a solid ionic compound

A current can flow through a molten ionic compound

Figure 4 The electrical conductivity of ionic compounds in **(a)** solid state and **(b)** molten state

However, if an ionic compound is in its **molten** (liquid) state or dissolved in water (aqueous), its ionic lattice is broken, resulting in individual cations and anions that are able to move freely. Because of this, an ionic compound in liquid or aqueous state can conduct electricity. As shown in figure 4b, when the ionic compound is now dissolved in water, individual cations and anions can move around in the solution. Consequently, the ions are able to carry electric charge through the solution, allowing the electric current to continue moving around the circuit. This makes the light bulb in the circuit shine, demonstrating that the ionic compound is now conducting electricity.

USEFUL TIP

In general, the greater the concentration of an ionic compound dissolved in a solution, the greater the electrical conductivity of the solution.

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Progress questions

Question 5

Properties of ionic compounds are determined by the nature of ionic bonds and the structure of 3D ionic crystal lattices.

A. True

B. False

Question 6

Ionic compounds are hard because

A. repulsive forces between positive and negative ions are strong.

B. ionic bonds between cations and anions are strong.

Question 7

Which of the following best explains the brittleness of an ionic compound?

A. When a strong force is exerted on an ionic compound, repulsive forces between like-charged ions will shatter and break the lattice.

B. When a strong force is exerted on an ionic compound, repulsive forces between delocalised ions separate the lattice.

Question 8

The majority of ionic compounds have ________ melting points due to the ________ ionic bonds between ________ and ________.

A. low, weak, cations, cations

B. high, strong, cations, anions

Question 9

Which of the following is true about the electrical conductivity of ionic compounds in solid state?

A. Ionic compounds are able to conduct electricity when in solid state due to the movement of delocalised electrons within ionic lattices.

B. Ionic compounds are not able to conduct electricity when in solid state because cations and anions are fixed in ionic lattices.

Question 10

In an aqueous state, ionic compounds ________ conduct electricity because ________.

A. cannot, cations and anions cannot move freely

B. can, cations and anions can move freely

Theory summary

- Ionic compounds form 3D crystal lattices.
- Ionic compounds are hard and brittle.
- Ionic compounds have high melting points.
- Ionic compounds cannot conduct electricity in solid states.
- Ionic compounds can conduct electricity in molten and aqueous states.

The content in this lesson is considered fundamental prior knowledge to primary galvanic cells and fuel cells as sources of energy (Unit 3 AOS 1).

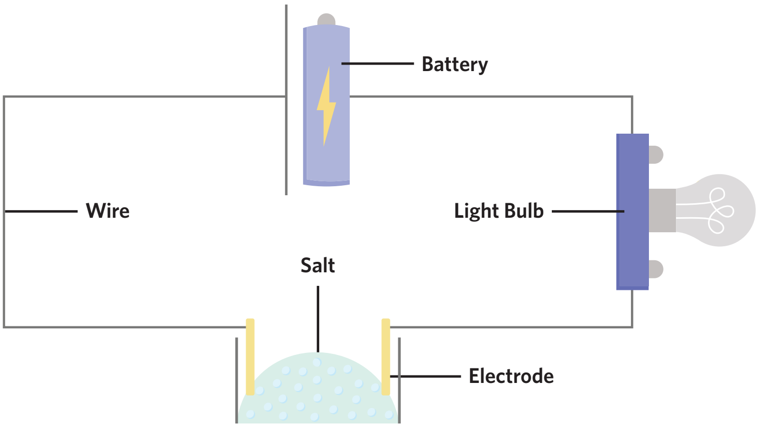
4A Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 11–13.

Jason, Angus and Liam go camping in the Flinders Ranges in South Australia. After staying there for 3 days, all of their flashlights have run out of power. Luckily, Angus brought a battery, a big box of table salt, a light bulb, 2 electrodes and some wires. Jason uses the solid table salt to set up a simple electrical circuit as shown. However, the light does not shine.



Question 11 (1 MARK) 🌶

Table salt is composed primarily of

A. sodium oxide.

B. sodium chloride.

C. magnesium hydroxide.

D. magnesium oxide.

Question 12 (1 MARK) 🌶

What is the key reason why the light bulb does not shine?

A. There is not enough salt in the box.

B. Table salt never conducts electricity.

C. The salt is in solid form where all ions are strongly held to each other so there is no movement of ions.

D. The salt is in solid form where plenty of ions are moving, which results in big lumps of ions.

Question 13 (3 MARKS) 🌶🌶

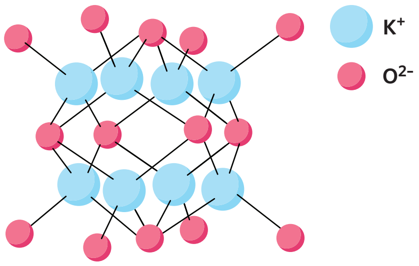
Suggest one change that can be made to the electrical circuit to allow the light bulb to shine. Justify your answer.

Exam-style

Question 14 (2 MARKS) 🌶

Consider the crystal lattice of potassium oxide (K_2O). Each potassium ion bonds to four neighbouring oxide ions as shown.

A student stated that the bond between each potassium ion and the oxide ion it donates its electrons to is an ionic bond, whilst the bonds between that potassium ion and the other three neighbouring oxide ions are not ionic bonds. Is the student's statement accurate? Justify your answer.



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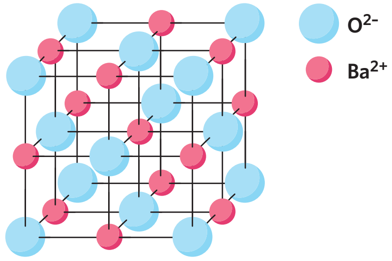
Question 15 (3 MARKS)

Carrie is a geologist investigating the properties of limestone, a sedimentary rock made of calcium carbonate. She uses a hammer to break the limestone. After striking the piece of limestone three times, it remains unbroken. On the fourth attempt, she uses all of her strength to hammer the limestone, finally shattering the rock. Explain why the limestone was not broken in the first three attempts but broke when Carrie applied a stronger force.

Question 16 (2 MARKS)

Barium oxide (BaO), whose lattice structure is given, has a melting point of 2000 °C. A chemistry student attributed the high melting point of barium oxide to the strong repulsive forces between Ba^{2+} cations. Is the student's explanation correct? Explain your answer.



Question 17 (2 MARKS)

The following table shows the ion content of tap water.

Ion	Percentage (%)
K^+	0.08
Na^+	1.07
Mg^{2+}	1.2
Ca^{2+}	1.4
Cl^-	0.77
SO_4^{2-}	0.7
HCO_3^-	1.8

Using the data presented in the table, explain why people can be electrocuted if a hair dryer is dropped into a bathtub filled with tap water.

Key science skills

Question 18 (6 MARKS)

Ella conducted an experiment to investigate the electrical conductivity of sodium chloride (NaCl).

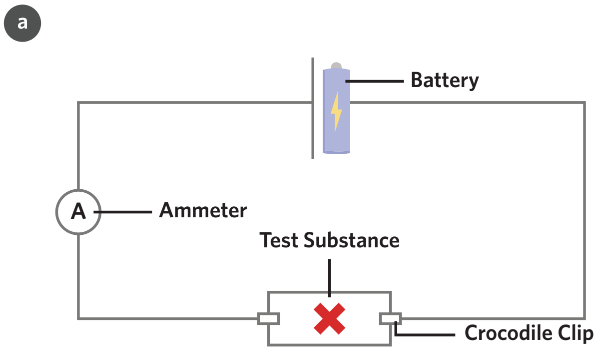
Materials

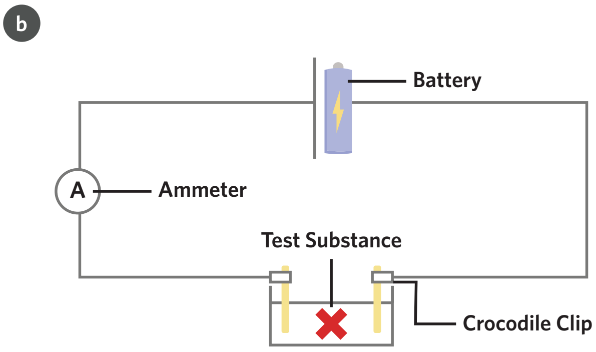
- 15 g of solid sodium chloride (NaCl)
- Distilled water
- 1 battery
- Crocodile clips
- 2 carbon rods
- Conducting wires
- 1 ammeter – the instrument used to measure the current in an electrical circuit.
- 2 beakers

Method

1. Use conducting wires to connect the battery and the ammeter as shown in diagram A.
2. Attach each crocodile clip to each end of 5 g of solid NaCl as seen in diagram (a) below.
3. Leave for 30 seconds.
4. Read the ammeter and record the current value.
5. Remove the solid sodium chloride.
6. Dissolve 4 g of solid NaCl into beaker I containing 500 mL of distilled water as in diagram (b) below.

7. Dissolve 6 g of solid NaCl into beaker II containing 500 mL of distilled water.
8. Use two carbon rods to connect the circuit to the solution in beaker I as shown in diagram (b).
9. Leave for 30 seconds.
10. Read the ammeter and record the current value.
11. Clean carbon rods.
12. Repeat steps 8, 9, 10 and 11 for beaker II.





Ella's results are presented in the given table.

Substance	Current (amperes)
Solid NaCl	0
Solution in beaker I	20
Solution in beaker II	23.5

a. Identify the independent and dependent variables. (2 MARKS)

b. What can be concluded about the electrical conductivities of solid NaCl, and NaCl solution in beaker I and beaker II? (2 MARKS)

c. Explain your conclusion about the electrical conductivities of NaCl solution in beaker I and beaker II. (2 MARKS)

FROM LESSONS 16A & 16C

Questions from multiple lessons

Question 19 (4 MARKS)

A sodium lattice has a high value of electrical conductivity, while a sodium chloride lattice cannot conduct electricity. Explain why this is the case with reference to the structure of each lattice.

FROM LESSON 3A

Question 20 (4 MARKS)

Lithium fluoride is widely used to make components of engines due to its high melting point of 848.2 °C.

a. Use Bohr models to demonstrate the formation of lithium and fluoride ions. (2 MARKS)

b. Explain why lithium fluoride has a high melting point. (2 MARKS)

FROM LESSONS 1B & 4A

Hints

14. Metals donate electrons to non-metals in such a way that both metals and non-metals gain a complete valence electron shell.

15. Disruptions in the ionic lattice cause the substance to break.

16. A high melting point is an indication that a lot of energy is required to disrupt the bonds present within a substance.

17. Moving ions can carry electrical current.

18a. An independent variable is changed by the experimenter. A dependent variable is the variable being measured or tested.

18b. A positive current value of ammeter indicates that a substance can conduct electricity.

18c. For ionic compounds in solid state, cations and anions are held together by ionic bonds. In aqueous solution, ions can move.

19. Sodium is a metallic substance.

20a. Metals donate electrons to non-metals in such a way that both metals and non-metals gain a complete valence electron shell.

20b. Ionic bonds present in ionic compounds are strong.

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4B Ionic compounds and equations

Image: Pixel-Shot/Shutterstock.com

What is the protective layer of our teeth made of?

Ionic compounds like calcium phosphate are key components of tooth enamel. In this lesson, we will learn how to name and write the formulas and reaction equations of ionic compounds like calcium phosphate, as well as how they form.

KEY TERMS AND DEFINITIONS

Ionic bond electrostatic force of attraction between cations and anions

Ionic compound compound made up of cations and anions held together by electrostatic forces of attraction

Monatomic ion ion made up of only one atom

Polyatomic ion ion made up of a group of atoms covalently bonded together

STUDY DESIGN DOT POINTS

- deduction of the formula and name of an ionic compound from its component ions, including polyatomic ions (NH_4^+, OH^-, NO_3^-, HCO_3^-, CO_3^{2-}, SO_4^{2-} and PO_4^{3-})
- the formation of ionic compounds through the transfer of electrons from metals to non-metals, and the writing of ionic compound formulas, including those containing polyatomic ions and transition metal ions

4A 4B 4C

1.1.14.1 Writing ionic formulas

1.1.15.1 Formation of ionic compounds

1.1.15.2 Writing balanced equations for the reaction between ionic compounds

ESSENTIAL PRIOR KNOWLEDGE

1B	The Bohr model
1C	Ionisation energy
1C	Transition metals
4A	Crystal structure of ionic compounds

See questions 21–24.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Writing ionic formulas 1.1.14.1

Despite being made up of charged particles, ionic compounds do not have an overall charge.

How can charges be used to balance ionic formulas?

As we learnt in lesson 4A, ionic compounds consist of cations and anions arranged in a three-dimensional crystal lattice. Since the lattice structure can go on indefinitely, it would be difficult to write a formula that represents the total number of ions involved in the structure. Unlike molecular formulas, which represent the total number of atoms in a molecule, ionic formulas represent the ratio of ions found in the compound.

When writing ionic formulas for **monatomic ions**, there are a few key ideas that we need to keep in mind:

- The cation is written first, followed by the anion.
- Ionic compounds are neutral compounds – they do not have an overall charge.
- Subscripts are used to denote multiples of the same ion.
- Superscripts are used to denote the charge of an ion e.g. Mg^{2+}

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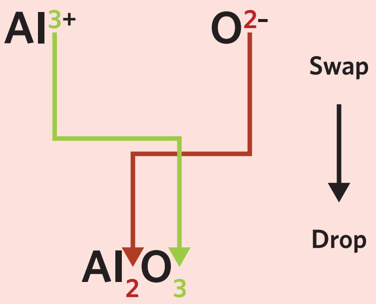
USEFUL TIP

The ‘swap’n’drop’ method can be used to balance ionic formulas when the charges on each ion are not the same. To do so, take the charge of each ion as the (multiplying) subscript for the partnering ion. Where necessary, the formula can then be reduced to the lowest ratio. If the charge on each ion is the same, we don’t need to do anything they are already balanced!



MISCONCEPTION

‘Ca^{+2}’

When describing the charge of a calcium ion, we say and write +2; however, when writing the charge on an ion that gains or loses more than one electron, the number is written before the + or − sign e.g. Ca^{2+}.

Since we know the charges of both the cation and the anion, and that the overall charge of the compound must sum to zero, we can write the ionic formula for the compound by balancing the charges. This process is shown in figure 1.

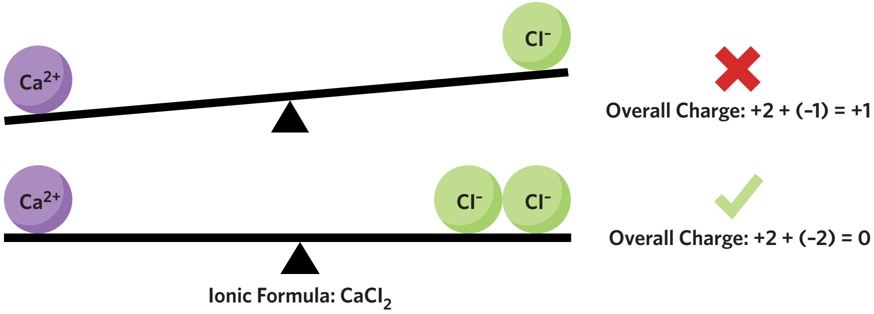


Figure 1 The balanced ionic formula for calcium chloride

So far, we have focused on monatomic ions; ions composed of only one atom. However, ions sometimes exist as **polyatomic ions**; ions that are made from multiple atoms covalently bonded together. For example, sulfate ions (SO_4^{2-}) have an overall charge of −2. When writing ionic formulas involving polyatomic ions, it’s important to remember that the polyatomic ion’s overall charge is the cumulative charge of each atom.

MISCONCEPTION

‘The formula for magnesium hydroxide is $MgOH_2$.’

When we have multiple polyatomic ions in an ionic compound, we write the polyatomic ion in brackets with its subscript multiple outside the brackets, to prevent confusion with any subscript numbers within the polyatomic ion The charge on Mg is +2 and the charge on OH is −1 and therefore the correct formula is $Mg(OH)_2$.

WORKED EXAMPLE 1

An ionic compound is made up of SO_4^{2-} anions and Na^+ cations. Write the ionic formula for this compound.

What information is presented in the question?

The ions found in the ionic compound.

What is the question asking us to do?

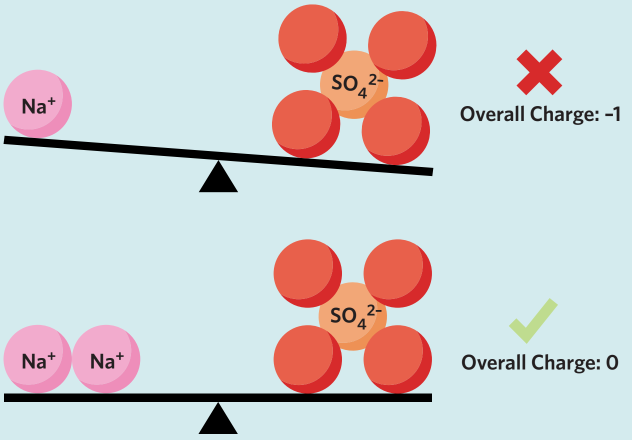
Write an ionic formula for the ionic compound.

What strategies do we need in order to answer the question?

1. Find the ratio of ions required to form a neutral ionic compound.
2. Use the established ratio to write the balanced formula.

Answer

Based on the information given, we know that there is an imbalance of charge between both ions. We also know that ionic compounds have an overall charge of zero, therefore we would need to balance the charges of each ion so that they cancel out.



Two Na^+ ions are required to balance the charge of a single SO_4^{2-} ion. As such, the ionic formula is Na_2SO_4.

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WORKED EXAMPLE 2

Write the ionic formula for the ionic compound consisting of CO_3^{2-} and Al^{3+} ions.

What information is presented in the question?

The ions found in the ionic compound.

What is the question asking us to do?

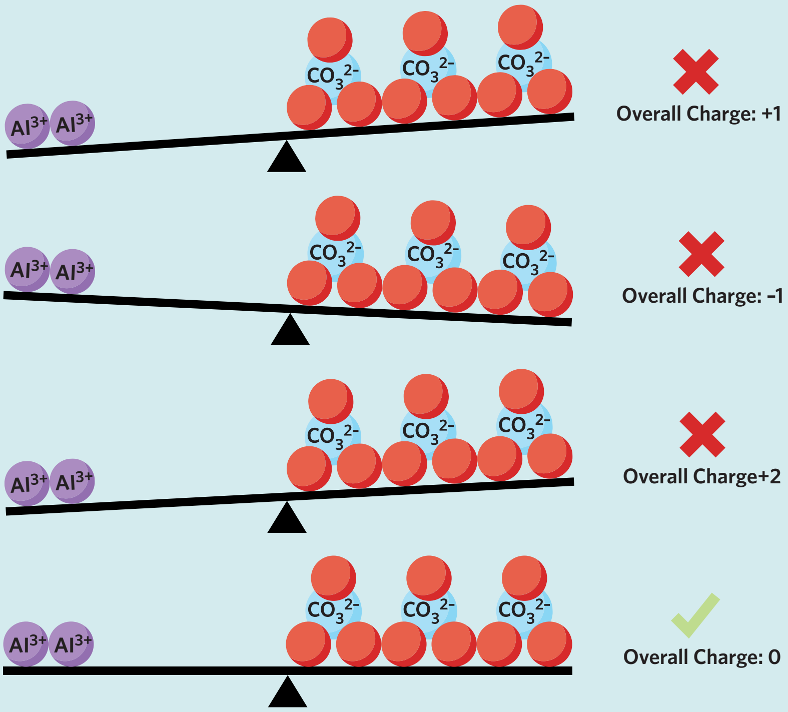
Write an ionic formula for the ionic compound formed between carbonate and aluminium ions.

What strategies do we need in order to answer the question?

1. Balance the charge of the ions.
2. Use the established ratio to write the balanced formula.

Answer

Based on the information given, we know that there is an imbalance of charge between both ions. Therefore, we need to find the appropriate ratio of both ions to ensure that the overall charge of the ionic compound is zero.



The balanced ionic formula is $Al_2(CO_3)_3$.

Table 1 List of common ions and their charges

Cations			Anions		
Charge	**Name of ion**	**Formula**	**Charge**	**Name of ion**	**Formula**
+1	Ammonium	NH_4^+	−1	Bromide	Br^-
	Copper (I)	Cu^+		Chloride	Cl^-
	Hydrogen	H^+		Hydroxide	OH^-
	Lithium	Li^+		Fluoride	F^-
	Potassium	K^+		Iodide	I^-
	Silver	Ag^+		Nitrate	NO_3^-
	Sodium	Na^+		Hydrogen–carbonate	HCO_3^-
				Hydride	H^-

Continues →

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Table 1 Continued

Cations			Anions		
Charge	**Name of ion**	**Formula**	**Charge**	**Name of ion**	**Formula**
+2	Calcium Barium Copper (II) Iron (II) Lead (II) Magnesium Nickel (II) Strontium Zinc	Ca^{2+} Ba^{2+} Cu^{2+} Fe^{2+} Pb^{2+} Mg^{2+} Ni^{2+} Sr^{2+} Zn^{2+}	−2	Carbonate Sulfate Sulfide Oxide	CO_3^{2-} SO_4^{2-} S^{2-} O^{2-}
+3	Aluminium Iron (III)	Al^{3+} Fe^{3+}	−3	Nitride Phosphate	N^{3-} PO_4^{3-}

Table 1 shows the charges of a number of common ions. If we look closely, we can see that for most of the monatomic ions (excluding the transition metals), the charge of the ions can be determined by looking at the periodic table. For example, Na^+, K^+ and Li^+ are metallic elements found in group 1 on the periodic table. As we know, elements in group 1 contain one valence electron and therefore these atoms would prefer to lose one electron to have a full outer shell. Consequently, these atoms become cations with a charge of +1.

MISCONCEPTION

'All transition metals can form only one ion.'

Transition metals, unlike the metallic elements found in groups 1 and 2, can form multiple ions with different charges. For the ions of these transition metals, we use Roman numerals to show the charge of the ion. For example, iron (III) has a charge of +3, whilst iron (II) has a charge of +2.

Progress questions

Question 1

When writing the formula of an ionic compound, the ________ comes first.

A. cation

B. anion

Question 2

Ionic bonds can only form between ions of like charge.

A. True

B. False

Question 3

What is the charge on a sulfate ion?

A. +2

B. −2

Question 4

Which of the following ionic compounds possesses a positively charged polyatomic ion?

A. NH_4Cl

B. $BaSO_4$

Question 5

Which of the following ionic compounds is not written according to convention?

A. $(H)_2PO_4$

B. $Al_2(SO_4)_3$

Continues →

Question 6

Which of the following ionic compounds contains the vanadium (III) ion?

A. V_2O_3

B. V_2O_5

How are ionic compounds named?

There are a few simple guidelines when it comes to naming ionic compounds:

- The name consists of two words.
- The first word is the full, unchanged name of the element or molecule from which the cation is formed.
- The second word is derived from the name of the element or molecule from which the anion is formed. For elemental (i.e. monoatomic) anions, the ending of the element's name is replaced with – ide e.g. phosphide P^{3-}. The names of a number of common anions can be found in table 1.

WORKED EXAMPLE 3

Give the name of the ionic compound KBr.

What information is presented in the question?

The ionic formula for the compound.

What is the question asking us to do?

Name the ionic compound with the formula KBr.

What strategies do we need in order to answer the question?

1. Identify the cation and the anion in the formula.
2. Identify the name of the ionic compound.

Answer

According to the formula, the cation has the elemental symbol K and the anion has the elemental symbol Br. The periodic table tells us that K is the symbol for potassium and Br is the symbol for bromine. To name the ionic compound, we use the full name of the cation, which is potassium, and change the name of the element of the anion to bromide since it is an elemental ion. Thus, the name of this compound is potassium bromide.

By knowing how to name ionic compounds, we can use our knowledge of balanced ionic compounds to determine the ionic formula for these compounds.

WORKED EXAMPLE 4

Write the formula for copper (I) oxide.

What information is presented in the question?

The name of the ionic compound.

What is the question asking us to do?

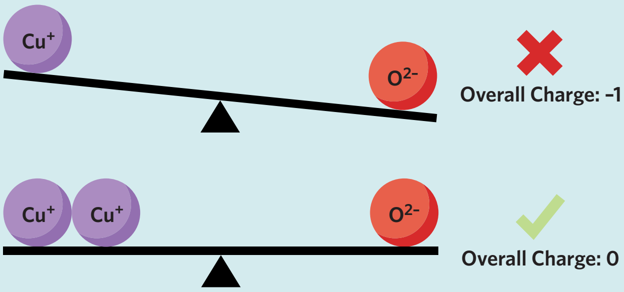
Give the ionic formula for copper (I) oxide.

What strategies do we need in order to answer the question?

1. Identify the charge of each ion.
2. Find the ratio of ions required for a neutral ionic compound.
3. Use the established ratio to write the balanced formula.

Answer

Based on table 1, copper (I) is represented as Cu^+ and oxide is represented as O^{2-}.



The ionic formula for copper (I) oxide is Cu_2O.

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Progress questions

Question 7

What is the name of the ionic compound with formula Ag_2SO_4?

A. Sulfate disilver

B. Silver sulfate

Question 8

Which of the following ionic compounds is ammonium nitrate?

A. $(NH_4)_3N$

B. NH_4NO_3

Formation of ionic compounds 1.1.15.1

The majority of ionic compounds are formed from metal atoms and non-metal atoms.

How are ionic compounds formed?

Based on our understanding of trends in the periodic table from lesson 1C, we know that metal atoms have low first ionisation energies and tend to lose electrons readily, forming cations. Non-metal atoms, on the other hand, tend to attract electrons to form a full valence shell.

Given metal atoms tend to lose electrons while non-metal atoms tend to gain electrons, the two resultant ions are of opposite charge. The electrostatic attraction between the oppositely charged ions forms an **ionic bond** as shown in figure 2.

STRATEGY

On the Pauling scale, a difference in electronegativity of 2.1 or higher creates an ionic bond.

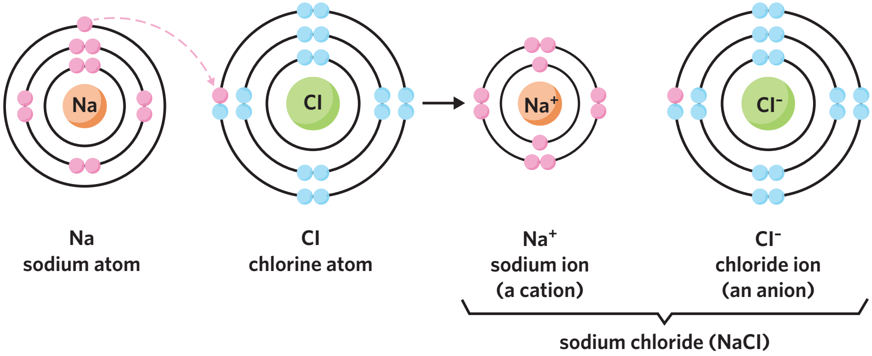

Figure 2 Transfer of an electron from a sodium atom to a chlorine atom

MISCONCEPTION

'Ammonia is a metal.'

Ammonia is made of two non-metal elements (nitrogen and hydrogen) covalently bonded together to form a molecule. Although its ion, ammonium, behaves similarly to metals in the formation of ionic compounds by acting as a cation, it's not a metal cation.

In a crystal lattice, individual ions will form ionic bonds with as many oppositely charged ions as possible, generating maximum attraction and minimum repulsion between ions. Figure 3 shows the bonds formed between a single chloride ion and 6 sodium ions in a sodium chloride lattice.

The ratio of ions in an ionic compound is dictated by the number of electrons that can be lost by the cation-forming element or compound, and the number accepted by the anion-forming element or compound. In figure 4, magnesium loses two electrons from its valence shell to obtain a full valence shell. However, chlorine only needs to accept one electron to obtain a full valence shell. Therefore, the two electrons lost by a single magnesium atom can be accepted by two chlorine atoms.

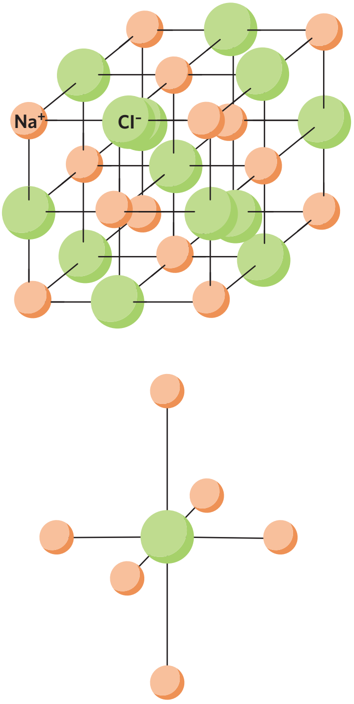

Figure 3 In NaCl, each ion is surrounded by six oppositely charged ions.

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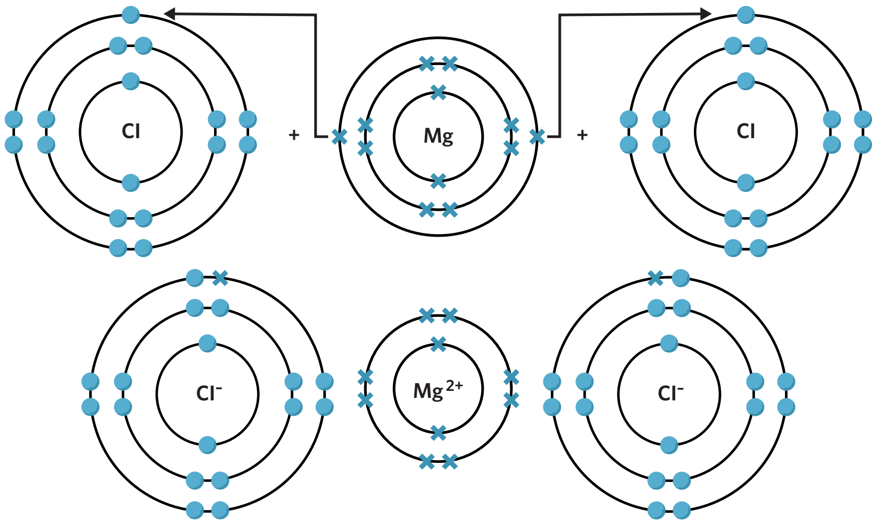


Figure 4 Formation of magnesium chloride from one magnesium and two chlorine atoms

USEFUL TIP

Since ions in a compound are arranged in such a way that attraction is maximised and repulsion is minimised, ionic compounds have strong bonds and thus high melting points.

WORKED EXAMPLE 5

Use Bohr models to explain the formation of the ionic compound Al_2S_3.

What information is presented in the question?

The ionic formula Al_2S_3.

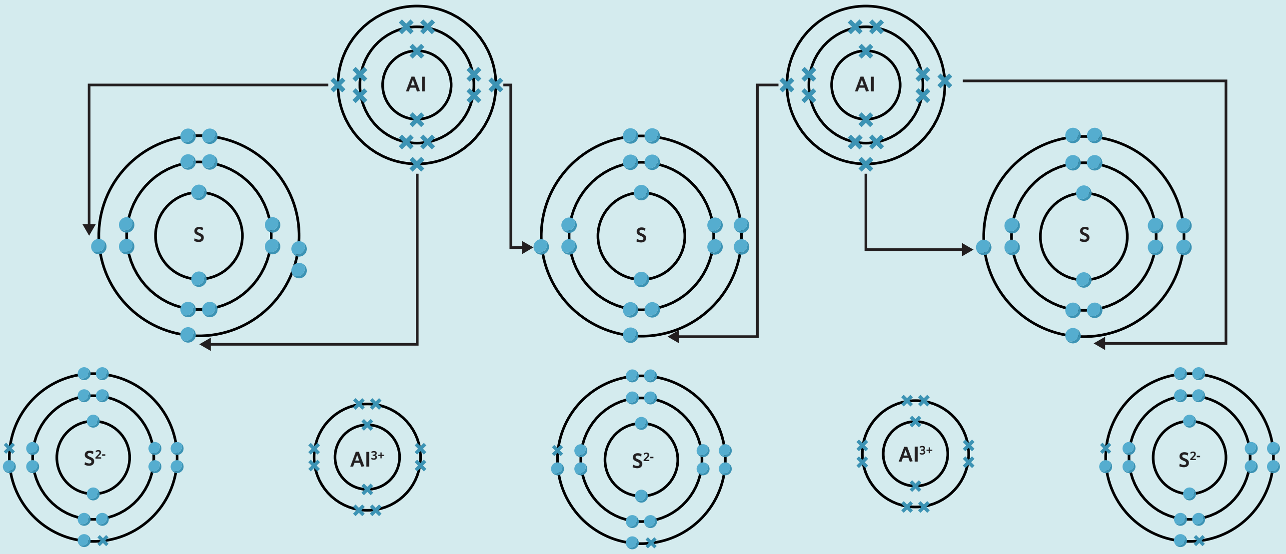
What is the question asking us to do?

Use Bohr models to explain the formation of the ionic compound.

What strategies do we need in order to answer the question?

1. Write the Bohr model of one Al atom.
2. Write the Bohr model of one S atom.
3. Determine the number of electrons that can be removed from one Al atom.
4. Determine the number of electrons that can be accepted by one S atom.
5. Determine the number of Al and S atoms required for a neutral compound.
6. Write the formation of Al_2S_3 from Al atoms and S atoms.

Answer



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Progress questions

Question 9

An ionic bond is formed by oppositely charged ions.

A. True

B. False

Question 10

Barium can lose two valence electrons to form a cation. Which of the following shows the correct symbol for a barium ion?

A. $2Ba^{+}$

B. Ba^{2+}

Question 11

A charged particle cannot form an ionic bond with more than one other charged particle.

A. True

B. False

Writing balanced equations for the reaction between ionic compounds 1.1.15.2

Some reactions between ionic compounds result in the 'swapping' of ion pairs.

How are ions exchanged between compounds?

Ionic compounds can take part in reactions where the ions of two different compounds are exchanged to create two new compounds. An example of this exchange is the reaction between silver chloride, AgCl, and lithium iodide, LiI, shown in figure 5.

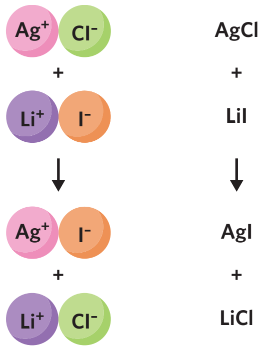

Figure 5 The reaction between silver chloride and lithium iodide

Depending on which perspective we look at it, we can say that there was a swapping of either cations or anions (but not both), resulting in the formation of two new substances. When writing the complete ionic equation for these reactions, we need to:

1. Establish the products formed as a result of the exchanging of ions between the reactants.
2. Write the ionic formula for each product.
3. Ensure that both the reactant and product side have the same number of atoms of each element by balancing the full equation.

WORKED EXAMPLE 6

Write the balanced equation for the reaction between magnesium chloride, $MgCl_2$, and potassium fluoride, KF.

What information is presented in the question?

The names and ionic formulas of each reactant involved in the reaction.

What is the question asking us to do?

Write a balanced equation.

What strategies do we need in order to answer the question?

1. Establish the products formed as a result of the swapping of ions between both reactants.
2. Identify the ionic formula for each product.
3. Write an unbalanced equation for the reaction.
4. Balance the equation.

Continues →

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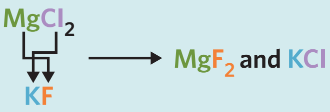
Answer

The ions are swapped between compounds, meaning magnesium becomes bound to fluoride, and potassium becomes bound to chloride. Since magnesium ions have a charge of +2, and fluoride ions have a charge of −1, two fluoride ions are needed to balance the compound, producing MgF_2. Potassium ions have a charge of +1, and chloride ions have a charge of −1, so only one chloride ion is needed to balance the compound, producing KCl.



Now, we need to write the equation for this reaction:
$MgCl_2 + KF \rightarrow MgF_2 + KCl$

Currently, this equation is not balanced. To balance the equation, we need to make sure that there are the same number of atoms of each element on both the reactant and product side of the reaction:
$MgCl_2 + 2KF \rightarrow MgF_2 + 2KCl$

WORKED EXAMPLE 7

Write the balanced equation for the reaction between sodium nitrate, $NaNO_3$, and barium iodide, BaI_2.

What information is presented in the question?

The names and ionic formulas of each reactant involved in the reaction.

What is the question asking us to do?

Write a balanced equation.

What strategies do we need in order to answer the question?

1. Establish the products formed as a result of the swapping of ions between both reactants.
2. Identify the ionic formula for each product.
3. Write an unbalanced equation for the reaction.
4. Balance the equation.

Answer

The ions are swapped between compounds, meaning sodium becomes bound to iodide and barium becomes bound to nitrate. Since sodium ions have a charge of +1, and iodide ions have a charge of −1, one iodide ion is needed to balance the compound, producing NaI. Barium ions have a charge of +2, and nitrate ions have a charge of −1, so two nitrate ions are needed to balance the compound, producing $Ba(NO_3)_2$.

Now we need to write the equation for this reaction:
$NaNO_3 + BaI_2 \rightarrow NaI + Ba(NO_3)_2$

Currently, this equation is not balanced. To balance the equation, we need to make sure that there are the same number of atoms of each element on both the reactant and product side of the reaction:
$2NaNO_3 + BaI_2 \rightarrow 2NaI + Ba(NO_3)_2$

Progress questions

Question 12

Which of the following represents a balanced equation?

A. $Na_2SO_4 + AgCl \rightarrow NaCl + Ag_2SO_4$

B. $Na_2SO_4 + 2AgCl \rightarrow 2NaCl + Ag_2SO_4$

Question 13

Which of the following represents a reaction between lead (II) nitrate and iron (III) sulfate?

A. $3Pb(NO_3)_2 + Fe_2(SO_4)_3 \rightarrow 3PbSO_4 + 2Fe(NO_3)_3$

B. $Pb(NO_3)_4 + 2FeSO_4 \rightarrow Pb(SO_4)_2 + 2Fe(NO_3)_2$

Theory summary

- Ionic compounds are neutral compounds.
- When establishing the formula for ionic compounds, the charges of cations and anions present must sum to zero.
- The 'swap'n'drop' method can be used to quickly determine the formula of an ionic compound.
- Polyatomic ions are considered as a single unit when writing an ionic formula.
- The ratio of atoms of each element in an ionic compound depends on the number of electrons removed from a metal and accepted by a non-metal.
- Most ionic compounds are formed from metal atoms and non-metal atoms.
- In an ionic compound, the bond between metal ions and non-metal ions is called an ionic bond.
- A single ion is surrounded by, and therefore bonded to, multiple oppositely charged ions.
- The strength of the ionic bond is greater than the forces of repulsion between ions of the same charge.
- Ionic compounds can take part in reactions that result in the 'swapping' of ion pairs.
- When writing balanced ionic equations, ionic formulas must be written accurately to ensure that there are the same number of atoms of each element on the reactant and product side.

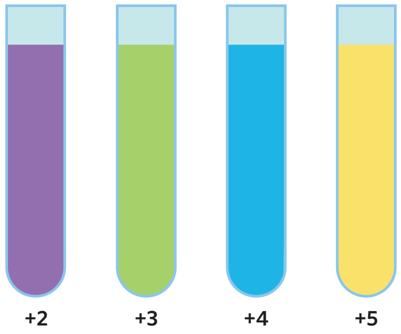
4B Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 14-16.

The ions of transition metals can have a number of different charges. For example, vanadium can form ions with charges of +2, +3, +4, +5, each with different colours.



Question 14 (1 MARK) 🌶

Which of the following correctly represents vanadium with a charge of +5?

A. VCl_5 **B.** V_2O_3 **C.** VCl_4 **D.** VO

Question 15 (1 MARK) 🌶

Which of the following correctly represents vanadium with a charge of +2?

A. VCl_5 **B.** V_2O_3 **C.** VCl_4 **D.** VO

Question 16 (2 MARKS) 🌶🌶

Write the balanced equation for the reaction between vanadium (V) chloride and vanadium (II) oxide. States are not required.

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Exam-style

Question 17 (9 MARKS)

Complete the following table.

	Name of Ion	Formula
a.	Silver	
b.		Cu^{2+}
c.	Hydroxide	
d.		F^-
e.	Nitrate	
f.		NH_4^+
g.	Sulfate	
h.		Fe^{3+}
i.	Phosphate	

Question 18 (9 MARKS)

Give the name and ionic formula of the ionic compounds created from the following cation and anion pairs.

		Name	Formula
a.	Na^+ and NO_3^-		
b.	Silver ion and sulfate ion		
c.	NH_4^+ and F^-		
d.	Hydrogencarbonate ion and strontium ion		
e.	K^+ and O^{2-}		
f.	Calcium ion and carbonate ion		
g.	Hydroxide ion and aluminium ion		
h.	Fe^{3+} and S^{2-}		
i.	Phosphate ion and Zn^{2+}		

Question 19 (8 MARKS)

Write the balanced equations for the reactions between the following pairs of ionic compounds. States are not required.

a. Copper (I) sulfate and nickel (II) iodide. (2 MARKS)

b. K_2O and $FeSO_4$ (2 MARKS)

c. Magnesium nitride and sodium hydrogencarbonate. (2 MARKS)

d. $Ba_3(PO_4)_2$ and NH_4F (2 MARKS)

Key science skills

Question 20 (5 MARKS)

Megan is investigating three newly discovered elements, X, Y and Z, which behave as cations in ionic compounds. She knows that the charge of different ions can be determined by the ratio of atoms of each element in an ionic compound. In order to find out the charges, she weighs the mass of each element present and, after a few calculations, determines the unknown charge of the cation.

a. What type of data is collected? (1 MARK)

b. What is the independent and dependent variable? (2 MARKS)

c. The masses she recorded in her logbook were based on a scale that wasn't calibrated. What type of error is this and how could her data be corrected? (2 MARKS)

FROM LESSONS 16A, 16B & 16D

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Questions from multiple lessons

Question 21 (5 MARKS)

Fluoride ions are essential for the production of fluorapatite which is found in tooth enamel. The ionic formula of fluorapatite is $Ca_5(PO_4)_3F$.

a. Name and identify all ions present in fluorapatite and their charges. (3 MARKS)

b. What properties of fluorapatite make it suitable to function as tooth enamel? Justify your answer. (2 MARKS)

FROM LESSON 4A

Hints

17a. Silver loses 1 electron to form an ion.

17b. Transition metals can form ions of different charges.

17c. Oxygen gains 2 electrons to form an ion, whilst hydrogen loses one electron to form an ion.

17d. Group 17 elements form anions by gaining an electron.

17e. Nitrogen loses 5 electrons to form an ion, whilst each oxygen atom gains 2 electrons to form an ion.

17f. This polyatomic ion is neutral in the form NH_3.

17g. Sulfur loses 6 electrons to form an ion, whilst each oxygen atom gains 2 electrons to form an ion.

17h. Transition metals can form ions of different charges.

17i. Phosphorus loses 5 electrons to form an ion, whilst each oxygen atom gains 2 electrons to form an ion.

18a. Balance the charges to create a neutral compound.

18b. The silver ion has a charge of +1, and the sulfate ion has a charge of −2.

18c. Balance the charges to create a neutral compound.

18d. The hydrogencarbonate ion has a charge of −1, and the strontium ion has a charge of +2.

18e. Balance the charges to create a neutral compound.

18f. The calcium ion has a charge of +2, and the carbonate ion has a charge of −2.

18g. The hydroxide ion has a charge of −1, and the aluminium ion has a charge of +3.

18h. Balance the charges to create a neutral compound.

18i. The phosphate ion has a charge of −3.

19a. The copper (I) ion has a charge of +1.

19b. In this reaction, iron is present as Fe^{2+}.

19c. 3 magnesium ions and 2 nitride ions are required to balance the first compound.

19d. NH_4^+ is a polyatomic ion.

20a. Qualitative data relates to observations while quantitative data relates to numbers.

20b. The independent variable is deliberately manipulated by the experimenter to observe a change in the dependent variable.

20c. Errors can either affect a value once, or they can occur consistently throughout the experiment.

21a. There are two monatomic ions and one polyatomic ion present.

21b. Ionic compounds contain very strong bonds that allow them to be used as durable substances.

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4C Precipitation and ionic equations

Image: Mia Garrett/Shutterstock.com

Why do some compounds precipitate in water?

Whilst many ionic compounds exist as **dissolved** ions in water, others remain solid. In this lesson, we will learn about how ionic compounds can react in solution to form solid products (**precipitates**), and how we can represent these reactions with equations.

STUDY DESIGN DOT POINT

- the use of solubility tables to predict and identify precipitation reactions between ions in solution, represented by balanced full and ionic equations including the state symbols: (s), (l), (aq) and (g)

4A 4B 4C

1.1.16.1 Use of solubility tables

1.1.16.2 Writing full and ionic equations for precipitation reactions

ESSENTIAL PRIOR KNOWLEDGE

4B	Ionic formulas
4B	Balancing equations

See questions 25–26.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

KEY TERMS AND DEFINITIONS

Aqueous	dissolved in water
Dissolved	incorporated into a liquid to form a solution
Full equation	chemical equation with all chemical species present
Insoluble	unable to be dissolved in a given solvent
Ionic equation	equation showing the net state changes of ions
Phase	distinct form in which matter can exist
Precipitate	solid compound formed in a precipitation reaction
Precipitation reaction	chemical reaction resulting in the formation of a solid from an aqueous solution
Solubility	ability to be dissolved in a given solvent
Spectator ions	aqueous ions that do not change state throughout a reaction

Use of solubility tables 1.1.16.1

A substance's **solubility** can be used to predict the outcome of a chemical reaction.

How can we predict and identify precipitates?

In a chemical reaction, the products formed may have different properties to the reactants they are composed of. An example of this is a **precipitation reaction**, in which two ionic compounds react in solution to form a solid product, or precipitate, which is **insoluble**: it cannot dissolve in the solution. Figure 1 shows the bright yellow lead iodide precipitate formed when lead nitrate is pipetted into sodium iodide solution.

Image: Lindsey Moore/shutterstock.com

Figure 1 Formation of a lead iodide precipitate

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KEEN TO INVESTIGATE?

[1] **How are precipitation reactions used in industrial chemistry?**
Search: What Type of Reaction Produces a Precipitate

Given a precipitate will not dissolve in the solution in which it is formed, precipitation reactions are a useful tool in the real world.[1] To use precipitation reactions to our advantage, we must be able to predict the products of a reaction. Solubility tables help us achieve this by giving us information about the solubility of common ions and compounds, such as in table 1.

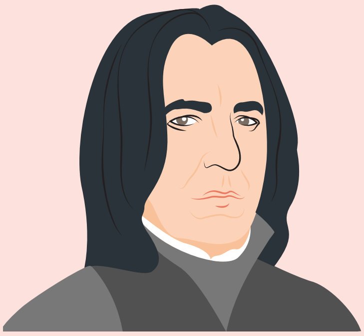
Table 1 The solubility of salts in water formed from common ions

Ion	Solubility	Exceptions
Sodium, Na^+	Soluble	None
Potassium, K^+	Soluble	None
Ammonium, NH_4^+	Soluble	None
Chloride, Cl^-	Soluble	Insoluble in a compound with Ag^+, Pb^{2+}
Bromide, Br^-	Soluble	Insoluble in a compound with Ag^+, Pb^{2+}
Iodide, I^-	Soluble	Insoluble in a compound with Ag^+, Pb^{2+}
Nitrate, NO_3^-	Soluble	None
Perchlorate, ClO_4^-	Soluble	None
Ethanoate, CH_3COO^-	Soluble	None
Sulfate, SO_4^{2-}	Soluble	Insoluble in a compound with Ca^{2+}, Ba^{2+}, Sr^{2+}, Ag^+, Pb^{2+}
Carbonate, CO_3^{2-}	Insoluble	Soluble in a compound with group 1 metal ions and NH_4^+
Phosphate, PO_4^{3-}	Insoluble	Soluble in a compound with group 1 metal ions and NH_4^+
Hydroxide, OH^-	Insoluble	Soluble in a compound with group 1 metal ions, Ca^{2+}, Ba^{2+}, Sr^{2+} and NH_4^+
Sulfide, S^{2-}	Insoluble	Soluble in a compound with group 1 and 2 metal ions, as well as NH_4^+

USEFUL TIP

The **SNAPE** rule can be useful to quickly recall whether a compound is soluble. The rule states that if a compound contains one of the following ions, then it will be soluble in water:

- **S**odium (Na^+)
- **N**itrate (NO_3^-)
- **A**mmonium (NH_4^+)
- **P**otassium (K^+)
- **E**thanoate (CH_3COO^-)



MISCONCEPTION

'Precipitation reactions can occur between any two ionic compounds.'

Precipitation reactions will only occur between two ionic compounds that are soluble in water and form an insoluble product.

To predict whether a precipitation reaction will occur, we need to ask ourselves two questions:

- Are both reactants soluble in water? If yes,
- is one of the products insoluble in water?

If the answer to both of these questions is yes, then a precipitation reaction will occur.

WORKED EXAMPLE 1

Will a precipitation reaction occur between NaCl and $AgNO_3$ in solution?

What information is presented in the question?

The molecular formulas of the reactants.

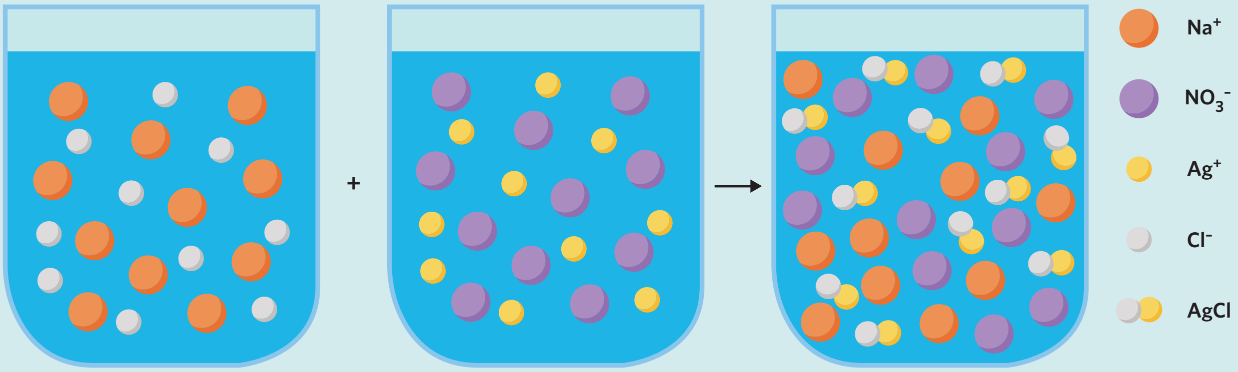
What is the question asking us to do?

Determine whether or not a precipitate will be formed between sodium chloride and silver nitrate.

What strategies do we need in order to answer the question?

1. Determine the solubility of the reactants using the solubility table.
2. Determine the products of the reaction.
3. Determine the solubility of the products using the solubility table.

Answer



According to the solubility table, all Na^+ salts are soluble without exception, and all NO_3^- salts are soluble without exception, so both reactants are soluble. The products of the reaction will be $NaNO_3$ and AgCl. Using the solubility table, we know $NaNO_3$ is soluble, however, AgCl is not. Therefore, a precipitation reaction will occur.

Once we know that a precipitation reaction will occur, we can utilise our solubility table to identify which of the products in a reaction is insoluble – this is the precipitate.

WORKED EXAMPLE 2

A student is investigating precipitation reactions, and adds a solution of calcium chloride to sodium carbonate solution. The student notices that the solution becomes cloudy. Explain the student's observation using solubility rules.

What information is presented in the question?

The reaction between calcium chloride and sodium carbonate solutions creates a cloudy solution.

What is the question asking us to do?

Use solubility rules to explain the cloudy solution that forms.

What strategies do we need in order to answer the question?

1. Identify the products formed in the reaction.
2. Determine the solubility of the products using solubility rules.
3. Determine which product is responsible for the cloudy solution.

Answer

The products formed in the reaction between calcium chloride and sodium carbonate solutions are calcium carbonate and sodium chloride. According to the solubility table, sodium chloride is soluble, however, calcium carbonate is insoluble. Hence, calcium carbonate is a precipitate, and is responsible for the cloudy solution observed by the student.

Progress questions

Question 1

Precipitates cannot dissolve in the solution they are formed in.

A. True

B. False

Question 2

Precipitation reactions occur between

A. compounds.

B. molecules.

Question 3

$MgCl_2(aq)$ and $NaOH(aq)$ are mixed together in a test tube. Which option correctly identifies the precipitate formed?

A. $NaCl$

B. $Mg(OH)_2$

Question 4

An industrial chemist wanted to remove the valuable Ag^+ ions from solution. Which one of the following solutions would produce a precipitate when mixed with a solution containing Ag^+ ions?

A. $NaCl(aq)$

B. $NaNO_3(aq)$

Writing full and ionic equations for precipitation reactions 1.1.16.2

Precipitation reactions can be expressed using equations.

How do we represent precipitation equations?

Like all other chemical reactions, precipitation reactions can be shown using chemical equations. However, all the equations we have come across thus far have occurred without **phase** changes; that is, the states of substances have not changed. Given precipitation reactions are defined by the formation of a solid, it is important to include all the states associated with each chemical species to identify the precipitate.

Consider the reaction between potassium chloride, KCl, and silver nitrate, $AgNO_3$, solutions. This reaction yields solid silver chloride – the precipitate – and **aqueous** potassium nitrate. This chemical reaction would be written as:

$KCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + KNO_3(aq)$

Here, (s) denotes the solid precipitate formed, silver chloride, and (aq) denotes the aqueous, or soluble, species potassium nitrate.

Since this reaction occurs in solution, the reactants are not in their ionic lattice structure, nor are the ions actually bonded together anymore. They are simply a collection of positive and negative ions floating around in an 'aqueous ion soup', written as:

$K^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \rightarrow AgCl(s) + K^+(aq) + NO_3^-(aq)$

In the silver nitrate example, whilst the **full equation** is written above, only certain ions change states during the reaction. The ions whose states remain unchanged are known as **spectator ions**. The spectator ions for this example are shown in figure 2.

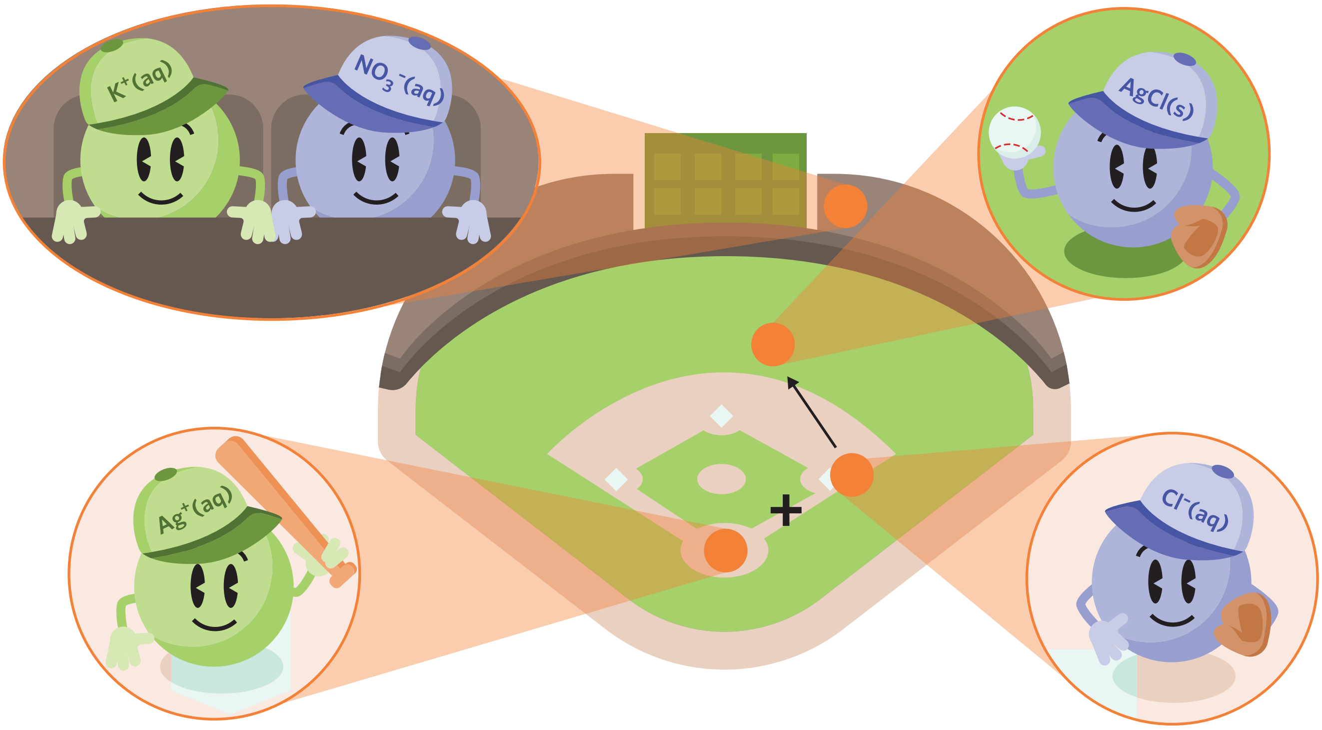


Figure 2 Spectator ions

If we omit K^+ and NO_3^-, the spectator ions, the **ionic equation** is:

$Cl^- (aq) + Ag^+(aq) \rightarrow AgCl(s)$

USEFUL TIP

A shortcut for generating the ionic equation for a precipitation reaction can be taken by identifying the precipitate formed. The reactants are simply the two aqueous ions that make up the precipitate, and the only product is the solid precipitate itself.

WORKED EXAMPLE 3

Sodium hydroxide, $NaOH(aq)$, reacts with copper (II) sulfate, $CuSO_4(aq)$, to produce a pale blue precipitate. Identify this precipitate, and write both the full and ionic equations for this reaction, including states.

What information is presented in the question?

The reactants are $NaOH(aq)$ and $CuSO_4(aq)$. A precipitate is formed.

What is the question asking us to do?

Determine the formula for the precipitate and write the full and ionic equations for the reaction.

What strategies do we need in order to answer the question?

1. Determine the products that can be formed from the ions of the reactants.
2. Use solubility rules to identify the precipitate.
3. Write the full equation, including states.
4. Identify and omit spectator ions to write the ionic equation.

Answer

In solution, sodium hydroxide and copper (II) sulfate will react to produce sodium sulfate and copper (II) hydroxide. According to the solubility table, sodium sulfate is soluble, however, copper (II) hydroxide is not. Therefore copper (II) hydroxide, $Cu(OH)_2(s)$, is the precipitate formed, and the full equation is:

$CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$

Given sodium and sulfate are aqueous on both sides of the equation, they are the spectator ions. Thus, the ionic equation is:

$Cu^{2+}(aq) + 2OH^- (aq) \rightarrow Cu(OH)_2(s)$

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Progress questions

Question 5

Which of the following chemical equations represents a precipitation reaction?

A. $Pb(NO_3)_2(aq) + 2NaI(aq) \rightarrow PbI_2(s) + 2NaNO_3(aq)$

B. $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

Use the following information to answer questions 6–8.

$BaCl_2(aq)$ and $Na_2SO_4(aq)$ are mixed together in a test tube.

Question 6

What is the full equation?

A. $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

B. $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

Question 7

What is the ionic equation?

A. $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

B. $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

Question 8

What are the spectator ions?

A. $Ba^{2+}(aq)$, $SO_4^{2-}(aq)$

B. $Na^+(aq)$, $Cl^-(aq)$

Theory summary

- Precipitation reactions involve the formation of a solid from aqueous solutions.
- The solubility of various ionic compounds depends on the anions and cations present.
- The full equation shows all the ions present, including spectator ions.
- Spectator ions are not involved in precipitation reactions and remain dissolved in solution throughout the reaction.
- The ionic equation has the spectator ions omitted.

4C Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 9–11.

Eloise is a chemistry student investigating the properties of different aqueous solutions and whether or not they form precipitates when mixed. At her disposal are solutions of $Co(NO_3)_2(aq)$, $NaOH(aq)$, $KNO_3(aq)$ and $CaCl_2(aq)$.

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Question 9 (1 MARK)

Which of the following statements is incorrect regarding spectator ions?

A. Spectator ions are in the aqueous state on both sides of the chemical equation.

B. Spectator ions are not involved in precipitation reactions.

C. The chemical species in the precipitation reaction have no effect on spectator ion identification.

D. The chemical composition of spectator ions remains unchanged in a precipitation reaction.

Question 10 (1 MARK)

Which solutions will form a precipitate when mixed together?

A. $Co(NO_3)_2(aq)$ and $NaOH(aq)$

B. $Co(NO_3)_2(aq)$ and $KNO_3(aq)$

C. $Co(NO_3)_2(aq)$ and $CaCl_2(aq)$

D. $KNO_3(aq)$ and $CaCl_2(aq)$

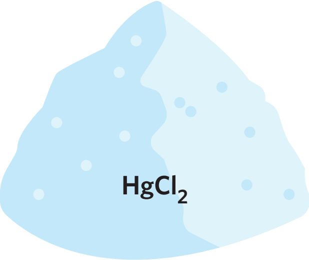
Question 11 (3 MARKS)

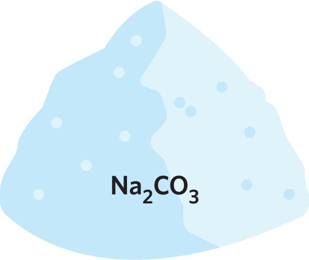
For the combination of Eloise's solutions that results in the formation of a precipitate, write out the full equation (including states) and the ionic equation. Identify the spectator ions.

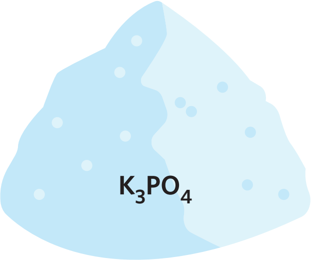
Exam-style

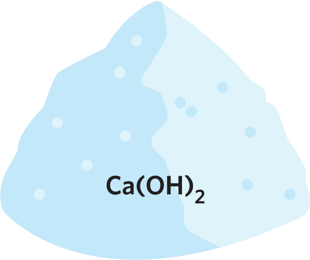
Question 12 (4 MARKS)

Solubility tables are useful when trying to find the solubility of various substances. Using a solubility table, determine the solubility of each of these salts.

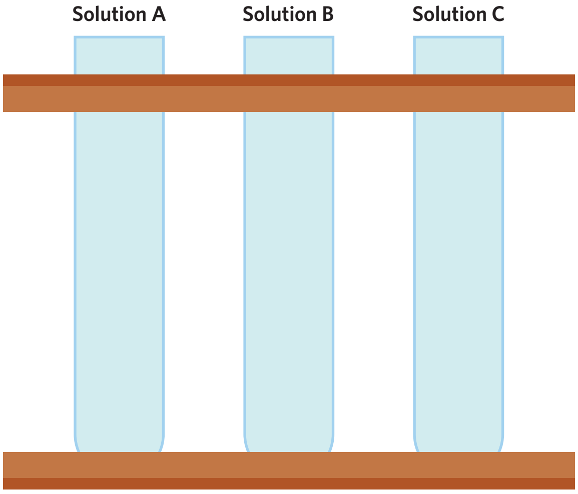








Question 13 (11 MARKS)



Some solutions that contain ionic compounds are colourless. One way of determining the identity of a solution is to use precipitation reactions. An industrial chemist has set up three test tubes, each containing an ionic solution of unknown identity.

The chemist knows that one solution contains $Cl^-(aq)$ ions, one contains $NO_3^-(aq)$ ions and one contains $OH^-(aq)$ ions.

a. What two observations might the industrial chemist make when a precipitation reaction occurs in a test tube? (2 MARK)

b. Using a dropper, the chemist puts a few drops of $Cu^{2+}(aq)$ into each test tube. The test tube containing solution B becomes cloudy. Identify the anion present in solution B. Justify your answer. (3 MARKS)

c. Write the ionic equation for the reaction that occurs in the test tube containing solution B. (1 MARK)

d. Using a dropper, the chemist puts a few drops of $Ag^+(aq)$ into the remaining two test tubes. A sediment begins to form at the bottom of the test tube containing solution A, whilst no change is observed in solution C. Identify the anions present in solutions A and C. Justify your answers. (4 MARKS)

e. Write the ionic equation occurring in the test tube containing solution A. (1 MARK)

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Question 14 (3 MARKS)

Limescale is a chalky and solid substance that often accumulates inside pipes and appliances like kettles. It results from a precipitation reaction that produces calcium carbonate ($CaCO_3(s)$).

a. Write out the ionic equation for the formation of the calcium carbonate precipitate. (1 MARK)

b. Pipes are often more frequently blocked in areas with water supplies that have a high concentration of calcium ions. Suggest a reason for this observation. (2 MARKS)

Key science skills

Question 15 (5 MARKS)

The colours of precipitates formed by different metals ions are given in the table:

Cobalt (II)	Copper (II)	Iron (II)	Iron (III)	Manganese (II)	Nickel (II)	Lead (II)
Pink	Blue	Green	Reddish brown	Pink	Green	Yellow

A mining company wanted to use the different colours of precipitates formed to determine the presence or absence of different metals in various solutions.

a. What type of data is being collected? (1 MARK)

b. Using this experimental technique, why might it be difficult to differentiate between cobalt and manganese? What type of error is this? (2 MARKS)

c. The mining company wanted to accurately determine the concentration of metal cations. To do so, they decided to approximate the concentration by having one scientist observe the cloudiness of the solution. Comment on the accuracy of the results obtained from technique. (2 MARKS)

FROM LESSONS 16B & 16D

Questions from multiple lessons

Question 16 (4 MARKS)

Epsom salt ($MgSO_4$) is believed to reduce muscle cramping and soreness. Therefore, it is often used by athletes after a competition.

a. Write a balanced chemical equation to show the reaction between solutions of epsom salt and barium chloride ($BaCl_2$). (1 MARK)

b. Would epsom salt be able to dissolve in water? Explain your answer. (2 MARKS)

c. In which group on the periodic table would magnesium be found? (1 MARK)

FROM LESSONS 1B & 4B

Hints

12. A solubility table gives information about the solubility of compounds formed from particular ions.
13a. Observations need to include colour or state changes.
13b. The cloudy solution indicates that a reaction has occurred.
13c. Ionic equations do not contain spectator ions.
13d. The cloudy solution indicates that a reaction has occurred.
13e. Ionic equations do not contain spectator ions.
14a. Ionic equations do not contain spectator ions.
14b. More precipitate can form when more reactant is present.
15a. Data is categorised based on whether it is numeric or descriptive.
15b. Consider the colours of the precipitates formed by each metal.
15c. Human error typically reduces accuracy.
16a. Consider the solubility of the products that would form.
16b. A solubility table gives information about the solubility of compounds formed from particular ions.
16c. Each column of the periodic table is assigned a different group number.

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Chapter 4 review

Multiple choice (10 MARKS)

Question 1 (1 MARK) 🌶

An ionic compound

A. is a three-dimensional structure formed from the interaction between one cation and one anion.

B. can involve single ions forming multiple ionic bonds with multiple other ions.

C. contains strong covalent bonds between ions.

D. is a two-dimensional structure formed from a single cation and anion.

Question 2 (1 MARK) 🌶

Most ionic compounds are formed from

A. metal and non-metal cations.

B. non-metal cations and metal anions.

C. metal anions.

D. metal cations and non-metal anions.

Question 3 (1 MARK) 🌶

Ionic compounds are

A. always effective conductors of electricity.

B. insulators of electricity in the dissolved state.

C. effective conductors of electricity in the molten or dissolved state.

D. effective conductors of electricity in the solid state.

Question 4 (1 MARK) 🌶

Which of the following best describes the physical properties of solid ionic compounds?

A. Ionic compounds are soft and malleable.

B. Ionic compounds are hard and brittle.

C. Ionic compounds are soft and flexible.

D. Ionic compounds are hard and malleable.

Question 5 (1 MARK) 🌶🌶

Lithium fluoride (LiF) is used to produce components of engines due to its

A. high melting point.

B. hardness.

C. low melting point.

D. tensile strength.

Question 6 (1 MARK) 🌶🌶

Which of the following is the correct formula for an ionic compound made from magnesium and fluoride ions?

A. Mg_2F

B. FMg_2

C. MgF_2

D. F_2Mg

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Question 7 (1 MARK)

Which of the following is the correct ionic formula for iron (III) hydroxide?

A. $FeOH$

B. Fe_3OH

C. $FeOH_3$

D. $Fe(OH)_3$

Question 8 (1 MARK)

The force between chloride ions is ___________, but the force between sodium and chloride ions is ___________.

A. repulsive, attractive

B. attractive, repulsive

C. electrostatic, repulsive

D. attractive, hydrostatic

Question 9 (1 MARK)

Which of the following represents the balanced equation of the reaction between calcium carbonate, $CaCO_3$, and aluminium chloride, $AlCl_3$?

A. $3CaCO_3 + 2AlCl_3 \rightarrow Ca_3Al_2 + 3Cl_2CO_3$

B. $CaCO_3 + AlCl_3 \rightarrow CaCl_3 + AlCO_3$

C. $3CaCO_3 + 2AlCl_3 \rightarrow 3CaCl_2 + Al_2(CO_3)_3$

D. $2CaCO_3 + 3AlCl_3 \rightarrow CaCl_2 + Al_2(CO_3)_3$

Question 10 (1 MARK)

Which of the following reactions is not a precipitation reaction?

A. $Pb(NO_3)_2(aq) + 2NaI(aq) \rightarrow PbI_2(s) + 2NaNO_3(aq)$

B. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

C. $2Na^+(aq) + SO_4^{2-}(aq) + Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s) + Na_2SO_4(aq)$

D. $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

Short answer (30 MARKS)

Question 11 (8 MARKS)

When determining the equation for reactions between ionic compounds, it is important to take the charges of compounds into account.

a. Provide the ionic formula and name of the ionic compound consisting of Ba^{2+} and PO_4^{3-} ions. (2 MARKS)

b. Write the balanced full equation for the reactions between the following ionic compounds. States are not required.

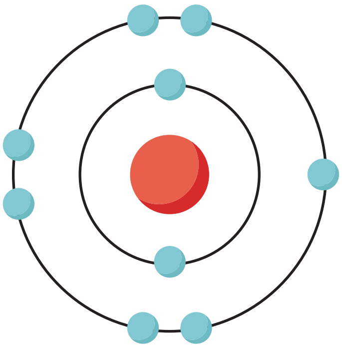
  i. Calcium bromide ($CaBr_2$) and copper (I) sulfate (Cu_2SO_4). (1 MARK)

  ii. Silver carbonate and potassium fluoride. (1 MARK)

  iii. Copper (II) nitride and sodium sulfide. (1 MARK)

c. Determine the names of the products of the reactions in part b. (3 MARKS)

Question 12 (4 MARKS)



A new element, J, was discovered, however scientists were unsure about J's properties.

a. In a reaction of J with oxygen, the product formed had a formula of J_2O.

- **i.** Determine the charge on the J ion present in J_2O. (1 MARK)
- **ii.** Is J more likely to be a metal or nonmetal element? (1 MARK)

b. J was then reacted with the element 'X' shown. Identify the formula of the compound formed. (1 MARK)

c. If a J ion formed an ionic bond with a nitrate ion, determine whether the compound formed would be soluble in water (assume the J ion is not an exception to the existing solubility rules). (1 MARK)

Question 13 (8 MARKS)

Magnesium chloride is a compound used to minimise the production of ice on roads in alpine regions. Silver nitrate is a compound with several medical applications, such as wound treatment or cauterisation.

a. Describe the conditions required for electrical conductivity, and hence comment on the electrical conductivity of magnesium chloride in solid, molten, and dissolved states. (3 MARKS)

b. Solutions of magnesium chloride and silver nitrate are mixed together. Identify the ionic formula of the precipitate formed. (1 MARK)

c. Identify the spectator ions, and subsequently give the ionic equation for this reaction, including states. (2 MARKS)

d. Identify two experimental observations that could be made during the formation of a precipitate. (2 MARKS)

Question 14 (10 MARKS)

Ionic compounds are found abundantly in the natural world, and serve a number of purposes in our daily lives. An example of a common ionic compound is sodium chloride, which is found in high concentrations in seawater and is also used as table salt.

a. Write the ionic formula of sodium chloride. (1 MARK)

b. The current Australian dietary guidelines state that no more than six grams of salt should be consumed daily due to the negative effects of excessive sodium on blood pressure. A scientist hypothesises that by heating food containing salt to a moderate temperature, the salt will break down and, thus the health consequences will not be as severe. Comment on the scientist's hypothesis, referring to the properties of ionic compounds. (3 MARKS)

c. The scientist obtains a block of pure sodium chloride and uses a hammer to hit the salt block as hard as they can. Explain, using the structure of ionic compounds, what the scientist is likely to observe. (3 MARKS)

d. Complete the table by identifying the properties of the given ionic compounds that enable their everyday uses. (3 MARKS)

Ionic compound	Use	Properties enabling use
Magnesium oxide (MgO)	Lining surface of engines	
Calcium phosphate ($Ca_3(PO_4)_2$)	Human bones and teeth	
Ammonium chloride (NH_4Cl)	Electrolyte enabling current flow in cell batteries	

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Key science skills (10 MARKS)

Question 15 (10 MARKS)

A student who has recently learned about ionic compounds has prepared an experiment to compare the properties of different ionic compounds. The experiment involves exposing a series of different samples of compounds to a lit match for a set amount of time and observing the interactions of each with the flame.

a. Draw a labelled diagram of the experimental setup for observing solid magnesium hydroxide's interaction with the flame. (2 MARKS)

b. Identify the independent and dependent variables in this experiment. (2 MARKS)

c. Identify at least two variables that should be controlled in this experiment. (2 MARKS)

d. The student gave two other students a copy of the method and each of them carried out the experiment using the same samples, all accounting for controlled variables. The results are shown in the table:

	Student 1	**Student 2**	**Student 3**
Magnesium hydroxide ($Mg(OH)_2$)	Melts	No change	Melts
Sodium chloride ($NaCl$)	No change	Hardens	Melts
Lead (II) sulfide (PbS)	Hardens	Hardens	Melts

i. Is this experimental method repeatable? Explain your answer. (3 MARKS)

ii. It was found that each of the students had a different standard metric that they were using to determine the change the sample had undergone. What type of error is this? (1 MARK)

FROM LESSONS 16B, 16D & 16E

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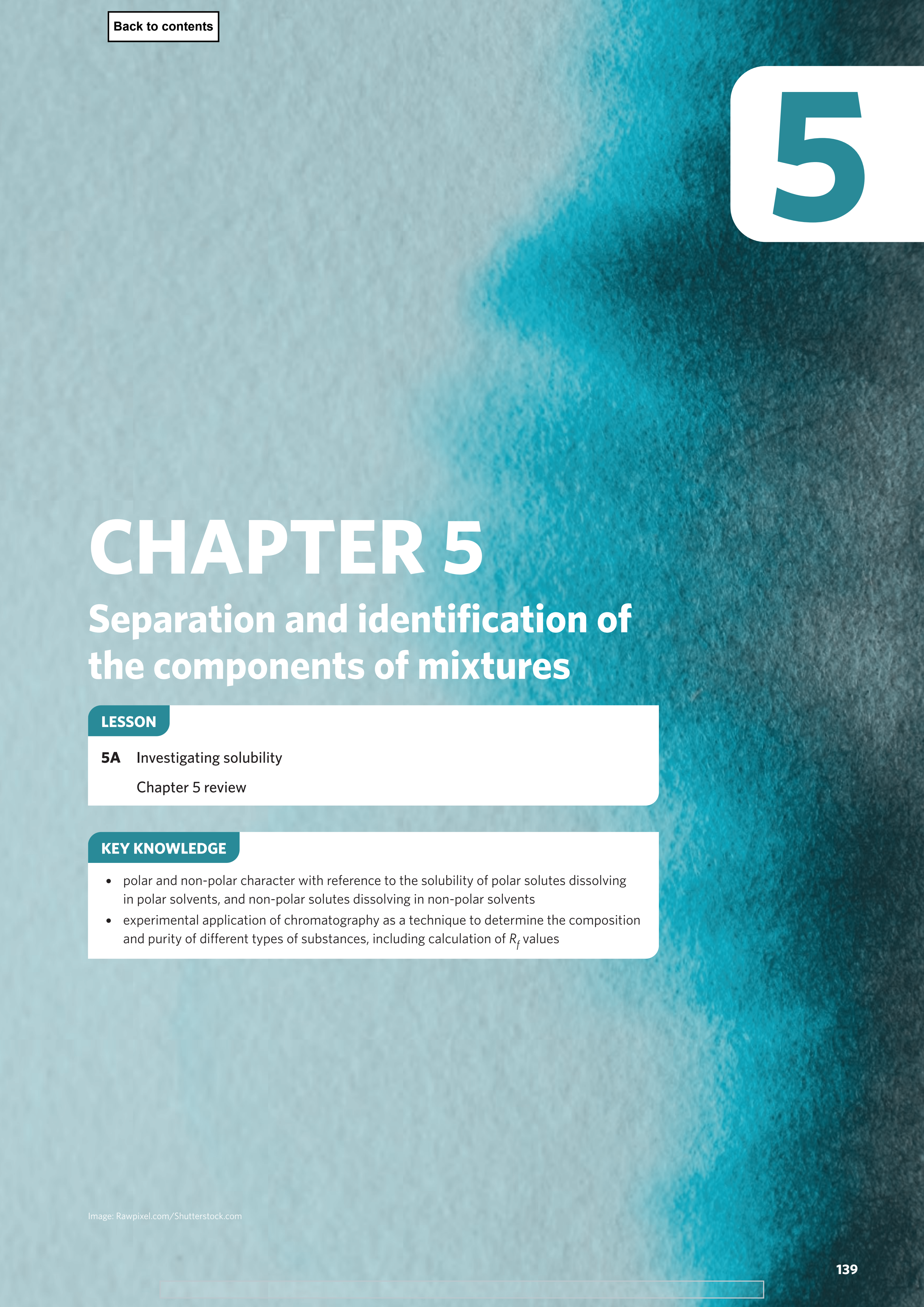

CHAPTER 5

Separation and identification of the components of mixtures

LESSON

5A Investigating solubility

Chapter 5 review

KEY KNOWLEDGE

- polar and non-polar character with reference to the solubility of polar solutes dissolving in polar solvents, and non-polar solutes dissolving in non-polar solvents
- experimental application of chromatography as a technique to determine the composition and purity of different types of substances, including calculation of R_f values

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5A Investigating solubility

STUDY DESIGN DOT POINTS

- polar and non-polar character with reference to the solubility of polar solutes dissolving in polar solvents, and non-polar solutes dissolving in non-polar solvents
- experimental application of chromatography as a technique to determine the composition and purity of different types of substances, including calculation of R_f values

5A

1.1.17.1 Solutes, solvents and solubility
1.1.18.1 Chromatography
1.1.18.2 Experimental applications

ESSENTIAL PRIOR KNOWLEDGE

2A Polar and non-polar molecules
2B Hydrogen bonding
2B Dipole-dipole forces
See questions 27–29.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: nevodka/Shutterstock.com

How much sugar can actually dissolve in a fizzy drink?

Substances like sugar are highly soluble in water. In this lesson, we will learn why compounds like glucose – the compound that sugar is formed from – are soluble in water, and why other compounds are not. We will also learn how to apply our understanding of solubility to scientific investigations.

KEY TERMS AND DEFINITIONS

Adsorption adhering of a component onto the stationary phase

Aqueous solution solution in which the solvent is water

Chromatogram visual output of chromatography

Desorption release of a component from the stationary phase by dissolving

Dissolved incorporated into a liquid to form a solution

Eluent solvent fluid that moves through a chromatography system

Homogeneous a mixture with uniform composition and properties

HPLC chromatography technique in which the mobile phase is pumped through a tightly packed stationary phase under pressure

Hydrated surrounded by water molecules

Ion-dipole attraction electrostatic attraction between a permanent dipole and an ion

Ionisation process by which a chemical species gains or loses an electron to gain an electric charge

Mobile phase fluid in a chromatography system that carries the sample over the stationary phase

Origin line at which samples are placed in paper and thin-layer chromatography

Paper chromatography chromatography technique using absorbent paper

Peak area the area enclosed between the peak and the baseline on a chromatogram

Retention factor the ratio of the distance moved by a component to the distance moved by the solvent front

Retention time time taken for a component to pass through a HPLC column

Solute substance that dissolves in a solvent

Solution homogeneous mixture in which a substance is dissolved in another

Solvent substance that dissolves another substance

Solvent front the wet moving edge of the solvent in chromatography

Stationary phase solid onto which the components of a sample adsorb

Thin-layer chromatography chromatography technique using a thin stationary phase supported by an inert backing

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Solutes, solvents and solubility 1.1.17.1

Substances can be described as **solutes** or **solvents** based on their solubility and state of matter.

How can polarity be used to determine solubility?

A **solution** is a **homogeneous** mixture of two or more different substances, where one is **dissolved** in another. When water is used to dissolve a solid, liquid, or gaseous substance, an **aqueous solution** is produced. In an aqueous solution, particles dissociate and can move freely. The chemical substance that is dissolved in the solution is known as the solute, and the substance in which the solute dissolves is the solvent. In aqueous solutions, water is always the solvent. However, there are other types of solutions formed from different solvents.

One of the defining characteristics of a solution is that it is homogeneous. This means that the solute and solvent are mixed together and cannot be differentiated from each other. For example, sea water is a solution of salt (solute) and water (solvent) because they are evenly mixed. A mixture of sand and water, however, is not a solution because sand and water can be easily separated from one another.

When a substance dissolves, there are three different kinds of intermolecular forces that need to be considered:

1. Intermolecular forces between solvent particles.
2. Intermolecular forces between solute particles.
3. Intermolecular forces between solute and solvent particles.

This is because in order for a substance to dissolve, the intermolecular forces between solute particles and between solvent particles need to be overcome by the intermolecular forces between the solute and solvent. If the solute–solvent forces are very weak, or the solute–solute and/or solvent–solvent forces are much stronger, a substance will not dissolve. Figure 1 shows the solute-solute and solvent-solvent forces being overcome, allowing a substance to dissolve.

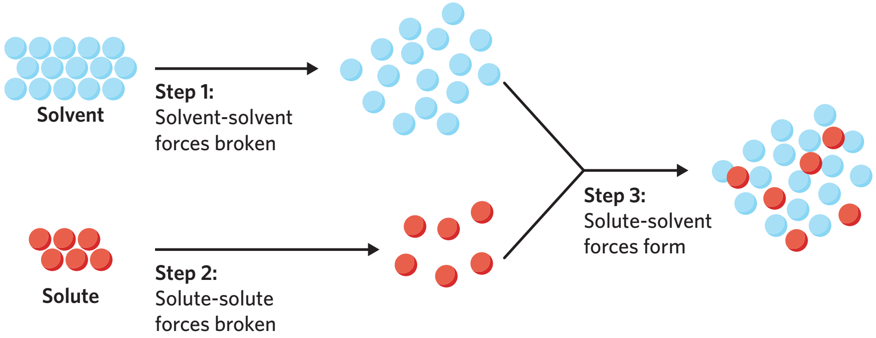

Figure 1 Solute–solute and solvent–solvent forces need to be overcome for a molecule to dissolve.

As a general principle, 'like dissolves like'. This means that polar compounds like water will dissolve compounds that are polar, but will not dissolve non-polar compounds. Similarly, non-polar compounds will dissolve compounds that are non-polar, but will not dissolve polar compounds.

Most molecular substances are insoluble in water. This is because the solute-solvent forces between water and these molecules cannot overcome the stronger hydrogen bonding present between the water molecules and dipole-dipole forces between solute molecules. Despite this, there are several molecules that are soluble in water. There are two key reasons why these select few molecules will dissolve in water:

- The molecules form hydrogen bonds with water molecules, or
- the molecules ionise in water.

Some molecules have a functional group that is capable of hydrogen bonding with water molecules (for example, the hydroxyl, carboxyl and ester functional groups). However, the influence of a functional group on a molecule decreases as the chain length increases (we will cover this in more detail in lesson 7D). Therefore, only small molecules with these functional groups, or large molecules with a large number of these functional groups, tend to dissolve in water.

STRATEGY

When determining whether one substance will dissolve in another, consider the polarity of each substance. If both substances are polar, then one will dissolve in the other and if both substances are non–polar then one will dissolve in the other.

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For example, one glucose ($C_6H_{12}O_6$) molecule contains 5 polar hydroxyl (–OH) groups and one polar ether (C–O–C) group, making it a highly polar molecule that is able to dissolve in water. The intermolecular forces that exist between glucose molecules are dispersion forces, permanent dipole–dipole forces and hydrogen bonds, as shown in figure 2.

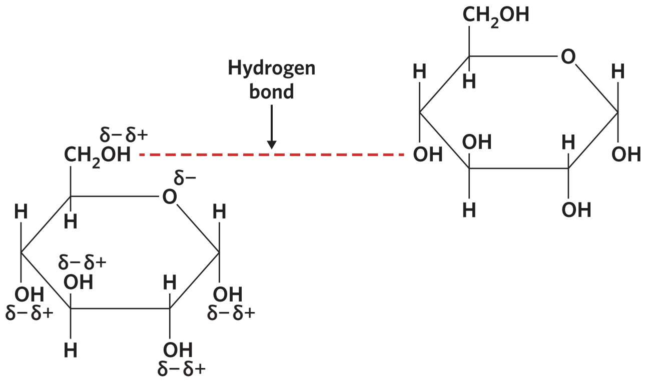


Figure 2 Hydrogen bonding between two glucose molecules, and other possible hydrogen bonding sites

In order for glucose to dissolve, the intermolecular hydrogen bonds formed between glucose and water molecules need to be stronger than those between glucose molecules and those between water molecules.

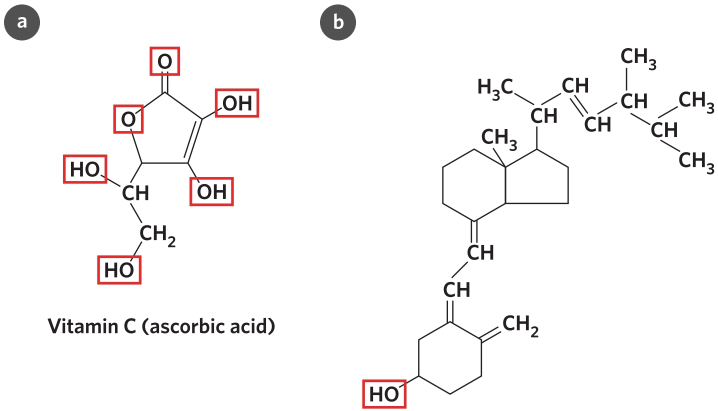
Therefore, in order for glucose to be dissolved,[1] the hydrogen bonds between water molecules have been broken, the hydrogen bonds between glucose molecules have been broken, and new hydrogen bonds between water and glucose molecules have formed. This is represented by the following chemical equation:

$$C_6H_{12}O_6(s) \xrightarrow{H_2O(l)} C_6H_{12}O_6(aq)$$

KEEN TO INVESTIGATE?

[1] **How much sugar is dissolved in different drinks?**
Search YouTube: How much sugar is in your drink?

Consider the structures of Vitamin C and Vitamin D_2 shown in figure 3. Whilst both molecules have at least one polar functional group (circled in red), Vitamin C has 5 more polar groups than Vitamin D_2, and is therefore significantly more polar than Vitamin D_2. In turn, Vitamin C will dissolve in polar solvents like water, whilst Vitamin D_2 will dissolve in non–polar solvents such as hexane.



Vitamin C (ascorbic acid)

Vitamin D_2 (ergocalciferol)

Figure 3 Vitamin C **(a)** is more polar overall whilst Vitamin D_2 **(b)** is more non–polar.

USEFUL TIP

In an asymmetrical molecule the more polar functional groups present on a molecule, the more polar the overall molecule is, and therefore, the more soluble the molecule is in polar solvents.

How can molecules dissolve through ionisation?

Whilst the majority of polar molecules dissolve through the formation of new hydrogen bonds, or dipole-dipole interactions, there are a small number of molecular compounds that can dissolve in water through **ionisation**, the process by which ions are formed from a chemical reaction. Some covalent bonds are sufficiently polar that when placed in a polar compound like water, they ionise; that is, they break into a positively and negatively charged ion. For example, hydrochloric acid, HCl, has a very electronegative chlorine atom bonded to a weakly electronegative hydrogen atom. When placed in water, the electronegative oxygen atom in H_2O is attracted to the partially positive hydrogen atom in HCl. Similarly, the two partially positive hydrogen atoms in H_2O are attracted to the electronegative chlorine atom in HCl. These attractions cause the hydrogen and chlorine atoms from HCl to ionise, as shown in figure 4.

MISCONCEPTION

'Any substance with at least one polar functional group will dissolve in a polar solvent.'

A substance is not absolutely polar or absolutely non–polar. Every substance has a degree of polarity, and therefore, a substance with one polar functional group and a large non–polar region would be relatively non–polar, and therefore insoluble in a polar solvent.

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However, unlike in the dissolving of ionic compounds, H^+ and Cl^- do not exist isolated in solution – they become **hydrated**. This means the polar ions are dissolved in water with water molecules surrounding them, as shown in figure 5. The force of attraction that keeps the ions in solution is called an **ion–dipole attraction**, as the permanent dipoles of water – due to either the partially positively charged hydrogen atoms or the partially negatively charged oxygen atom – are attracted to the positive H^+ and negative Cl^- ions.

Other molecular compounds that can dissolve in water through ionisation like hydrochloric acid include nitric acid, HNO_3, sulfuric acid, H_2SO_4 and ethanoic acid, CH_3COOH.

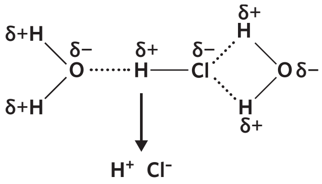


Figure 4 Ionisation of HCl in water

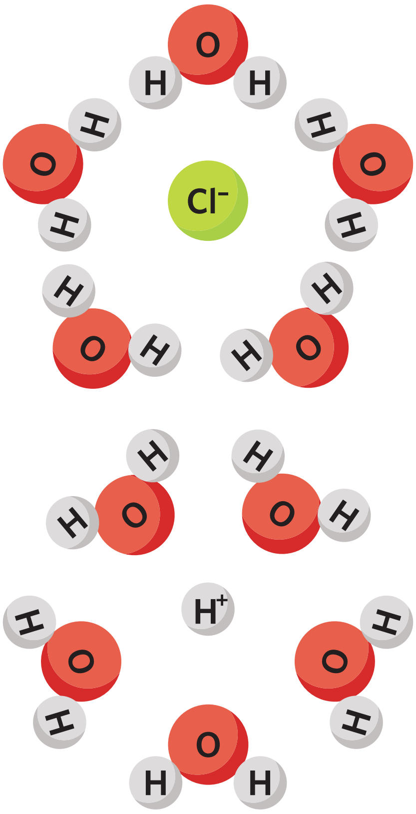


Figure 5 Hydrated H^+ and Cl^- ions

Progress questions

Question 1

Solutions consist of ___________ particles which are evenly dispersed in a(n) ___________ mixture.

- **A.** solute, homogeneous
- **B.** solvent, inhomogeneous

Question 2

What needs to be overcome before a substance is able to dissolve?

- **A.** Solute–solvent attractive forces
- **B.** Solute–solute attractive forces

Question 3

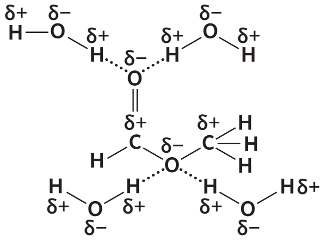
Which of the following molecules will dissolve in water by forming hydrogen bonds with water molecules?

- **A.** HCl
- **B.** CH_3OH

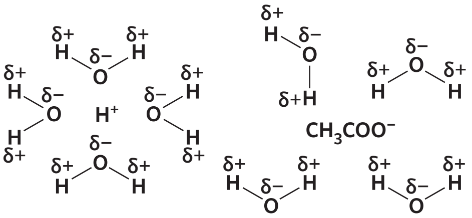
Question 4

Which of the following represents a compound that has dissolved in water through ionisation?

A.



B.



Question 5

If water is used as a solvent, which of the following compounds would not be able to dissolve to any significant extent?

- **A.** $C_{12}H_{25}OH$
- **B.** HNO_3

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Chromatography 1.1.18.1

The components of a mixture can be separated by chromatography.

How can substances be separated based on solubility?

Chromatography is an analysis technique used to separate the components of a mixture. A solvent, known as the **mobile phase**, is passed through a mixture over a medium known as the **stationary phase**. As the components in the mobile phase pass over the stationary phase, they undergo a continual process of **adsorption** to and **desorption** from the stationary phase. The polarity of the components of the mobile phase and of the stationary phase will dictate how strongly each component adheres to the stationary phase, and how readily each component desorbs by redissolving in the mobile phase.

Paper chromatography is one of the cheapest and most common chromatography techniques. It involves the placing of highly absorbent paper (similar to filter paper) in a solvent. In paper chromatography, the paper is the stationary phase over which the components of the solvent, the mobile phase, can migrate. **Thin–layer chromatography** is a very similar technique, however, the stationary phase is instead a thin layer of adsorbent material - such as aluminium oxide, cellulose, or silica gel-covering a glass or plastic plate. Figure 6 shows a typical paper chromatography setup.

MISCONCEPTION

'The position of the origin doesn't matter.'

In paper and thin–layer chromatography, it is very important that the origin is not initially immersed in the solvent. If the origin is immersed in the solvent to begin with, then the components of the sample mixture may dissolve in the solvent at the bottom of the vessel, hindering the reliability of the results.

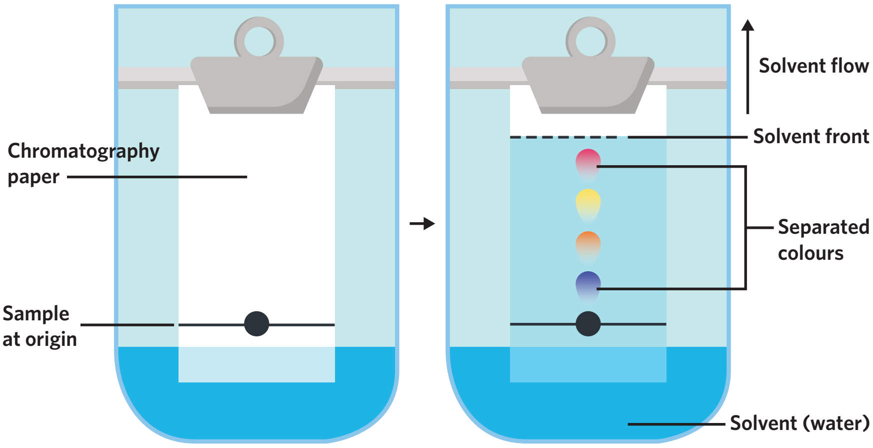

Figure 6 Paper chromatography analysis

As shown in figure 6, the sample being analysed is placed at the **origin**. As the solvent – in this case water – passes through the filter paper, the components of the mixture migrate through the stationary phase to different extents. Given water is polar, we can say that the component represented by the pink colour is the most polar of the 4 components as it migrated furthest up the paper, and therefore must dissolve most readily in water. On the other hand, since the blue colour did not migrate very far up the paper, it is reasonable to conclude that the blue component is the least polar of the 4 components, as it has desorbed from the stationary phase to the smallest extent.

The spread of colours on the stationary phase is known as a **chromatogram**. We can use the chromatogram to identify the components of our mixture through two different methods:

- By including standards of known chemicals on the same chromatogram as the unknown sample
- By calculating the **retention factor**, R_f, of each component in the unknown sample

USEFUL TIP

The term retention factor comes from the fact that an R_f value is an indication of a stationary phases ability to 'retain' or adsorb a given compound.

To determine the composition of the mixture by using known chemicals, it's important to first have an idea of the possible chemical composition of the sample. If a spot from a known chemical travels the same distance from the originas a spot from the unknown sample, then the sample is likely to contain that unknown chemical. Consider, for example, a certain medication that is believed to contain glucose. To test for this, we can run a thin–layer chromatography analysis with a spot of the unknown sample and a spot of glucose at the origin.

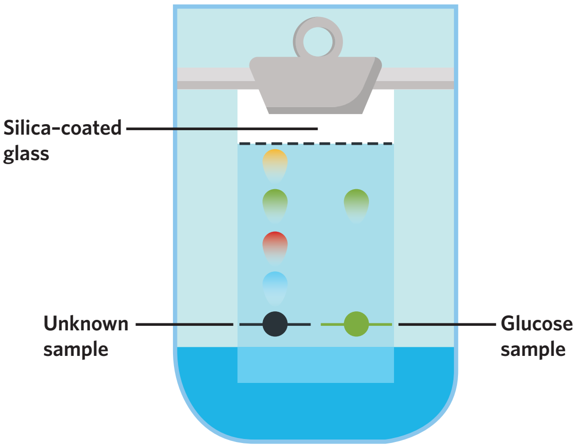

Figure 7 Thin-layer chromatography with glucose and unknown sample

As shown in figure 7, once the **solvent front** is allowed to travel a significant way across the stationary phase, the known glucose sample has migrated the same distance as the green component of the unknown sample. Therefore, we can conclude that our unknown medication contains glucose.

The second method for identifying the components of a mixture from a chromatogram is to compare the distances travelled by each component to the distance travelled by the solvent front. This is known as the retention factor, R_f, which is calculated as:

$$R_f = \frac{\text{distance travelled by the compound}}{\text{distance travelled by the solvent front}}$$

In figure 8, we see a chromatogram of a sample composed of two compounds. The red compound travelled 8 cm, and the solvent front travelled 13 cm. Therefore, the R_f value for the red compound can be calculated as follows:

$$R_f = \frac{8}{13}$$

$$R_f = 0.6$$

USEFUL TIP

Since the retention factor for a particular compound is simply a ratio of distances, R_f does not have units.

To complete the analysis of a sample, once the R_f values of the components have been calculated, they can be compared against the R_f values of known chemicals under the same conditions. However, variables such as temperature, type of stationary phase, and type of solvent all need to be considered when comparing R_f values of unknown and known compounds. If any of these are different, the R_f value for the given chemical will change.

USEFUL TIP

Since the distance travelled by the solvent front will always be greater than the distance travelled by a particular compound, R_f values will always be less than one.

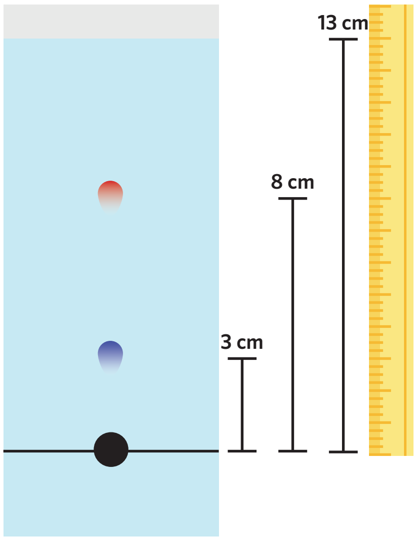

Figure 8 Chromatogram of a sample with two components

MISCONCEPTION

'Calculating the R_f value provides quantitative data.'

Whilst it is a numeric value, the result that can be deduced from the R_f value is the identity of a given component, which is in fact qualitative data.

WORKED EXAMPLE 1

Calculate the R_f value of the blue compound in figure 8.

What information is presented in the question?

The blue compound in figure 8.

What is the question asking us to do?

Calculate the R_f value.

What strategies do we need in order to answer the question?

1. Determine the distance travelled by the blue compound.
2. Determine the distance travelled by the solvent front.
3. Substitute these values into the formula for R_f.

Answer

Distance travelled by blue compound = 3 cm

Distance travelled by solvent front = 13 cm

$$R_f = \frac{3}{13}$$

$$R_f = 0.2$$

Progress questions

Question 6

In the process of chromatography, components

A. desorb into the mobile phase once and adsorb to the stationary phase once.

B. are continuously desorbing into the mobile phase and adsorbing to the stationary phase.

Question 7

If a polar stationary phase is being used, components which are more polar will experience ___________ attraction to the stationary phase, leading to a ___________ R_f value.

A. less, larger

B. greater, smaller

Question 8

A chromatogram obtained from paper chromatography at 25 °C can be compared to a chromatogram of known chemical substances obtained at 41 °C to determine the identities of unknown substances.

A. True

B. False

Experimental applications 1.1.18.2

The principles of chromatography can be applied to laboratory investigations.

How can chromatography be used to determine the composition and purity of different substances?

As we have learned so far, chromatography can be used to determine the composition of different types of substances. In industry, the techniques of thin-layer and paper chromatography have been extrapolated to develop several more refined techniques. Column chromatography, for example, involves a glass column containing a tightly packed, solid stationary phase. The sample is placed at the top of the column, followed by the **eluent**, or solvent.

The first component of the sample to pass through the column is the one that has adhered least strongly to the stationary phase, and is hence the component with a polarity closest to the polarity of the solvent. Since the components will migrate down the column at varying rates, the components of a mixture can be separated by the control of a tap at the bottom of the column. A column chromatography setup is shown in figure 9.

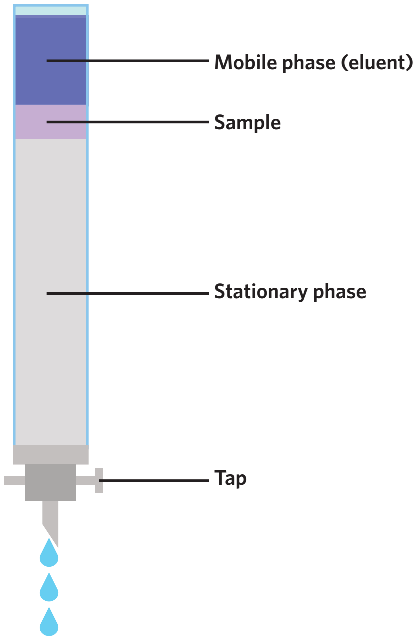

Figure 9 Column chromatography setup

High–performance (formerly high-pressure) liquid chromatography, **HPLC**, works in a similar manner to column chromatography. Again, a column is packed tightly with a stationary phase and a mobile solvent is allowed to pass through. However, HPLC is a much more sensitive technique than column chromatography, as the particles used are approximately 20 times smaller, resulting in more frequent adsorption and desorption of the components of a given sample. Therefore, mixtures composed of compounds of similar polarity can be better separated.

Figure 10 shows a HPLC setup. Given the significantly smaller particle size, the mobile phase experiences a much greater resistance as it passes through. To counteract this resistance, a pump is used to create a pressurised environment inside the column, thereby forcing the mobile phase and sample through.

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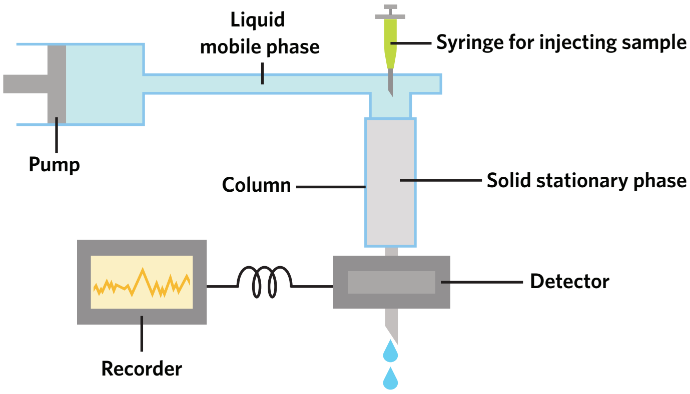


Figure 10 High-performance liquid chromatography setup

As they pass through the stationary phase, the components of a sample are typically detected based on UV light absorption. The amount of light detected corresponds to a peak that forms on a digitally recorded chromatogram, shown in figure 11.

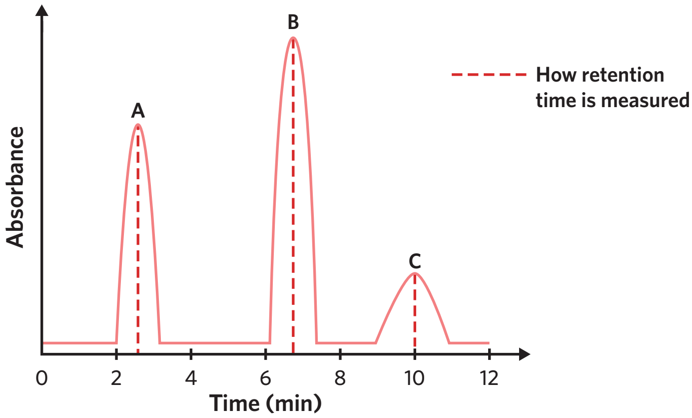


Figure 11 Chromatogram of a sample containing 3 components

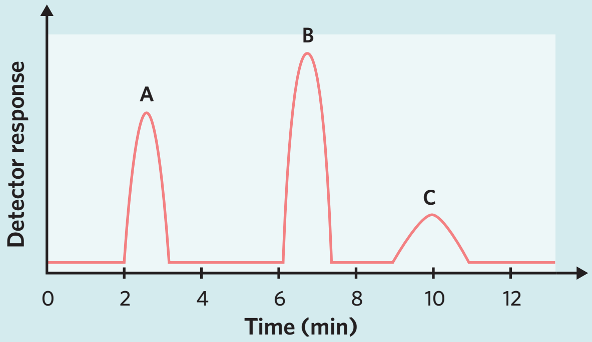
In HPLC, the time taken for a component to pass through the column and reach the detector is known as the **retention time**, R_t. The R_t value is similar to the R_f value for paper and thin-layer chromatography; it is a standardised value under particular conditions that can be used to determine the identity of a component in a sample. However, unlike in paper and thin–layer chromatography, a HPLC chromatogram can also provide quantitative information about the concentration of components in a sample. The **peak area** – the area under a peak in a HPLC chromatogram – can be cross–referenced against the peak areas produced from a solution at several known concentrations. We can plot the peak area produced from a component of a sample on the calibration curve of the given component, and interpolate its concentration.

USEFUL TIP

The retention time, R_t, can be used to obtain qualitative information about a sample, whilst the peak area can be used to obtain quantitative information about a sample.

WORKED EXAMPLE 2

A sample known to contain dopamine, norepinephrine, and epinephrine was analysed using HPLC under standard laboratory conditions. The retention time, R_t, for each compound under standard laboratory conditions is given in the table. In the chromatogram shown, which of the peaks is likely to have been caused by epinephrine?



Identity of component	Retention time (minutes)
Dopamine	10
Norepinephrine	2.5
Epinephrine	6.6

Continues →

What information is presented in the question?

The retention time for each component of the sample. The chromatogram produced when the sample was run through HPLC under standard laboratory conditions.

What is the question asking us to do?

Identify the peak on the chromatogram caused by epinephrine.

What strategies do we need in order to answer the question?

1. Identify the retention time of epinephrine.
2. Locate the peak that occurs closest to the retention time of epinephrine.

Answer

The unique retention time for epinephrine under the given conditions is 6.6 minutes. Therefore, peak B is likely to have been caused by epinephrine, as it occurs at approximately 6.6 minutes.

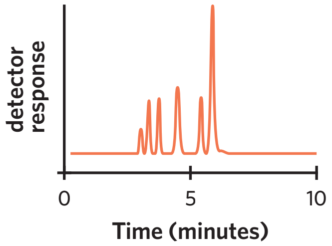
Because HPLC provides such sensitive results, it is used in situations requiring the analysis of chemical substances, such as food and drug-testing industries. For example, HPLC can be used to detect the cancer-causing chemical aflatoxin which is formed by a mould present on peanuts.

Progress questions

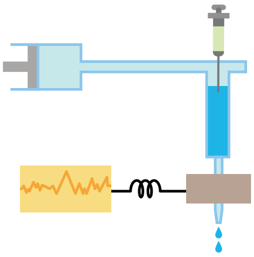
Question 9

Which of the following represents a chromatogram?

A.



B.



Question 10

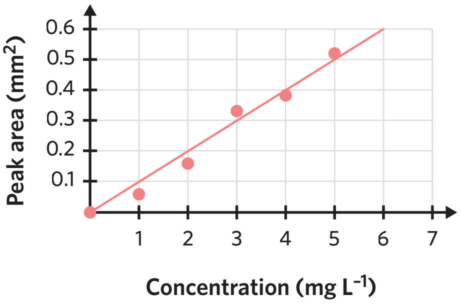
In HPLC, the solvent (liquid) is the __________ phase and the solid in the glass column is the __________ phase.

A. mobile, stationary

B. stationary, mobile

Question 11

The calibration curve for a component in a chromatography column is shown in the following graph.



If a sample was run and a peak area of 0.30 mm^2 was produced, which of the following concentrations would be closest to the true value?

A. 2.5 mg L^{-1}

B. 3.0 mg L^{-1}

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Theory Summary

- A solution is a homogeneous mixture formed when a solute is dissolved in a solvent.
- For a substance to dissolve, the intermolecular forces between solvent and solute particles must be stronger than the intermolecular forces between solvent particles and between solute particles.
- When determining the solubility of a substance, remember that 'like dissolves like'.
- Molecules can dissolve through ionisation and become hydrated.
- Chromatography is a technique used to separate substances based on their solubility.
- The retention factor, R_f, is calculated by:

$$R_f = \frac{\text{distance travelled by the compound}}{\text{distance travelled by the solvent front}}$$

- Experimentally obtained R_f values can be compared to R_f values of known chemicals to identify components of a sample.
- Column chromatography can be used to physically separate the components of a sample.
- HPLC can separate the components of a sample, in addition to providing information about the purity of a sample i.e. the concentration of a component in a sample.
- The retention time, R_t, is the time taken for a component of a sample to pass through the HPLC column.

The content in this lesson is considered fundamental prior knowledge to the instrumental analysis of organic compounds (Unit 4 AOS 2).

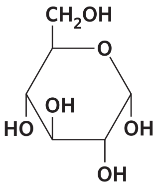
5A Questions

Mild Medium Spicy

Deconstructed

Use the following information to answer questions 12–14.

Glucose, $C_6H_{12}O_6$, is an important source of energy for the body and is transported through the bloodstream.



α - glucose

Question 12 (1 MARK)

Which of the following is a correct statement regarding the solubility of glucose?

A. Glucose is soluble in non–polar solvents because 'opposites attract'.

B. Glucose is soluble in polar solvents because 'like dissolves like'.

C. Glucose is insoluble in polar solvents because glucose can form hydrogen bonds with the solvent.

D. Glucose is insoluble in non–polar solvents because glucose can form hydrogen bonds with the solvent.

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Question 13 (1 MARK)

In order for glucose to dissolve, which of the following correctly lists all of the forces that need to be broken and formed?

I. Solute–solute forces broken
II. Solute–solute forces formed
III. Solvent–solvent forces broken
IV. Solvent–solvent forces formed
V. Solute–solvent forces broken
VI. Solute–solvent forces formed

A. I, IV, VI
B. II, III, V
C. I, III, VI
D. I, III, V

Question 14 (4 MARKS)

a. Draw the interaction between glucose and water molecules. (1 MARK)

b. Explain why the interactions between molecules enable solid glucose to dissolve in water, including the chemical equation that represents this process in your answer. (3 MARKS)

Exam-style

Question 15 (1 MARK)

Which of the following substances will be able to dissolve in water without forming ions?

A. Sulfuric acid (H_2SO_4)
B. Nitric acid (HNO_3)
C. Hydrochloric acid (HCl)
D. Formaldehyde (CH_2O)

Question 16 (1 MARK)

Consider the following statements about a high performance liquid chromatography (HPLC) column that uses a non-polar mobile phase and a polar stationary phase to analyse a solution:

I. Polar molecules in the solution will be attracted to the solvent particles by dipole-dipole attraction.
II. Non-polar molecules in the solution will be attracted to the stationary phase by dispersion forces.
III. Non-polar molecules in the solution will travel through the HPLC column more rapidly than polar molecules.

Which of these statements are true?

A. I and II only **B.** I and III only **C.** II and III only **D.** I, II and III

Adapted from VCAA 2015 Exam 2 Multiple choice Q8

Question 17 (1 MARK)

A chemistry student has inadvertently mixed benzene, a non–polar compound, and water together. Which of the following HPLC columns would lead to water eluting from the column in a shorter time than benzene?

	Stationary phase	**Mobile phase**
A.	Octane (C_8H_{18})	Ethanol (C_2H_5OH)
B.	Silica	Propane (C_3H_8)
C.	Silica	Hexane (C_6H_{14})
D.	Silica	CCl_4

Question 18 (7 MARKS)

A scientific researcher has just purchased a new HPLC device and wanted to start using it to determine the concentration of a specialised immune cell in the blood of a patient recently diagnosed with COVID-19.

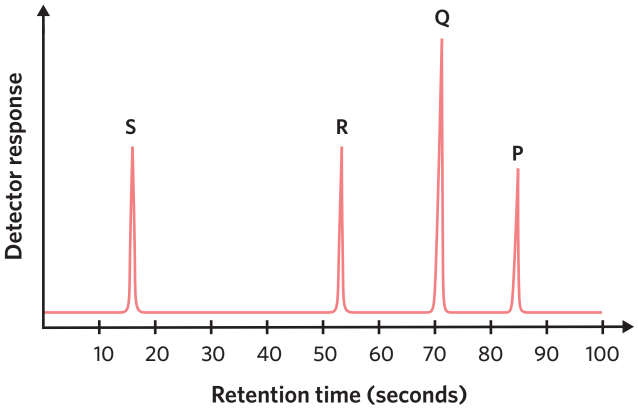
a. Using your knowledge of chromatography, explain how components of a sample are separated by HPLC. (3 MARKS)

b. Outline the steps required to produce a calibration curve for this experiment. (2 MARKS)

c. A calibration curve was then generated. However, a few moments later, a laboratory assistant opened a window in the lab, and the temperature of the room increased. If the concentration of this immune cell needs to be determined accurately, will a new calibration curve have to be generated? Explain why or why not. (2 MARKS)

Question 19 (6 MARKS)

A small sample of blood was taken from a crime scene and sent off to the laboratory for forensic analysis. The coroner decided to use HPLC with a polar stationary phase to analyse the sample. The results are shown below.



a. Based on the chromatogram, how many components are in the sample? (1 MARK)

b. Which of the components, S, P, Q or R, is the most polar? Explain your answer. (2 MARKS)

c. Suggest how the components could be accurately identified in order to help the police solve the crime. (1 MARK)

Police soon suspect that alcohol was involved in the crime and instructed the coroner to determine the concentration of alcohol in the sample of blood.

d. Given they know which peak corresponds to ethanol, explain how the coroner could determine the concentration of alcohol in the sample of blood. (2 MARKS)

Adapted from VCAA 2016 Exam 2 Short answer Q1

Key science skills

Question 20 (5 MARKS)

Margaret is testing whether certain compounds can be dissolved in water. She believes that she can do this visually by placing different solid compounds into water, and then observing whether they are visible or not after 30 seconds.

a. What type of data is being collected? (1 MARK)

b. What is the independent variable in this experiment? (1 MARK)

For one experiment, Margaret times how long it takes for 5 g of a solid substance to dissolve. She places 5 g of substance A into a beaker and 5 g of substance B into another beaker, each with the same volume of water. However, she stirs substance A but not substance B. Both substances are completely dissolved after exactly 20 seconds.

c. Is this experiment valid? Explain why or why not. (2 MARKS)

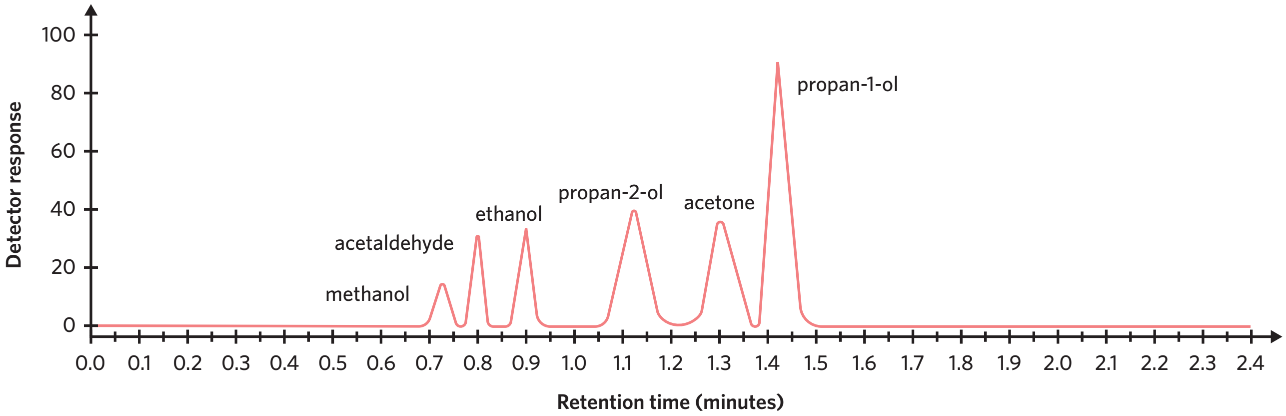
d. Suggest one improvement to the experimental design to improve its validity. (1 MARK)

FROM LESSONS 16A, 16B & 16C

Questions from multiple lessons

Question 21 (5 MARKS)

A forensic chemist wants to test the accuracy of a HPLC column that is to be used for the analysis of blood alcohol content. A blood sample may contain a number of volatile chemicals that can interfere with the identification and measurement of ethanol in the blood. A sample containing a mixture of ethanol and several other volatile chemicals was injected into the column. The following chromatogram was obtained.



a. The forensic chemist claims that the presence of these volatile chemicals would not affect the qualitative analysis of ethanol.

- **i.** What evidence is presented in the chromatogram to support this claim? (1 MARK)
- **ii.** To determine the percentage of alcohol in a blood sample only the peak at a retention time of 0.9 minutes is measured. Explain why. (2 MARKS)

b. Consider a sample of methanol. What is the strongest type of bonding that would be present between methanol molecules? (1 MARK)

c. Draw a diagram showing which atoms the bonds mentioned in part b form between. (1 MARK)

Adapted from VCAA 2010 Exam 1 Short answer Q5

FROM LESSON 2B

Hints

15. Molecular substances are able to dissolve in water either through hydrogen bonding or ionisation.

16. All molecules are able to form instantaneous dipoles.

17. Retention time depends on the polarities of the mobile phase, stationary phase and components in the sample.

18a. Chromatography relies on different strengths of attraction to the stationary and mobile phase

18b. A calibration curve links the concentration of a component to its absorbance.

18c. Values can only be compared if obtained in identical conditions.

19a. The peaks in a chromatogram reflect the components in a sample.

19b. The polarities of the stationary phase, mobile phase, and a given component all affect the retention time of a component.

19c. Different compounds have different retention times.

19d. A calibration curve links the concentration of a component to its absorbance.

20a. Data can be either numeric or observational

20b. The independent variable is the variable that is changed by the experimenter.

20c. Validity refers to whether an experiment accurately measures the aim.

20d. All variables except the independent variable should be controlled in an experiment.

21ai. The presence of a compound in a sample can be determined from the chromatogram produced.

21aii. Each compound has a unique retention time under given conditions.

21b. The forces between methanol molecules are very strong.

21c. Hydrogen bonding occurs between a hydrogen atom and a very electronegative atom.

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Chapter 5 review

Multiple choice (10 MARKS)

Question 1 (1 MARK)

Which of the following compounds is not soluble in water?

A. C_4H_{10}

B. CH_3OH

C. KNO_3

D. HCl

Question 2 (1 MARK)

A thin layer chromatography plate was set up with a non-polar solvent, hexane, and a polar stationary phase, silica gel. The chromatogram shown was obtained. A ruler was then placed next to the plate.

The R_f value for the compound would be

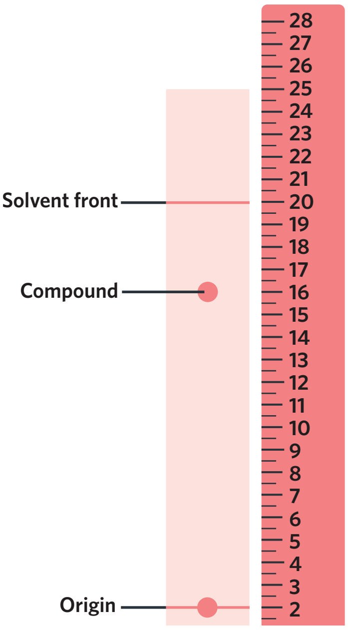
A. 0.80

B. 0.78

C. 0.64

D. 0.61

VCAA 2014 exam Multiple choice Q14



Question 3 (1 MARK)

A paper chromatograph of four dyes, G, P, R and Y, is shown.

The R_f value of the dye most strongly adsorbed onto the stationary phase is

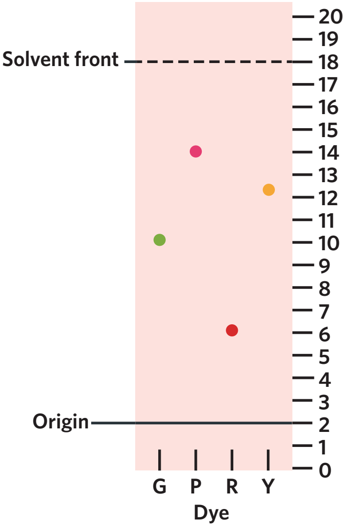
A. 0.25

B. 0.33

C. 0.75

D. 0.78

VCAA 2016 exam Multiple choice Q4



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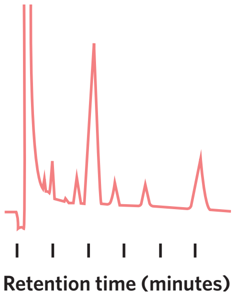
Use the following information to answer questions 4–5.

Deadly diseases such as tetanus, botulism and gangrene are caused by related groups of bacteria (Clostridium genus) found in the soil. Each group of bacteria produces specific fatty acids. These fatty acids can be identified using HPLC chromatography by comparison with a control chromatogram of known standards. The chromatogram shown is of the fatty acids produced by one such group of bacteria.



Question 4 (1 MARK)

The identity of the fatty acids can be determined by measuring

A. their retention times.

B. the temperature of the column.

C. the flow rate of the solvent used.

D. the area under each of the peaks.

Adapted from VCAA 2008 Exam 1 Multiple choice Q12

Question 5 (1 MARK)

The relative amount of each of the fatty acids can be determined by measuring

A. their retention times.

B. the temperature of the column

C. the flow rate of the solvent used.

D. the area under each of the peaks.

Adapted from VCAA 2008 Exam 1 Multiple choice Q13

Question 6 (1 MARK)

Hexane is a non-polar solvent used to extract edible oils from seeds and vegetables. Which of the following organic compounds would be most soluble in hexane?

A. Methanol, CH_3OH

B. Butanol, C_4H_9OH

C. Glucose, $C_6H_{12}O_6$

D. Octanol, $C_8H_{17}OH$

Question 7 (1 MARK)

The thin layer chromatography plate shown has a polar stationary phase. It was developed using the hydrocarbon benzene (C_6H_6) as the solvent.

Which sample has the most polar molecules?

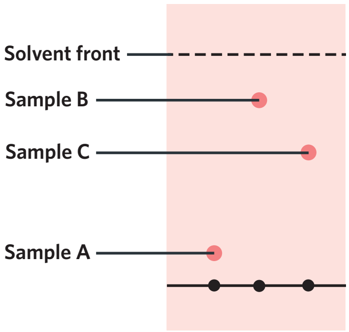
A. Sample A

B. Sample B

C. Sample C

D. There is not enough information to determine which sample has the most polar molecules.

Adapted from VCAA 2013 exam Multiple choice Q7



Question 8 (1 MARK)

Which of the following would not ionise when dissolving in water?

A. H_2SO_4

B. CH_3COOH

C. C_2H_5OH

D. HNO_3

Question 9 (1 MARK)

Consider the following statements about a high-performance liquid chromatography (HPLC) column that uses a polar solvent and a non-polar stationary phase to analyse a solution:

Statement I – Polar molecules in the solution will be attracted to the solvent particles by dipole-dipole attraction.

Statement II – Non-polar molecules in the solution will be attracted to the stationary phase by dispersion forces.

Statement III – Polar molecules in the solution will travel through the HPLC column more rapidly than non-polar molecules.

Which of these statements are true?

A. I and II only

B. I and III only

C. II and III only

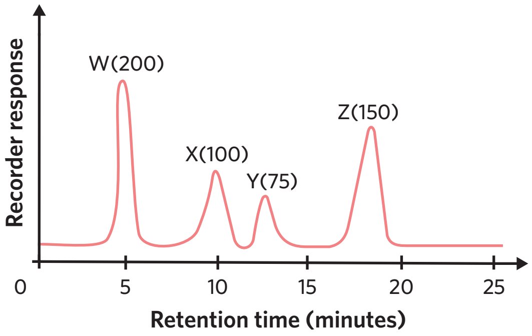
D. I, II and III

VCAA 2015 exam Multiple choice Q8

Question 10 (1 MARK)

A polar solvent was passed through a HPLC column to analyse a sample containing four organic compounds. The diagram given shows the HPLC chromatogram produced.

Relative peak areas are given in brackets.



The following statements refer to this chromatogram.

Statement I – The retention times will stay the same if the temperature at which the chromatogram is recorded is increased, with all other conditions remaining constant.

Statement II – The polarity of the compounds arranged from least to most polar is Z > Y > X > W.

Statement III – Compound Z is present in the highest concentration.

Which of the above statements are true?

A. I only

B. II only

C. I and II only

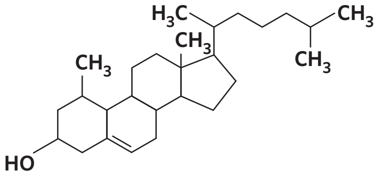
D. II and III only

Adapted from VCAA 2009 Exam 1 Multiple choice Q6

Short answer (30 MARKS)

Question 11 (3 MARKS)

Cholesterol, $C_{27}H_{45}OH$, is a biological compound that aids in the production of cell membranes, hormones and vitamin D. Cholesterol's structure is shown.

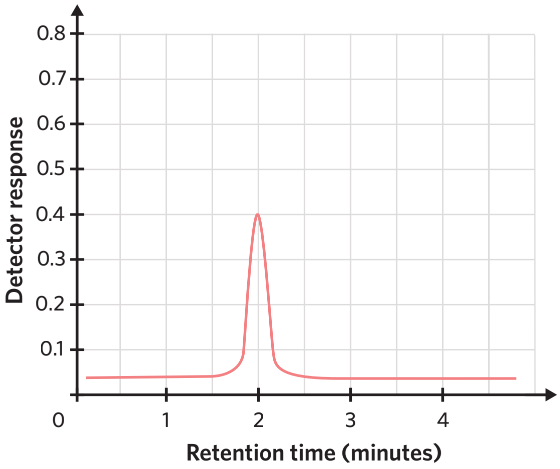


Cholesterol

Comment on the solubility of cholesterol in water, giving reference to the polarity of each compound in your answer.

Question 12 (2 MARKS)

High-performance liquid chromatography is used to determine the amount of glucose in a sample of a soft drink. The chromatogram given shows the detector response when a standard solution of glucose with a concentration of 200 mg L^{-1} is measured using the instrument.



a. What is the retention time of glucose in this experiment? (1 MARK)

b. On a chromatogram similar to the one shown, sketch the detector response when a commercial soft drink with a glucose content of 350 mg L^{-1} is measured using the same instrument. (1 MARK)

VCAA 2013 exam Short answer Q1

Question 13 (5 MARKS)

Researchers have developed a technique that allows high-performance liquid chromatography (HPLC) to simultaneously determine the concentration of sugar, organic acids and alcohols in fermentation products and food samples. The table shown contains the retention times for some common organic molecules using this technique.

Organic molecule	Retention time (min)
Maltose	12.50
Lactose	12.70
Glucose	14.45
Succinic acid	18.25
Glycerol	20.33
Ethanol	30.63

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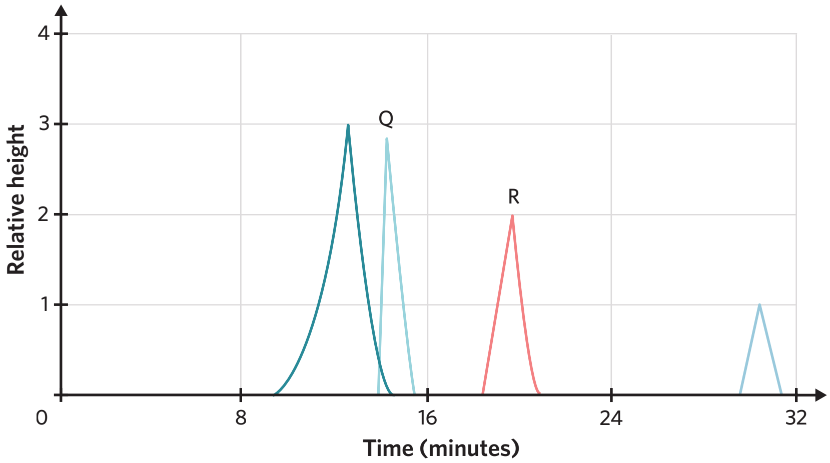
Shown is a chromatogram of four organic molecules.



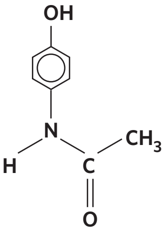
a. Identify the substances responsible for each of the peaks Q and R. (2 MARKS)

b. Two substances have different retention times under identical conditions using the same HPLC equipment. With reference to solubility, explain this difference in retention times. (3 MARKS)

Adapted from VCAA 2018 exam Short answer Q5a,b

Question 14 (4 MARKS)

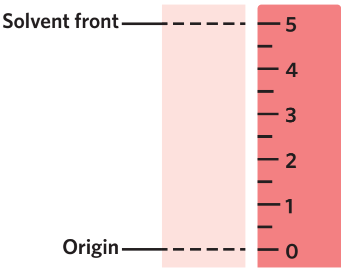
Paracetamol is a commonly used painkiller. The structure of paracetamol is shown.



Paracetamol

A student uses thin layer chromatography (TLC) to analyse the products of the preparation of paracetamol. For the stationary and mobile phases used for this analysis, the R_f of paracetamol is 0.4.

a. On a diagram of a TLC plate similar to the one shown, use a horizontal line to mark the spot where paracetamol would appear in such an analysis. (1 MARK)

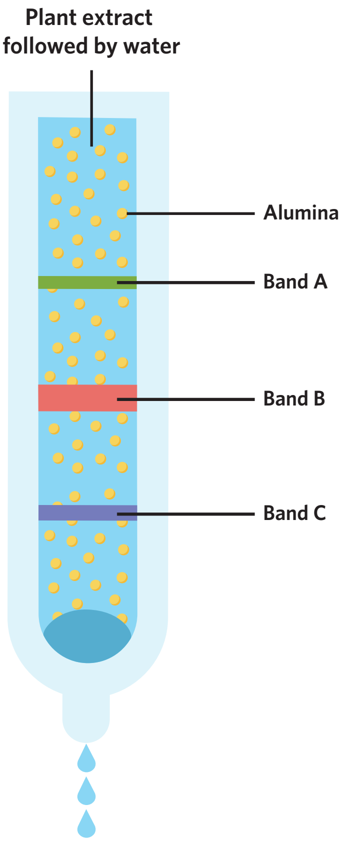


b. 4-aminophenol adsorbs less strongly than paracetamol onto the stationary phase of this TLC plate. Predict whether the R_f value of 4-aminophenol in this analysis is greater or smaller than that of paracetamol, giving a reason for your choice. (3 MARKS)

Adapted from VCAA 2008 Exam 1 Short answer Q3b

Question 15 (10 MARKS)

A burette packed with finely divided alumina powder, Al_2O_3, was used to separate the components in a plant extract by column chromatography. The alumina acts as the stationary phase and water was used as the mobile phase. A solution of the plant extract was placed at the top of the alumina and, once it had been adsorbed, further water was added periodically. The components separated as three coloured bands, as shown in the diagram.



a. In which one of the three bands, A, B or C, were the components most strongly adsorbed to the stationary phase?
Justify your answer. (2 MARKS)

b. Band B starts to show signs of separating into two different bands just before it emerges from the bottom of the burette. Suggest two possible modifications that could be made to this experiment to more effectively separate band B into its separate components. (2 MARKS)

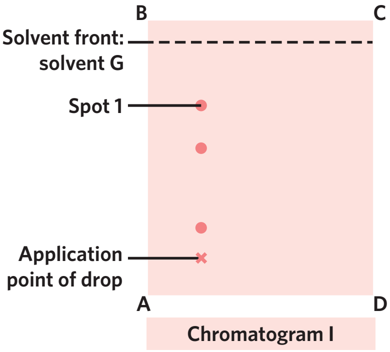
c. In another experiment, the components in the plant extract were separated in a similar way to paper chromatography, using a glass sheet coated with alumina and water as the mobile phase. Which one of the three bands, A, B or C, contains the component that would be most likely to have the largest R_f value? Explain your answer. (3 MARKS)

d. Which one of the three bands, A, B or C, contains the component that would have the largest retention time if this separation was conducted using high-performance liquid chromatography with alumina as the stationary phase? Explain your answer. (3 MARKS)

Adapted from VCAA 2006 Exam 1 Short answer Q2

Question 16 (6 MARKS)

A drop that contains a mixture of four amino acids was applied to a thin layer chromatography plate. The plate was placed in solvent G and the following chromatogram was obtained.



The R_f values for each of the amino acids in solvent G are provided in table 1.

Table 1 R_f values in solvent G

amino acid	R_f (solvent G)
alanine	0.51
arginine	0.16
threonine	0.51
tyrosine	0.68

a. Threonine is known to be significantly more polar than alanine.
Compare the solubilities of threonine and alanine in water. (2 MARKS)

b. Name the amino acid that corresponds to spot 1. (1 MARK)

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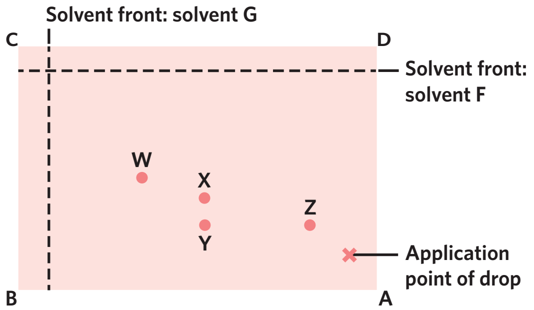
c. The plate was dried, rotated through 90° in an anticlockwise direction, and then placed in solvent F. The R_f values for each of the amino acids in solvent F are provided in table 2.

Table 2 R_f values in solvent F

Amino acid	R_f (solvent F)
Alanine	0.21
Arginine	0.21
Threonine	0.34
Tyrosine	0.43

The given chromatogram was obtained.



Chromatogram II

Identify the spot on chromatogram II that represents alanine. (1 MARK)

d. Explain, in terms of the data provided, why only three spots are present in chromatogram I while four spots are present in chromatogram II. (2 MARKS)

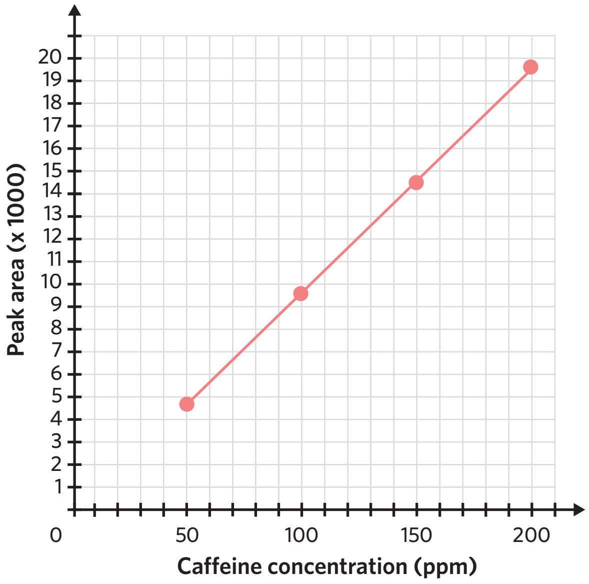
Adapted from VCAA 2012 Exam 1 Short answer Q2

Key science skills (10 MARKS)

Use the following information to answer questions 17–23.

Caffeine is a stimulant drug that is found in coffee, tea, energy drinks and some soft drinks. The concentration of caffeine in a sample can be determined using HPLC. Four solutions containing 50 ppm, 100 ppm, 150 ppm and 200 ppm of caffeine were prepared. 25 microlitres of each sample was injected into the HPLC column. The peak areas were measured and used to construct the calibration curve shown.

Peak area of caffeine solutions: retention time = 96 seconds



25 microlitre samples of various drinks thought to contain caffeine were then separately passed through the HPLC column. The results are shown below.

Sample	Retention time of major peak (seconds)	Peak area of largest peak
Soft drink A	96	12 000
Soft drink B	32	8500
Espresso coffee	96	211 000

Question 17 (1 MARK)

What is the independent variable in this investigation?

Question 18 (1 MARK)

What type of data is being collected in this investigation?

Question 19 (1 MARK)

Based on the results of this investigation, which of the following samples has the highest concentration of caffeine?

A. 50 ppm caffeine solution

B. 100 ppm caffeine solution

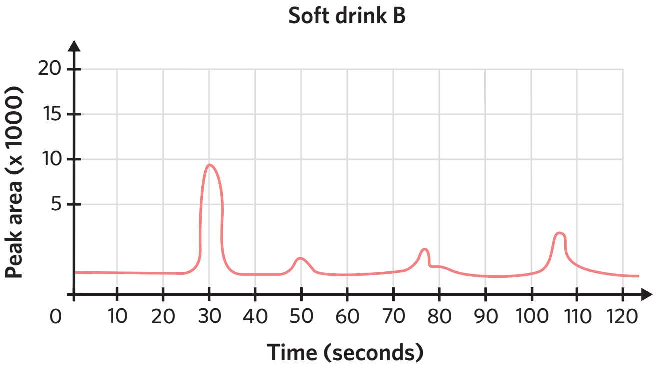
C. Soft drink B

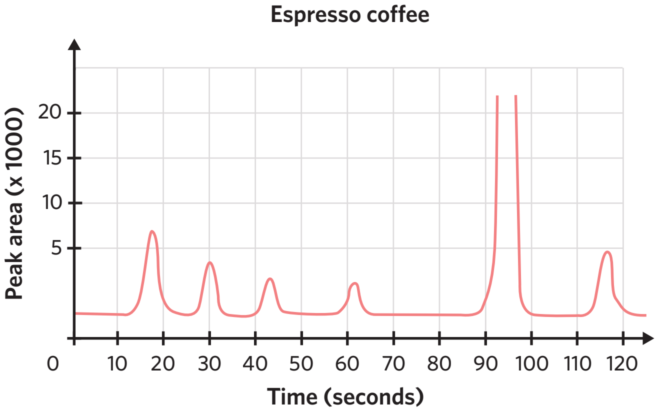
D. Espresso coffee

Question 20 (1 MARK)

Determine the caffeine concentration, in ppm, of soft drink A.

The chromatograms obtained from the samples of soft drink B and espresso coffee are shown.





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Question 21 (2 MARKS)

What evidence is presented in the chromatogram that supports the conclusion that soft drink B does not contain any caffeine?

Question 22 (2 MARKS)

Explain why the caffeine content of the espresso coffee sample cannot be reliably determined using the information provided.

Question 23 (2 MARKS)

Describe what could be done to the espresso coffee sample so that its caffeine content can be reliably determined using the information provided.

Adapted from VCAA 2011 Exam 1 Short answer Q3

FROM LESSONS 16B & 16C

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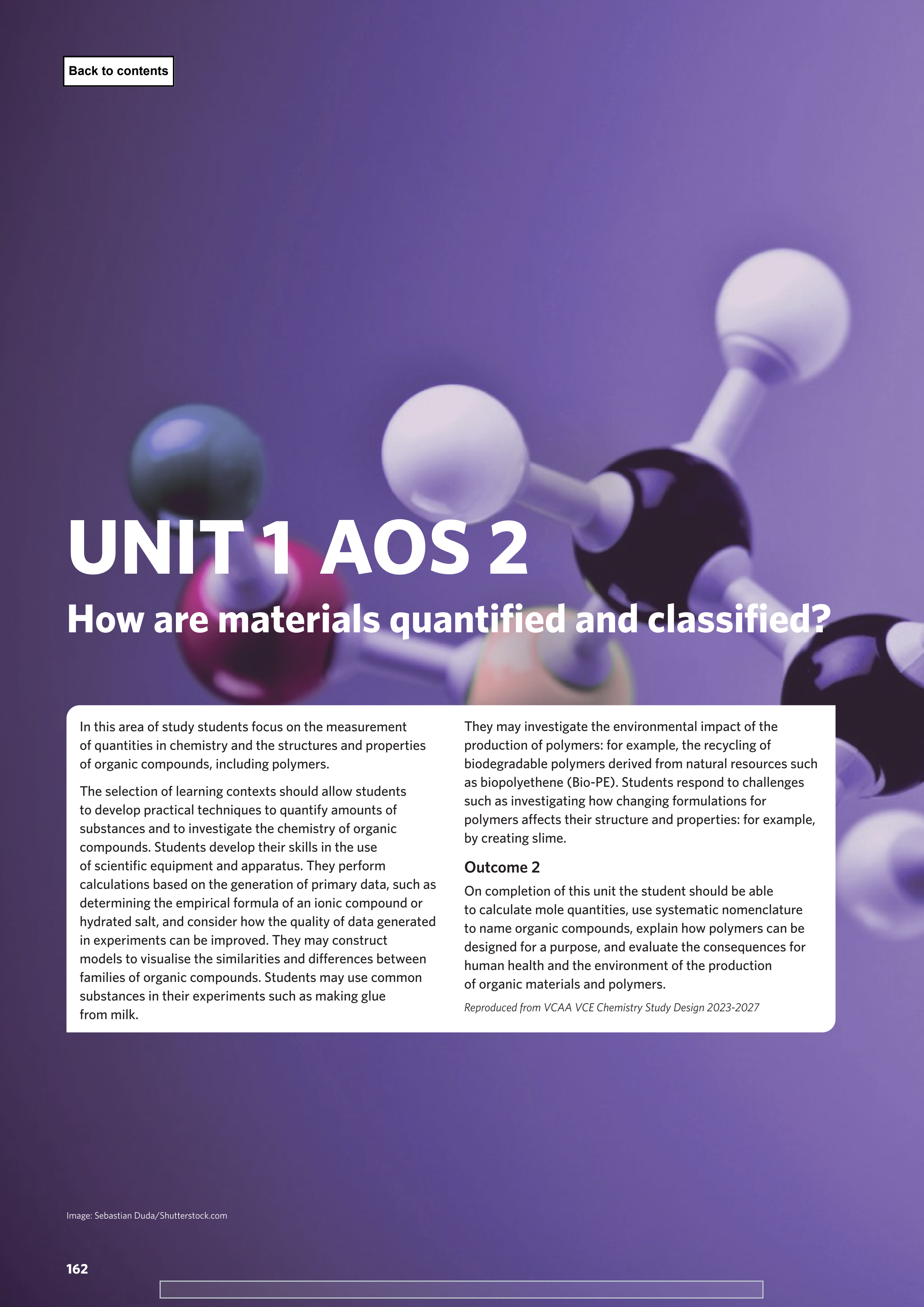

UNIT 1 AOS 2

How are materials quantified and classified?

In this area of study students focus on the measurement of quantities in chemistry and the structures and properties of organic compounds, including polymers.

The selection of learning contexts should allow students to develop practical techniques to quantify amounts of substances and to investigate the chemistry of organic compounds. Students develop their skills in the use of scientific equipment and apparatus. They perform calculations based on the generation of primary data, such as determining the empirical formula of an ionic compound or hydrated salt, and consider how the quality of data generated in experiments can be improved. They may construct models to visualise the similarities and differences between families of organic compounds. Students may use common substances in their experiments such as making glue from milk.

They may investigate the environmental impact of the production of polymers: for example, the recycling of biodegradable polymers derived from natural resources such as biopolyethene (Bio-PE). Students respond to challenges such as investigating how changing formulations for polymers affects their structure and properties: for example, by creating slime.

Outcome 2

On completion of this unit the student should be able to calculate mole quantities, use systematic nomenclature to name organic compounds, explain how polymers can be designed for a purpose, and evaluate the consequences for human health and the environment of the production of organic materials and polymers.

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Image: Sebastian Duda/Shutterstock.com

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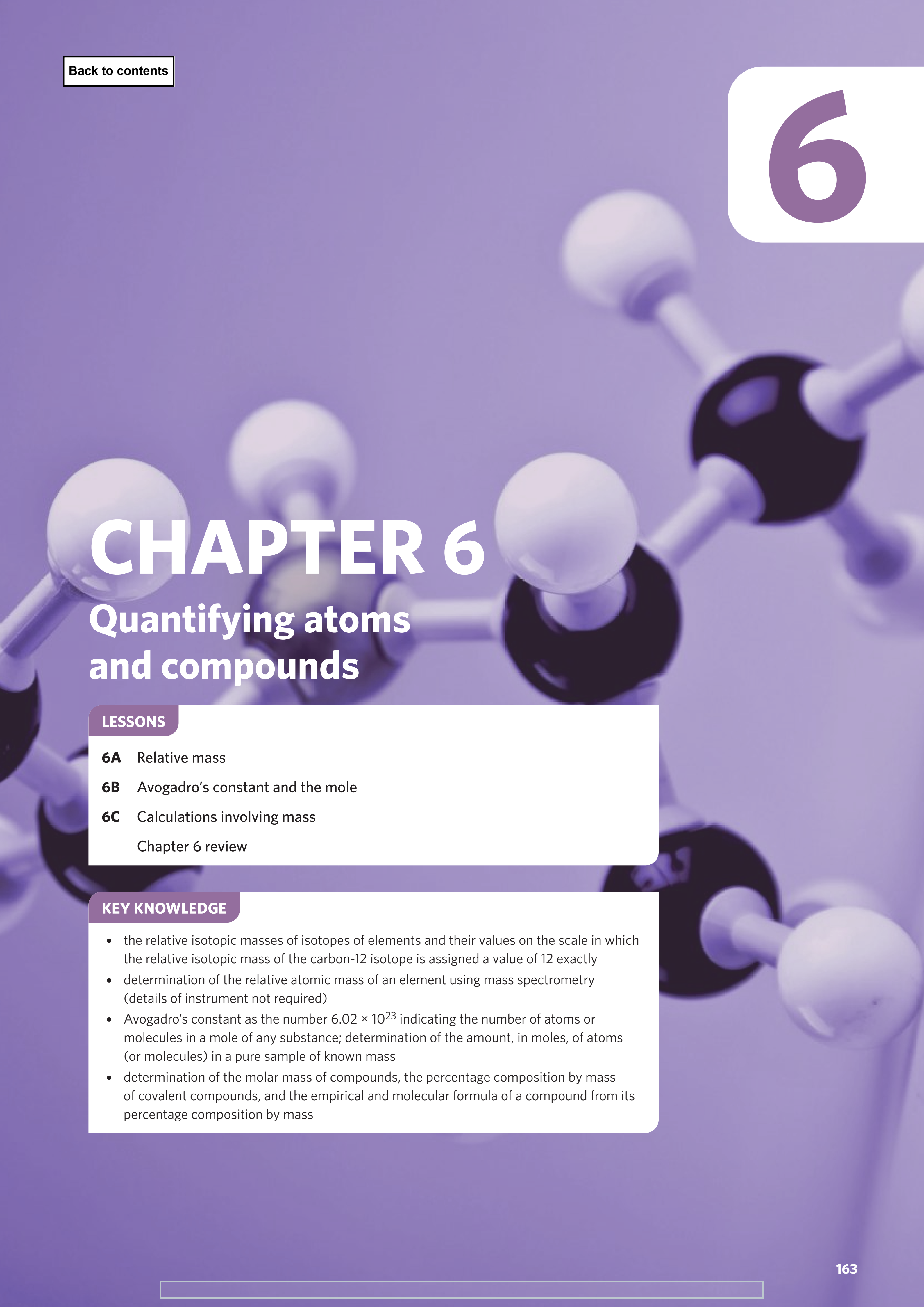

6

CHAPTER 6

Quantifying atoms and compounds

LESSONS

6A Relative mass
6B Avogadro’s constant and the mole
6C Calculations involving mass
Chapter 6 review

KEY KNOWLEDGE

- the relative isotopic masses of isotopes of elements and their values on the scale in which the relative isotopic mass of the carbon-12 isotope is assigned a value of 12 exactly
- determination of the relative atomic mass of an element using mass spectrometry (details of instrument not required)
- Avogadro’s constant as the number 6.02×10^{23} indicating the number of atoms or molecules in a mole of any substance; determination of the amount, in moles, of atoms (or molecules) in a pure sample of known mass
- determination of the molar mass of compounds, the percentage composition by mass of covalent compounds, and the empirical and molecular formula of a compound from its percentage composition by mass

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6A Relative mass

STUDY DESIGN DOT POINTS

- the relative isotopic masses of isotopes of elements and their values on the scale in which the relative isotopic mass of the carbon-12 isotope is assigned a value of 12 exactly
- determination of the relative atomic mass of an element using mass spectrometry (details of instrument not required)

6A 6B 6C

1.2.1.1 Relative masses of elements

1.2.1.2 Relative molecular and formula mass

1.2.2.1 Mass spectrometry

ESSENTIAL PRIOR KNOWLEDGE

1A Isotopes
2A Molecular formula
4A Ionic crystals
See questions 30–32.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

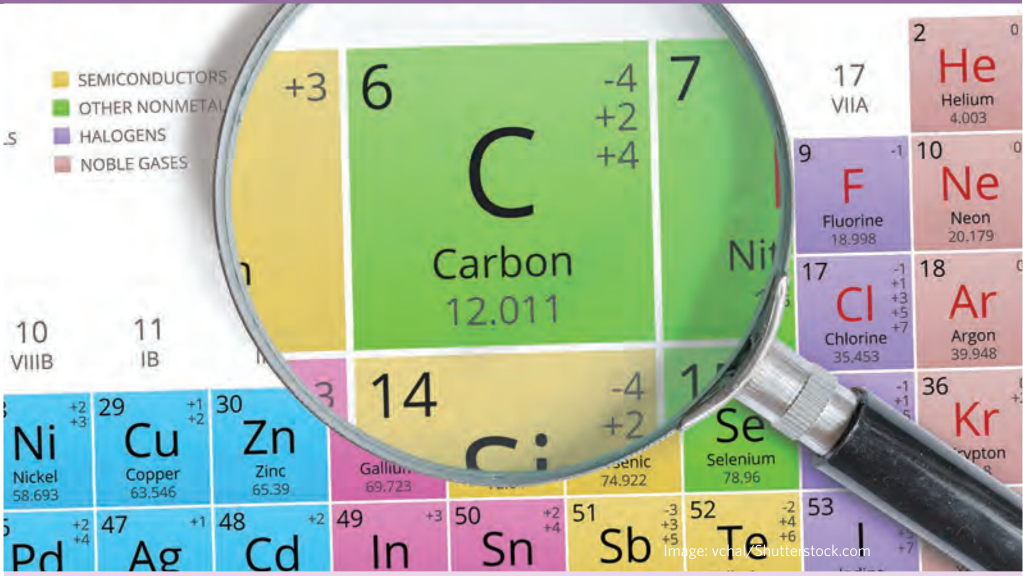

Image: vchal/Shutterstock.com

How is carbon-12 used to calculate the relative atomic masses of all elements?

Carbon-12 is used as the standard on the scale of atomic masses. In this lesson, we will learn about how carbon-12 is used to calculate the relative masses of elements, compounds and molecules, as well as how the **relative abundance of isotopes** can be determined using **mass spectrometry**.

KEY TERMS AND DEFINITIONS

Isotope atoms of the same element that have a different mass due to a different number of neutrons

Mass spectrometry analytical technique used to measure the mass of ions relative to their charge

Mass-to-charge ratio (m/z) the mass of an ion divided by its charge

Molecular formula actual number of atoms in a molecule

Relative abundance percentage of a particular isotope found in a naturally occurring sample of an element

Relative atomic mass (A_r) weighted average of the masses of an element's isotopes on a scale on which a carbon-12 atom is assigned a value of 12 exactly

Relative formula mass (M_r) weighted average of a substance's masses of the formula units on a scale on which the mass of a carbon-12 atom is assigned a value of 12 exactly

Relative isotopic mass the mass of the isotope on the scale in which the relative isotopic mass of the carbon-12 atom is assigned a value of 12 exactly

Relative molecular mass (M_r) weighted average of a molecule's masses of the formula units on a scale on which the mass of a carbon-12 atom is assigned a value of 12 exactly

Relative masses of elements 1.2.1.1

Despite being invisible to the naked eye, individual atoms of each element still have a mass.

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e relative masses of all elements compared to the mass of an atom of carbon-12?

Given that atoms are so small, it is difficult (and inconvenient) for chemists to measure them in kilograms, grams, or other SI units. Instead, chemists use a relative measurement that measures the mass of every atom in comparison to the mass of a single atom of the carbon-12 isotope, ^{12}C. The ^{12}C isotope has six protons and six neutrons in the nucleus, with the mass of electrons negligible in comparison, and is, therefore, assigned an atomic mass of exactly 12.

How can the relative mass of an element be calculated from the relative abundance of its isotopes?

As we learned in lesson 1A, there are different forms of the same element called isotopes. Each isotope of an element contains the same number of protons but a different number of neutrons in its nucleus. This means that each isotope of an element has a slightly different mass, which can be compared with the mass of a single carbon-12 atom. This property is referred to as the **relative isotopic mass**. Each isotope of an element takes up a different percentage of the overall mass of an element, otherwise known as the relative abundance of each isotope. A rarer isotope with a lower relative abundance contributes less to the **relative atomic mass (A_r)** – the weighted average of all the individual relative isotopic masses – of an element than a more common isotope with a higher relative abundance.

As shown in figure 1, the relative atomic mass of each element is displayed the element's symbol on the periodic table.

MISCONCEPTION

'The heaviest isotope of an element is always the most abundant.'

It's important to make the distinction between relative abundance and relative isotopic mass. Just because an isotope is heavier, it does not necessarily mean that it is the most abundant!

USEFUL TIP

The relative atomic masses of all elements can be found on the periodic table in the VCE Data Book and as they are relative masses, they have no units.

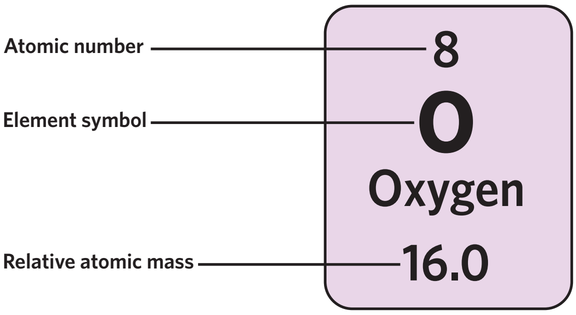

Figure 1 The relative atomic mass of oxygen on the periodic table

USEFUL TIP

The relative atomic mass of most elements on the periodic table appears as a decimal, rather than a whole number. This is because the relative atomic mass is found by averaging the masses – the combined number of protons and neutrons – of all naturally occurring isotopes.

As the relative atomic mass (A_r) is the weighted mean of the relative isotopic masses of an element, we can calculate the relative atomic mass of an element using the formula:

$$A_r = \frac{(\text{relative abundance }\% \times \text{relative isotopic mass}) + (\text{relative abundance }\% \times \text{relative isotopic mass}) + \ldots}{100\%}$$

MISCONCEPTION

'All elements have at least two naturally occurring isotopes.'

Some elements have only one naturally occurring isotope. For example, sodium-23 is the only naturally occurring isotope of sodium.

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PLE 1

The element chlorine has two stable isotopes – ^{35}Cl and ^{37}Cl. Chlorine-35 has a relative abundance of 75%, and chlorine-37 has a relative abundance of 25%. Calculate the relative atomic mass of chlorine.

What information is presented in the question?

The two isotopes of chlorine have relative isotopic masses of 35 and 37, which have a relative abundance of 75% and 25% respectively.

What is the question asking us to do?

Calculate the relative atomic mass.

What strategies do we need in order to answer the question?

1. Identify the formula for relative atomic mass.
2. Substitute the values given from the question into the formula.
3. Calculate the relative atomic mass.

Answer

Relative isotopic mass	Relative abundance (%)
35	75
37	25

The formula for relative atomic mass is:

$$A_r = \frac{(\text{relative abundance } \% \times \text{relative isotopic mass}) + (\text{relative abundance } \% \times \text{relative isotopic mass}) + \ldots}{100\%}$$

We can substitute the values into the equation.

$$A_r = \frac{(75\% \times 35) + (25\% \times 37)}{100\%}$$

$$A_r = 35.5$$

How can the relative abundance of an element's isotopes be calculated from its relative atomic mass?

Sometimes, the relative abundance of each isotope of an element may be unknown. However, if the relative atomic mass of the element is known, we can calculate the relative abundance.

Consider the element bromine, which has a relative atomic mass of 79.9. It exists as two isotopes, ^{79}Br and ^{81}Br. Although we don't know the relative abundance of either bromine isotope, we know that the total relative abundance of isotopes must always add up to 100%. As such, the relative abundance of ^{79}Br can be denoted as x and the relative abundance of ^{81}Br can be denoted as $100 - x$, as shown in table 1.

Table 1 Calculating the relative abundance of bromine isotopes

Relative isotopic mass		Relative abundance (%)
^{79}Br	79	x
^{81}Br	81	$100 - x$

If we substitute the information from table 1 and the relative atomic mass for bromine (from the periodic table) into the formula for relative atomic mass, we can calculate the value of x as shown.

$$A_r = \frac{(\text{relative abundance } \% \times \text{relative isotopic mass}) + (\text{relative abundance } \% \times \text{relative isotopic mass}) + \ldots}{100\%}$$

$$79.9 = \frac{(x\% \times 79) + ((100 - x)\% \times 81)}{100\%}$$

$$79.9 = \frac{(79x + 8100 - x)\%}{100\%}$$

$$79.9 \times 100\% = (79x + 8100 - 81x)\%$$

$$7990\% = (79x + 8100 - 81x)\%$$

$$7990 = -2x + 8100$$

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$$\frac{-110}{-2} = x$$

$$x = 55$$

If we then substitute the value obtained for x into table 2 below, we can determine the relative abundance of each isotope of bromine.

Table 2 Calculating the relative abundance of bromine isotopes

Isotope	Relative isotopic mass	Relative abundance (%)
^{79}Br	79	55
^{81}Br	81	100 – 55 = 45

Therefore, the relative abundance of ^{79}Br is 55%, and the relative abundance of ^{81}Br is 45%.

USEFUL TIP

In VCE chemistry only *M* is used to represent the relative masses of all compounds (both ionic and covalent).

WORKED EXAMPLE 2

Europium was discovered by French chemist Eugène–Antole Demarçay in the late 1800s. Europium is known to have two naturally occurring isotopes: ^{151}Eu and ^{153}Eu. If the relative atomic mass of europium is 151.96, calculate the relative abundance of each isotope.

What information is presented in the question?

There are two isotopes of europium.

The relative atomic mass of europium is 151.96.

What is the question asking us to do?

Calculate the relative abundance of each isotope.

What strategies do we need in order to answer the question?

1. Identify the formula for relative atomic mass and let x represent one of the unknown relative isotopic abundances.
2. Substitute the values given from the question into the formula.
3. Calculate the relative abundance of one isotope.
4. Use the value obtained in step 3 to calculate the abundance of the second isotope.

Answer

Isotope	Relative isotopic mass	Relative abundance (%)
^{151}Eu	151	x
^{153}Eu	153	$100 - x$

The formula for relative atomic mass is:

$$A_r = \frac{(\text{relative abundance \%} \times \text{relative isotopic mass}) + (\text{relative abundance \%} \times \text{relative isotopic mass}) + \ldots}{100\%}$$

We can substitute the values into the equation.

$$151.96 = \frac{(x\% \times 151) + ((100 - x)\% \times 153)}{100\%}$$

$$151.96 = \frac{(151x + 15\,300 - 153x)\%}{100\%}$$

$$151.96 \times 100\% = (151x + 15\,300 - 153x)\%$$

$$15\,196\% = (151x + 15\,300 - 153x)\%$$

$$15\,196 = -2x + 15\,300$$

$$-104 = -2x$$

$$\frac{-104}{-2} = x$$

$$x = 52.0$$

Therefore, the relative abundance of ^{151}Eu is 52.0%, and the relative abundance of ^{153}Eu is:

100 – 52.0 = 48.0%.

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Progress questions

Question 1

The masses of elements on the periodic table are given relative to

A. the mass of the carbon-12 isotope.

B. the mass of the oxygen-16 isotope.

Question 2

The relative atomic mass of an element

A. depends on the abundance of neutrons in an element.

B. depends on the abundance of each isotope of an element.

Question 3

Element Z exists in two forms. ^{14}Z has a relative abundance of 78% and ^{15}Z has a relative abundance of 22%. The relative atomic mass of element Z will be closest to

A. 14.

B. 15.

Relative molecular and formula mass 1.2.1.2

Using the relative atomic mass of elements, we can calculate the relative mass of molecules.

How can the relative mass of a molecule be calculated?

As we learned in lesson 2A, multiple atoms can bond together to form molecules. We can use the relative atomic mass of the atoms of elements to calculate the **relative molecular mass (M_r)** of molecules. To calculate the relative molecular mass of a molecule, we first need to know how to use the **molecular formula** of a molecule to determine how many atoms of each element are present in the molecule.

For example, the molecular formula for water is H_2O. We can see that there are two elements present: hydrogen and oxygen. The subscript after the chemical symbol H indicates that there are two hydrogen atoms present, and only one oxygen atom present in every water molecule. As such, in order to calculate the relative molecular mass for water, we would add the relative atomic masses of all atoms present in the molecule, as shown in figure 2.

The A_r of two hydrogen atoms

$$M_r\,(H_2O) = (1.0 \times 2) + 16.0 = 18.0 \text{ (no units)}$$

The A_r of a single oxygen atom

The molecular formula of a water molecule

Figure 2 Calculating the relative molecular mass of water

USEFUL TIP

In chemistry, we have a special way of communicating ideas in calculations. For example, in this lesson we use the M_r symbol to indicate that we are calculating the relative mass of whatever is inside the brackets – this sits on the left-hand side of the equation. On the right-hand side, we write out the numerical calculations that lead to our solution.

$$M_r(H_2O) = (1.01 \times 2) + 16.0 = 18.0$$

The strategy to calculate the molecular mass of H_2O

“The relative molecular mass of H_2O is”

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AMPLE 3

Hydrogen peroxide, H_2O_2, is used as a hair dye. What is the relative molecular mass of hydrogen peroxide?

What information is presented in the question?

The molecular formula of hydrogen peroxide, H_2O_2.

What is the question asking us to do?

Calculate the relative molecular mass of hydrogen peroxide.

What strategies do we need in order to answer the question?

1. Determine the relative atomic mass of each element present in the molecule.
2. Sum the relative atomic masses of all atoms present in the molecule.

Answer

Element	Number of atoms of the element	Relative atomic mass
H	2	1.0
O	2	16.0

The formula for relative atomic mass is:

$M_r(H_2O_2) = (2 \times 1.0) + (2 \times 16.0)$

$M_r(H_2O_2) = 34.0$ (no units)

How can the relative mass of larger compounds be calculated?

As we learned in lesson 4A, ionic compounds don't exist as molecules, but rather as continuous crystal lattices. In this case, it is inaccurate to use the term relative molecular mass. Instead, we use the term **relative formula mass (M_r)** to illustrate that we are finding the relative mass based on the formula of the compound. However, it is calculated in the same way as the relative molecular mass.

For example, consider the ionic compound lithium chloride, LiCl. As shown in figure 3, the lattice structure of lithium chloride is continuous, and there are many lithium and chloride ions present. Therefore, we cannot find the relative molecular mass (M_r) of lithium chloride, but we can calculate the relative formula mass (M_r) as shown.

$M_r(LiCl) = (1 \times 6.9) + (1 \times 35.5)$

$M_r(LiCl) = 42.4$ (no units)

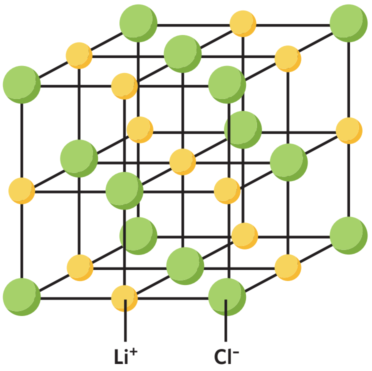

Figure 3 The simplest ratio of ions in the lattice structure of LiCl

WORKED EXAMPLE 4

Calculate the relative formula mass of calcium phosphide, Ca_3P_2.

What information is presented in the question?

The relative ionic formula of calcium phosphide, Ca_3P_2.

What is the question asking us to do?

Calculate the relative formula mass of calcium phosphide.

What strategies do we need in order to answer the question?

1. Determine the relative atomic mass of each element present in the compound.
2. Sum the relative atomic masses of all atoms present in the relative ionic formula of the compound.

Answer

Element	Number of atoms of the element	Relative atomic mass
Ca	3	40.1
P	2	31.0

The formula for relative atomic mass is:

$M_r(Ca_3P_2) = (3 \times 40.1) + (2 \times 31.0)$

$M_r(Ca_3P_2) = 182.3$ (no units)

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Progress questions

Question 4

Carbon dioxide, CO_2, and oxygen, O_2, are both molecules. Therefore, the relative molecular mass of both of these molecules

A. is the same.

B. is different.

Question 5

The relative molecular mass of a molecule is

A. the sum of the total number of protons and neutrons found in the molecule.

B. the sum of the relative atomic masses of each atom in the molecule.

Question 6

What is the symbol used to represent relative molecular and relative formula mass?

A. M_r

B. M

Question 7

What is the relative molecular mass of Cl_2?

A. 35.5

B. 71.0

Mass spectrometry 1.2.2.1

Mass spectrometry is an experimental technique used to determine the relative atomic mass of elements.

How can the relative atomic mass of an element be determined experimentally?

So far, we have learned that the isotopes of elements exist in different proportions in nature. We can measure the relative abundance of isotopes using an analytical technique known as mass spectrometry.[1] In a mass spectrometer – the instrument used to carry out mass spectrometry – samples are converted into cations which then move through the instrument. They move in a particular way depending on the relationship between the mass of the sample and its charge, also known as the **mass–to–charge ratio (m/z)**. The results of a mass spectrometry experiment are visualised on a mass spectrum, as shown in figure 4.

KEEN TO INVESTIGATE?

[1] **How does a mass spectrometer work?**
Search YouTube:
2.2 The Mass Spectrometer

USEFUL TIP

It's commonplace to see m/e used to represent mass-to-charge ratio, instead of m/z.

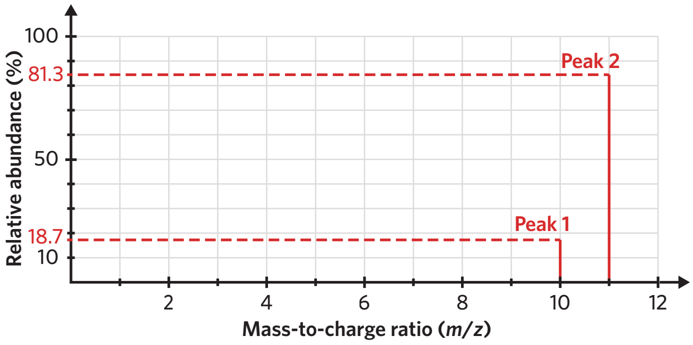

Figure 4 The mass spectrum of boron

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the spectrum in figure 4 corresponds to an isotope of boron. The height of a peak indicates how abundant the isotope is as a percentage of the total mass of the sample, whilst the horizontal axis indicates the mass-to-charge ratio. Given each isotope of an element has the same charge, the mass-to-charge ratio of each isotope can be taken as the relative isotopic mass of each isotope. Using the mass spectrum shown in figure 4, we can calculate the A_r of boron as follows.

$$A_r = \frac{(\text{relative abundance \%} \times \text{relative isotopic mass}) + (\text{relative abundance \%} \times \text{relative isotopic mass})}{100\%}$$

$$A_r = \frac{(18.7\% \times 10) + (81.3\% \times 11)}{100\%}$$

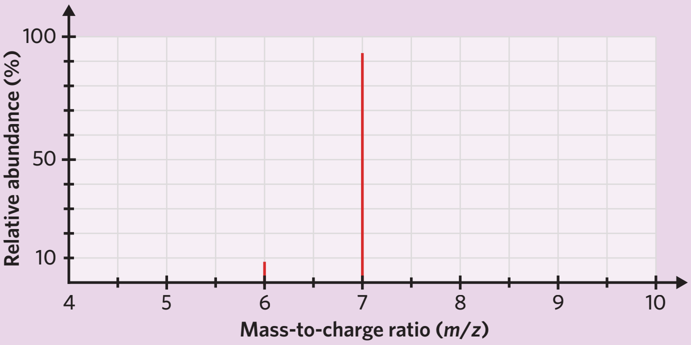
$$A_r = 10.8 \text{ (no units)}$$

USEFUL TIP

In the case of elements with two isotopes, we can quickly determine which of the two isotopes has the higher relative abundance. If the A_r of an element is closer to one isotope than the other, the isotope with a relative isotopic mass closest to the A_r of the element has a higher relative abundance.

WORKED EXAMPLE 5

A sample of a single monoatomic element was analysed using mass spectrometry to produce the spectrum given. Using the results, determine which element was analysed.



Mass–to–charge ratio (*m*/*z*)	Relative abundance (%)
6.02	7.5
7.02	92.5

What information is presented in the question?

A mass spectrum of a single monatomic element that produced two peaks.

A table containing the mass–to–charge ratios and relative abundance of each isotope.

What is the question asking us to do?

Determine the identity of the element that was analysed.

What strategies do we need in order to answer the question?

1. Substitute the mass–to–charge ratio and relative abundance of each isotope into the formula for the relative atomic mass of an element.
2. Use the relative atomic mass and the periodic table to determine which element was analysed.

Answer

$$A_r = \frac{(\text{relative abundance \%} \times \text{relative isotopic mass}) + (\text{relative abundance \%} \times \text{relative isotopic mass})}{100\%}$$

$$A_r = \frac{(7.5\% \times 6.02) + (92.5\% \times 7.02)}{100\%}$$

$$A_r = 6.9 \text{ (no units)}$$

Therefore, the relative atomic mass of the unknown element is 6.9. Using the periodic table, the element with a relative atomic mass of 6.9 is lithium and therefore, it can be concluded that lithium was analysed.

How can mass spectra be used to calculate the relative abundance of isotopes?

Now that we know how to interpret the information presented on a mass spectrum, we can use this understanding to determine the relative abundance of different isotopes. Depending on the information given, there are a couple of different ways to do this.

Consider the mass spectrum shown in figure 5. Notice that there is no scale for the y–axis. However, we know that the peaks in the spectrum are relative to each other, and together add up to a total of 100%.

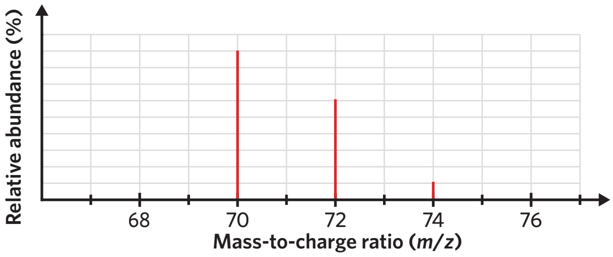


Figure 5 Mass spectrum of an unknown element

Therefore, we can use a ruler to determine the length of each peak as shown in figure 6, and use these values to determine the relative abundance of each isotope.

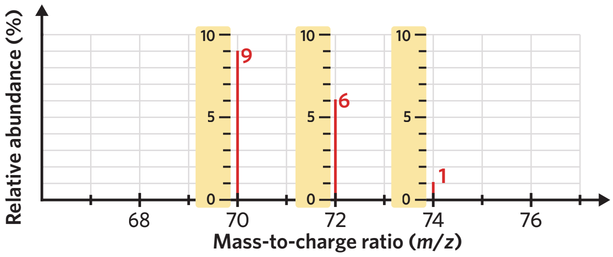
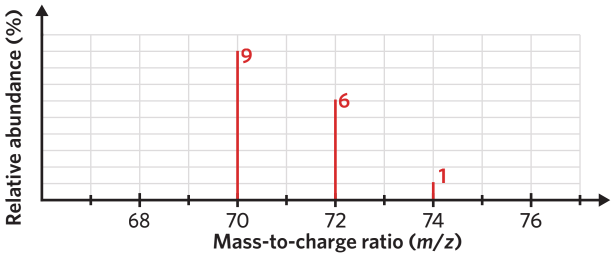


Figure 6 Using a ruler to measure the peaks in a mass spectrum

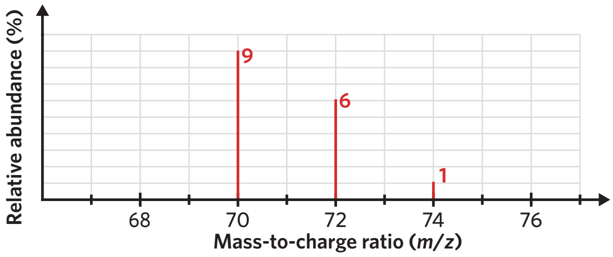
As these peaks are a percentage of the total, we first need to determine the total peak height – the sum of each individual peak height – as shown in figure 7.



Total peak height 9 + 6 + 1 = 16

Figure 7 Calculation of the total peak height in a mass spectrum

Once we have calculated the total height, we can calculate the relative abundance of the isotope represented by each peak, as shown in figure 8.



% abundance of peak 70 = $\frac{9}{16} \times 100\% = 56.25\%$

% abundance of peak 72 = $\frac{6}{16} \times 100\% = 37.50\%$

% abundance of peak 74 = $\frac{1}{16} \times 100\% = 6.25\%$

Figure 8 Calculation of the percentage abundance of each peak

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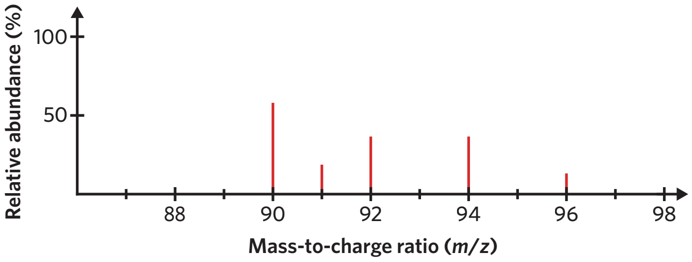
Progress questions

Use the following mass spectrum of Zirconium, Zr, to answer questions 8–10.



Question 8

According to the mass spectrum, Zr has

A. 96 isotopes.

B. 5 isotopes.

Question 9

According to the mass spectrum, the least abundant isotope of Zr is

A. ^{90}Zr.

B. ^{96}Zr.

Question 10

According to the mass spectrum, the relative atomic mass of Zr would be closest to

A. 90.

B. 95.

Theory summary

- The mass of every atom, molecule, and compound is measured relative to the mass of a single carbon–12 atom.
- The relative isotopic mass refers to the mass of a single atom of a particular isotope relative to the mass of a single carbon–12 atom.
- The relative abundance of an isotope refers to the percentage of that isotope in a naturally occurring sample of an element.
- The mass of each element on the periodic table is referred to as the relative atomic mass (A_r).
- The relative atomic mass is calculated by taking into account the relative abundance and relative isotopic masses of each isotope of an element.
- Relative masses have no units.
- $A_r = \frac{(\text{relative abundance \%} \times \text{relative isotopic mass}) + (\text{relative abundance \%} \times \text{relative isotopic mass})}{100\%}$
- A molecular formula gives information about all the elements and the number of atoms of each in one molecule of a substance.
- The relative molecular mass (M_r) is calculated by summing all of the relative atomic masses for each atom present in a single molecule.
- The relative formula mass (M_r) is calculated by summing all of the relative atomic masses of each atom in the formula of a compound.
- Mass spectrometry is used to determine the relative abundance of different isotopes of an element.
- The relative abundance of isotopes can be calculated using mass spectra by measuring and comparing the lengths of the peaks present.

The content in this lesson is considered fundamental prior knowledge to instrumental analysis of organic compounds (Unit 4 AOS 2).

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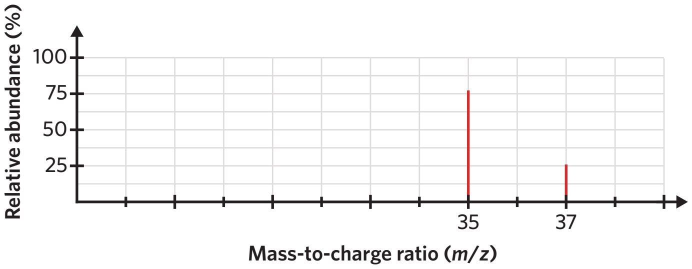
6A Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 11–13.

Mass spectroscopy is used to determine the mass of isotopes.



Question 11 (1 MARK) 🌶

Based on the mass spectrum, the element in question exists in how many forms?

A. 35

B. 37

C. 1

D. 2

Question 12 (1 MARK) 🌶🌶

Which of the following formulas would be used to calculate the relative atomic mass of the element?

A. $\frac{(35 \times 25\%) + (37 \times 75\%)}{100\%}$

B. $\frac{(35 \times 37\%) + (25 \times 75\%)}{100\%}$

C. $\frac{(35 \times 75\%) + (37 \times 25\%)}{100\%}$

D. $\frac{(100\% \times 25) + (37 \times 75\%)}{35\%}$

Question 13 (2 MARKS) 🌶🌶

Use the information in the mass spectrum to determine the identity of the element it represents. Justify your answer.

Exam-style

Question 14 (3 MARKS) 🌶🌶

An element contains three different isotopes. The relative abundances of two of the isotopes are shown in the table.

Relative isotopic mass	Relative abundance (%)
19.99	90.92
20.99	
21.99	8.82

a. What is the relative abundance of the isotope with a relative isotopic mass of 20.99? (1 MARK)

b. Calculate the relative atomic mass of the element. (2 MARKS)

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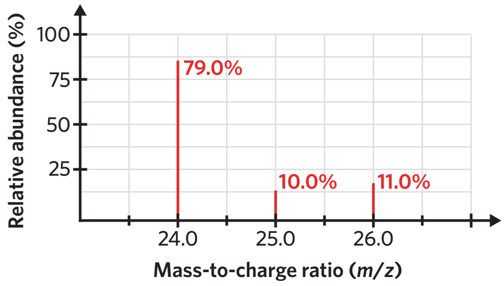
(3 MARKS)

A sample of an unknown element was placed into a mass spectrometer and the mass spectrum showed a total of three peaks.

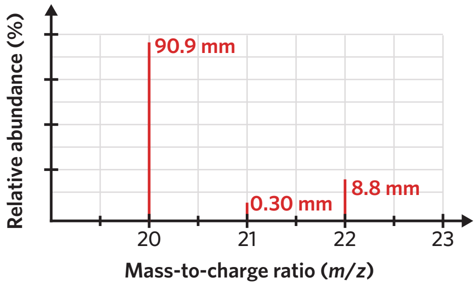


a. Calculate the relative atomic mass of the element. (2 MARKS)

b. What is the name of the element in the sample? (1 MARK)

Question 16 (6 MARKS)

Students from a university lab came across a sample containing an unknown element. To determine the nature of the unknown element, the students performed mass spectrometry and obtained the mass spectrum provided.



Unfortunately, the program responsible for the development of the mass spectrum was not set to include the percentage abundance of each isotope. Instead, students measured the length of the peaks using a ruler and included this as part of the spectrum.

a. Identify the number of isotopes for the unknown element. (1 MARK)

b. Calculate the relative abundance of each isotope. (3 MARKS)

c. Identify the name of the unknown element. (2 MARKS)

Question 17 (8 MARKS)

Lutetium is used in radiation to treat cancer and is known to have two isotopes, ^{175}Lu and ^{176}Lu. When researching Lu, students were unable to find the relative abundance of each isotope however they did find that Lu has an A_r of 174.97 according to the periodic table.

a. Based on the relative atomic mass, which isotope would most likely be the most abundant? (1 MARK)

b. Strangely, they found that the isotopic masses of ^{175}Lu and ^{176}Lu were 174.94 and 175.94 respectively. Calculate the relative abundance of each isotope. (3 MARKS)

c. Draw a mass spectrum for Lu. (3 MARKS)

d. Lutetium (III) can form a crystal structure with bromide ions to form $LuBr_3$. What is the relative formula mass of this compound? (1 MARK)

Question 18 (5 MARKS)

Rafa was on the verge of discovering a new element. To determine the nature of the element, he conducted multiple rounds of mass spectrometry on his sample. Rafa's results are shown in the three tables provided.

Test 1

Characteristic	**Data**
Number of peaks	2
m/*z* ratio	71 and 73
Relative abundance of each peak	24.5% for peak 71 and 75.5% for peak 73

Test 2

Characteristic	**Data**
Number of peaks	2
m/*z* ratio	71 and 73
Relative abundance of each peak	73.2% for peak 71 and 27.8% for peak 73

Test 3

Characteristic	**Data**
Number of peaks	2
m/*z* ratio	71 and 73
Relative abundance of each peak	24.9% for peak 71 and 75.1% for peak 73

a. If Rafa stopped the experiment after test two, comment on the repeatability of the experiment. (3 MARKS)

b. Calculate the relative abundance for each peak. (2 MARKS)

FROM LESSON 16D

Questions from multiple lessons

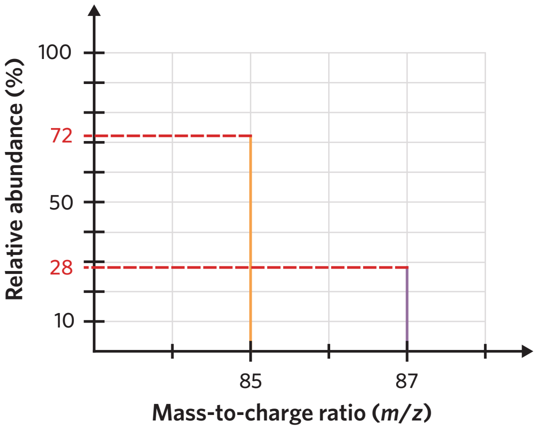
Question 19 (4 MARKS)

Rubidium is an extremely reactive element and must therefore be stored in very strict conditions.

a. In which block of the periodic table is rubidium found? (1 MARK)

b. What is the electron configuration of rubidium's outermost subshell? (1 MARK)

The mass spectrum of rubidium is shown.



c. Calculate the relative atomic mass of rubidium. (2 MARKS)

FROM LESSON 1B

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(6 MARKS) 🌶🌶🌶

Consider the following element:

29
Cu
Copper
63.5

a. How many protons does Cu have in its nucleus? (1 MARK)

b. A given atom of Cu is known to have 35 neutrons in its nucleus. What would be the mass number for this atm of Cu? (1 MARK)

c. Cu is known to exist as two isotopes: ^{63}Cu and ^{65}Cu. Calculate the relative abundance of each isotope. (3 MARKS)

d. Copper is able to bond with oxygen to form CuO. What is the relative formula mass? (1 MARK)

FROM LESSON 1A

Hints

14a. Relative abundance of all isotopes of an element adds up to 100%.

14b. Relative atomic mass takes into account the relative isotopic mass of each isotope and its relative abundance.

15a. Mass/charge ratio of each peak represents the relative mass of the isotope.

15b. Use the periodic table.

16a. A mass spectrum shows one peak per isotope of an element.

16b. Peak heights can be used as a point of comparison.

16c. Use the periodic table.

17a. Relative atomic mass of an element is skewed towards the most abundant isotope.

17b. Relative abundance of all isotopes of an element adds up to 100%.

17c. The height of each peak in a mass spectrum represents the relative abundance of an isotope.

17d. Relative formula mass is the sum of the relative atomic masses in the ionic formula.

18a. Repeatability of experiments depends on the similarity of the experimental conditions.

18b. Not all data obtained should be included when calculating values.

19a. The periodic table is arranged in blocks based on the subshell notation.

19b. Blocks in which elements are found represent their outermost subshell.

19c. Relative atomic mass takes into account the relative isotopic mass of each isotope and its relative abundance.

20a. The atomic number of an element indicates the number of protons in the nucleus.

20b. Mass number depends on the number of nucleons.

20c. Relative abundance of isotopes can be calculated from the relative atomic mass.

20d. Relative formula and molecular masses are calculated in the same way.

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6B Avogadro's constant and the mole

STUDY DESIGN DOT POINT

- Avogadro's constant as the number 6.02×10^{23} indicating the number of atoms or molecules in a mole of any substance; determination of the amount, in moles, of atoms (or molecules) in a pure sample of known mass

6A 6B 6C

1.2.3.1 & 1.2.3.4 The mole & relationship between number and amount of particles

1.2.3.2 Molar mass

1.2.3.3 Mass calculations involving the mole

ESSENTIAL PRIOR KNOWLEDGE

2A Molecular formulas

6A Carbon-12

See questions 33–34.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: Ekaterina_Minaeva/Shutterstock.com

Why is the 23rd of October international mole day?

The 23rd of October was chosen to be international **mole** day based on the value of **Avogadro's constant.** In this lesson, we will explore how the mole concept is related to the mass and number of **entities** in a substance.

KEY TERMS AND DEFINITIONS

Avogadro's constant (N_A)	the number of atoms in exactly 12 g of ^{12}C, 6.02×10^{23} mol^{-1}
Entities	atoms, molecules, compounds, ions, electrons etc.
Molar mass (*M*)	the mass, in grams, per mol of substance (g mol^{-1})
Mole (*n*)	amount of substance, in mol

The mole & relationship between number and amount of particles 1.2.3.1 & 1.2.3.4

The mole can allow us to express large quantities in a more accessible way.

Why do we use the mole in chemistry?

As we learned in lesson 6A, Carbon-12 serves as a reference point for the relative atomic masses of other elements. Experimentation has shown that 12 g of ^{12}C contains approximately 6.02×10^{23} ^{12}C atoms. This value is known as Avogadro's constant (N_A), and it defines the value of one mole of substance. Given Avogadro's constant is so large, the mole concept was created to make quantities of atoms and molecules more accessible.[1] One mole of a substance – denoted by the symbol n – is equal to 6.02×10^{23} atoms or molecules of the substance. The relationship between the mole (expressed in the unit mol) and Avogadro's constant is shown in figure 1.

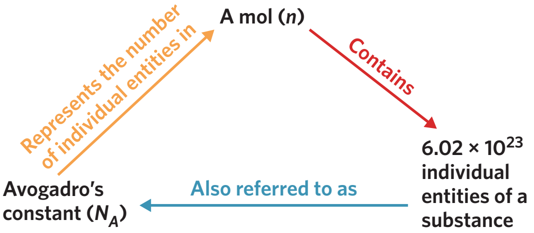

Figure 1 Relationship between the mole and Avogadro's constant

USEFUL TIP

One mol of any atom or molecule will always contain the same number of individual atoms or molecules.

KEEN TO INVESTIGATE?

[1] **How big is one mole?**
Search YouTube: How big is a mole?
(Not the animal, the other one.)

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we can deduce that a greater number of mol of substance will contain a greater number of particles, and a smaller number of mol of substance will contain fewer particles. Therefore, we can condense the relationship between Avogadro's constant and the mole into the formula shown in figure 2.

USEFUL TIP

As shown in the VCE Data Book, *L* can also be used to express Avogadro's constant. However, in this textbook, Avogadro's constant will always be expressed as N_A.

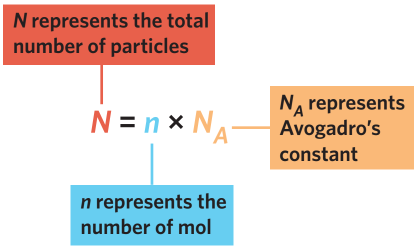

Figure 2 Formula for the total number of particles (*N*) in a sample

WORKED EXAMPLE 1

How many atoms of copper are present in 1.5 mol of copper?

What information is presented in the question?

The amount, in mol, of copper.

What is the question asking us to do?

Calculate the number of copper atoms present.

What strategies do we need in order to answer the question?

1. Substitute the known values into the equation.
2. Calculate the number of copper atoms present.

Answer

$n(\text{Cu}) = 1.5 \text{ mol}$

$N_\text{A} = 6.02 \times 10^{23} \text{ mol}^{-1}$

$N(\text{Cu}) = n(\text{Cu}) \times N_\text{A}$

$N(\text{Cu}) = 1.5 \text{ mol} \times (6.02 \times 10^{23} \text{ mol}^{-1})$

$N(\text{Cu}) = 9.03 \times 10^{23}$

Therefore, there are 9.0×10^{23} copper atoms present in 1.5 mol of copper.

Given the mole represents a specific quantity of particles, we can work backwards to find the amount of mol of a substance when we know the number of particles in a sample. The amount of mol can be calculated using the formula shown in figure 3.

USEFUL TIP

Avogadro's constant, $N_A = 6.02 \times 10^{23}$ mol^{-1}, is a proportionality constant between the quantity of substance (*n*, with the unit 'mol') and the quantity for counting entities (*N*).

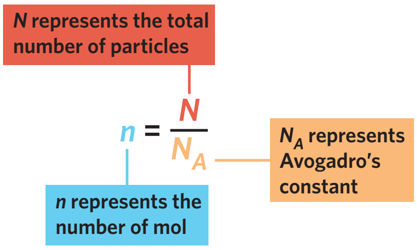

Figure 3 Formula for the amount in mol (*n*) of a substance

WORKED EXAMPLE 2

In a chemical reaction, scientists determined that 4.5×10^{15} molecules of water (H_2O) were produced. What amount, in mol, of water was produced?

What information is presented in the question?

The number of water molecules.

What is the question asking us to do?

Calculate the amount, in mol, of water produced.

What strategies do we need in order to answer the question?

1. Substitute the known values into the equation.
2. Calculate the amount of water produced, in mol.

Answer

$N(\text{H}_2\text{O}) = 4.5 \times 10^{15}$

$N_\text{A} = 6.02 \times 10^{23} \text{ mol}^{-1}$

$$n(\text{H}_2\text{O}) = \frac{N(\text{H}_2\text{O})}{N_A}$$

$$n(\text{H}_2\text{O}) = \frac{4.5 \times 10^{15}}{6.02 \times 10^{23} \text{ mol}^{-1}}$$

$n(\text{H}_2\text{O}) = 7.5 \times 10^{-9} \text{ mol}$

Therefore, 7.5×10^{-9} mol of water was produced.

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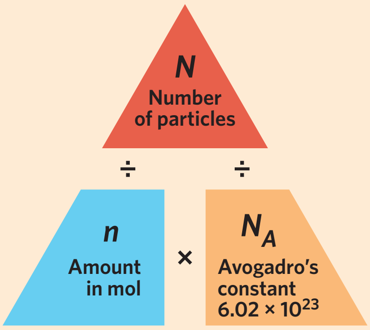
When doing calculations involving the number of particles, the amount in mol, and Avogadro's constant, we can use the following concept triangle to work out what formula to use.



To use this triangle, cover up the value you want to find with your finger, and calculate using the remaining variables and sign. For example, to find the amount in mol (*n*), cover the 'amount' section of the triangle and calculate $\frac{N}{N_A}$.

MISCONCEPTION

'CO_2 and CO contain the same number of moles of oxygen.'

The amount of a given element in a compound depends on the number of atoms of that element in the compound. Therefore, 1 mol of CO_2 molecules would contain 2 mol of oxygen atoms, whilst 1 mol of CO would only contain 1 mol of oxygen atoms.

USEFUL TIP

Relative masses and molar masses have the same value for an element or compound, however, molar mass has units. For example, the relative formula mass of limestone, $CaCO_3$, is 100.1, whilst its molar mass is 100.1 g mol^{-1}.

As we learned in lesson 2A, molecular formulas give us information about the type and number of atoms in a molecule. We can apply the mole concept to molecular formulas, such as that of water. The molecular formula for water, H_2O, tells us that in every water molecule, there are two hydrogen atoms and one oxygen atom. Therefore, we would say that in 1 mol of water molecules, there are 2 mol of hydrogen atoms and 1 mol of oxygen atoms.

As shown in table 1, we can use the information given in molecular formulas to determine the amount, in mol, of each element present in a molecule.

Table 1 The relationship between the mole concept and molecular formulas

Substance	Number of moles of the compound	Number of moles of each element
Water H_2O	1	Hydrogen: 2 mol Oxygen: 1 mol
Oxygen O_2	1	Oxygen: 2 mol
Carbon dioxide CO_2	1	Carbon: 1 mol Oxygen: 2 mol

Progress questions

Question 1

The amount of particles in one mole is equivalent to

A. Avogadro's constant.

B. the mass of 1 atom of the ^{12}C isotope.

Question 2

If one mole of X contains 6.02×10^{23} particles, then two moles of X would contain

A. 6.02×10^{23} particles.

B. 1.20×10^{24} particles.

Question 3

One mole of CO_2 and one mole of NH_3 would contain the same number of

A. atoms.

B. molecules.

Question 4

In 1 mol of fluorine gas, F_2, there is 0.5 mol of fluorine atoms.

A. True

B. False

Molar mass 1.2.3.2

Each substance has a unique **molar mass**.

How can the molar mass of an element or compound be calculated?

The molar mass (*M*) is the mass, in grams, of one mole of an atom or molecule. As we know, one mole of ^{12}C atoms has a mass of 12.0 g. Therefore, we can say that ^{12}C has a molar mass of 12.0 g mol^{-1}.

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s of any element is taken to be the relative atomic mass given on the periodic table. Therefore, we can use the molar mass of individual atoms of elements to calculate the molar mass of larger molecules. Consider the molecular formula for ammonia, NH_3. We can calculate the molar mass of ammonia as follows.

$$M(NH_3) = M(N) + (3 \times M(H))$$

$$M(NH_3) = 14.0 \text{ g mol}^{-1} + \left(3 \times 1.0 \text{ g mol}^{-1}\right)$$

$$M(NH_3) = 17.0 \text{ g mol}^{-1}$$

USEFUL TIP

When calculating the molar mass of compounds involving brackets, such as $Ca(OH)_2$, the subscript outside of the brackets represents a multiple of what's inside the brackets. For example, $(OH)_2$ represents 2 × O and 2 × H.

WORKED EXAMPLE 3

Nitric oxide, NO, is produced in the body to relax the blood vessels. Calculate the molar mass of nitric oxide.

What information is presented in the question?

The formula for nitric oxide, NO.

What is the question asking us to do?

Calculate the molar mass of NO.

What strategies do we need in order to answer the question?

1. Use the periodic table to determine the relative atomic mass of each element, N and O.
2. Determine the molar mass of each element.
3. Add the molar mass of both elements together.

Answer

According to the periodic table, the relative atomic masses of N and O are 14.0 and 16.0 respectively. Therefore, the molar masses of N and O are 14.0 g mol^{-1} and 16.0 g mol^{-1} respectively.

$$M(NO) = M(N) + M(O)$$

$$M(NO) = 14.0 \text{ g mol}^{-1} + 16.0 \text{ g mol}^{-1}$$

$$M(NO) = 30.0 \text{ g mol}^{-1}$$

Therefore, the molar mass of NO is 30.0 g mol^{-1}.

WORKED EXAMPLE 4

Copper (II) nitrate, $Cu(NO_3)_2$, is sometimes used as a pigment for glass and ceramics. Calculate the molar mass of copper (II) nitrate.

What information is presented in the question?

The formula for copper (II) nitrate, $Cu(NO_3)_2$.

What is the question asking us to do?

Calculate the molar mass of $Cu(NO_3)_2$.

What strategies do we need in order to answer the question?

1. Use the periodic table to determine the relative atomic mass of each element, Cu, N and O.
2. Determine the molar mass of each element.
3. Add the molar masses of all the elements together.

Answer

According to the periodic table, the relative atomic masses of Cu, N and O are 63.5, 14.0 and 16.0 respectively. Therefore, the molar masses of Cu, N and O are 63.5 g mol^{-1}, 14.0 g mol^{-1} and 16.0 g mol^{-1} respectively.

$$M(Cu(NO_3)_2) = M(Cu) + 2 \times (M(N) + (3 \times M(O)))$$

$$M(Cu(NO_3)_2) = 63.5 \text{ g mol}^{-1} + 2 \times (14.0 + (3 \times 16.0)) \text{ g mol}^{-1}$$

$$M(Cu(NO_3)_2) = 187.5 \text{ g mol}^{-1}$$

Hence, the molar mass of copper (II) nitrate is 187.5 g mol^{-1}.

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Progress questions

Question 5

The molar mass of an element is equivalent to

A. the mass of 6.02×10^{23} atoms of the element.

B. the mass of 12 atoms of the element.

Question 6

The numerical value for the molar mass of an element is equivalent to

A. the number of protons in the element.

B. the relative atomic mass of the element.

Question 7

The molar mass of BeO is

A. 25.0 g.

B. 25.0 g mol^{-1}.

Mass calculations involving the mole 1.2.3.3

We can use the molar mass of a substance to calculate the moles of substance present in a sample of known mass.

How can the mole concept be used in calculations involving mass?

So far, we have learned that we can calculate the molar mass of molecules using the individual molar masses of the atoms present. When given a sample of a substance of known mass, we can calculate the moles of the substance present in the sample using the molar mass of the substance, as shown in the formula in figure 4.

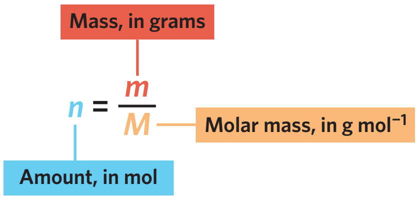

Figure 4 Formula for the amount (n) of a substance

WORKED EXAMPLE 5

A student was conducting an experiment that required 12.0 g of water, H_2O. Calculate the amount of water, in mol, required.

What information is presented in the question?

The mass of water is 12.0 g.

What is the question asking us to do?

Calculate the amount of water in moles.

What strategies do we need in order to answer the question?

1. Calculate the molar mass of water.
2. Use the formula $n = \frac{m}{M}$ to determine the amount of water required.

Answer

The molar masses of H and O are 1.0 g mol^{-1} and 16.0 g mol^{-1} respectively. Therefore,

$M(H_2O) = (2 \times M(H)) + M(O)$

$M(H_2O) = (2 \times 1.0)\ g\ mol^{-1} + 16.0\ g\ mol^{-1}$

$M(H_2O) = 18.0\ g\ mol^{-1}$

To calculate the amount:

$$n(H_2O) = \frac{m(H_2O)}{M(H_2O)}$$

$$n(H_2O) = \frac{12.0\ g}{18.0\ g\ mol^{-1}}$$

$n(H_2O) = 0.667\ mol$

Hence, 0.667 mol of water is required for the experiment.

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we may be required to calculate the mass of substance in a sample of known amount. We can do so using the formula shown in figure 5.

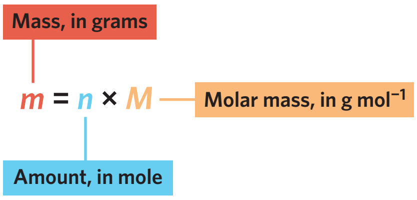


Figure 5 Formula for the mass (m) of a substance

WORKED EXAMPLE 6

Calculate the mass of a 1.66 mol sample of nitrogen dioxide, NO_2.

What information is presented in the question?

The amount of nitrogen dioxide, NO_2, which is 1.66 mol.

What is the question asking us to do?

Calculate the mass of nitrogen dioxide in a 1.66 mol sample.

What strategies do we need in order to answer the question?

1. Calculate the molar mass of nitrogen dioxide.
2. Use the formula $m = n \times M$ to calculate the mass of nitrogen dioxide.

Answer

The molar masses of N and O are 14.0 g mol^{-1} and 16.0 g mol^{-1} respectively. Therefore:

$M(NO_2) = M(N) + (2 \times M(O))$

$M(NO_2) = 14.0 \text{ g mol}^{-1} + (2 \times 16.0) \text{ g mol}^{-1}$

$M(NO_2) = 46.0 \text{ g mol}^{-1}$

To calculate the mass:

$m(NO_2) = n(NO_2) \times M(NO_2)$

$m(NO_2) = 1.66 \text{ mol} \times 46.0 \text{ g mol}^{-1}$

$m(NO_2) = 76.4 \text{ g}$

Hence, the mass of 1.66 mol of nitrogen dioxide, NO_2, is 76.4 g.

Progress questions

Question 8

Which of the following must be known to calculate the amount, in mol, of a substance?

A. The molar mass

B. Both the mass and molar mass

Question 9

The mass of a substance, in grams, will ___________ as the amount, in mol, of substance decreases.

A. increase

B. decrease

Question 10

If we are given the mass of a sample of SO_3^{2-} ions, we can calculate the amount, in mol, of SO_3^{2-} ions by

A. multiplying the mass by 80.1 g mol^{-1}.

B. dividing the mass by 80.1 g mol^{-1}.

Question 11

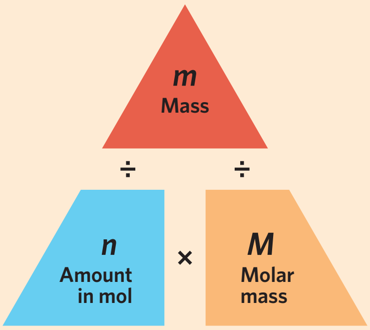
Consider the equation $\frac{N}{N_A} = \_\_\_\_\_\_\_\_ = \frac{m}{M}$. What is the missing symbol?

A. M

B. n

STRATEGY

When doing concentrations involving the mass, the amount in mol and the molar mass, we can use the following concept triangle to work out what formula to use.



To use this triangle, cover up the value you want to find with your finger, and calculate using the remaining variables and sign. For example, to find the amount, cover the 'amount' section of the triangle and calculate $\frac{m}{M}$.

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Theory summary

- Avogadro's constant is equal to the number of atoms found in exactly 12 g of ^{12}C.
  - $N_A = 6.02 \times 10^{23}\ mol^{-1}$
- One mole of any substance will always contain 6.02×10^{23} individual particles or entities.
- The mole (the amount of substance) is expressed in mol.
- Since the value of N_A is constant, we can calculate the amount (n) or number of particles (N) using variations of the equation $N = n \times N_A$.
- The molar mass is the mass, in grams, of one mole of a substance and has the units g mol^{-1}.
- The amount of substance in a sample of known mass and molar mass can be calculated using $n = \frac{m}{M}$.
- Given that the mole concept is a common variable in both the number of particles and mass, we can calculate the number of particles when given the mass and vice versa.

The content in this lesson is considered fundamental prior knowledge to carbon-based fuels (Unit 3 AOS 1).

6B Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 12–14.

Ammonia, NH_3, is found in many household cleaners. A particular bottle of household cleaner contains 6.11×10^{26} molecules of ammonia.

Question 12 (1 MARK) 🌶

The symbol N_A represents

A. the molar mass of a compound.

B. the number of protons in the nucleus.

C. 12 particles.

D. 6.02×10^{23} particles.

Question 13 (1 MARK) 🌶

When the amount of particles in a sample is known, the formula used to find the amount, in mol, present is

A. $N = n \times N_A$

B. $n = \frac{M}{m}$

C. $n = \frac{A_r}{M}$

D. $n = \frac{N}{N_A}$

Question 14 (2 MARKS) 🌶🌶

Calculate the mass, in grams, of ammonia (NH_3) present in the bottle of household cleaner.

Exam-style

Question 15 (3 MARKS) 🌶

Calculate the molar mass of each of the following compounds.

a. Hydroiodic acid, HI. (1 MARK)

b. Magnesium chloride, $MgCl_2$. (1 MARK)

c. Lactic acid, $C_3H_6O_3$. (1 MARK)

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(8 MARKS)

For each of the following comparisons, convert the relevant quantity(ies) into moles, and hence use the symbols > or < to identify the larger molar quantity.

a. 2.11 mol of iron (Fe) atoms __________ 3.33×10^{33} molecules of chlorine gas (Cl_2). (2 MARKS)

b. 5.17×10^{11} molecules of nitrogen dioxide (NO_2) __________ 2.98 mol of carbon tetrachloride (CCl). (2 MARKS)

c. 7.01×10^{9} molecules of carbon disulfide (CS_2) __________ number of hydrogen atoms in 2.70×10^{9} molecules of phosphine (PH_3). (4 MARKS)

Question 17 (4 MARKS)

In the process of converting fuel into energy, cars can produce carbon monoxide, CO. Carbon monoxide is a poisonous gas that can cause severe health complications.

a. If 201.3 mol of carbon monoxide was produced during a car trip, what mass, in kg, of carbon monoxide was produced? (1 MARK)

b. The manufacturers of two cars of similar size both claim that their cars have the lowest release of carbon monoxide into the atmosphere. Upon further testing, the results of both cars were recorded in the following table.

	Car A	**Car B**
Quantity of carbon monoxide released per 40 L tank of petrol	114 g	12.04×10^{33} molecules

Which car releases the least amount of CO for every tank of fuel? (3 MARKS)

Question 18 (6 MARKS)

Ethanol, C_2H_5OH, can be produced from a fermentation process involving glucose, $C_6H_{12}O_6$, and yeast.

$$C_6H_{12}O_6(aq) \xrightarrow{yeast} 2C_2H_5OH(aq) + 2CO_2(g)$$

a. Which of ethanol and glucose has the lower molar mass? Include calculations in your answer. (3 MARKS)

b. If 4.11 g of glucose was used to produce ethanol, calculate the amount, in mol, of glucose used. (1 MARK)

c. What mass of glucose needs to be added to the mass from part b such that a total of 3.00 mol of glucose is used in a fermentation reaction? (2 MARKS)

Question 19 (1 MARK)

A sample of the anticancer drug Taxol®, $C_{47}H_{51}NO_{14}$, contains 0.157 g of carbon. The mass, in grams, of oxygen in the sample is

A. 0.0468

B. 0.0624

C. 0.209

D. 0.703

VCAA 2012 Exam 1 Multiple choice Q15

Key science skills

Question 20 (4 MARKS)

Methane, CH_4, is a gaseous fuel that can be found trapped deep under the earth's surface from the decay of organic material. In order to extract methane, excavators need to drill deep into the ground. The gas is then collected in pipes and delivered to the appropriate area.

A group of miners wanted to see if the depth of drilling affected the amount of methane that can be collected. As such, they drilled holes of different depths in what was believed to be a methane-rich area.

a. Identify the dependent variable. (1 MARK)

b. Calculate the number of molecules present in 15.8 mol of methane gas. (1 MARK)

c. Identify the key environmental concern associated with the collection of methane. Explain your answer. (2 MARKS)

FROM LESSONS 16A & 16B

Question 21 (5 MARKS)

A cooking recipe displayed quantities of ingredients as shown.

1.22 mol of salt (NaCl)
0.98 mol of sugar ($C_6H_{12}O_6$)
0.11 mol of bicarbonate ($NaHCO_3$)

a. What mass of sugar was required for this recipe? (1 MARK)

b. Calculate the total mass of sugar and bicarbonate in the recipe. (2 MARKS)

c. Bicarbonate contains a metallic element. Identify the name of the metallic element and the period in which it would be found on the periodic table. (2 MARKS)

FROM LESSONS 1B & 3A

Question 22 (4 MARKS)

The black soot sometimes found on the outside of test tubes that have been placed over a bunsen burner flame is pure carbon.

a. Draw a Lewis dot diagram for carbon. (1 MARK)

b. Write the electronic configuration for carbon. (1 MARK)

c. A student wanted to heat 10 g of water in a 5 g test tube. During the experiment, she noticed some black substance forming on the test tube so decided to re-weigh the water-containing test tube.

Given that the mass of the test tube recorded was 17.41 g, calculate the amount, in mol, of carbon formed. (2 MARKS)

FROM LESSONS 1B & 2A

Question 23 (6 MARKS)

Argon is sometimes used in light bulbs to stop oxygen from corroding the filament inside the bulb.

a. Identify the atomic number for argon. (1 MARK)

b. Write the electron configuration for argon. (1 MARK)

c. Argon contains three different isotopes: ^{36}Ar, ^{38}Ar and ^{40}Ar. Which isotope of argon would be the most abundant? Explain. (3 MARKS)

d. In a 2.44 mol sample of argon, how many atoms of argon would be present? (1 MARK)

FROM LESSONS 1B & 6A

Hints

15a. The molar mass of each element present in the molecule can be determined from the periodic table.
15b. The molar mass of each element present in the compound can be determined from the periodic table.
15c. The molar mass of each element present in the molecule can be determined from the periodic table.
16a. Use the formula $n = \frac{N}{N_A}$ to convert the quantity of chlorine gas to moles.
16b. Use the formula $n = \frac{N}{N_A}$ to convert the quantity of nitrogen dioxide to moles.
16c. There are 3 moles of hydrogen atoms per mole of phosphine.
17a. The mass of carbon monoxide can be calculated using the formula $m = n \times M$.
17b. Convert both quantities to mol.
18a. The molar mass of each element present in each molecule can be determined from the periodic table.
18b. Use the formula $n = \frac{m}{M}$.
18c. Subtract your answer from part b from the total amount to be used.
19. The amount, in mol, of Taxol® present in the sample is equal to the amount, in mol, of carbon present in the sample divided by the number of carbon atoms in a molecule of Taxol®.
20a. The dependent variable is the variable being measured.
20b. Use the formula $N = n \times N_A$.
20c. Consider the effect of drilling on the environment.
21a. Use the formula $m = n \times M$.
21b. Use the formula $m = n \times M$.
21c. On the periodic table, periods are the horizontal rows.
22a. Carbon has four valence electrons.
22b. Carbon is located in the p-block of the periodic table.
22c. The mass of carbon formed is equal to the difference between the final mass and the original mass.
23a. The atomic number can be determined from the periodic table.
23b. Argon is located in the p-block of the periodic table.
23c. The relative atomic mass is nearest to the most abundant isotope.
23d. Use the formula $N = n \times N_A$.

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6C Calculations involving mass


Image: Alexandr Grant/Shutterstock.com

How can we use mass to determine the composition of compounds?

Ammonium dichromate,[1] $(NH_4)_2Cr_2O_7$, is an interesting and dangerous compound, where one mole contains four different elements and nineteen moles of elements in total. In this lesson, we will learn about the relationship between the mole concept and the development of chemical formulas.

STUDY DESIGN DOT POINT

- determination of the molar mass of compounds, the percentage composition by mass of covalent compounds, and the empirical and molecular formula of a compound from its percentage composition by mass

ESSENTIAL PRIOR KNOWLEDGE

6B The mole concept

6B Calculations involving the mole

See questions 35–36.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

KEEN TO INVESTIGATE?

[1] **What is the volcano reaction?**
Search YouTube:
Ammonium Dichromate Volcano – Cool Chemical Experiment

KEY TERMS AND DEFINITIONS

Empirical formula chemical formula depicting the lowest whole number ratio of atoms of different elements in a compound

Molecular formula actual number of atoms in a molecule

Percentage composition percentage by mass of an element in a compound

Ratio numerical relationship between the amounts of two or more elements

Determining empirical formulas 1.2.4.1

Empirical formulas show the lowest whole number **ratio** of atoms present.

How do we calculate percentage composition?

Compounds are chemical substances that consist of atoms of two or more different elements. Since compounds contain different quantities and ratios of each element, chemists have created a number of different ways to represent the names, structures, and formulas of molecules.

The composition of a compound can be expressed by looking at the percentage of mass that each element present contributes. This is known as **percentage composition**, and it shows the proportion of each element present, expressed as a percentage with respect to the mass of the whole compound.

In order to calculate the percentage by mass of an element in a compound, the following formula is used.

% by mass of an element in a compound

$$= \frac{\text{mass of element in 1 mol of the compound}}{\text{molar mass of the compound}} \times 100\%$$

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For example, in the molecule methane (CH_4), there is one mol of carbon atoms and therefore, the mass of carbon in the molecule will be:

$m(C) = n(C) \times M(C)$

$m(C) = 1\ mol \times 12.0\ g\ mol^{-1}$

$m(C) = 12.0\ g$

The molar mass of the compound is:

$M(CH_4) = (1 \times 12.0\ g\ mol^{-1}) + (4 \times 1.0\ g\ mol^{-1})$

$M(CH_4) = 16.0\ g\ mol^{-1}$

Using the formula above, calculate the % by mass of carbon in methane:

$$\%\ \text{by mass of carbon in methane} = \frac{12.0\ g\ mol^{-1}}{16.0\ g\ mol^{-1}} \times 100\%$$

$\%$ by mass of carbon in methane $= 75.0\%$

In one mol of methane (CH_4), there are four moles of hydrogen atoms and therefore, the mass will be:

$m(H) = n(H) \times M(H)$

$m(H) = 4\ mol \times 1.0\ g\ mol^{-1}$

$m(H) = 4.0\ g$

$$\%\ \text{by mass of hydrogen in methane} = \frac{4.0\ g\ mol^{-1}}{16.0\ g\ mol^{-1}} \times 100\%$$

$\%$ by mass of hydrogen in methane $= 25\%$

STRATEGY

When calculating the percentage composition of a compound, check your answer by ensuring all the individual percentages by mass of elements add up to 100% $\pm$ 0.1% (error due to resolution of masses used for calculations).

WORKED EXAMPLE 1

Calculate the percentage by mass of each element in the compound H_2SO_4.

What information is presented in the question?

The compound is H_2SO_4.

What is the question asking us to do?

Determine the percentage by mass of hydrogen, sulfur, and oxygen in the compound.

What strategies do we need in order to answer the question?

1. Determine the mass of hydrogen, sulfur, and oxygen in one mol of the compound.
2. Determine the molar mass of the compound.
3. Use the percentage by mass formula to calculate the percentage composition of each element in the compound.

Answer

In one mol of H_2SO_4, there are two moles of hydrogen atoms. Therefore, the mass of hydrogen atoms present in one mol of the compound is:

$m(H) = n(H) \times M(H)$

$m(H) = 2\ mol \times 1.0\ g\ mol^{-1}$

$m(H) = 2.0\ g$

In one mol of H_2SO_4, there is one mol of sulfur atoms. Therefore, the mass of sulfur atoms present in one mol of the compound is:

$m(S) = n(S) \times M(S)$

$m(S) = 1\ mol \times 32.1\ g\ mol^{-1}$

$m(S) = 32.1\ g$

Continues →

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f H_2SO_4, there are four moles of oxygen atoms. Therefore, the mass of oxygen atoms present in one mol of the compound is:

$m(\text{O}) = n(\text{O}) \times M(\text{O})$

$m(\text{O}) = 4 \text{ mol} \times 16.0 \text{ g mol}^{-1}$

$m(\text{O}) = 64.0 \text{ g}$

The molar mass of the compound is:

$M(\text{H}_2\text{SO}_4) = (2 \times 1.0 \text{ g mol}^{-1}) + (1 \times 32.1 \text{ g mol}^{-1})$

$+ (4 \times 16.0 \text{ g mol}^{-1})$

$M(\text{H}_2\text{SO}_4) = 98.1 \text{ g mol}^{-1}$

$$\% \text{ by mass of hydrogen in H}_2\text{SO}_4 = \frac{2.0 \text{ g mol}^{-1}}{98.1 \text{ g mol}^{-1}} \times 100\%$$

$\% \text{ by mass of hydrogen in H}_2\text{SO}_4 = 2.0\%$

$$\% \text{ by mass of sulfur in H}_2\text{SO}_4 = \frac{32.1 \text{ g mol}^{-1}}{98.1 \text{ g mol}^{-1}} \times 100\%$$

$\% \text{ by mass of sulfur in H}_2\text{SO}_4 = 32.7\%$

$$\% \text{ by mass of oxygen in H}_2\text{SO}_4 = \frac{64.0 \text{ g mol}^{-1}}{98.1 \text{ g mol}^{-1}} \times 100\%$$

$\% \text{ by mass of oxygen in H}_2\text{SO}_4 = 65.2\%$

How do we determine and interpret empirical formulas?

Atoms of different elements in molecules exist in fixed whole number ratios. In chemistry, this ratio represents how many atoms of one element there are compared to one or more other elements. The lowest whole number form of this ratio is called the empirical formula, in which elements are represented by their chemical symbol found on the periodic table and their respective ratios with subscript notation. The fact that the ratio is in its lowest form means that any common factors have been cancelled out.

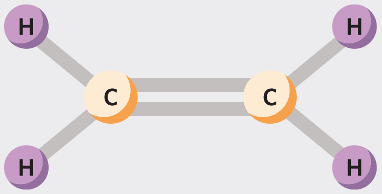
Name	Ethene
Ball-and-stick model	
Empirical formula	CH_2
Ratio	C : H 2 : 4 1 : 2

Figure 1 Structure and empirical formula of ethene

For example, if a compound has two carbon atoms and four hydrogen atoms as shown in figure 1, the empirical formula is CH_2, not C_2H_4. This is because the lowest whole number ratio is 1:2, not 2:4, as both the number of carbon atoms and the number of hydrogen atoms are divisible by two.

Table 1 contains the structures and different types of formulas of ethanoic acid and ethene.

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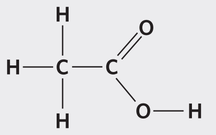
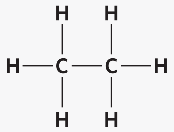
Table 1 Different compounds and their empirical formulas

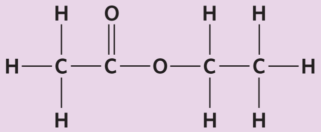
Name and structure	Formula based on the molecular structure	Lowest whole number ratio of elements in the compound	Empirical formula
Ethanoic acid 	$C_2H_4O_2$	C : H : O 1 : 2 : 1	CH_2O
Ethane 	C_2H_6	C : H 1 : 3	CH_3

USEFUL TIP

The structural formulas shown in table 1, and throughout the rest of the lesson, will be explored in detail in chapter 7. For now, just count the atoms that can be seen in the formula to determine the number of various atoms present.

WORKED EXAMPLE 2

What is the empirical formula of the molecule presented?



What information is presented in the question?

The diagram of the molecule.

What is the question asking us to do?

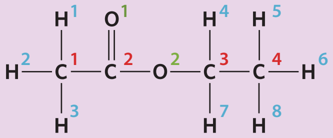
Determine the empirical formula of the compound which is the lowest whole number ratio of atoms in a compound.

What strategies do we need in order to answer the question?

1. Identify the elements present.
2. Count the number of atoms of each element present.
3. Express this as a ratio.
4. Determine if a common factor is present which means that the ratio can be further simplified.

Answer

From the structure, there are four carbon atoms, eight hydrogen atoms and two oxygen atoms.



Therefore, the actual ratio of atoms of different elements in the compound is:

C : H : O

4 : 8 : 2

2 is a common factor and therefore, the lowest whole number ratio is:

C : H : O

$\frac{4}{2} : \frac{8}{2} : \frac{2}{2}$

2 : 4 : 1

Therefore, the empirical formula is C_2H_4O.

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Since empirical formulas are the lowest whole number ratio of elements in a compound, different compounds can have the same empirical formula (e.g. C_5H_{10} and C_6H_{12} both have the empirical formula CH_2 as the lowest whole number ratio is 1:2).

Empirical formulas can also be determined from quantitative data. Different types of experiments can be carried out to determine the mass of each element present in a compound. Using the mole concept, we can convert masses into moles to determine the mole ratio of atoms of particular elements in a compound.

Let's take a look at an example. A chemist is investigating an oxide of nitrogen (a compound containing only nitrogen and oxygen) which consists of 30.4% nitrogen. She would like to determine the empirical formula of the compound. First of all, since we know that the percentage by mass of elements in a compound must add up to 100%, and the compound contains only nitrogen and oxygen, the percentage by mass of oxygen in the compound is 100% − 30.4% = 69.6%. Secondly, if a compound consists of 30.4% nitrogen by mass and 69.6% oxygen by mass, in 100 g of the compound, there would be 30.4 g of nitrogen and 69.6 g of oxygen. Therefore:

$$n(\mathrm{N}) = \frac{m(\mathrm{N})}{M(\mathrm{N})}$$

$$n(\mathrm{N}) = \frac{30.4\ \mathrm{g}}{14.0\ \mathrm{g\ mol^{-1}}}$$

$$n(\mathrm{N}) = 2.17\ \mathrm{mol}$$

$$n(\mathrm{O}) = \frac{m(\mathrm{O})}{M(\mathrm{O})}$$

$$n(\mathrm{O}) = \frac{69.6\ \mathrm{g}}{16.0\ \mathrm{g\ mol^{-1}}}$$

$$n(\mathrm{O}) = 4.35\ \mathrm{mol}$$

This means that for every 2.17 mol of nitrogen in the compound, there is 4.35 mol of oxygen. In order to convert this to the empirical formula, we divide both quantities by the lowest amount, which in this case is 2.17 mol.

$$\begin{array}{ccc} \mathrm{N} & : & \mathrm{O} \\ \frac{2.17\ \mathrm{mol}}{2.17\ \mathrm{mol}} & : & \frac{4.35\ \mathrm{mol}}{2.17\ \mathrm{mol}} \\ 1.00 & : & 2.00 \end{array}$$

Therefore, the empirical formula is NO_2 as the lowest whole number ratio of nitrogen to oxygen is 1:2. An overview of the process outlined above is shown in figure 2.

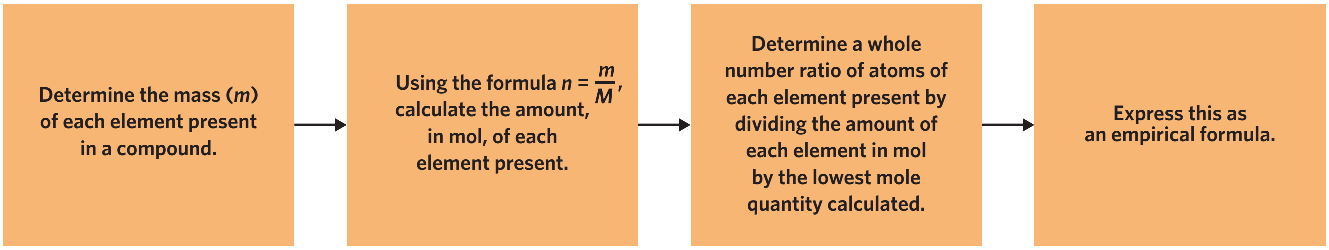

Figure 2 Flowchart of the steps required to determine the empirical formula

Sometimes the mole ratios calculated are not whole numbers. When this happens, we need to multiply the ratio by a factor to find the lowest whole number ratio between the elements.

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PLE 3

Determine the empirical formula of a compound with only carbon, hydrogen, and oxygen atoms, consisting of 47.37% carbon and 10.59% hydrogen.

What information is presented in the question?

The compound contains only carbon, hydrogen, and oxygen. It consists of 47.37% carbon and 10.59% hydrogen.

What is the question asking us to do?

To determine the empirical formula of the compound, which is the lowest whole number ratio of atoms in a compound.

What strategies do we need in order to answer the question?

1. Determine the mass of carbon, hydrogen, and oxygen present in 100 g of the compound.
2. Calculate the amount in mol of each element present in 100 g of the compound.
3. Divide each amount by the lowest mole quantity.
4. Determine the lowest whole number ratio and express this as an empirical formula.

Answer

Since the compound contains only carbon, hydrogen, and oxygen in 100 g of the compound, there will be 47.37 g of carbon, 10.59 g of hydrogen, and the mass of oxygen can be calculated accordingly:

$m(\text{O}) = 100\text{ g} - 47.37\text{ g} - 10.59\text{ g}$

$m(\text{O}) = 42.04\text{ g}$

$$\begin{array}{ccccc} \text{C} & : & \text{H} & : & \text{O} \\ 47.37\text{ g} & : & 10.59\text{ g} & : & 42.04\text{ g} \end{array}$$

Therefore,

$$n(\text{C}) = \frac{47.37\text{ g}}{12.0\text{ g mol}^{-1}}$$

$n(\text{C}) = 3.95\text{ mol}$

$$n(\text{H}) = \frac{10.59\text{ g}}{1.0\text{ g mol}^{-1}}$$

$n(\text{H}) = 10.59\text{ mol}$

$$n(\text{O}) = \frac{42.04\text{ g}}{16.0\text{ g mol}^{-1}}$$

$n(\text{O}) = 2.63\text{ mol}$

$$\begin{array}{ccccc} \text{C} & : & \text{H} & : & \text{O} \\ 3.95\text{ mol} & : & 10.59\text{ mol} & : & 2.63\text{ mol} \\ \dfrac{3.95\text{ mol}}{2.63\text{ mol}} & : & \dfrac{10.59\text{ mol}}{2.63\text{ mol}} & : & \dfrac{2.63\text{ mol}}{2.63\text{ mol}} \\ 1.50 & : & 4.03 & : & 1.00 \end{array}$$

Multiply each number by 2 to convert to a whole number ratio.

$$\begin{array}{ccccc} \text{C} & : & \text{H} & : & \text{O} \\ 3 & : & 8 & : & 2 \end{array}$$

Therefore, the empirical formula is $C_3H_8O_2$.

WORKED EXAMPLE 4

Polyethylene terephthalate (PET) is a compound found in plastic bottles, containing only carbon, hydrogen and oxygen atoms. 3.80 g of the compound was burnt in oxygen which produced 1.32 g of H_2O and 8.04 g of CO_2. What is the empirical formula of the compound?

What information is presented in the question?

The compound consists of only carbon, hydrogen, and oxygen.

Mass of compound burnt: 3.08 g

Mass of H_2O produced: 1.32 g

Mass of CO_2 produced: 8.04 g

What is the question asking us to do?

Determine the empirical formula of the compound.

What strategies do we need in order to answer the question?

1. Determine the amount, in mol, of hydrogen present in the compound.
2. Determine the amount, in mol, of carbon present in the compound.
3. Determine the mass of carbon and hydrogen in the compound.
4. Determine the mass of oxygen in the compound.
5. Determine the amount, in mol, of oxygen in the compound.
6. Divide each amount in mol by the lowest amount to determine the lowest whole number ratio of elements in the compound.
7. Express this ratio as a whole number and in empirical formula notation.

Continues →

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$$n(H_2O) = \frac{1.32\text{ g}}{(2 \times 1.0\text{ g mol}^{-1}) + 16.0\text{ g mol}^{-1}}$$

$n(H_2O) = 0.073\text{ mol}$

$n(H) = 2 \times n(H_2O)$

$n(H) = 2 \times 0.073\text{ mol}$

$n(H) = 0.1466\text{ mol}$

$m(H)$ in compound $= n(H) \times M(H)$

$m(H)$ in compound $= 0.1466\text{ mol} \times 1.0\text{ g mol}^{-1}$

$m(H) = 0.1466\text{ g}$

$$n(CO_2) = \frac{8.04\text{ g}}{12.0\text{ g mol}^{-1} + (2 \times 16.0\text{ g mol}^{-1})}$$

$n(CO_2) = 0.1827\text{ mol}$

$n(C) = 1 \times n(CO_2)$

$n(C) = 1 \times 0.1827\text{ mol}$

$n(C) = 0.1827\text{ mol}$

$m(C)$ in compound $= n(C) \times M(C)$

$m(C)$ in compound $= 0.1827\text{ mol} \times 12.0\text{ g mol}^{-1}$

$m(C) = 2.192\text{ g}$

$m(O)$ in compound = total mass of compound – $m(C)$ – $m(H)$

$m(O)$ in compound $= 3.80\text{ g} - 2.192\text{ g} - 0.1466\text{ g} = 1.461\text{ g}$

$$n(O) = \frac{1.461\text{ g}}{16.0\text{ g mol}^{-1}}$$

$n(O) = 0.09133\text{ mol}$

$$\begin{array}{ccccc}
C & : & H & : & O \\
0.1827\text{ mol} & : & 0.1466\text{ mol} & : & 0.09133\text{ mol} \\
\frac{0.1827\text{ mol}}{0.09133\text{ mol}} & : & \frac{0.1466\text{ mol}}{0.09133\text{ mol}} & : & \frac{0.09133\text{ mol}}{0.09133\text{ mol}} \\
2.00 & : & 1.60 & : & 1.00
\end{array}$$

Multiply each number by 5 to convert to a whole number ratio.

C : H : O

10 : 8 : 5

Therefore, the empirical formula is $C_{10}H_8O_5$.

Progress questions

Question 1

What is the percentage by mass of nitrogen in $C_5H_{10}N_2O_2$?

A. 22%

B. 28%

Question 2

What is the percentage by mass of carbon in C_2H_5OH?

A. 26%

B. 52%

Continues →

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Question 3

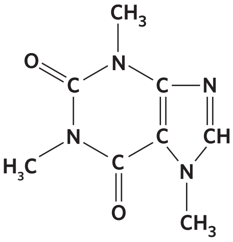
Molecule X is composed of three carbon atoms, eight hydrogen atoms and two oxygen atoms. Which element has the highest percentage by mass in Molecule X?

A. Oxygen

B. Carbon

Question 4

What is the empirical formula of the pictured compound?



A. $C_8H_{10}N_4O_2$

B. $C_4H_5N_2O$

Question 5

Which of the following is an empirical formula?

A. $C_6H_{12}O_6$

B. $C_{12}H_{22}O_{11}$

Determining molecular formulas 1.2.4.2

Molecular formulas depict the actual number of atoms present in a molecule or covalent compound.

How can we determine and interpret molecular formulas?

While the empirical formula is the lowest whole number ratio of atoms of each element in a compound, the molecular formula depicts the actual number of each type of element present in the molecule as shown in figure 3.

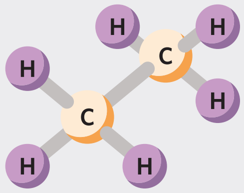
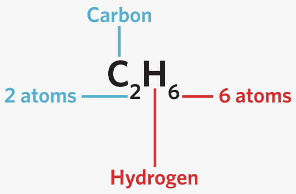
Name	Ethane
Ball-and-stick model	
Molecular formula	

Figure 3 A breakdown of the formulas of ethane

As we can see in figure 3, the molecular formula C_2H_6 indicates that two carbon atoms and six hydrogen atoms are present. However, the empirical formula is CH_3 because the lowest whole number ratio of carbon atoms to hydrogen atoms is 1:3.

USEFUL TIP

The term ‘molecular formula’ is used in calculations involving molecules or covalent compounds – the term ‘ionic formula’ can be used in calculations for ionic compounds.

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e, the molecular formula and empirical formula are different. However, this is not always the case. For example, water's molecular formula is H_2O because there are two atoms of hydrogen and one atom of oxygen in each molecule. Furthermore, the lowest whole number ratio of hydrogen atoms to oxygen atoms is 2:1 and therefore, water's empirical formula is also H_2O.

The molecular formula of a compound can be determined from its empirical formula if the molar mass of the compound is known. This is because the molar mass of a compound's molecular formula will always be some whole number multiple of the molar mass of the empirical formula. The number that multiplies the empirical formula to get to the molecular formula is calculated by the following equation:

Empirical formula multiplier

$$= \frac{\text{molar mass of the compound (molar mass of the molecular formula)}}{\text{molar mass of the empirical formula}}$$

The process of determining the molecular formula of a molecule from its empirical formula is shown in figure 4.

MISCONCEPTION

'A molecule's molecular formula is always different to its empirical formula.'

The molecular formula and empirical formula of a molecule can be the same if the atoms in a molecule are already in the lowest whole number ratio, e.g. C_3H_8.

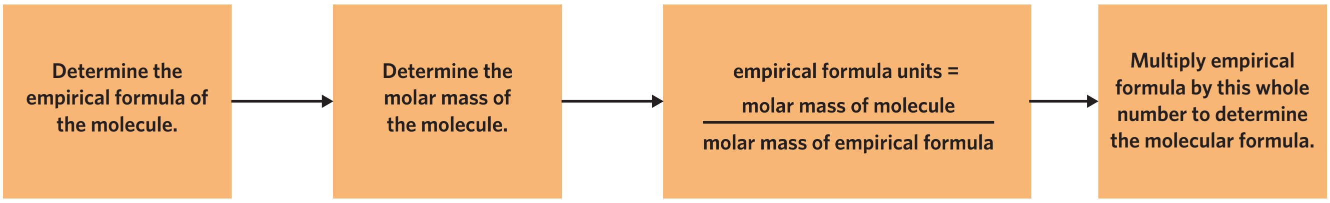


Figure 4 Flowchart of the process required to determine the molecular formula of a molecule from an empirical formula

WORKED EXAMPLE 5

The empirical formula of a molecule is CH_2O. Through mass spectrometry, the molar mass of the compound is shown to be 60.0 g mol^{-1}. What is the compound's molecular formula?

What information is presented in the question?

Empirical formula: CH_2O

Molar mass: 60.0 g mol^{-1}

What is the question asking us to do?

Determine the molecular formula.

What strategies do we need in order to answer the question?

1. Determine the molar mass of the empirical formula.
2. Determine the number of empirical formula units in the molar mass of the compound.
3. Multiply the empirical formula by the number of empirical formula units to determine the molecular formula.

Answer

Molar mass of empirical formula = $(1 \times 12.0\ \text{g mol}^{-1}) + (2 \times 1.0\ \text{g mol}^{-1}) + (1 \times 16.0\ \text{g mol}^{-1})$

Molar mass of empirical formula = 30.0 g mol^{-1}

$$\text{Empirical formula multiplier} = \frac{\text{molar mass of the molecule (molar mass of the molecular formula)}}{\text{molar mass of the empirical formula}}$$

$$\text{Empirical formula multiplier} = \frac{60.0\ \text{g mol}^{-1}}{30.0\ \text{g mol}^{-1}}$$

Empirical formula multiplier = 2

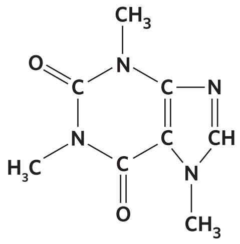
Molecular formula = $2 \times CH_2O$

Molecular formula = $C_2H_4O_2$

Progress questions

Question 6

What is the molecular formula of the pictured compound?



A. $C_8H_{10}N_4O_2$

B. $C_4H_5N_2O$

Question 7

The empirical and molecular formulas of a compound are always different.

A. True

B. False

Question 8

The molecular formula of a compound is known. What piece of information, if any, is required to determine the empirical formula?

A. The molar mass of the compound.

B. No other piece of information is required.

Question 9

The empirical formula of a compound is known. What piece of information, if any, is required to determine the molecular formula of the compound?

A. The molar mass of the compound

B. No other piece of information is required

Question 10

A molecule with the empirical formula CH_2O has a relative molecular mass of 180.2. What is the molecular formula of the molecule?

A. $C_6H_{12}O_6$

B. $C_3H_6O_3$

The content in this lesson is considered fundamental prior knowledge to carbon-based fuels (Unit 3 AOS 1).

Theory summary

- Percentage composition by mass reveals the percentage that the atom(s) of each element contributes to the total mass of the compound.
- Empirical formulas reveal the lowest whole-number ratio of atom(s) of each element in a compound.
- Molecular formulas reveal the actual number of atom(s) of each element in a compound.

6C Questions

Deconstructed

Use the following information to answer questions 11–13.

Iron oxides are chemical compounds which are present in large quantities in Australia. A chemist working for a mining company was investigating the properties of these compounds. She burnt 0.424 g of iron in the presence of excess oxygen and produced 0.606 g of a reddish-brown oxide.

Question 11 (1 MARK) 🌶

What is the amount of iron, in mol, present in the initial sample of iron?

A. 0.424 mol

B. 55.8 mol

C. 0.00760 mol

D. 0.0109 mol

Question 12 (1 MARK) 🌶🌶

What is the amount of oxygen, in mol, present in the reddish-brown oxide formed?

A. 0.0379 mol

B. 0.0265 mol

C. 0.182 mol

D. 0.0114 mol

Question 13 (4 MARKS) 🌶🌶

Based on the above answers, determine the empirical formula of the compound. The chemist also discovered that the molar mass of the compound is 159.6 g mol^{-1}. Is the empirical formula the same as the ionic formula of the compound? Justify your answer.

Exam-style

Question 14 (9 MARKS) 🌶🌶

Calculate the percentage by mass of the following elements in their compounds.

a. H in HCl (2 MARKS)

b. Mn in $KMnO_4$ (2 MARKS)

c. H in $(NH_4)_3PO_4$ (2 MARKS)

d. Fe in Fe_2O_3 (2 MARKS)

e. Using your answer to part d, calculate the mass of iron that can be recovered from a 36.2 g sample of Fe_2O_3. (1 MARK)

Question 15 (10 MARKS) 🌶🌶

Based on the following pieces of information, calculate the empirical formula of the compounds given.

a. H: 2.8% Cl: 97.2% (2 MARKS)

b. Pb: 38.43 g C: 17.83 g H: 3.74 g (2 MARKS)

c. Al: 15.8% S: 28.1% O: 56.1% (2 MARKS)

d. A 6.4 g compound containing hydrogen and carbon that contains 4.8 g of carbon. (2 MARKS)

e. A 1.996 g compound composed of only nitrogen and oxygen contains 0.608 g of nitrogen. (2 MARKS)

Back to contents (10 MARKS)

Based on the following pieces of information, determine the molecular formula of the compounds given.

a. Empirical formula: CH_2, molar mass: 70.0 g mol^{-1}. (2 MARKS)

b. Empirical formula: CH_2O, molar mass: 180.0 g mol^{-1}. (2 MARKS)

c. Empirical formula: $C_6H_{10}S_2O$, molar mass: 162.2 g mol^{-1}. (2 MARKS)

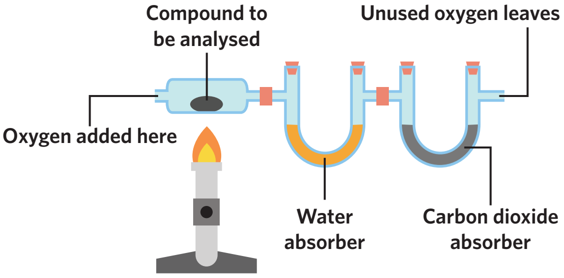
d. Empirical formula: CH_4N, molar mass: 60.0 g mol^{-1}. (2 MARKS)

e. Empirical formula: C_3H_8N, molar mass: 116.0 g mol^{-1}. (2 MARKS)

Key science skills

Question 17 (6 MARKS)

The apparatus shown is used in combustion analysis to determine the empirical formula of compounds containing carbon, hydrogen and oxygen.



a. The experimenter unseals the water absorber, exposing it to the atmosphere, but leaves the carbon dioxide absorber sealed. Comment on possible effects on the accuracy of the results. (2 MARKS)

b. The scale that is used is not calibrated. What type of error does this introduce? (2 MARKS)

c. The experimenter says that due to the uncalibrated scale, the results won't be precise but will be accurate. Is this correct? (2 MARKS)

FROM LESSON 16D

Questions from multiple lessons

Question 18 (11 MARKS)

Lithium (Li(s)) is a highly reactive metal and in the presence of water ($H_2O(l)$), it will spontaneously react to produce lithium hydroxide (LiOH(aq)) and flammable hydrogen gas ($H_2(g)$).

a. Write the unbalanced chemical equation for the reaction between lithium and water. (1 MARK)

b. Balance this chemical equation. Explain your reasoning. (2 MARKS)

c. What is the electron configuration of lithium according to Schrödinger's model of the atom? (1 MARK)

d. Compare the electronegativity of lithium and fluorine. (3 MARKS)

e. Compare the reactivity of lithium and potassium. (4 MARKS)

FROM LESSONS 1C & 3B

Hints

14. Percentage by mass is the percentage that an element contributes to the total mass of the compound.

15. The empirical formula is the lowest whole-number ratio of atoms of each element present.

16. The empirical formula and molecular formula are multiples of each other.

17a. Accuracy is a measure of how close the measured value is to the true value.

17b. Different errors have different effects on the results.

17c. Precision is a measure of how close measured values are to each other, not the true value.

18a. The reactants are on the left-hand side of the chemical equation and the products are on the right-hand side.

18b. Balanced chemical equations have the same number of atoms of each element on each side of the chemical equation.

18c. Schrödinger's model of the atom involves shells and subshells.

18d. Electronegativity is a measure of how strongly an atom attracts electrons to itself.

18e. Reactivity of group 1 elements is how easily they lose valence electrons.

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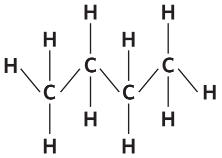
Chapter 6 review

Multiple choice (10 MARKS)

Question 1 (1 MARK) 🌶

Butane is the structure shown and it is often used as a fuel for cooking while camping.



What is the relative molecular mass, empirical formula and molecular formula of butane respectively?

A. 34.0, C_2H_5, C_4H_{10}

B. 29.0, C_4H_{10}, C_2H_5

C. 58.0, C_2H_5, C_4H_{10}

D. 12.0, C_4H_{10}, C_2H_5

Question 2 (1 MARK) 🌶🌶

How many moles of oxygen are there in 25.1 g of $CaCO_3(s)$?

A. 0.250 mol

B. 0.251 mol

C. 0.522 mol

D. 0.752 mol

Question 3 (1 MARK) 🌶🌶

How many moles of oxygen atoms are there in 2.020×10^{23} $Ba(NO_3)_2(aq)$?

A. 0.336 mol

B. 0.671 mol

C. 2.01 mol

D. 5.37 mol

Question 4 (1 MARK) 🌶🌶

The units of relative atomic mass are

A. g mol^{-1}.

B. no units.

C. g.

D. mol.

Question 5 (1 MARK) 🌶🌶

A molecule has the empirical formula CH_3, and a molar mass of 30 g mol^{-1}. For this molecule, the empirical formula and the molecular formula

A. are the same.

B. may be the same or different, there is not enough information to tell.

C. are different.

D. are sometimes the same, and sometimes different.

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Use the following information to answer questions 6–8.

Element X exists as five isotopes, the relative abundances of which are shown in the table below.

Isotope	Relative abundance (%)
^{46}X	8.25
^{47}X	7.44
^{48}X	73.7
^{49}X	5.41
^{50}X	5.18

Question 6 (1 MARK) 🌶🌶

The least abundant isotope of Element X is

A. ^{46}X.

B. ^{50}X.

C. ^{47}X.

D. ^{49}X.

Question 7 (1 MARK) 🌶🌶

The relative atomic mass of Element X will be closest to

A. 48.

B. 74.

C. 5.

D. 50.

Question 8 (1 MARK) 🌶🌶

What is the most likely identity of Element X?

A. Cadmium

B. Carbon

C. Sodium

D. Titanium

Question 9 (1 MARK) 🌶🌶

In 2 moles of glucose ($C_6H_{12}O_6$), there are ___________ moles of oxygen atoms.

A. 2

B. 12

C. 6

D. 24

Question 10 (1 MARK) 🌶🌶🌶

A compound is composed of one iron atom, five carbon atoms and five oxygen atoms. Which element has the highest percentage by mass in the compound?

A. Iron

B. Oxygen

C. Carbon

D. Not enough information to tell

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Short answer (30 MARKS)

Question 11 (5 MARKS)

A sample containing 30.0 g of ethanoic acid (CH_3COOH), also known as vinegar acid, is being analysed.

a. Calculate the number of moles of ethanoic acid in the sample. (1 MARK)

b. How many hydrogen atoms are present in 30 g of ethanoic acid? (2 MARK)

c. What is the percentage by mass of oxygen in ethanoic acid? (2 MARKS)

Question 12 (4 MARKS)

For the following compounds, calculate the percentage by mass of

a. K in $KMnO_4$ (2 MARKS)

b. N in $(NH_4)_3PO_4$ (2 MARKS)

Question 13 (3 MARKS)

Bioethanol, C_2H_5OH, is produced by the fermentation of glucose, $C_6H_{12}O_6$, according to the following equation.

$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$

Calculate the mass of C_2H_5OH obtained when 5.68 g of carbon dioxide, CO_2, is produced.

Adapted from VCAA 2018 exam Multiple choice Q10

Question 14 (4 MARKS)

A sample of the insecticide dichlorodiphenyltrichloroethane (DDT), $C_{14}H_9Cl_5$, was found to contain 0.120 g of carbon. Calculate the mass of chlorine present in the sample.

Adapted from VCAA 2010 Exam 1 Multiple choice Q3

Question 15 (4 MARKS)

Magnesium has three naturally occurring isotopes. Their relative abundances and masses are given in the table below.

	Relative abundance (%)	Relative isotopic mass
^{24}Mg	78.99	23.985
^{25}Mg	10.00	24.986
^{26}Mg	11.01	25.983

a. The abundances and relative isotopic masses have been determined experimentally. What instrument is commonly used to obtain this information? (1 MARK)

b. Which isotope of magnesium is the least abundant in nature? (1 MARK)

c. Using the information above, calculate the relative atomic mass of magnesium. (2 MARKS)

Adapted from VCAA 2005 Exam 2 Short answer Q2

Question 16 (10 MARKS)

Esters are molecules containing only carbon, hydrogen and oxygen, and are responsible for the smell of many fruits, including oranges. A confectionery company wanted to extract the ester from an orange in order to make orange smelling lollies.

a. Combustion analysis of 3.07 g of this molecule produced 7.85 g of carbon dioxide and 3.21 g of water vapour. What is the empirical formula of the compound? (4 MARKS)

b. The following is a mass spectrum of the organic molecule. It reveals that the relative molecular mass of the compound is 172.0 g mol^{-1}.

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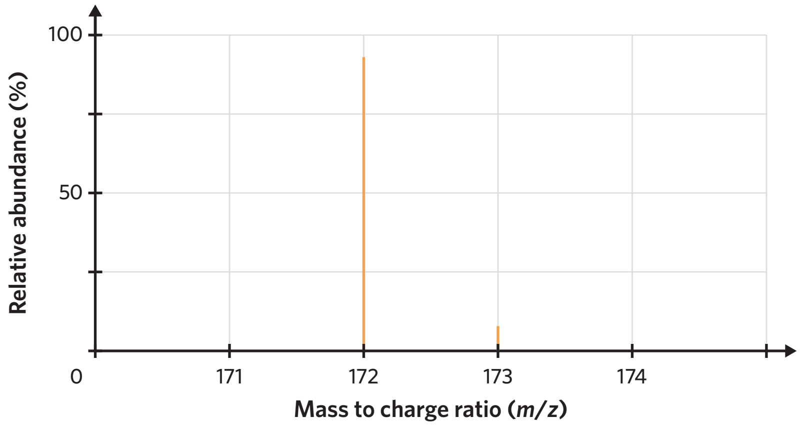

What is the molecular formula of this compound?

c. The peak at m/z 173 is due to an isotope of an element. What is an isotope? (1 MARK)

d. The relative abundances of the ^{13}C isotope, ^{2}H isotope and ^{17}O isotope are 1.1%, 0.02% and 0.04% respectively. Based on these values, explain why there is one peak present on the mass spectrum that is greater than the relative molecular mass of the compound. (2 MARKS)

Key science skills (8 MARKS)

Question 17 (8 MARKS)

There are two naturally occurring isotopes of copper, ^{63}Cu and ^{65}Cu. Using mass spectrometry, a school laboratory technician was able to determine the relative atomic mass of copper to be 63.6.

a. In terms of abundance, which isotope is more commonly found in nature? Justify your answer. (2 MARKS)

b. Solid copper (I) oxide (Cu_2O) decomposes when heated into solid copper (Cu) and oxygen gas (O_2). Write the balanced chemical equation for this reaction. (1 MARK)

c. How many molecules of $O_2(g)$ will be produced from the decomposition of 1.00 mole of Cu_2O? (3 MARKS)

d. Are the results obtained from a mass spectrometer qualitative or quantitative? (1 MARK)

e. Name one limitation associated with the use of mass spectrometry. (1 MARK)

FROM LESSONS 16B & 16D

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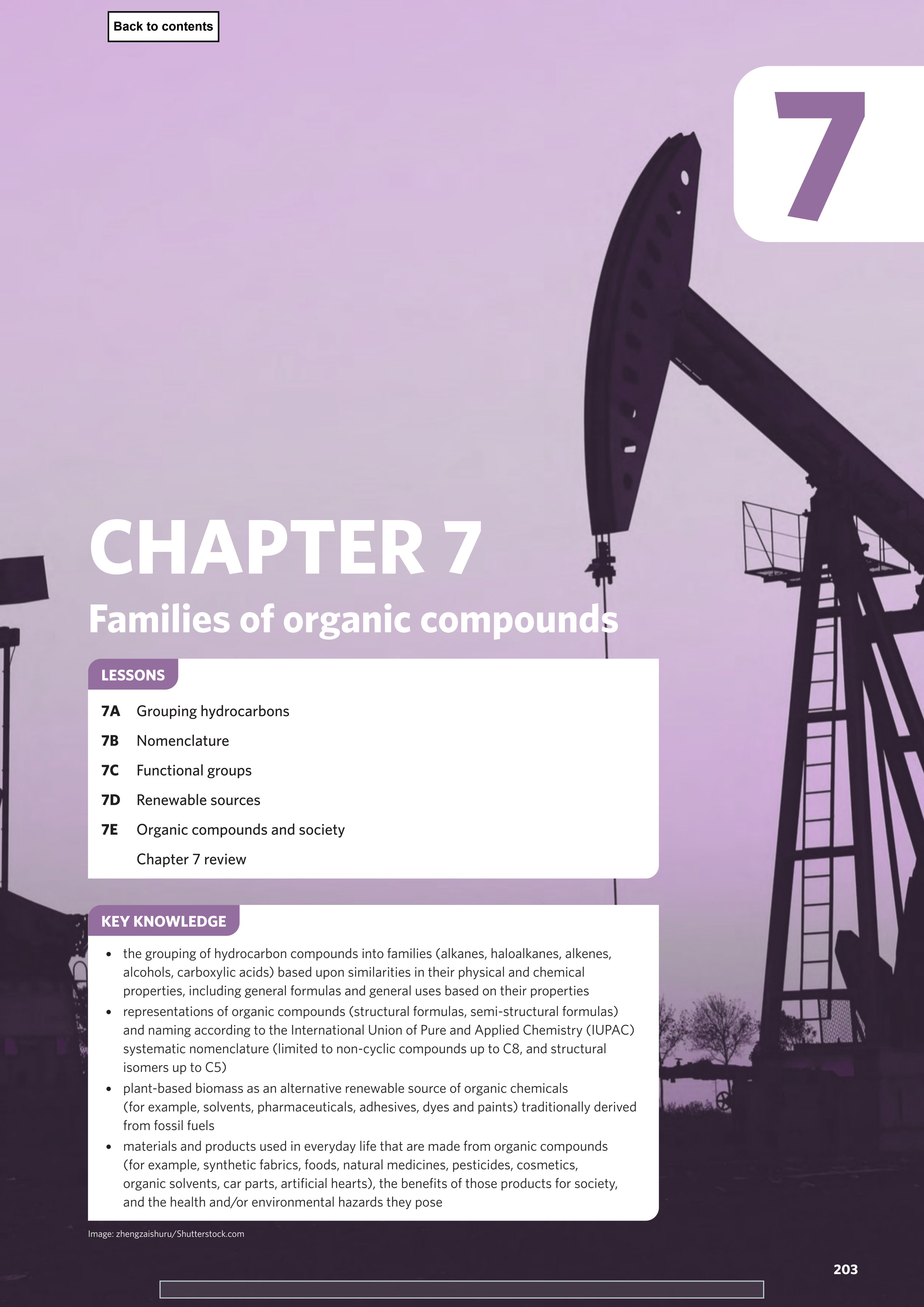

CHAPTER 7

Families of organic compounds

LESSONS

KEY KNOWLEDGE

- the grouping of hydrocarbon compounds into families (alkanes, haloalkanes, alkenes, alcohols, carboxylic acids) based upon similarities in their physical and chemical properties, including general formulas and general uses based on their properties
- representations of organic compounds (structural formulas, semi-structural formulas) and naming according to the International Union of Pure and Applied Chemistry (IUPAC) systematic nomenclature (limited to non-cyclic compounds up to C8, and structural isomers up to C5)
- plant-based biomass as an alternative renewable source of organic chemicals (for example, solvents, pharmaceuticals, adhesives, dyes and paints) traditionally derived from fossil fuels
- materials and products used in everyday life that are made from organic compounds (for example, synthetic fabrics, foods, natural medicines, pesticides, cosmetics, organic solvents, car parts, artificial hearts), the benefits of those products for society, and the health and/or environmental hazards they pose

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7A Grouping hydrocarbons

STUDY DESIGN DOT POINT

- the grouping of hydrocarbon compounds into families (alkanes, haloalkanes, alkenes, alcohols, carboxylic acids) based upon similarities in their physical and chemical properties, including general formulas and general uses based on their properties

7A 7B 7C 7D 7E

1.2.5.1 Introduction to hydrocarbons

1.2.5.2 Hydrocarbon bonding

ESSENTIAL PRIOR KNOWLEDGE

2A Covalent bonding

2B Intermolecular forces

See questions 37–38.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.


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Why is this 'black gold' so important to society?

Crude oil, also known as 'black gold', is a rich source of **hydrocarbons**, an abundant and useful class of molecules with applications ranging from fuels to foods. In this lesson we will discuss some properties of these molecules and how they can be applied.

KEY TERMS AND DEFINITIONS

Addition reaction	a reaction in which one molecule combines with another molecule to form a larger molecule
Alkane	hydrocarbon that contains only single carbon to carbon bonds
Alkene	hydrocarbon with at least one double carbon to carbon bond
Carbon chain	linear set of carbon atoms in a hydrocarbon
Homologous series	series of hydrocarbons that have a similar chemical structure and chemical properties
Hydrocarbon	compound consisting of hydrogen and carbon only
Organic compounds	compounds consisting of carbon atoms most commonly covalently bonded to hydrogen, nitrogen, phosphorus or oxygen atoms
Saturated hydrocarbon	a hydrocarbon possessing only single bonds between carbon atoms
Substitution reaction	one atom or group of atoms replacing a part of another molecule
Unsaturated hydrocarbon	a hydrocarbon possessing at least one double bond between carbon atoms

Introduction to hydrocarbons 1.2.5.1

Hydrocarbons are organic molecules found abundantly around the world that serve many different purposes.

Why are hydrocarbons important to society?

The functioning of modern human society relies on various forms of energy, whether that be in the form of electricity to power an electrical device, or perhaps thermal energy to warm our homes. However, the use of these fuels can have detrimental environmental effects, an idea that has become a prominent point of contemporary political debate.

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One primary source of energy for human consumption is hydrocarbons. These are **organic compounds** consisting of carbon and hydrogen atoms, organised with the carbon atoms in a linear **carbon chain**, as shown in figure 1.

As well as generally being excellent fuels, hydrocarbons are commonly used as cooking oils and solvents or precursors in chemical manufacturing. Organic compounds also make up the majority of our bodies and the food we consume – all living things are composed primarily of carbon-based molecules, and so an understanding of hydrocarbons is essential in fields such as biology, biochemistry and medicine.

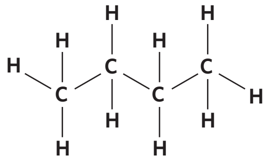


Figure 1 An example of a 'carbon chain'

MISCONCEPTION

'In chemistry, 'organic' means naturally occuring.'

Although the term organic is often used to describe things as natural, such as 'organic' food, this is not what the term means in chemistry. When a molecule is described as an organic molecule, this means that the molecule is carbon-based, and may be either naturally or synthetically produced.

Progress questions

Question 1

Hydrocarbons are inorganic molecules.

A. True

B. False

Question 2

Organic molecules are always naturally-occurring.

A. True

B. False

Question 3

Hydrocarbons are usually arranged in a ___________ sequence.

A. chain-like

B. box-like

Hydrocarbon bonding 1.2.5.2

The unique bonding characteristics of carbon atoms give rise to a diverse range of carbon-based molecules.

How are hydrocarbons grouped?

The basic structure of a single chain hydrocarbon consists of a linear chain of carbon atoms bonded together. Hydrocarbons can vary in the number of carbon and hydrogen atoms they have. Those that differ from the previous hydrocarbons in the series through the addition of a carbon atom to the chain are described as belonging to the same **homologous series**. As a result, hydrocarbons in the same homologous series generally have:

- similar physical and chemical properties,
- similar structure,
- and the same general formula.

Each carbon atom contains four valence electrons and therefore, to satisfy the octet rule, would need to share four electrons with other atoms. Hence, each carbon atom can form four covalent bonds. As we will see later in this chapter, a carbon atom can bond to a multitude of different elements to gain a full valence shell.

Perhaps the most notable difference between hydrocarbons is the number of carbon atoms in the chain, ranging from methane with 1 carbon atom, to molecules with thousands. An illustration of this can be seen in figure 2, in which a 1-carbon hydrocarbon (methane) and a 5-carbon chain (propane) are shown. Nomenclature will be discussed in greater detail in lesson 7B.

Carbon atoms are also able to form single or double bonds with other atoms. This means that these carbon atoms can form one or two covalent bonds with another carbon. As we learned in lesson 2A, a covalent bond involves two atoms sharing an electron each to fill their shells.

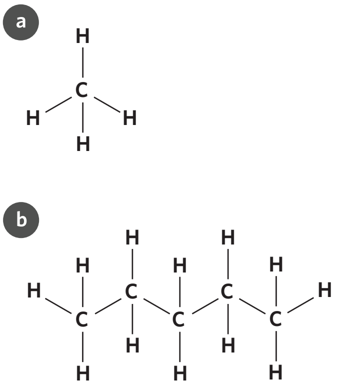


Figure 2 Structural formulas of methane **(a)** and pentane **(b)**

Therefore, in single bonds, the atoms are sharing two electrons, and in double bonds, the atoms are sharing four electrons, as shown in figure 3.

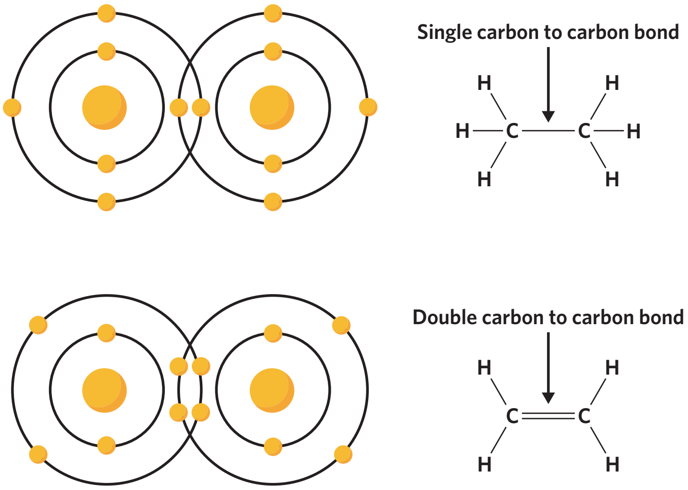
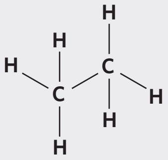
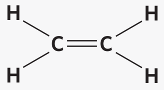


Figure 3 Single and double carbon-carbon bonding

Different groups of hydrocarbons have different names based on the type of bonds present in their carbon chain. Hydrocarbons that only contain single carbon to carbon bonds in their carbon chain are known as **alkanes** and have the general molecular formula C_nH_{2n+2} where *n* represents the number of carbons in the parent chain. If there is a double bond between two of the carbon atoms then the hydrocarbon is called an **alkene** and has a general formula of C_nH_{2n}. This is summarised in table 1.

Table 1 Grouping of different types of hydrocarbons

Description	Representation	Type of hydrocarbon	General formula (where *n* represents the number of carbons in the parent chain)
Carbon chain contains only single carbon to carbon bonds		Alkane	C_nH_{2n+2}
Carbon chain contains at least one carbon to carbon double bond		Alkene	C_nH_{2n}

WORKED EXAMPLE 1

A hydrocarbon has the molecular formula C_6H_{12}. Would this molecule be considered an alkane or alkene?

What information is presented in the question?

The molecular formula for a molecule.

What is the question asking us to do?

Classify the hydrocarbon as an alkane or alkene.

What strategies do we need in order to answer the question?

1. Identify the two different formulas for each hydrocarbon type.
2. Determine the formula that is able to generate the formula given in the question.
3. Classify the molecule.

Continues →

Answer

General formula:

Alkane: C_nH_{2n+2}

Alkene: C_nH_{2n}

The formula given is C_6H_{12}, therefore we can see that $n = 6$

If the molecule was an alkane, it would have the molecular formula:

$= C_6H_{(2 \times 6) + 2}$

$= C_6H_{14}$

As we can see, this formula is different from the given formula.

If the molecule was an alkene, it would have the molecular formula:

$= C_6H_{(2 \times 6)}$

$= C_6H_{12}$

This molecular formula matches the general formula. As a result, C_6H_{12} is an alkene.

Another way in which we can describe hydrocarbons based on their bonding patterns is through saturation. **Saturated hydrocarbons** are those that only have single bonds between the carbon atoms, whereas **unsaturated hydrocarbons** have at least one double bond between carbon atoms in a carbon chain. The term 'saturated' refers to the number of hydrogens bonded to each carbon atom in the molecule. An organic molecule with no double bonds (and no other functional groups) is able to accommodate the maximum number of hydrogen atoms. When double bonds are added, the molecule loses sites for bonding to hydrogen, thus becoming 'unsaturated'.

How do the physical and chemical properties of hydrocarbons differ?

Homologous series of hydrocarbons have similar physical properties. However, as the carbon chain length in the hydrocarbon increases or decreases, this alters the extent to which these properties are exhibited.

Hydrocarbons without functional groups, such as alkanes, are non-polar molecules which cannot dissolve in water. The strongest intermolecular forces between alkanes are weak dispersion forces. Dispersion forces arise as a result of interactions between instantaneous dipoles between atoms in adjacent molecules, and therefore, as the length of a carbon chain increases, the strength of the intermolecular forces between the hydrocarbons also increases due to the stronger forces between the adjacent molecules.

As a result, different alkanes exhibit different properties based on their differing carbon chain lengths, which makes them useful for many different purposes. For example, at room temperature methane is found in a gaseous form, whereas octane is found in a liquid form. The differences in state of these hydrocarbons reflects the strength of the intermolecular forces.[1] This is illustrated in table 2.

STRATEGY

By using general formulas, one can distinguish between alkanes and alkenes without having to draw or see the whole structure. This is particularly useful when dealing with large molecules that would take a long time to draw.

KEEN TO INVESTIGATE?

[1] **How can we use the differences in physical properties of alkanes to our advantage?**
Search YouTube: Fractional Distillation | Organic Chemistry | Chemistry

Table 2 Relationship between the number of carbons in chain and boiling point

Number of carbons in chain	1	2	3	4	5	6	8	10
Boiling point (°C)	−164	−89	−42	−1	36	69	125	174
State at 25 °C	gas	gas	gas	gas	liquid	liquid	liquid	liquid

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Table 3 Physical and chemical properties of alkanes and alkenes

Alkanes		Alkenes	
Physical properties	**Chemical properties**	**Physical properties**	**Chemical properties**
• Boiling/ melting point depends on chain length • Less dense than water • Colourless • Insoluble in water	• Relatively unreactive • Can undergo **substitution reactions** in presence of UV light • Combustible	Similar to alkanes – can differ depending on number of double bonds present	• Reactive • Undergo **addition reactions** readily • Can undergo incomplete combustion

As seen in table 3, physical properties such as intermolecular forces are largely similar between alkanes and alkenes. Alkenes are more easily distinguished from their alkane counterparts by their differing chemical properties.

While alkanes undergo substitution reactions, wherein they have a hydrogen atom in their structure replaced by another atom, alkenes participate in addition reactions. Rather than replacing an atom in the structure of the hydrocarbon, addition reactions involve the addition of a molecule onto the hydrocarbon without the loss of any existing atoms. This occurs due to the presence of reactive, electron-rich double bonds. When a bond in a double bond is broken, this frees up an additional electron on each carbon involved in the carbon to carbon double bond, enabling the bonding of additional atoms to its structure. This ability to participate in addition reactions makes alkenes a useful precursor in many chemical manufacturing processes.

Progress questions

Question 4

All hydrocarbons, no matter their carbon chain length, have the same properties.

A. True

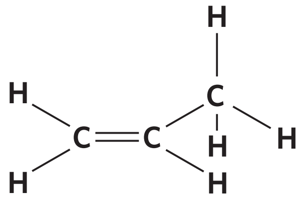
B. False

Question 5

The molecule shown is a(n)

A. saturated hydrocarbon.

B. unsaturated hydrocarbon.



Question 6

The molecule shown in Question 5 is more likely to undergo

A. an addition reaction.

B. a substitution reaction.

Question 7

The main force of attraction affecting the boiling point of the molecule shown in Question 5 is

A. hydrogen bonding.

B. dispersion forces.

Continues →

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Question 8

An alkane contains 14 hydrogen atoms. How many carbon atoms will it contain?

A. 8

B. 6

Question 9

Consider two alkanes with the formulas C_2H_6 and C_8H_{18} respectively. Which compound would be expected to have the lower melting point?

A. C_2H_6

B. C_8H_{18}

Question 10

Alkanes and alkenes are easily distinguishable from each other by their

A. physical properties.

B. chemical properties.

The content in this lesson is considered fundamental prior knowledge to structure, nomenclature and properties of organic compounds (Unit 4 AOS 1).

Theory summary

Alkanes: General formula: C_nH_{2n+2}

Structure	Uses	Physical properties	Chemical properties
• Chain-like sequence of carbons • Single bonds only (saturated)	• Fuels • Cooking oils • Solvents for chemical manufacturing processes	• Boiling/melting point depends on chain length • Less dense than water • Colourless • Insoluble in water	• Relatively unreactive • Can undergo substitution reactions in presence of UV light • Combustible

Alkenes: General formula: C_nH_{2n}

Structure	Use	Physical properties	Chemical properties
• Chain-like sequence of carbons • Contains one or more double bonds (unsaturated)	• Precursors in many chemical manufacturing processes	• Similar to alkanes – can differ depending on number of double bonds present	• Reactive • Undergo addition reactions readily • Can undergo incomplete combustion

7A Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 11–13.

Hydrocarbons are very useful in everyday life. For example, octane (an alkane with 8 carbons) is widely used as a transport fuel whereas fuels like methane (an alkane with 1 carbon) can be used for cooking purposes.

Question 11 (1 MARK) 🌶

Octane has the molecular formula

A. C_8H_{16}

B. C_8H_{20}

C. C_8H_{14}

D. C_8H_{18}

Question 12 (1 MARK) 🌶

In terms of the main intermolecular forces responsible for the boiling point of each molecule,

A. octane is predominantly held together by hydrogen bonds whereas methane is predominantly held together by dispersion forces.

B. methane is predominantly held together by hydrogen bonds whereas octane is predominantly held together by dispersion forces.

C. both are held together by hydrogen bonds only.

D. both are held together by dispersion forces only.

Question 13 (4 MARKS) 🌶🌶

Methane is a much cheaper fuel than octane. A car company wanted to test whether or not they could build a fuel tank that was suitable for methane as part of a new, more cost effective car. One of the biggest concerns is to be able to develop a tank that is able to insulate the fuel from heat generated by the car while it is functioning. After conducting tests, they were able to determine the following boiling points of both fuel sources.

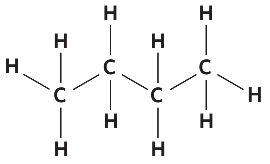
	Fuel A	**Fuel B**
Boiling point (°C)	−162	126

Which of the fuels, octane or methane, would be able to withstand the most heat before boiling? Based on your answer, identify the name of fuel A and B.

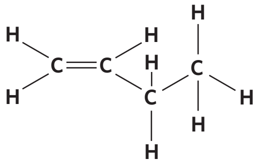
Exam-style

Question 14 (5 MARKS) 🌶

Consider the molecules shown.



Molecule A



Molecule B

a. Identify which of the molecules is considered 'saturated', and explain what the term 'saturated' refers to. (2 MARKS)

b. A student wanted to select one of the molecules to take part in an addition reaction. Suggest the appropriate molecule to meet this requirement, and identify a property that allows the molecule to undergo such a reaction. (2 MARKS)

c. Identify which of the molecules is more likely to be used as a fuel. (1 MARK)

Key science skills

Question 15 (4 MARKS) 🌶🌶

A scientist on an oil rig has been called to investigate an oil spill that has taken place in the Bass Strait, south of Melbourne. Having noticed oil floating on the water, workers called in the scientist to determine the composition of the substance and the severity of the spill.

a. Suggest a type of compound likely to be found in the oil spill. (1 MARK)

b. Suggest a reason why the oil does not mix with the water. (2 MARKS)

c. When measuring the amount of oil in the water, a worker accidentally knocks a barrel of oil over, with its contents all spilling into the oil spill.

What type of error has taken place? (1 MARK)

FROM LESSON 16D

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Questions from multiple lessons

Question 16 (6 MARKS)

There are a large number of different types of hydrocarbons that exist. At the most basic level, hydrocarbons can be categorised into two main categories.

a. For saturated hydrocarbons:

- **i.** Do they contain covalent bonds?
- **ii.** What bonds do they have within the carbon chain?
- **iii.** Can they participate in addition reactions?

b. For unsaturated hydrocarbons:

- **i.** Do they contain covalent bonds?
- **ii.** What bonds do they have within the carbon chain?
- **iii.** Can they participate in addition reactions?

FROM LESSON 2A

Question 17 (6 MARKS)

A chemistry student is investigating the chemical structures present in bonding between atoms.

a. The student observes a carbon-carbon double bond. How many electrons are being shared between the two carbon atoms? Justify your answer. (3 MARKS)

b. Despite both being carbon-based, diamond is a solid with the highest hardness of any naturally occurring substance, while most hydrocarbons are in gaseous or liquid states at room temperature. Explain this difference. (3 MARKS)

FROM LESSON 2C

Question 18 (6 MARKS)

For a road trip, Grantham wanted to figure out how much fuel he would need to make sure that he reached his destination. His new car will only take the best petrol, which consists of the hydrocarbon octane C_8H_{18}, which has a density of 0.703 g mL^{-1}.

a. Prior to the trip, Grantham filled his car so that he would have a full tank of petrol. If a full tank is 40.0 L, how many grams of octane fuel did he have in total? (2 MARKS)

b. What amount, in mol, of octane fuel, does he have when his tank is full? (2 MARKS)

c. If one mol of octane fuel is able to generate 5 460 kJ of energy when burnt, how much energy would a full tank of fuel generate? (2 MARKS)

FROM LESSON 6C

Hints

14a. Saturation relates to the number of hydrogen atoms per carbon atom.

14b. Addition reactions require unsaturated molecules.

14c. Alkanes are highly combustible, while alkenes generally undergo incomplete combustion.

15a. Oil is considered a mixture of compounds.

15b. The ability of two substances to mix is closely related to polarity.

15c. Errors can be classified based on the effect they have on the experiment.

16. Hydrocarbons can be classified into saturated and unsaturated hydrocarbons.

17a. Covalent bonds are a result of atoms sharing electrons.

17b. Diamonds are carbon lattice allotropes.

18a. Density in g mL^{-1} helps us understand the mass of a substance per millilitre.

18b. The amount of substance depends on the mass of substance present.

18c. The energy generated depends on how much of the fuel is used.

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7B Nomenclature

STUDY DESIGN DOT POINT

- representations of organic compounds (structural formulas, semi-structural formulas) and naming according to the International Union of Pure and Applied Chemistry (IUPAC) systematic nomenclature (limited to non-cyclic compounds up to C8, and structural isomers up to C5)

7A **7B** 7C 7D 7E

ESSENTIAL PRIOR KNOWLEDGE

7A	Hydrocarbon structure
7A	Saturation

See questions 39–40.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: PENpics Studio/Shutterstock.com

How is caffeine assigned the name 1,3,7-trimethylpurine-2, 6-dione?

All organic molecules, no matter how complicated, can be assigned a **systematic name.** In this chapter, we will discuss the IUPAC (International Union of Pure and Applied Chemistry) **nomenclature** guidelines, and how they can be used to identify molecules like caffeine.

KEY TERMS AND DEFINITIONS

Alkyl group	a side-chain attached to the parent carbon chain, containing only carbon and hydrogen atoms
Nomenclature	system developed in order to identify organic compounds
Parent chain	longest continuous chain of carbon-carbon bonds
Parent name	component of systematic name that refers to the length of the parent chain
Structural isomers	isomers that have the same molecular formula but the atoms are arranged in a different spatial arrangement
Systematic name	standardised name used to identify organic compounds

Naming saturated hydrocarbons 1.2.6.1

A set of naming conventions determined by the International Union of Pure and Applied Chemistry (IUPAC) provides a systematic method of naming any organic molecule.

How are saturated hydrocarbons named?

In the last lesson we introduced hydrocarbons, exploring the structure, uses and variations of these molecules. Since there is such a large variety of different hydrocarbons, a standard naming convention is required in order to distinguish between hydrocarbons through their names alone.

To ensure that organic molecules are named in a consistent manner, IUPAC developed nomenclature that is the standard for naming organic molecules worldwide. The standardised name for organic compounds prescribed by IUPAC is also referred to as the systematic name. There are many rules involved in this naming system, which we will work through in this chapter.

When naming hydrocarbons, we need to first consider the length of the carbon chain, as the longest carbon chain in the molecule is the basis of the name of the molecule. The longest continuous carbon chain in an organic molecule is known as the **parent chain**, and this provides the **parent name** of the molecule. For basic hydrocarbons, the naming convention is shown in figure 1.

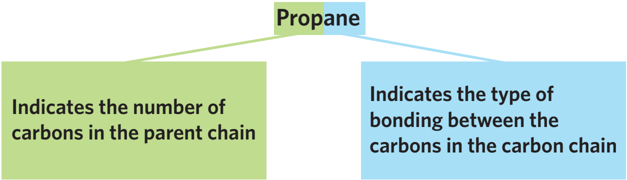

Figure 1 Naming convention of simple hydrocarbons

As seen above, every component of the name of a hydrocarbon signifies something about the structure of the hydrocarbon itself.

A hydrocarbon containing only carbon to carbon single bonds in the parent chain is marked by the suffix '-ane'.

The parent name of a hydrocarbon is another important component as it identifies the number of carbons in the parent chain. Table 1 summarises the parent names used to indicate the chain length of a hydrocarbon.

Table 1 Hydrocarbon parent names

Carbon chain length	Parent name	Example
1	'Meth-'	Methane
2	'Eth-'	Ethane
3	'Prop-'	Propane
4	'But-'	Butane
5	'Pent-'	Pentane
6	'Hex-'	Hexane
7	'Hept-'	Heptane
8	'Oct-'	Octane

Through combining the parent name (denoting the carbon chain length) and the suffix (denoting the types of bonds in the molecule), it is possible to come up with the systematic name of many basic hydrocarbons. Consider the molecule shown in figure 2.

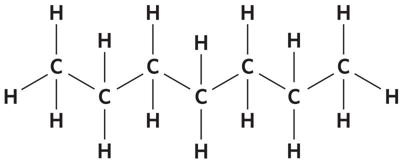

Figure 2 Structural formula of a simple hydrocarbon

The parent chain can be seen as having a total of seven carbon atoms bonded together in an unbroken chain, as shown in figure 3. As such, the parent name for the molecule will be 'hept-'.

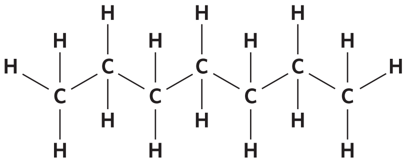

Figure 3 Longest unbroken carbon chain in the hydrocarbon

Since there are only carbon to carbon single bonds in the parent chain, we classify it as an alkane. Therefore the suffix for this molecule will be '-ane'. Combining these components ('hept-' and '-ane'), the systematic name of this molecule is heptane.

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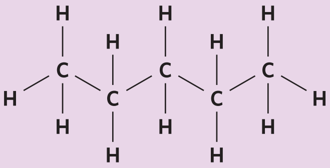
WORKED EXAMPLE 1

What is the systematic name of the molecule with the following structural formula?



What information is presented in the question?

The structural formula for a molecule.

What is the question asking us to do?

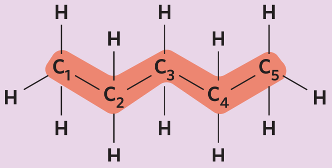
Provide the systematic name of the molecule provided.

What strategies do we need in order to answer the question?

1. Identify the number of carbons in the parent chain.
2. Identify the type(s) of bonds in the parent chain.
3. Apply this information using correct IUPAC naming conventions.

Answer

This molecule is an alkane, and the longest carbon chain length is five carbon atoms.



The parent name 'pent-' is used to denote five carbon atoms in the parent chain. The molecule itself contains only carbon to carbon single bonds so the suffix '-ane' is used.

Therefore the name of this molecule is pentane.

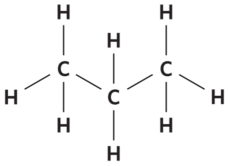
Progress questions

Question 1

The parent name of this molecule is

A. 'prop-'.

B. 'but-'.

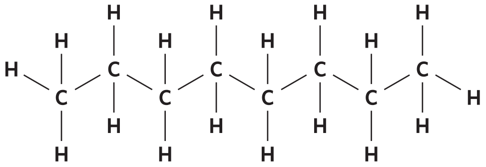


Question 2

The systematic name of the molecule shown is

A. heptane.

B. octane.



Naming unsaturated hydrocarbons 1.2.6.2

Hydrocarbons with carbon to carbon double bonds have a more in-depth naming system compared to simple hydrocarbons.

How are unsaturated hydrocarbons named?

Unsaturated hydrocarbons have at least one carbon to carbon double bond. The way in which we identify this is to change the suffix of the name. A hydrocarbon containing carbon to carbon double bonds in the parent chain is marked by the suffix '-ene'. To be more specific, we also need to indicate where the carbon to carbon double bond is located in the parent chain. Consider the molecule shown in figure 4.

We can see that the longest carbon chain contains five carbons. There is a carbon to carbon double bond in the structure so the suffix for the molecule would be '-ene'. However, we need to be able to indicate to the reader the location of the double bond.

To do so, we need to number the carbon atoms in the parent chain. As shown in figure 5, there are two ways in which we can do this, depending on which end we start from.

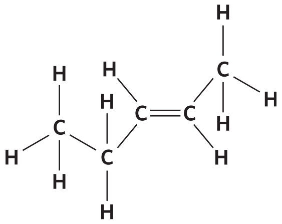


Figure 4 Structural formula of an unsaturated hydrocarbon

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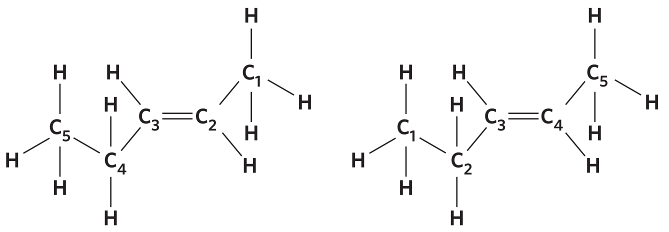


Figure 5 Numbering of the parent chain of a hydrocarbon

The way in which we decide which numbering system to use is to determine which allows for the double bond to occur at the lowest numbered carbon. Based on the two options, we can see that the double bond can occur at either carbon number 3 or carbon number 2. As such, we will proceed with the numbering system where the double bond occurs at carbon number 2. The name of this hydrocarbon is shown in figure 6.

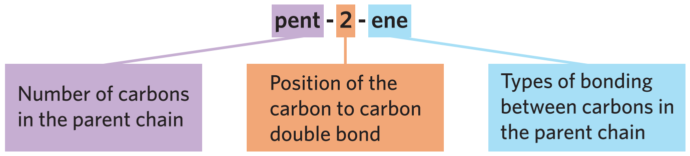


Figure 6 Naming convention of unsaturated hydrocarbons

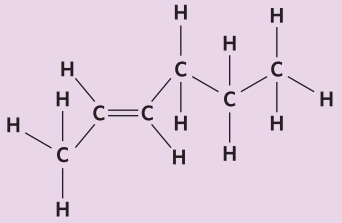
As we can see, the positioning of the carbon to carbon double is indicated just before the suffix.

USEFUL TIP

Whenever a number is present in a systematic name, it should be separated from any words around it by a hyphen '-'.

WORKED EXAMPLE 2

What is the systematic name of the molecule with the following structural formula?



What information is presented in the question?

The structural formula for a molecule.

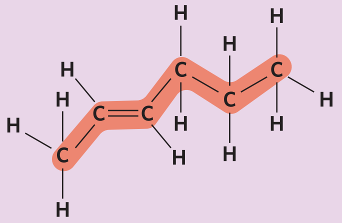
What is the question asking us to do?

Provide the systematic name of the molecule provided.

What strategies do we need in order to answer the question?

1. Determine the longest carbon chain length.
2. Determine the types of bonds within the parent carbon chain.
3. Number the carbons in the parent chain.
4. Apply this information using correct IUPAC naming conventions.

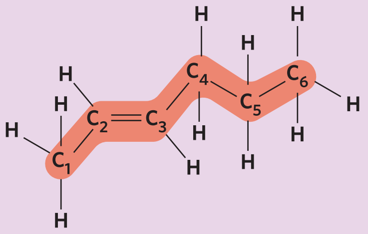
Answer



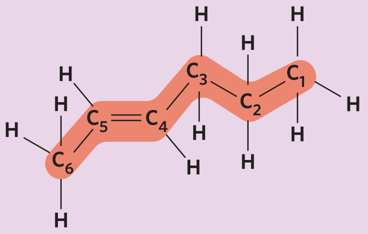
The longest carbon chain contains six carbons.
The structure also contains a double bond, and therefore the suffix will be '-ene'.

There are two different ways we can number the carbons in the parent chain:

Option 1:



Option 2:



Option 1 allows for the carbon to carbon double bond to occur at the lowest numbered carbon, carbon 2. Therefore the name of this compound is hex-2-ene.

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Progress questions

Question 3

The suffix used to indicate a carbon to carbon double bond is

A. '-ene'.

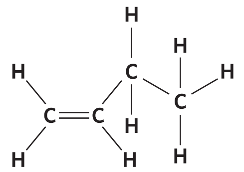
B. '-ane'.

Question 4

The name of the given molecule is

A. but-1-ene.

B. but-2-ene.



Naming branched hydrocarbons 1.2.6.3

Some hydrocarbons have branches associated with the parent chain and therefore require a more in-depth naming system. This results in compounds, called **structural isomers**, which have the same molecular formula but a different structural formula and therefore, a different systematic name.

How are branched hydrocarbons named?

Hydrocarbons can contain side groups/side chains, known as **alkyl groups**, which are composed of carbon and hydrogen atoms, much like the parent chain of the hydrocarbon. They can sometimes also be referred to as branches. These groups can be found bonded to carbons within the parent chain of a hydrocarbon.

The presence of an alkyl group is shown through a prefix added to the name of a hydrocarbon. Similar to the conventions used to name alkanes, alkyl groups have different names depending on the number of carbon atoms in the alkyl group. This is summarised in table 2.

Table 2 Names of common alkyl groups

Alkyl group	Number of carbons in the group	Parent name	Name of alkyl group
-CH_3	1	meth-	methyl
-CH_2CH_3	2	eth-	ethyl
-$CH_2CH_2CH_3$	3	prop-	propyl
-$CH_2CH_2CH_2CH_3$	4	but-	butyl
-$CH_2CH_2CH_2CH_2CH_3$	5	pent-	pentyl

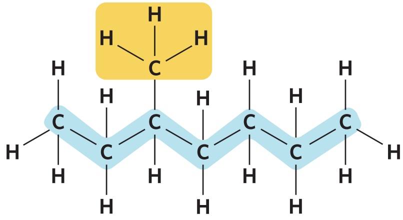


Figure 7 Identification of the parent chain in a hydrocarbon

When naming hydrocarbon chains with alkyl groups, we must first identify the number of carbons on the parent chain. Consider the molecule shown in figure 7.

The longest carbon chain, shown in blue, contains seven carbons. We can also see that the parent chain contains only carbon to carbon single bonds. As such, we can identify that the name of the hydrocarbon is heptane. However, there is another portion of the molecule that we are yet to include in the name of our molecule; the alkyl group in orange.

As we can see, the alkyl group contains only one carbon and therefore, as can be seen in table 2, this would be called a methyl group. With this, we can update the name of this compound to methylheptane.

The final piece of information that is missing is where the alkyl group is found. When naming compounds, we need to be able to signal the exact location of side groups within a molecule. Similar to the way we numbered the carbons in the parent chain when naming unsaturated hydrocarbons, we need to number the carbons in a way that allows the alkyl group to occur at the lowest numbered carbon.

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As can be seen in figure 8, the methyl group projects out from position 3 on the parent chain.

We can therefore identify the name of the compound as ‘3-methylheptane’, as shown in figure 9.

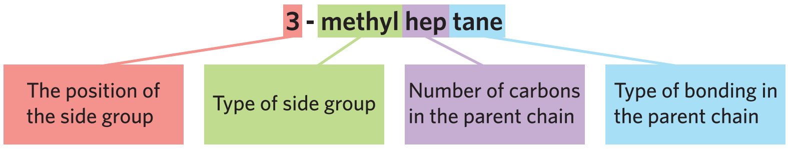


Figure 9 Naming convention of branched hydrocarbons

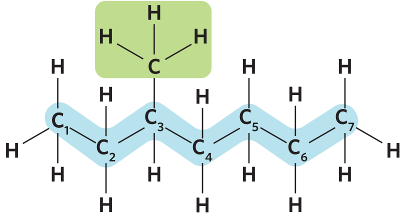
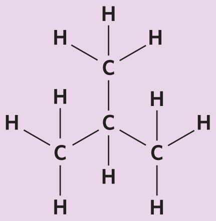


Figure 8 Numbering of the parent chain in a hydrocarbon with a side group

WORKED EXAMPLE 3

Identify the systematic name of the following molecule and a structural isomer of this molecule.



What information is presented in the question?

The structural formula for a molecule.

What is the question asking us to do?

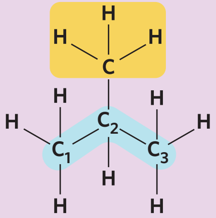
Provide the systematic name of the molecule and an isomer of this molecule.

What strategies do we need in order to answer the question?

1. Determine the length of the longest carbon chain.
2. Determine the type of bonding within the parent carbon chain.
3. Identify any side groups present.
4. Identify the carbon number to which the side group is bonded.
5. Name the molecule using IUPAC naming conventions.
6. Name an isomer of this molecule using IUPAC naming conventions.

Answer

The longest carbon chain contains three carbons, and the only bonds within the parent chain are carbon to carbon single bonds. This means that this molecule is an alkane with a parent name of ‘prop-’.



There is a methyl side group found on carbon number 2.

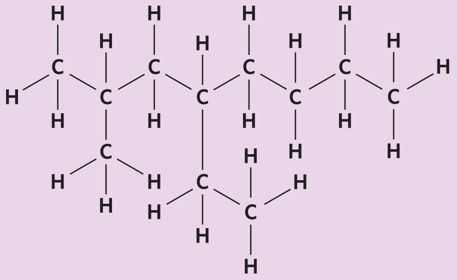
The name of this molecule is 2-methylpropane and the only isomer of this molecule is butane which also has the molecular formula C_4H_{10}.

Sometimes, these compounds will have more than one side group bonded to the parent chain.

When this occurs, the same rules apply. The only difference is that the alkyl groups are named in alphabetical order, as seen in worked example 4.

WORKED EXAMPLE 4

Use the IUPAC naming system to determine the name of the molecule provided.



What information is presented in the question?

The structural formula for a molecule.

What is the question asking us to do?

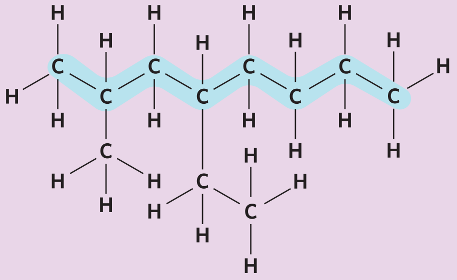
Provide the systematic name of the molecule provided.

What strategies do we need in order to answer the question?

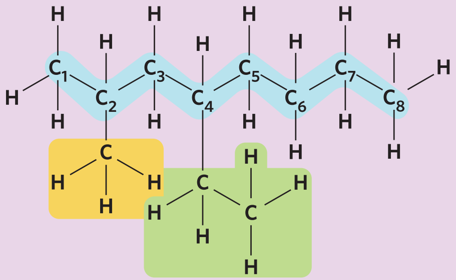
1. Determine the longest carbon chain length.
2. Determine the type of bonding within the parent carbon chain.
3. Identify any side groups/double bonds present.
4. Identify the carbon number to which the side group/double bond is bonded.
5. Name the molecule using IUPAC naming conventions.

Answer

The longest carbon chain contains eight carbons and within the parent chain is only carbon to carbon single bonds. This means that this molecule is an alkane with a parent name of 'oct-'.



There are two alkyl groups present in the molecule; one with 1 carbon and one with 2 carbons. As such, this molecule contains methyl and ethyl side groups.



Based on the numbering of the carbons in the parent chain, we can see that the methyl group is bonded to carbon number 2 and the ethyl group is bonded to carbon number 4. 'E' comes before 'm' in the alphabet, so ethyl is named before methyl.

Therefore, the name of this molecule is 4-ethyl-2-methyloctane. This name allows for the side groups to be located at the lowest carbon numbers.

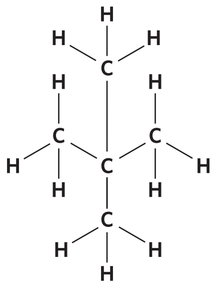


Figure 10 Structural formula of a hydrocarbon with multiple side groups

At times, there may be multiple of the same side groups present on the same hydrocarbon. When this is the case, we use prefixes such as di/tri to indicate multiple of the same groups, with di- indicating two of the side group and tri-indicating three of the side group being present. For example, consider the molecule shown in figure 10.

MISCONCEPTION

'Dimethyl would come before ethyl in the name.'

When using "di", "tri" or "tetra" this prefix is ignored when putting side chains in alphabetical order - e.g. dimethyl would come after ethyl as 'di' is ignored.

Based on our understanding of the naming conventions, we can see that the parent chain contains three carbons. The molecule also contains two of the same side groups located on the same carbon as shown in figure 11.

Since the side groups contain only one carbon, they are both methyl groups. The name of this hydrocarbon is shown in figure 12.

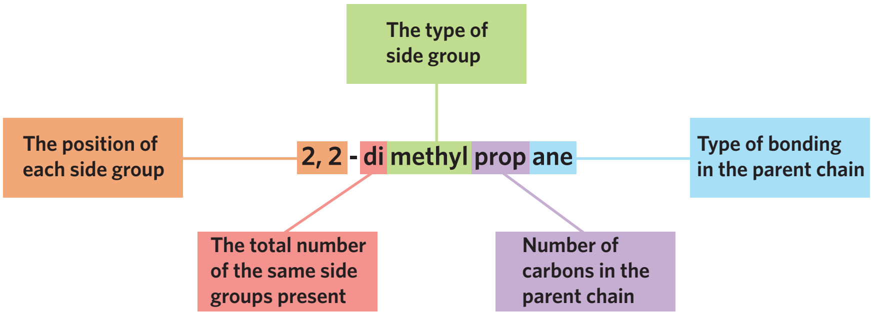


Figure 12 Naming convention of a hydrocarbon containing multiple side groups

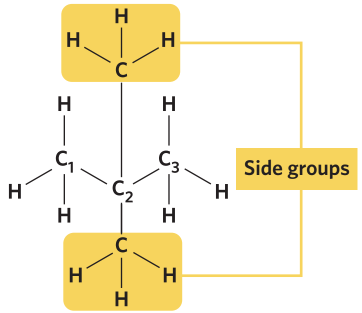


Figure 11 Identification of the side groups found in a branched hydrocarbon

As we can see, the position of every side group present needs to be indicated in the name of the compound. Given that this hydrocarbon has five carbons it can also have two other structural isomers (pentane and 2-methylbutane).

USEFUL TIP

Whenever two or more numbers are adjacent in a systematic name, they should be separated from each other by a comma ','.

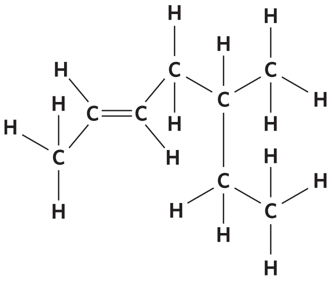
Progress questions

Question 5

The name of the side group in this molecule is

A. ethyl.

B. methyl.



Question 6

Which of the following pairs of molecules could be structural isomers?

A. 2-methylbutane and 2,2-dimethylpropane

B. 2-methylpropane and pentane

Naming complex hydrocarbons 1.2.6.4

Some hydrocarbons have multiple different components that need to be indicated by the name of the compound.

How are complex hydrocarbons named?

Some hydrocarbons can be unsaturated as well as containing branches. Consider the compound in figure 13.

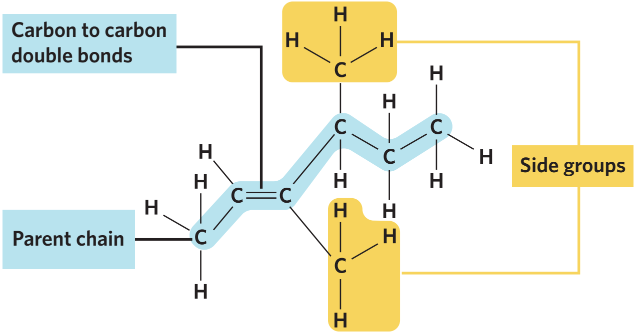


Figure 13 Structural formula of a compound

The parent chain contains five carbon atoms, one carbon to carbon double bond, and two methyl groups associated with the parent chain.

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In previous examples, we have numbered the carbons in the carbon chain based on the location where the side groups can be found. However, in this example, there is a carbon to carbon double bond in addition to the side groups. Double bonds take priority over side groups when naming, so the chain should be numbered to allow the double bond to occur at the lower numbered carbon.

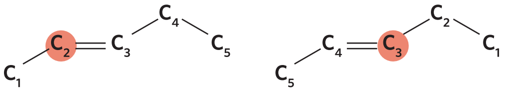


Figure 14 Assigning the lowest number of the chain - 2 wins

If we just consider the parent chain, we can number the carbons from both directions as shown in figure 14. One naming system allows the carbon to carbon double bond to occur at carbon 2, whilst the other at carbon 3. As such, we will choose the numbering system where the carbon to carbon double bond occurs at carbon 2 as shown in figure 15.

Numbering this way, the two methyl groups present in the molecule fall on carbon numbers 3 and 4. As a result, the name of this hydrocarbon would be 3,4-dimethylpent-2-ene, as shown in figure 16.[1]

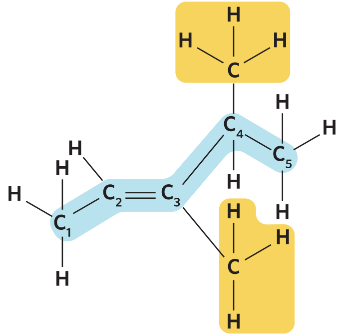


Figure 15 Breakdown of a structural formula

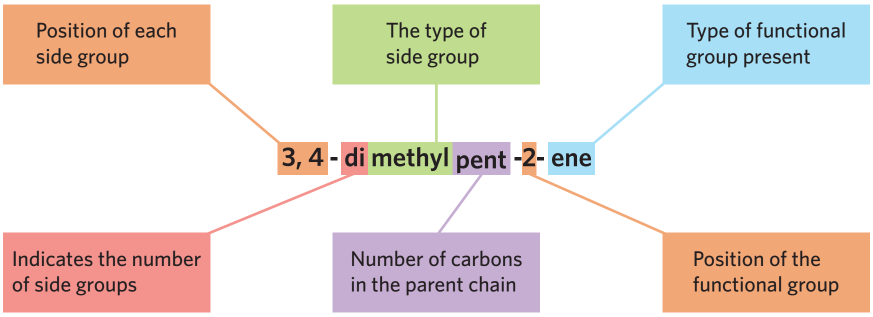


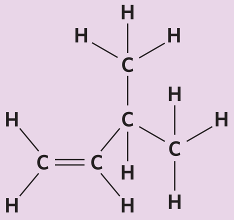
Figure 16 Breakdown of the naming of complex organic compounds

KEEN TO INVESTIGATE?

[1] **Where can I get more practice at naming complex alkanes?**
Search YouTube: Alkane naming quiz

WORKED EXAMPLE 5

Identify the IUPAC name for the organic compound below and give the name of a possible structural isomer.



What information is presented in the question?

The structure of a hydrocarbon.

What is the question asking us to do?

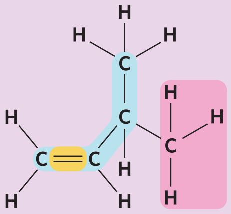
Provide the systematic name of the molecule and an isomer of this molecule.

What strategies do we need in order to answer the question?

1. Identify the number of carbons in the parent chain.
2. Identify the type(s) of bonds in the parent chain.
3. Identify any branches.
4. Number the carbons in the parent chain.
5. Apply this information using correct IUPAC naming conventions.
6. Name an isomer of this molecule using IUPAC naming conventions.

Continues →

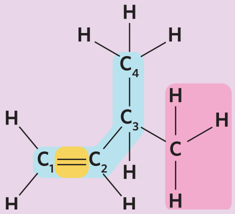
Answer



As shown by the blue line, the parent chain contains four carbons, hence will have a prefix of 'but-'.

Shown in the orange line is a carbon to carbon double bond, making this compound an alkene. Therefore the suffix for this compound is '-ene'.

Shown in the pink is a branch/alkyl group containing one carbon. Therefore the alkyl's name is methyl.



Based on the numbering system that allows the carbon to carbon double bond to occur at the lowest numbered carbon, the carbon to carbon double bond occurs at carbon 1 and the methyl group occurs at carbon 3.

Therefore the name of this organic compound is 3-methylbut-1-ene.

Possible isomers include: 2-methylbut-1-ene, 2-methylbut-2-ene, pent-1-ene and pent-2-ene.

Progress questions

Question 7

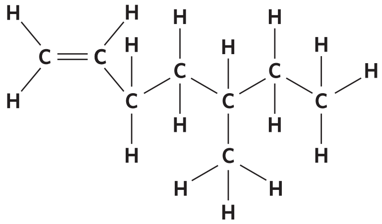
When numbering carbons in an unsaturated hydrocarbon,

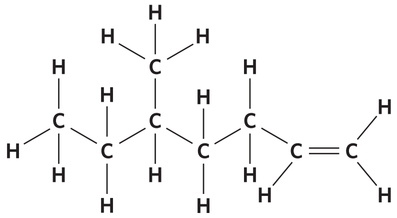
A. alkyl groups are prioritised.

B. carbon to carbon double bonds are prioritised.

Question 8

The molecules shown have





A. the same IUPAC names.

B. different IUPAC names.

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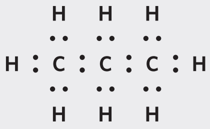
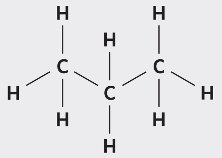
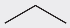
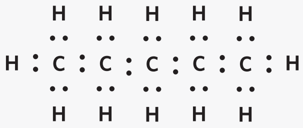
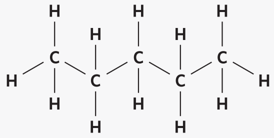
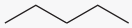
Representing hydrocarbons 1.2.6.5

Hydrocarbons can be represented using different types of formulas.

How can hydrocarbons be represented in different ways?

There are a variety of ways to write hydrocarbon compound formulas. These are shown in table 3.

Table 3 Different representations of hydrocarbons

Hydrocarbon	Molecular formula	Semi-structural formula	Electron dot formula	Structural formula	Skeletal formula
Propane	C_3H_8	$CH_3CH_2CH_3$			
Pentane	C_5H_{12}	$CH_3CH_2CH_2CH_2CH_3$			

USEFUL TIP

Skeletal formulas are not in the year 11 study design, however they will be used extensively in year 12 (and beyond).

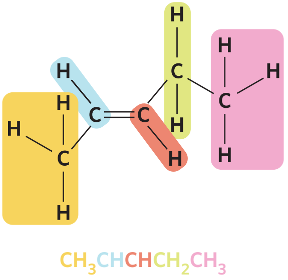


Figure 17 Relationship between structural and semi-structural formulas

We have come across most of these formulas in previous lessons, particularly structural formulas. Although semi-structural formulas seem long, they are very useful in helping to identify the structural formula of the compound as shown in figure 17.

As we can see, the order in which the semi-structural formula is written reflects the way in which the structural formula would be drawn. Though it doesn't include the single bonds in the structure, it may sometimes include any double bonds present (e.g. $CH_2 = CH_2$). If the structure contains a branch, the alkyl group is generally written in brackets.

Consider the molecule shown in figure 18. The semi-structural formula would be written as $CH_3CH(CH_3)CH_2CH_3$, where the methyl group is written in brackets.

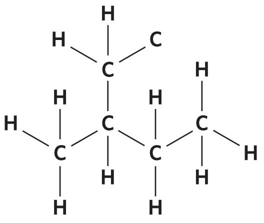


Figure 18 Structural formula for 2-methylbutane

USEFUL TIP

When brackets are included in a semi-structural formula, $CH_3CH(CH_3)CH_2CH_3$, this indicates that this group is branching from the main carbon chain at carbon number 2.

In a skeletal formula, as shown in figure 19, the explicit notation of carbon atoms and hydrogen atoms is removed from the structure, leaving only the backbone of the organic molecule. When interpreting this type of formula, we can imagine the lines as carbon to carbon bonds, and the intersections or ends of lines as carbon atoms. Double bonds can be represented by a double line. The number of hydrogen atoms present can be deduced from our understanding of carbon atoms' ability to form 4 covalent bonds each.

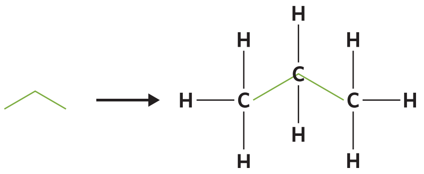


Figure 19 The relationship between skeletal and structural formulas

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WORKED EXAMPLE 6

What is the systematic name of the molecule with the following skeletal formula?

What information is presented in the question?

The skeletal formula for a molecule.

What is the question asking us to do?

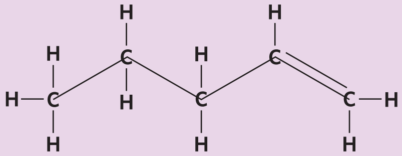
Provide the systematic name of the molecule provided.

What strategies do we need in order to answer the question?

1. Determine the longest carbon chain length.
2. Determine the type of bonding within the parent carbon chain.
3. Identify any side groups/double bonds present.
4. Identify the carbon number to which the side group/double bond is found.
5. Name the molecule.

Answer

If we imagine the full structural formula of this molecule, it would look like this:



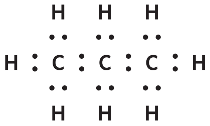
The parent chain contains five carbons. There is also a carbon to carbon double bond located at carbon number 1.

The systematic name of this molecule is pent-1-ene.

Progress questions

Question 9

What type of organic formula is shown?

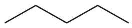


A. Skeletal formula

B. Electron dot formula

Question 10

The semi-structural formula for the molecule shown is



A. $CH_3CH_2CH_2CH_2CH_3$.

B. $CH_3CH_2CH_2CH_3$.

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Theory summary

- There are many different aspects of a molecule and observations that must be taken into account when naming hydrocarbons. The following is a summary of all of the steps that should be taken when naming a complex hydrocarbon:

1. Identify the longest unbranched carbon chain, and count the number of carbon atoms to identify the parent name. If the parent chain contains carbon to carbon double bonds, this must be included in the parent chain.
2. Identify the type of bonding between the carbon atoms within the parent chain to determine the suffix of the hydrocarbon.
3. Identify the side group(s) present (if any) and hence the prefix of the name.
4. Number the positions of the carbons in the parent chain. This will be affected by the relative priority of any carbon to carbon double bonds and side groups present.
5. Identify the number and position(s) of any present double bonds.
6. Identify the position(s) and name(s) of the side group(s), using the prefix 'di-/tri-' in the case that it is present multiple times.
7. Combine the above to write the complete name of the organic compound. Ensure that commas are used to separate adjacent numbers and that hyphens are used to separate words and numbers.

The content in this lesson is considered fundamental prior knowledge to structure, nomenclature and properties of organic compounds (Unit 4 AOS 1).

7B Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 11–13.

Dorian has an interest in oil paints, and wishes to better understand the hydrocarbons found in them. In order to achieve this, he must understand the naming conventions of hydrocarbons, as designed by The International Union of Pure and Applied Chemistry.

Question 11 (1 MARK) 🌶

Which of the following options has the parent names associated with the correct number of carbon atoms in the carbon chain?

A. Pent-: 5, Oct-: 8, Prop-: 4

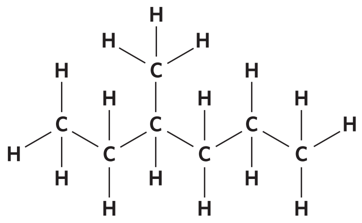
B. Pent-: 3, Oct-: 8, Prop-: 5

C. Pent-: 3, Oct-: 7, Prop-: 5

D. Pent-: 5,: Oct-: 8, Prop-: 3

Question 12 (1 MARK) 🌶

Which of the following is the correct IUPAC name of the molecule shown?



A. 2-ethylpent-4-ene

B. 4-ethylpent-1-ene

C. 4-methylhexane

D. 3-methylhexane

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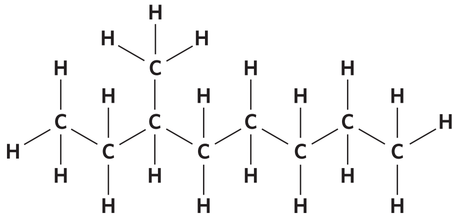
Question 13 (3 MARKS)

Having learnt about the naming of hydrocarbons, Dorian names the following molecule as 2-ethylheptane.

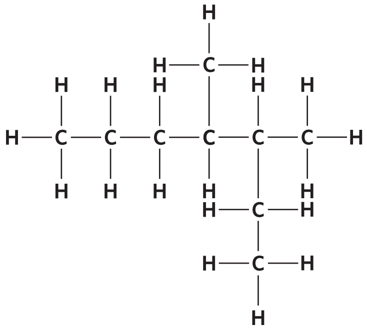
Explain why Dorian's suggested name for this molecule is incorrect, and provide the correct name.



Exam-style

Question 14 (1 MARK)

What is the correct systematic name for the compound shown?



A. 4-methyl-5-ethylhexane

B. 2-ethyl-3-methylhexane

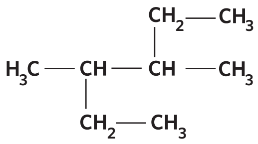
C. 4,5-dimethylheptane

D. 3,4-dimethylheptane

VCAA 2016 exam Multiple choice Q2

Question 15 (1 MARK)

What is the correct systematic name for the compound shown?



$\mathrm{H_3C{-}\underset{\displaystyle CH_2{-}CH_3}{CH}{-}\overset{\displaystyle CH_2{-}CH_3}{CH}{-}CH_3}$

A. 2-ethyl-3-methylpentane

B. 3-methyl-4-ethylpentane

C. 3,4-dimethylhexane

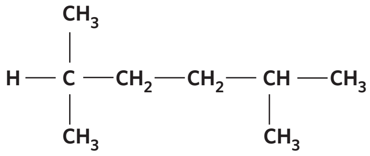
D. 2,3-diethylbutane

VCAA 2010 Exam 1 Multiple choice Q10

Question 16 (1 MARK)

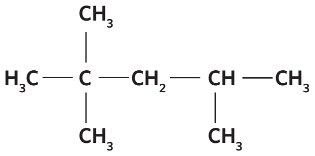
Consider the following structures.

I.



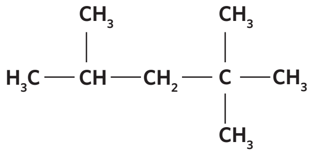
$\mathrm{H{-}\overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{C}}{-}CH_2{-}CH_2{-}\underset{\displaystyle CH_3}{CH}{-}CH_3}$

II.



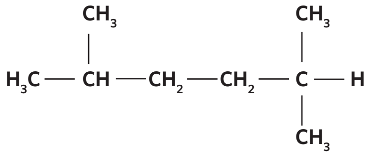
$\mathrm{H_3C{-}\overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{C}}{-}CH_2{-}\underset{\displaystyle CH_3}{CH}{-}CH_3}$

III.



$\mathrm{H_3C{-}\overset{\displaystyle CH_3}{CH}{-}CH_2{-}\overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{C}}{-}CH_3}$

IV.



$\mathrm{H_3C{-}\overset{\displaystyle CH_3}{CH}{-}CH_2{-}CH_2{-}\overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{C}}{-}H}$

Which of these structures is that of 2,2,4-trimethylpentane?

A. I and III only **B.** I and IV only **C.** II and III only **D.** II and IV only

VCAA 2011 Exam 1 Multiple choice Q3

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Question 17 (7 MARKS)

Lucinda is a year 11 chemistry student learning about the structures of hydrocarbons. Having learned the systematic IUPAC naming conventions of hydrocarbons, she now needs to connect this information to the ability to draw structural formulas of hydrocarbons.

a. Draw the structural formulas of the following hydrocarbons:

  i. But-2-ene (1 MARK)

  ii. 3-methylpentane (1 MARK)

  iii. 2,3-dimethylheptane (1 MARK)

b. While learning about the naming conventions of hydrocarbons, Lucinda came up with the molecule '6-methyloctane', however she has her doubts about this molecule. Explain why the molecule 6-methyloctane does not exist. (2 MARKS)

c. What is the correct name of the molecule Lucinda was attempting to name? (1 MARK)

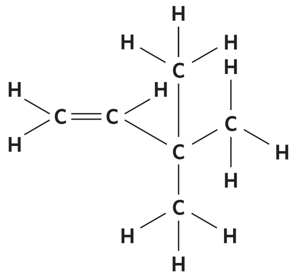
d. Draw the structural formula of the molecule Lucinda was attempting to name. (1 MARK)

Question 18 (4 MARKS)

Give the systematic name for each of the following compounds:

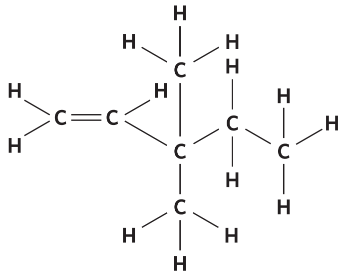
a. $CH_3CH_2CHC(CH_3)CH_3$ (1 MARK)

b.



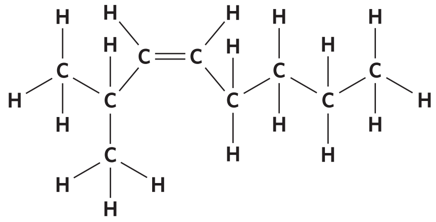
(1 MARK)

c.



(1 MARK)

d.



(1 MARK)

Key science skills

Question 19 (6 MARKS)

Theo is performing chemical analysis on a sample of octane. As part of his analysis, he wanted to assess the properties of it, including boiling point, viscosity and heat of combustion.

a. Upon calculating the boiling point of octane multiple times, Theo's results were not what he expected. Although all his measurements were close to each other, none of them were close to the boiling point he expected.

  i. Comment on the accuracy and precision of the results. (2 MARKS)

  ii. What type of error do these results indicate? Explain your reasoning. (2 MARKS)

b. Calculating the boiling point of a substance involves the use of an open flame. Suggest two safety precautions that should be taken to prevent injury. (2 MARKS)

FROM LESSONS 16B & 16D

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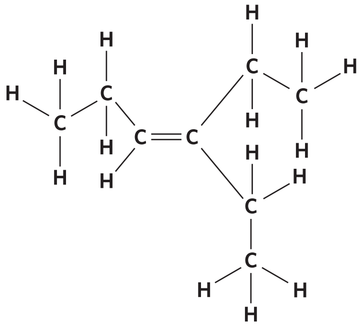
Questions from multiple lessons

Question 20 (5 MARKS)

Compared to many of the other isomers of this molecule, this structural isomer has a relatively low melting point.



a. Determine the molar mass for the molecule shown. (1 MARK)

b. Give the systematic name of the molecule shown. (1 MARK)

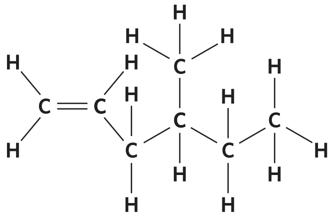
c. Would this molecule be considered a saturated or unsaturated hydrocarbon? Explain. (2 MARKS)

d. Why does this molecule have a lower melting point than its other isomers? (1 MARK)

FROM LESSONS 2B, 6B & 7A

Question 21 (5 MARKS)

Alkenes are able to take part in addition reactions with hydrogen gas.



a. Based on the structure of the molecule, what type(s) of intermolecular forces would need to be disrupted in order to change the substance from a liquid to gaseous form? Explain. (3 MARKS)

b. Give the semi-structural formula for the molecule shown. (1 MARK)

c. In a reaction with hydrogen, a total of 0.122 mol of the compound was used during the reaction. Calculate the mass, in grams, of the compound used. (1 MARK)

FROM LESSONS 2B & 6B

Hints

14.	The parent chain does not have to be written linearly.
15.	The parent chain does not have to be written linearly.
16.	The first step in systematic naming is to determine the number of carbons in the parent chain.
17a.	Each section of a hydrocarbon's name identifies a key structural feature of the molecule.
17b.	The position of a branch is determined based on the shortest distance from the end to main carbon.
17c.	The position of a branch is determined based on the shortest distance from the end to main carbon.
17d.	Structural formulas show all covalent bonds.
18.	The longest consecutive carbon chain is the parent chain.
19ai.	Accuracy and precision describe different qualities of a test's results.
19aii.	Precise results are all similar to each other.
19b.	An open flame can cause safety concerns.
20a.	The molar mass is the sum of the masses of atoms in a molecule.
20b.	Like carbon to carbon double bonds, the position of alkyl groups is on the lowest numbered carbon possible.
20c.	Saturation of hydrocarbons depends on the type of bonding in the parent chain.
20d.	Surface area determines melting points in hydrocarbons.
21a.	Intermolecular forces depend on the type of covalent bonding present within the molecule.
21b.	Semi-structural formulas do not include bonds.
21c.	The mass of the compound present depends on the amount of substance available.

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7C Functional groups

STUDY DESIGN DOT POINTS

- the grouping of hydrocarbon compounds into families (alkanes, haloalkanes, alkenes, alcohols, carboxylic acids) based upon similarities in their physical and chemical properties, including general formulas and general uses based on their properties
- representations of organic compounds (structural formulas, semi-structural formulas) and naming according to the International Union of Pure and Applied Chemistry (IUPAC) systematic nomenclature (limited to non-cyclic compounds up to C8, and structural isomers up to C5)

7A 7B **7C** 7D 7E

1.2.5.3 Haloalkanes

1.2.5.4 Alcohols

1.2.5.5 Carboxylic acids

1.2.6.6 Naming haloalkanes

1.2.6.7 Naming alcohols

1.2.6.8 Naming carboxylic acids

ESSENTIAL PRIOR KNOWLEDGE

1C Electronegativity
7A Hydrocarbon structure
7B Hydrocarbon nomenclature
See questions 41–43.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

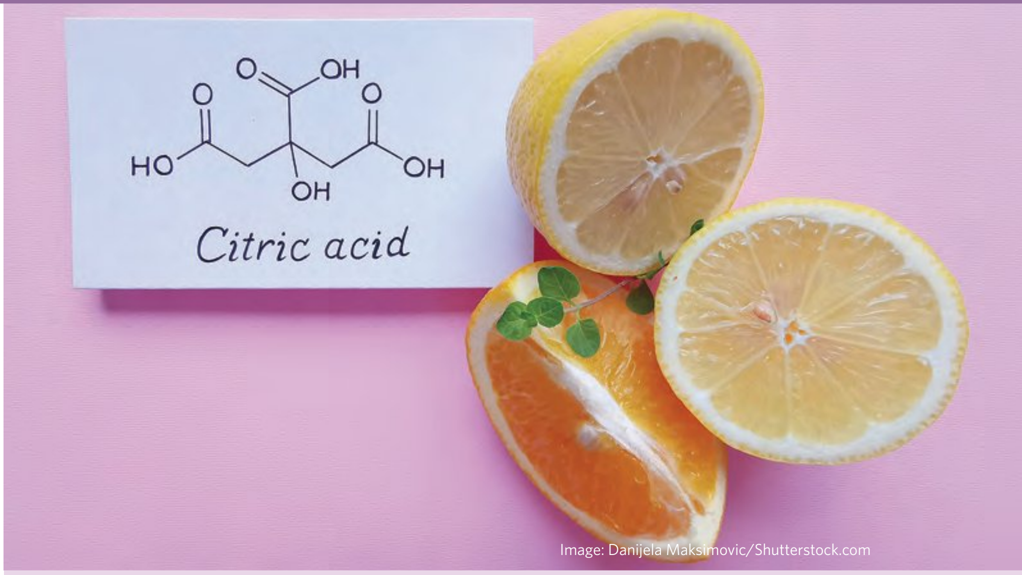

Image: Danijela Maksimovic/Shutterstock.com

How do functional groups determine the physical and chemical properties of compounds?

In the last lesson, we discussed the basic structure of hydrocarbons, including naming conventions as well as variations in bonding. These hydrocarbons can further diversify with the addition of **functional groups.** Many of these functional groups give compounds predictable properties.

KEY TERMS AND DEFINITIONS

Alcohol organic compound that has a hydroxyl (–OH) group bonded to a carbon atom in the parent chain

Carboxylic acid organic compound that contains a carboxyl functional group (–COOH)

Carboxyl group functional group that has a carbon double bonded to an oxygen and singly bonded to a hydroxyl group

Functional group specific groups of atoms within a compound that affect the properties of the compound

Haloalkane organic compound that has one or more halogen atom bonded to a carbon atom in the carbon chain

Halogen element in group 17 of the periodic table (fluorine, chlorine, bromine, iodine)

Hydroxyl group functional group consisting of an oxygen atom covalently bonded to a hydrogen atom (–OH)

Haloalkanes 1.2.5.3

Haloalkanes are organic compounds that contain at least one **halogen** atom.

What are haloalkanes?

In this lesson, we will be discussing various functional groups, starting with the halogen group, which produces **haloalkanes**. In haloalkanes, a halogen atom (F, Cl, Br, I) is covalently bonded to a carbon atom, as shown in figure 1 and table 1. In VCE chemistry, this functional group is commonly denoted as -X.

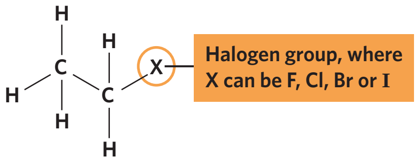

Figure 1 An example of a haloalkane

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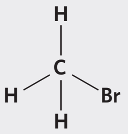
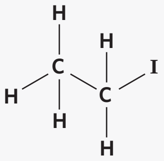
Table 1 Representations of haloalkanes

Systematic name	Molecular formula	Structural formula	Semi-structural formula
Bromomethane	CH_3Br		CH_3Br
Iodoethane	C_2H_5I		CH_3CH_2I

How does the structure of haloalkanes affect their properties?

As mentioned in lesson 1C, halogens are highly electronegative elements. The difference in electronegativity between halogen atoms and carbon atoms in haloalkanes causes a polar covalent bond to be formed between them, and hence allows for permanent dipole-dipole interactions between haloalkane molecules, as illustrated in figure 2. As a result, haloalkanes generally have a higher boiling point than their corresponding alkanes or alkenes, which only have weak dispersion forces present between molecules.

It is important to note that while haloalkanes are able to take part in stronger intermolecular forces than their counterpart alkanes, as the length of the nonpolar hydrocarbon section of the haloalkane increases, the effect of the halogen on the intermolecular force strength decreases.

Haloalkanes are somewhat soluble in water, as they are able to form dipole-dipole interactions with water molecules. However, due to the strength of the C-X bond, they do not dissolve in water as easily as some other organic molecules, which will be discussed later in this lesson.

Due to the high electronegativity of halogens, haloalkanes are relatively reactive molecules, and are commonly used in chemical manufacturing processes. Halogens are readily added to and removed from hydrocarbons, and are hence useful as an 'intermediate' in chemical reactions, allowing for the conversion of one molecule to another.

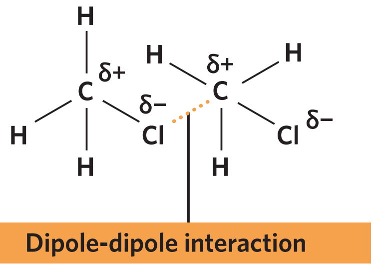


Figure 2 Dipole-dipole interactions between haloalkanes

Progress questions

Question 1

Haloalkanes are often used as ____________ in chemical processes due to their reactivity.

A. intermediates

B. products

Question 2

The boiling points of haloalkanes are generally ____________ than those of alkanes.

A. lower

B. higher

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Naming haloalkanes 1.2.6.6

As haloalkanes are structurally unique from other hydrocarbon compounds, they have their own naming conventions.

How are haloalkanes named?

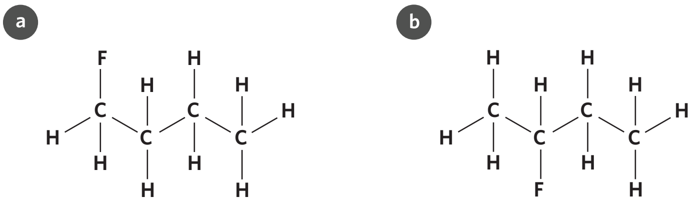
When naming a haloalkane, a prefix depicting a given halogen is added onto the hydrocarbon parent name. Each halogen has a designated form of its original name which is used as a prefix in the haloalkane naming process, as shown in table 2.

Table 2 Halogen prefixes

Halogen	Prefix when naming haloalkanes
Fluorine	'fluoro-'
Chlorine	'chloro-'
Bromine	'bromo-'
Iodine	'iodo-'

For example, a saturated hydrocarbon with four carbon atoms has the parent alkane name of butane. If there was a fluorine atom attached to this hydrocarbon, making it a haloalkane, the prefix 'fluoro-' would be added, leading to the name 'fluorobutane'.

However, for haloalkanes with more than two carbon atoms, it is also important to determine which carbon atom the halogen is bonded to. As such, the position of the functional group is signified by placing the carbon number it is attached to before the name, separated by a hyphen. Using the example mentioned, if the fluorine group was bonded to the first carbon atom, this molecule would be 1-fluorobutane, whereas if it were bonded to the second carbon atom it would be 2-fluorobutane. These structural isomers are shown in figure 3.



1-fluorobutane

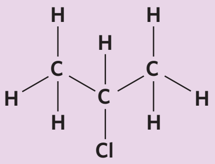
2-fluorobutane

Figure 3 Two structural isomers of C_4H_9F are 1-fluorobutane **(a)** and 2-fluorobutane **(b)**.

As always, we want to number the carbon chain such that the functional group occurs at the lowest numbered carbon. In figure 3a, the halogen group could be attached to the 1st carbon atom (if counting from left to right) or the 4th carbon atom (if counting from right to left). Since the halogen is the only functional group present, we want it to have the lowest numbered position possible and therefore, the molecule is called 1-fluorobutane and not 4-fluorobutane.

WORKED EXAMPLE 1

Identify the IUPAC name for the organic compound shown and give the name of a possible structural isomer.



What information is presented in the question?

The structural formula for a molecule.

What is the question asking us to do?

Provide the systematic name of the molecule and an isomer of this molecule.

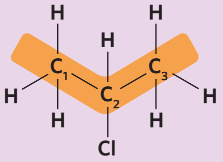
Continues →

What strategies do we need in order to answer the question?

1. Identify the number of carbons in the parent chain.
2. Identify the type(s) of bonds in the parent chain.
3. Identify any branches.
4. Number the carbons in the parent chain.
5. Apply this information using correct IUPAC naming conventions.
6. Name an isomer of this molecule using IUPAC naming conventions.

Answer

This parent chain contains 3 carbon atoms, all of which are bonded by single bonds.



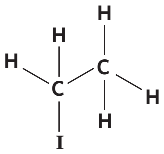
The parent alkane name is propane for a 3-carbon chain. The prefix ‘2-chloro-’ is used to indicate a chlorine atom bonded to the 2nd carbon atom.

The name of this molecule is 2-chloropropane and the only isomer is 1-chloropropane.

Progress questions

Question 3

What is the systematic name of the following molecule?



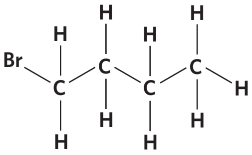
A. 1-iodoethane

B. Iodoethane

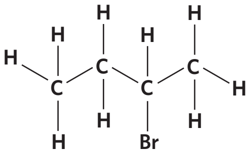
Question 4

The molecule 2-bromobutane has the structural formula:

A.



B.



USEFUL TIP

For molecules with only one or two carbons in the parent chain, the assignment of numbers is not necessary, as we can deduce the position of a functional group without specific notation.

Alcohols 1.2.5.4

Alcohols are organic compounds that contain at least one **hydroxyl group**.

What are alcohols?

Alcohols are organic compounds that contain a hydroxyl functional group. This hydroxyl functional group is made up of an oxygen atom covalently bonded to a hydrogen atom, denoted as –OH, and this replaces a hydrogen atom on the carbon chain. An example of an alcohol is shown in figure 4, as well as different representations of alcohols in table 3.

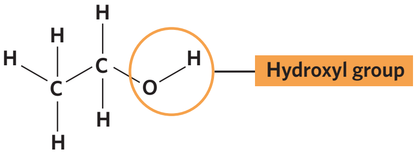


Figure 4 Structure of ethanol, an alcohol

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MISCONCEPTION

'An "alcohol" is a functional group.'

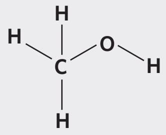
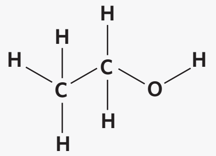
The term 'alcohol' describes the whole molecule. The functional group that defines an alcohol is called a hydroxyl group.

MISCONCEPTION

'When drawing the structural formula you do not need to show the bond between the O and H in the hydroxyl group.'

Structural formulas must show every bond in the molecule.

Table 3 Representations of alcohols

Systematic name	Molecular formula	Structural formula	Semi-structural formula
Methanol	CH_3OH		CH_3OH
Ethanol	C_2H_5OH		CH_3CH_2OH

How does the structure of alcohols affect their properties?

The presence of a hydroxyl group in an alcohol molecule results in the boiling point of alcohols being higher than that of hydrocarbons. Due to the difference in electronegativity between oxygen and hydrogen, the bond that occurs between the two atoms in the hydroxyl group is considered a polar covalent bond. Consequently, a hydroxyl group of an alcohol is able to take part in permanent dipole-dipole interactions as well as hydrogen bonding with a hydroxyl group from another alcohol. Such hydrogen bonding is shown in figure 5.

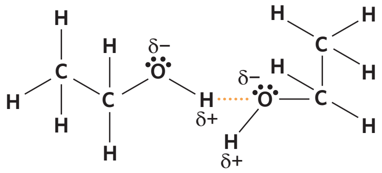


Figure 5 Hydrogen bonding between two ethanol molecules

Due to the fact that dipole-dipole interactions and hydrogen bonds are relatively strong intermolecular forces, a higher amount of energy is needed to overcome these forces compared to its corresponding alkane, which is only held together by weaker dispersion forces. Since the intermolecular forces between alcohols are stronger, the boiling points of alcohols are higher than their corresponding alkanes as more energy is required to overcome the intermolecular forces between alcohols than those of the corresponding alkane. This relationship can be seen in figure 6.

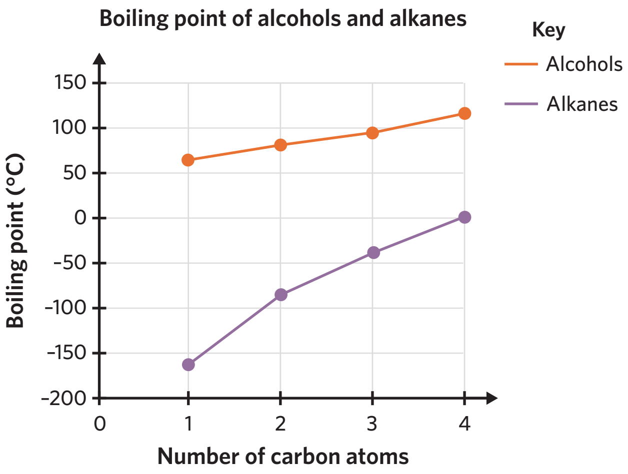


Figure 6 Boiling point of alkanes and alcohol

As with haloalkanes, it is important to note that as the non-polar section of the hydrocarbon chain increases, the effect of the hydroxyl group on the boiling point decreases. This phenomenon can be seen in figure 6 – as the number of carbon atoms in the molecule increases, the difference in boiling point between the alcohols and the alkanes decreases. This shows the diminishing influence of the hydroxyl functional group on the boiling point as the size of the non-polar section grows.

While the presence of a hydroxyl group has a significant influence on the intermolecular interactions between alcohol molecules, this also impacts the interactions of alcohols with other types of molecules. The hydroxyl group enables alcohols to form dipole-dipole interactions with other polar molecules, as well as strong hydrogen bonds with water molecules as shown in figure 7, enabling alcohols to dissolve in water.

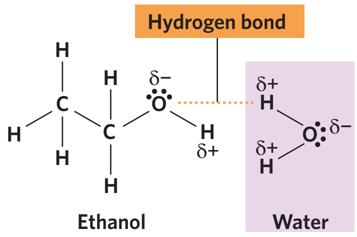


Figure 7 Intermolecular forces formed between water molecules and ethanol

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Similarly to the diminishing influence of the hydroxyl group on boiling point, as the carbon chain length of an alcohol increases, the solubility of the molecule in water also decreases. This is because the molecule becomes less polar as the non-polar carbon chain length increases, therefore reducing its ability to form hydrogen bonds with water molecules.

There are many different alcohols that exist, all of which have different uses, however the primary use of alcohols in modern society tends to be as a fuel. Bioethanol, a form of ethanol produced from the fermentation of glucose, is often added to petrol due to it being renewable. This concept will be covered in greater detail in Unit 3 of VCE chemistry.[1]

KEEN TO INVESTIGATE?

[1] **How are alcohols made and used?** Search YouTube: Making Alcohols by Fermentation and from Ethane

Progress questions

Question 5

The functional group of an alcohol consists of oxygen and hydrogen.

A. True

B. False

Question 6

As the length of a carbon chain increases, the influence of a hydroxyl group on the boiling point of the molecule

A. increases.

B. decreases.

Naming alcohols 1.2.6.7

As alcohols are structurally unique from other hydrocarbon compounds, they have their own naming conventions.

How are alcohols named?

When naming an alcohol, the ‘e’ at the end of the alkane’s name is removed and is replaced by ‘ol’.

For example, a saturated hydrocarbon with three carbon atoms, as shown in figure 8, has the parent alkane name of ‘propane’. If there was a hydroxyl group attached to this hydrocarbon, making it an alcohol, the ‘e’ at the end of ‘propane’ would be replaced with ‘ol’, leading to the name ‘propanol’.

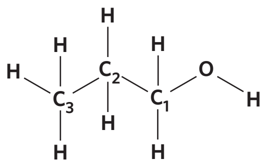


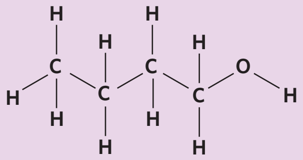
Figure 8 Structural formula of propan-1-ol

As with haloalkanes, we must signify which carbon the hydroxyl group is bonded to (for parent chains with more than 2 carbons). Again, we want to number the carbon chain such that the functional group occurs at the lowest numbered carbon.

The position of the hydroxyl group is signified by placing the carbon number it is attached to, in hyphens, between the parent alkane’s name and the ‘-ol’. Using the example above, this molecule would be ‘propan-1-ol’.

WORKED EXAMPLE 2

What is the systematic name of the following molecule and what are the names of two other structural isomers of this molecule?



What information is presented in the question?

The structural formula for a molecule.

What is the question asking us to do?

Provide the systematic names of the molecule and two isomers of this molecule.

Continues →

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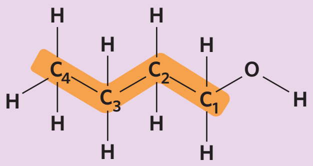
What strategies do we need in order to answer the question?

1. Identify the number of carbons in the parent chain.
2. Identify the type(s) of bonds in the parent chain.
3. Identify any branches.
4. Number the carbons in the parent chain.
5. Apply this information using correct IUPAC naming conventions.
6. Name two other isomers of this molecule using IUPAC naming conventions.

Answer



This parent chain contains 4 carbon atoms, all of which are bonded by single bonds.

Numbering the carbons from right to left allows the hydroxyl group to occur at the lowest numbered carbon.

The parent alkane name ‘butan-’ is used for a 4 carbon chain. Additionally, the suffix ‘-1-ol’ is used to indicate a hydroxyl group on the 1st carbon atom.

The name of this molecule is butan-1-ol.

Possible isomers include: 2-methylpropan-1-ol, 2-methylpropan-2-ol and butan-2-ol.

Progress questions

Question 7

The name of an alcohol with two carbon atoms is

A. ethanol.

B. ethanoic acid.

Question 8

The molecule pentan-3-ol has the semi-structural formula

A. $CH_3CH_2CH(OH)CH_2CH_3$.

B. $CH_3CH(OH)CH_2CH_2CH_3$.

USEFUL TIP

When brackets are included in a semi-structural formula, $CH_3CH(OH)CH_2CH_3$, this indicates that this group is branching from the main carbon chain, in this case from carbon number 2.

Carboxylic acids 1.2.5.5

Carboxylic acids are organic compounds that possess a **carboxyl group**.

What are carboxylic acids?

A carboxylic acid is an organic compound that contains a carboxyl functional group, –COOH. This functional group consists of a hydroxyl group, –OH, as well as a carbonyl group, which is a carbon atom double bonded to an oxygen atom, depicted as –C=O. It is important to understand that in carboxyl groups, the carbonyl and hydroxyl groups are considered to be a single group. We do not refer to them separately. Figure 9 and table 4 show the structural formulas of carboxylic acids.

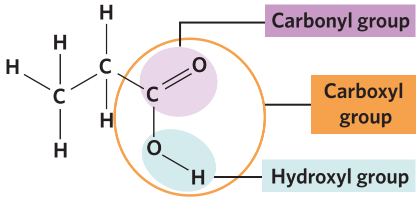
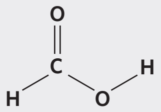
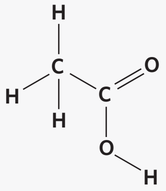


Figure 9 Propanoic acid

Table 4 Representations of carboxylic acids

Systematic name	Molecular formula	Structural formula	Semi-structural formula
Methanoic acid	CH_2O_2		HCOOH
Ethanoic acid	$C_2H_4O_2$		CH_3COOH

How does the structure of carboxylic acids affect their properties?

When comparing the electronegativities of the elements found in both the C=O bond and O–H bond, we can identify that these bonds are both polar. As a consequence, carboxylic acids are able to form strong dipole-dipole interactions and hydrogen bonds with other carboxylic acids and other polar molecules such as water. As seen in figure 10, the oxygen atom in the C=O bond of one carboxylic acid is able to form a hydrogen bond with the hydrogen atom from the O–H bond of another carboxylic acid to form a complex known as a dimer. A dimer is a molecular complex consisting of two molecules linked together. The formation of dimers between carboxylic acids is one of the reasons why the boiling point of these molecules is higher than alcohols, as the strength of the intermolecular forces takes a substantial amount of energy to overcome.

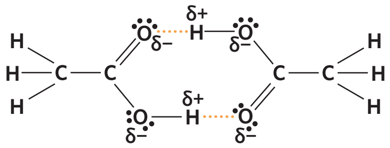


Figure 10 Hydrogen bonding between two carboxylic acids, forming a dimer

Due to the presence of an additional carbonyl group in carboxylic acids and the ability to form dimers, the boiling points of these molecules are even higher than those of haloalkanes, alcohols and alkanes with the same parent chain length. Additionally, since carboxylic acids are polar molecules and can form strong dipole-dipole interactions and hydrogen bonds with water, as shown in figure 11, these molecules are able to dissolve in water.

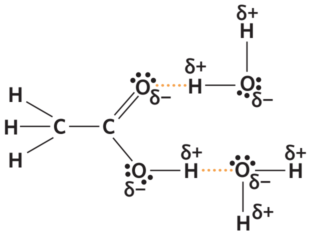


Figure 11 Ethanoic acid forming hydrogen bonds with two water molecules

However, much like with alcohols, as the carbon-chain length of the carboxylic acid increases, the effect that the functional group has on the properties of the molecule overall is reduced, since the molecule becomes less polar. It should also be noted that carboxylic acids are weak acids. The concept of acids will be discussed in a lesson 10A.

Progress questions

Question 9

A dimer is formed between two carboxylic acids through

A. covalent bonding.

B. hydrogen bonding.

Question 10

Alcohols have higher boiling points than carboxylic acids with the same number of carbon atoms.

A. True

B. False

Naming carboxylic acids 1.2.6.8

Carboxylic acids possess unique naming conventions due to their unique chemical structure.

How are carboxylic acids named?

When naming carboxylic acids, the parent name of the molecule remains the same as its corresponding alkane. The 'e' at the end of the name of the hydrocarbon is removed and replaced with the suffix '-oic acid'.

Consider figure 12. As we can see, the parent chain contains four carbons and consists only of carbon to carbon single bonds. The parent name of this molecule is therefore butane. Due to the presence of the carboxyl group, the 'e' is removed, and replaced with '-oic acid' so that the name of the compound is butanoic acid. It's important to note that when determining the parent chain, the carbon involved in the carboxyl group is always included.

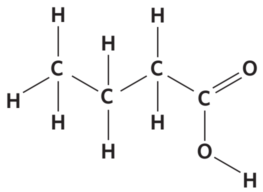


Figure 12 Structural formula of butanoic acid

It should also be noted that the carbon atom with carboxylic acid bonded to it is always considered the first carbon atom on the chain and so it is not required to specify the position of the carboxyl group.

USEFUL TIP

Remember to include the carbon in the '–COOH' group when identifying the parent chain of carboxylic acids.

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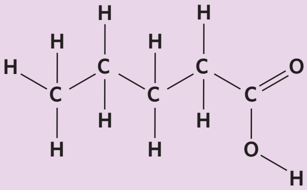
WORKED EXAMPLE 3

What is the systematic name of the molecule with the following structural formula?



What information is presented in the question?

The structural formula for a molecule.

What is the question asking us to do?

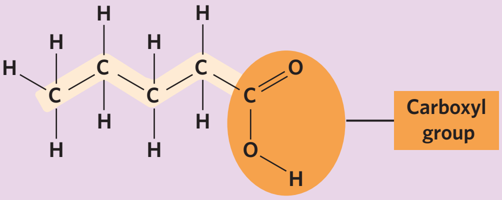
Give the systematic name of the molecule provided.

What strategies do we need in order to answer the question?

1. Identify the number of carbons in the parent chain.
2. Identify the type(s) of bonds in the parent chain.
3. Identify any branches.
4. Number the carbons in the parent chain.
5. Apply this information using correct IUPAC naming conventions.

Answer

This molecule's parent chain contains five carbon atoms and it is a saturated hydrocarbon.



It has a carboxyl group.

The parent name 'pentan-' is used for a 5 carbon chain. The suffix '-oic acid' is used to indicate a carboxyl group on the 1st carbon atom.

The name of the molecule is pentanoic acid.

Progress questions

Question 11

The name of a carboxylic acid with three carbon atoms is

A. propanoic acid.

B. propoxylic acid.

Question 12

When naming the parent chain of a carboxylic acid, the carbon in the COOH group ___________ be considered.

A. should

B. should not

Theory summary

- Functional groups are a set of atoms that provide a molecule with a characteristic set of properties.
- Structural isomers have the same molecular formula but a different structural formula (spatial arrangement of atoms) and therefore a different systematic name.
- Haloalkanes are organic compounds with a halogen functional group (–X).
  - Haloalkanes have strong intermolecular forces due to the electronegative halogen, leading to the properties of a high boiling point.
  - The influence of the halogen diminishes as the carbon chain length increases.
  - Naming involves identifying the parent chain and adding a prefix based on the halogen present.

Continues →

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- Alcohols are organic compounds with a hydroxyl functional group (–OH).
  - Alcohols have strong intermolecular forces due to the polar hydroxyl group, leading to the properties of a high boiling point and solubility in water.
  - The influence of the hydroxyl group diminishes as the carbon chain length increases.
  - Naming involves identifying the parent chain and adding the suffix '-ol'.
- Carboxylic acids are organic compounds with a carboxyl functional group (–COOH).
  - Carboxylic acids are able to have very strong intermolecular forces due to the polar carboxyl group.
  - The influence of the carboxyl group diminishes as the carbon chain length increases.
  - Carboxylic acids are able to form dimers and very strong hydrogen bonds with each other.
  - Carboxylic acids are named through the combination of the parent name and the suffix '-oic acid'.

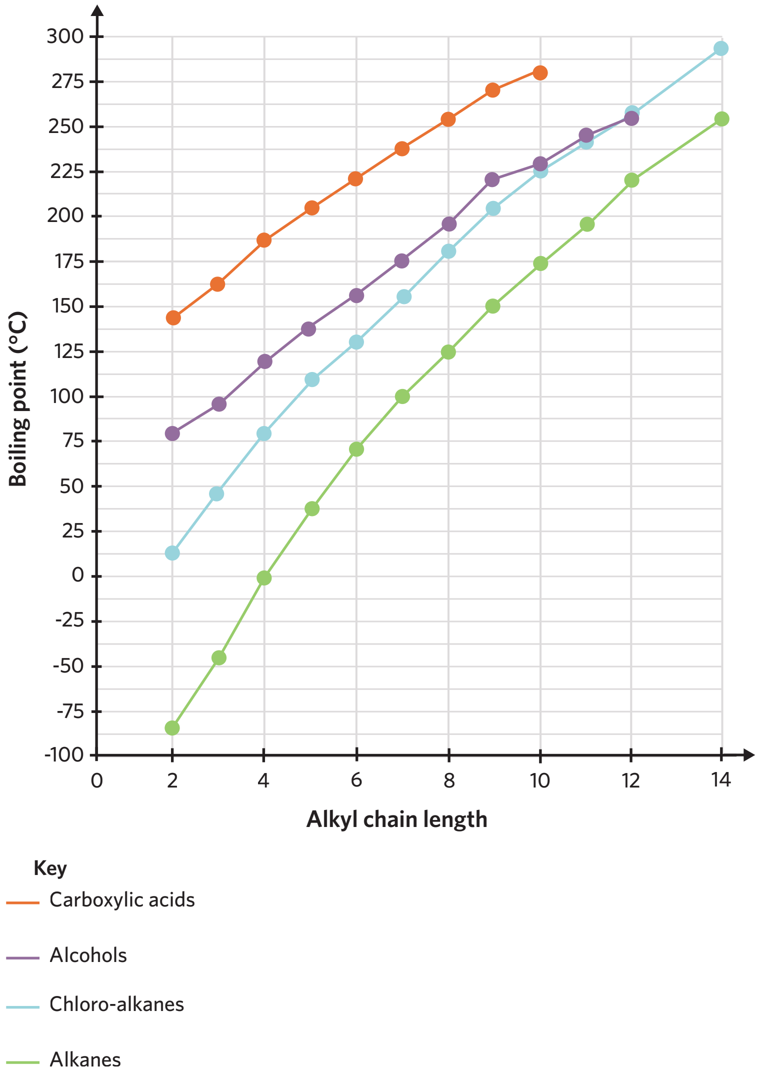

Figure 13 Various boiling points of organic molecule series

The content in this lesson is considered fundamental prior knowledge to structure, nomenclature and properties of organic compounds (Unit 4 AOS 1).

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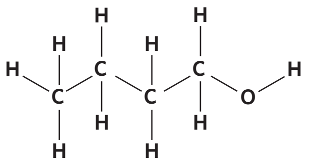
7C Questions

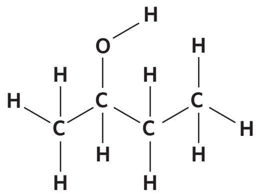
Mild Medium Spicy

Deconstructed

Use the following information to answer questions 13–15.

Timothée has recently learned about functional groups in chemistry class, and in his excitement has drawn the structural formulas of a pair of isomers and names them butan-4-ol and butan-2-ol to test his understanding.





Question 13 (1 MARK)

The molecules drawn by Timothée are ___________, defined by the ___________ functional groups.

A. carboxylic acids, hydroxyl

B. hydroxyls, alcohol

C. alcohols, hydroxyl

D. haloalkanes, halogen

Question 14 (1 MARK)

The semi-structural formulas of Timothée's molecules, respectively, are:

A. $CH_3CH_2CH_2CH_2OH$ and $CH_3CH(OH)CH_2CH_3$

B. $CH_3CH_2CH_2OH$ and $CH_3CH(OH)CH_2CH_3$

C. C_4H_9OH for both.

D. $CH_3CH_2CH(OH)CH_3$ and $CH_3CH(OH)CH_2CH_3$

Question 15 (3 MARKS)

Are the names of the molecules proposed by Timothée correct? Explain your answer, and if either of the proposed names are incorrect, provide the correct names.

Exam-style

Question 16 (1 MARK)

Which of the following compounds is least soluble in water at room temperature?

A. Ethane

B. Ethanol

C. Chloroethane

D. Ethanoic acid

Adapted from VCAA 2011 Exam 1 Multiple choice Q1

Question 17 (1 MARK)

A compound with the formula $C_6H_{12}O_2$ has the following features:

1. It is unbranched.
2. It has only one type of functional group.
3. All carbon-to-carbon bonds are single bonds.

The compound could be classified as an

A. alkane.

B. haloalkane.

C. alcohol.

D. carboxylic acid.

Adapted from VCAA (NHT) exam 2019 Multiple choice Q18

Question 18 (6 MARKS)

Jeremy is a chemistry student with a passion for anything related to ants. Driven by curiosity, Jeremy decides to research the composition of ant venom, and finds that the main molecule in the venom is methanoic acid. This molecule, which is injected at high concentration into the body through an ant stinger, creates a pain response.

a. What type of molecule is methanoic acid, and what functional group does it contain? (2 MARKS)

b. Draw the structural formula of methanoic acid. (1 MARK)

c. Having understood methanoic acid, Jeremy wishes to compare this molecule to another type of molecule he's familiar with.

- **i.** If the functional group of methanoic acid were replaced with a chlorine atom, what type of molecule would it be? (1 MARK)
- **ii.** What is the IUPAC name of the molecule identified in part ci? (1 MARK)
- **iii.** Draw the structural formula for the molecule identified in part ci. (1 MARK)

Question 19 (8 MARKS)

Hydrocarbons play a fundamental role in many aspects of modern life, making up the world around us. The addition of functional groups to these molecules can change the way that these molecules interact with each other, and with the world at large.

a. Explain how the addition of a hydroxyl group changes the intermolecular forces between hydrocarbons. (3 MARKS)

b. Explain how the intermolecular forces between alcohols change the boiling point, compared with hydrocarbons not possessing functional groups. (2 MARKS)

c. Carboxylic acids, like alcohols, possess functional groups which enable them to form stronger intermolecular forces.

- **i.** What is the name of the functional group found in carboxylic acids? (1 MARK)
- **ii.** Describe the structure of a carboxylic acid dimer, including the bonds involved. (2 MARKS)

Key science skills

Question 20 (8 MARKS)

Layla is performing an experiment in which she compares the properties of alcohols and carboxylic acids. In order to do this, Layla compared the boiling points of each of these substances by placing them in a test tube and exposing them to a direct flame from a bunsen burner.

a. Suggest at least two factors that must be taken into account for this experiment to be carried out safely. (2 MARKS)

b. What is the independent variable in this experiment? Explain your answer. (2 MARKS)

c. What is the dependent variable in this experiment? (1 MARK)

d. What are some variables that must be controlled for the result of this experiment to be valid? Suggest three variables. (3 MARKS)

FROM LESSONS 16A & 16B

Questions from multiple lessons

Question 21 (16 MARKS)

Eileen is a year 11 chemistry student practising interpretation of IUPAC nomenclature of alcohols and carboxylic acids in preparation for an upcoming assessment. In the course of her studies, she has come across a number of names which she has been unfamiliar with, and filled in her knowledge in the form of a table.

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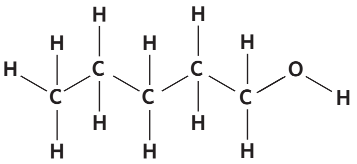
a. Fill in the partially completed table below. (8 MARKS)

IUPAC Name	Semi-structural formula	Structural formula
		
	$CH_3CH_2CH(Br)CH_3$	
	CH_3CH_2COOH	
3-iodohexane		

b. Which of these molecules has the greatest molar mass? (1 MARK)

c. When comparing each of these molecules to their alkane counterparts, Eileen noticed that the molecules possessing functional groups each had a higher boiling point than alkanes with the same parent chain but without functional groups.

- **i.** Why do each of the molecules in part a have a higher boiling point than their alkane counterparts? (3 MARKS)
- **ii.** Explain why the difference in boiling points between an alcohol and its alkane counterpart decreases as the carbon chain length of the molecule increases. (4 MARKS)

FROM LESSONS 6B & 7B

Hints

16. The presence of dipole-dipole interactions and/or hydrogen bonds in molecules generally increases their solubility.

17. Molecular formulas illustrate the number of various atom types in a molecule.

18a. Functional groups have specific naming conventions that display their presence in a molecule.

18b. Structural formulas explicitly show atoms, and the bonding between atoms, in a molecule.

18ci. The presence of a halogen defines haloalkanes.

18cii. IUPAC conventions outline ways to show the presence of various functional groups.

18ciii. Structural formulas explicitly show atoms, and the bonding between atoms, in a molecule.

19a. Functional groups affect the intermolecular forces between molecules.

19b Functional groups can enable different types of bonding.

19ci. Carboxylic acid is not a functional group.

19cii. Carboxylic acid dimers result from intermolecular interactions

20a. Safety protocol minimises exposure to hazardous chemicals or flames.

20b. The independent variable is changed by the experimenter.

20c. Changes to the dependent variable are observed to determine the effect of changing the independent variable.

20d. Controlled substances should remain constant and not change experimental conditions between trials.

21a. IUPAC nomenclature has specific conventions used to indicate particular functional groups.

21b. Molar mass is equal to the molecular weight of a molecule.

21ci. Functional groups affect the boiling point of molecules.

21cii. The type and magnitude of intermolecular forces present can affect the boiling point.

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7D Renewable sources

Image: Ovnigraphic/Shutterstock.com

How can plants be used to colour denim jeans?

Denim jeans get their blue colour from indigo dye, an **organic chemical** usually produced through **synthetic** means. The same dye can also be produced from plants like woad, a flowering plant native to Central Asia that has become a noxious weed in parts of North America. In this lesson we will discuss the production of organic chemicals from plants and the need to make this process as economical, environmentally-friendly, and **sustainable** as possible.

STUDY DESIGN DOT POINT

- plant-based biomass as an alternative renewable source of organic chemicals (for example, solvents, pharmaceuticals, adhesives, dyes and paints) traditionally derived from fossil fuels

7A 7B 7C **7D** 7E

1.2.7.1 Renewable organic chemicals

ESSENTIAL PRIOR KNOWLEDGE

7A Hydrocarbons

See question 44.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

KEY TERMS AND DEFINITIONS

Biomass	any organic material made of plant or animal matter
Crude oil	mixture of differently sized hydrocarbons found in underground deposits and used as a resource
Ecosystem	community of life forms interacting with their environment
Fossil fuel	hydrocarbon-containing material formed in the Earth's crust from plant and animal remains
Non-renewable	unable to be naturally replenished
Organic chemical	any compound that contains carbon-carbon and/or carbon-hydrogen bonds
Petrochemical	derived from crude oil
Renewable	comes from sources that naturally renew themselves at a rate that allows them not to become depleted
Sustainable	can be produced at a rate that is greater than consumption without compromising future generations
Synthesise	produce a chemical through reaction with other compounds
Synthetic	produced artificially by humans using chemical reactions

Renewable organic chemicals 1.2.6.1

The major source of organic chemicals used in society has traditionally been fossil fuels, which will not naturally replenish in the foreseeable future. However, some of these chemicals can also be produced from plants, and this is becoming an increasingly feasible option for the future of many industries.

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How have organic chemicals traditionally been sourced?

Organic chemicals are a component of nearly every item we use each day, including furniture, clothes, and phones. The vast majority of these organic chemicals are **synthesised** from hydrocarbons that have been extracted from **crude oil** (figure 1). Crude oil is a mixture of hydrocarbons[1] formed over millions of years from dead, ancient, marine organisms, and as a result is calleda **fossil fuel**. The process by which organic chemicals are produced from crude oil is shown in figure 2.

KEEN TO INVESTIGATE?

[1] **How are hydrocarbons in crude oil separated in refineries?**
Search YouTube: Fractional distillation of crude oil

Image: Anan Kaewkhammul/Shutterstock.com

Figure 1 Crude oil

MISCONCEPTION

‘If a chemical is ‘organic’, it is produced from plants.’

In chemistry, ‘organic’ describes any compounds that contain carbon-carbon or carbon-hydrogen bonds, regardless of their source.

MISCONCEPTION

‘Petroleum and crude oil are the same chemical.’

Petroleum is one of the many different chemical mixtures found in crude oil.

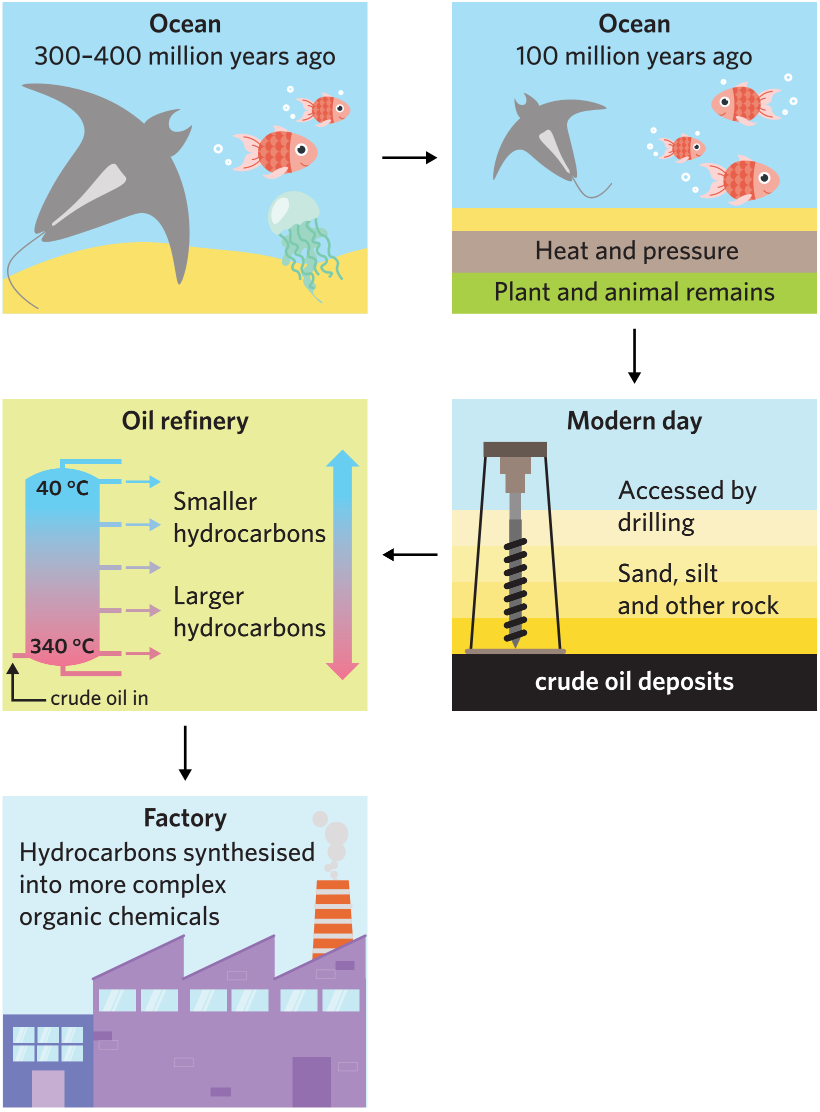

Figure 2 Formation, extraction, and separation of crude oil

Hydrocarbons from crude oil are used for fuels, solvents, dyes, adhesives, plastics and more. For example, PVA glue – which stands for polyvinyl acetate (systematic name ethenyl ethanoate) – is produced from ethene according to the process described in figure 3. PVA glue is non-toxic and suitable for glueing together porous materials like paper and wood, leading to its widespread use in schools, home repairs, and carpentry.

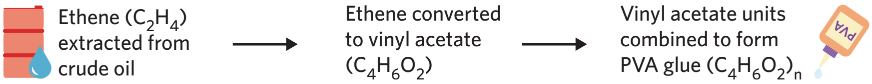

Image (left to right): Leyli Polayeva, My Flat Style/Shutterstock.com

Figure 3 Synthesis of PVA from crude oil

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Crude oil is not the only fossil fuel from which organic chemicals are sourced. For example, paracetamol – an extremely popular pain medication commonly marketed in Australia as Panadol© – can be synthesised from petroleum or from coal tar, itself a product of another major fossil fuel: coal. Paracetamol has the chemical formula $C_8H_9NO_2$ (figure 4) and is one of many pharmaceuticals traditionally derived from fossil fuels.

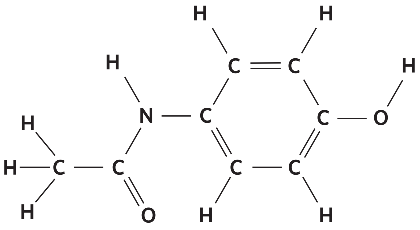


Figure 4 Molecular structure of paracetamol

Solvents and dyes are other important classes of organic chemicals. For example, hexane (C_6H_{14}) is a non-polar solvent widely used in industry and in a number of laboratory processes, including chromatography (explored in lesson 5A). As a hydrocarbon, hexane is extracted directly from crude oil. Most dyes used in foods and textiles are synthetic, such as the green dye 'Fast green' shown in figure 5. These synthetic dyes tend to have extremely complex chemical structures – the IUPAC name for 'Fast green' is…

ethyl-[4-[[4-[ethyl-[(3-sulfophenyl)methyl]amino]phenyl]-(4-hydroxy-2-sulfophenyl)methylidene]-1-cyclohexa-2,5-dienylidene]-[(3-sulfophenyl) methyl]azanium

… so the process of deriving these compounds is often complicated and energy-intensive.

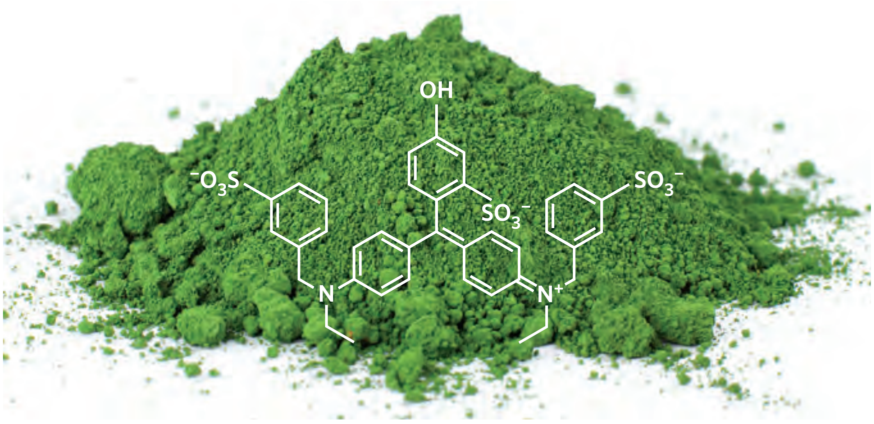
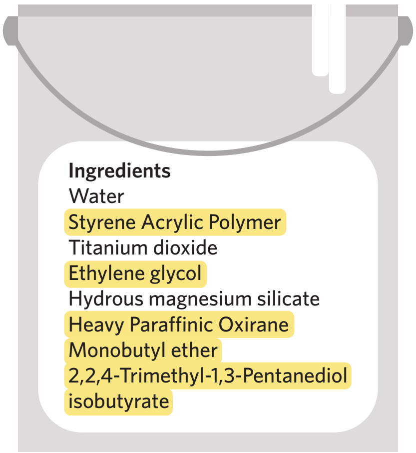


Image: liliya Vantsura/Shutterstock.com

Figure 5 Colour and structure of 'Fast Green'

If you are currently in a painted room, organic chemicals from fossil fuels are all around you. Paints are made up of a binder, solvents, pigments, and additives, the majority of which are typically sourced from fossil fuels. Figure 6 illustrates the typical components of a tin of white paint, highlighting those which are standardly derived from crude oil.



Ingredients
Water
Styrene Acrylic Polymer
Titanium dioxide
Ethylene glycol
Hydrous magnesium silicate
Heavy Paraffinic Oxirane
Monobutyl ether
2,2,4-Trimethyl-1,3-Pentanediol
isobutyrate

Figure 6 Typical ingredients in a tin of white paint; components highlighted in yellow are derived from crude oil.

However, there are a number of issues with society's reliance on deriving organic chemicals from fossil fuels. Extraction of crude oil is primarily achieved through onshore and offshore drilling, which can:

- disrupt **ecosystems** and destroy habitats of plants and animals,
- cause spillage of crude oil into water sources, potentially contaminating water supply (figure 7),
- and destroy sites of cultural and aesthetic significance.

Image: Korelidou Mila/Shutterstock.com

Figure 7 Negative impact of offshore crude oil drilling on the environment

Finally and most significantly, crude oil is a finite resource that takes millions of years to form; once it is used up, it cannot be replenished in the foreseeable future. Hence, it is classed as a **non-renewable** resource. Unless other sources of organic chemicals are used, these fossil fuel-derived products will become increasingly expensive and eventually inaccessible in years to come.

How can organic chemicals be renewably sourced?

To ensure that society has continuing access to these important organic chemicals in the future, **renewable** alternatives to crude oil are necessary. Plant-based **biomass** has attracted research and industrial attention as a renewable source for many organic chemicals.

Image: mailsonpignata/Shutterstock.com

Figure 8 Sugarcane farm where ethanol is produced

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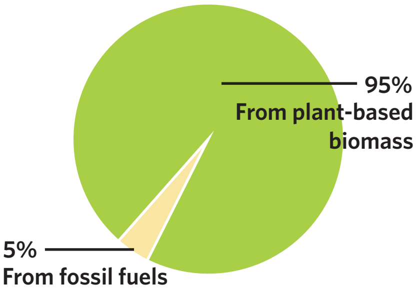

Figure 9 Proportion of industrial ethanol derived from plant-based biomass and from fossil fuels

Image: Peter K. Ziminski/Shutterstock.com

Figure 10 Pacific yew tree bark harvested to make Taxol©

MISCONCEPTION

'Renewable and sustainable mean the same thing.'

Renewable resources can be naturally replenished, whereas sustainable resources can be naturally replenished at a rate faster than the rate at which they are consumed.

A key example of this is ethanol (C_2H_5OH), which is used as a fuel and as a common solvent in products such as hand sanitisers. Ethanol can be produced from the fermentation of crops such as corn and sugarcane, and massive fields of crops are grown for this purpose (figure 8). In fact, this simple and inexpensive plant-based production method is now responsible for the vast majority of the world's industrial ethanol production, as shown in figure 9. Starch, produced by all green plants, can also be dry-roasted to produce dextrin glue, an inexpensive adhesive commonly used for cardboard boxes, bottle labels, and envelopes.

One example of an organic compound which is near-exclusively sourced from plant-based biomass is paclitaxel, sold as Taxol©: a major chemotherapy medication used to treat breast cancer, ovarian cancer, and others. It can be synthesised from crude oil products, but the process is prohibitively expensive on a commercial scale. Fortunately, paclitaxel is present in the bark of the Pacific yew tree (figure 10), and the extraction of paclitaxel from this treebark is the primary source of the drug used worldwide.

Unlike in the case of paclitaxel, renewable alternative sources for some organic chemicals have been difficult to find; for example, as of 2022, 95% of ingredients in paints are sourced from fossil fuels. This is nonetheless an active area of global research, with chemists recently developing a coating (the outer protective layer of paint) made from lignocellulose – a major component of plant biomass – and experimenting with natural pigments like anthocyanin – found in blueberries, raspberries, black rice and black soybeans. Additionally, linseed oil, sourced from the flax plant, is the most common binding agent used in oil paints. Figure 11 shows a number of renewable materials traditionally used in oil paintings.

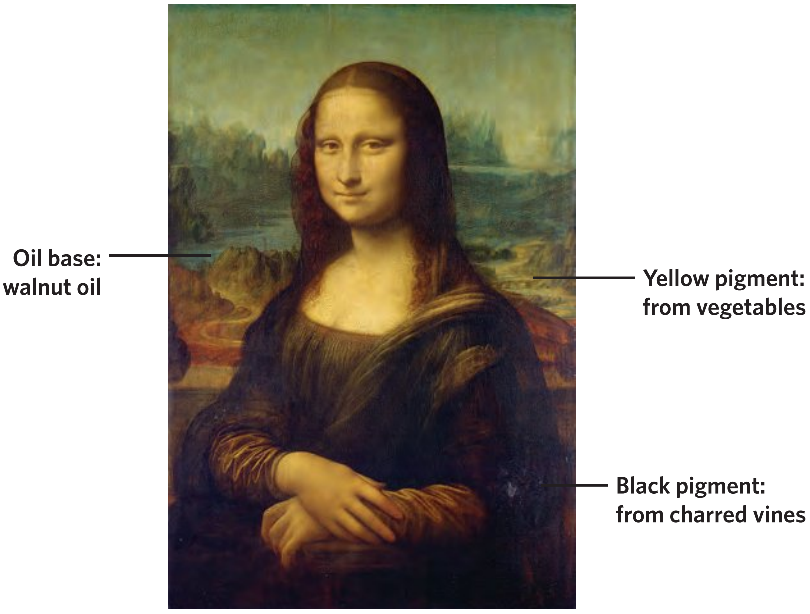

Figure 11 Plant-based organic chemicals used in Leonardo da Vinci's *Mona Lisa*

Some renewable alternatives to fossil fuels face other challenges (figure 12). For example, woad – a flowering plant native to Asia and Europe – has been used for thousands of years as a source of blue dye for textiles, and more recently in inkjet printers. Unlike synthetic dyes, it is biodegradable, and it can be used to dye denim jeans. However, to achieve this, the indigo dye needs petrochemical additives to convert it to a water-soluble form, and 100 kg of woad plant biomass produces just 1 kg of dye; as a result, this production method is prohibitively expensive, is not sustainable, and has not been widely adopted.

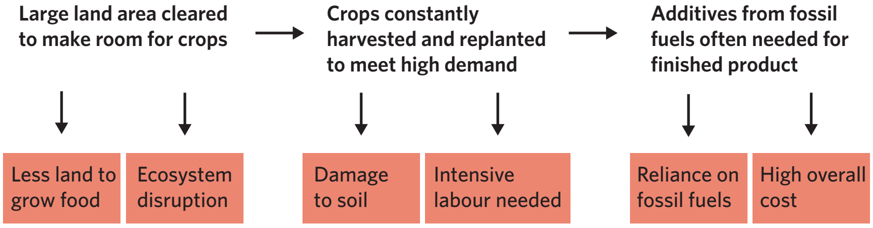

Figure 12 Typical process and impacts of producing organic chemicals from plant-based biomass at an industrial scale

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Therefore, despite the advantage of renewability, there are a number of barriers to the widespread adoption of plant-based organic chemicals in industry. These include:

- prohibitive production cost e.g. plant-based paracetamol,
- large areas of land often need to be cleared to grow feedstock e.g. ethanol,
- less generality of usage, unable to replace certain chemicals e.g. high-strength glue,
- often require **petrochemical** additives e.g. indigo dye,
- and, worker exploitation in farming chemicals for which there is high demand e.g. saffron dye.

Progress questions

Question 1

Crude oil comes from dead marine organisms.

A. True

B. False

Question 2

Most organic chemicals are currently produced from

A. renewable sources.

B. non-renewable sources.

Question 3

Biomass is

A. compacted remains of ancient marine life.

B. plant-based material used to produce organic chemicals.

Question 4

The majority of global ethanol is derived from

A. fermentation of plant-based biomass.

B. extraction from crude oil.

Question 5

All organic chemicals from renewable sources are sustainable to produce.

A. True

B. False

Theory summary

- Everyday organic chemicals, including dyes, paints, and adhesives, are sourced primarily from crude oil.
- Crude oil is a non-renewable fossil fuel, and so cannot be replenished.
- To ensure access to organic chemicals continues, renewable alternative sources to crude oil are needed.
- Plant-based biomass, such as from corn or tree bark, is a renewable source of many everyday organic chemicals.
- Cost, practicality, ethics, and sustainability are challenges currently faced by many plant-based alternatives.

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7D Questions

Deconstructed

Use the following information to answer questions 6–8.

A cardboard coffee cup company is searching for a way to glue their cups together without needing petroleum-based products.

Question 6 (1 MARK)

A renewable resource is defined as a resource

A. that can be replenished at a rate faster than the rate at which it is consumed.

B. formed from dead ancient marine life over millions of years.

C. that comes from naturally replenishing sources.

D. that is able to be improved.

Question 7 (1 MARK)

Dextrin glue is an example of

A. an adhesive.

B. a solvent.

C. a dye.

D. a paint binder.

Question 8 (2 MARKS)

Suggest an adhesive for the coffee cup company and explain how it can be produced from a renewable source.

Exam-style

Question 9 (3 MARKS)

A student claims that ‘PVA glue comes from sea animals that died 400 million years ago’. Evaluate the validity of the student’s statement.

Question 10 (7 MARKS)

Organic chemicals are used throughout society. In recent years, there has been a shift away from crude oil towards plant-based biomass as a source of these organic chemicals.

a. Explain why it is necessary for society to develop methods of producing organic chemicals from biomass. (3 MARKS)

b. Identify four different barriers to the widespread adoption of plant-based sources of organic chemicals. (4 MARKS)

Question 11 (5 MARKS)

Ethanol is an organic chemical that can be synthesised from ethene, a common hydrocarbon.

a. Explain the process by which ethanol is produced from a non-renewable source. (2 MARKS)

b. Identify the major renewable source of ethanol, and explain the production process. (2 MARKS)

c. Identify a use for ethanol by humans. (1 MARK)

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Key science skills

Question 12 (8 MARKS)

The high demand for chemical products in society means that organic chemicals must be produced in extremely large quantities every day. As a result, these products often come at a cost in more ways than one.

Compare fossil fuels with plant-based biomass as sources of organic chemicals, discussing issues under the following themes:

a. Environmental (2 MARKS)

b. Sustainable (2 MARKS)

c. Economic (2 MARKS)

d. Ethical (2 MARKS)

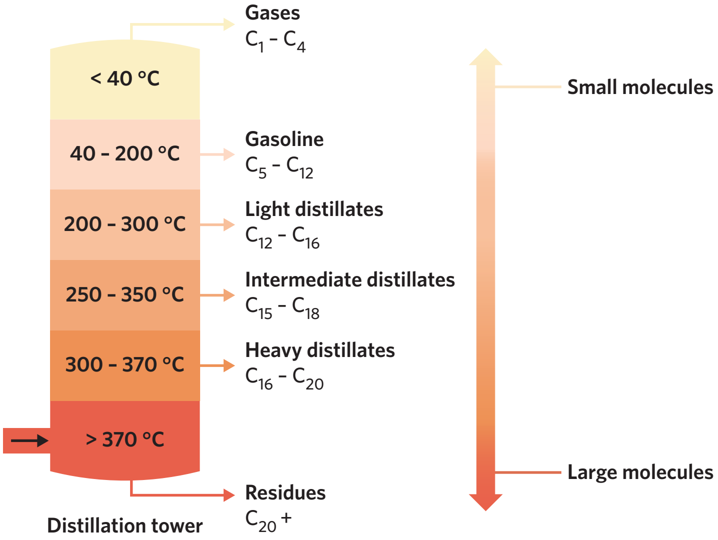
FROM LESSON 16B

Image: Christopher Halloran/Shutterstock.com

Questions from multiple lessons

Question 13 (7 MARKS)

Fractional distillation is the process by which hydrocarbons in crude oil, including alkanes, are separated according to their boiling point.



a. What intermolecular forces are present between alkanes? (1 MARK)

b. Explain why larger alkanes (e.g. $C_{16}H_{34}$) have higher boiling points than smaller alkanes (e.g. C_2H_6) with reference to the intermolecular forces present. (4 MARKS)

c. Is fractional distillation the final step in the production of most organic chemicals? Explain. (2 MARKS)

FROM LESSONS 2B & 7A

Hints

9.	Consider the origin and uses of crude oil.
10a.	Crude oil will not be available forever.
10b.	Demand for organic chemicals is extremely high, posing a number of barriers to plant-based production.
11a.	Ethanol has a similar formula to ethane.
11b.	A very large quantity of ethanol is produced from renewable sources each year, so its source must be abundant.
11c.	Ethanol-containing products are often labelled with the percentage of ethanol.
12a.	Environmental factors include impact on land, plants, and animals.
12b.	Sustainability can often be thought of as a resource's long-term availability.
12c.	Consider large-scale and long-term cost of production.
12d.	Ethical issues may involve workers or communities affected by production.
13a.	Alkanes are symmetrical molecules composed of carbon and hydrogen only.
13b.	Boiling point varies according to the strength of interactions between molecules.
13c.	Compare the components of crude oil with the chemical formulae of the organic chemicals studied.

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7E Organic compounds and society

STUDY DESIGN DOT POINT

- materials and products used in everyday life that are made from organic compounds (for example, synthetic fabrics, foods, natural medicines, pesticides, cosmetics, organic solvents, car parts, artificial hearts), the benefits of those products for society, and the health and/or environmental hazards they pose

7A 7B 7C 7D **7E**

1.2.8.1 & 1.2.8.2 Everyday organic compounds & benefits and hazards

ESSENTIAL PRIOR KNOWLEDGE

7C Types of functional groups

See question 45.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: Jollyx/Shutterstock.com

What actually is 'new car smell', and is it toxic?

The distinct, 'pleasant' smell associated with new cars comes from a mixture of gases produced by vinyl and other chemicals used in the interiors of cars. However, research is not conclusive as to whether this concoction of chemicals is harmful to breathe in or not. In this lesson, we will explore the organic compounds used in many aspects of everyday life, including the benefits and hazards they pose.

KEY TERMS AND DEFINITIONS

Adhesivity	property of being sticky
Bioaccumulation	phenomenon of increasing contaminant concentration in an organism over time
Biomagnification	phenomenon of higher contaminant concentration in organisms higher in the food chain
Carcinogenic	having the potential to cause cancer
Food chain	a series of organisms each dependent on the next as a source of food
Inertness	property of being unreactive
Inhalant	chemical whose vapours can be concentrated and breathed in to produce intoxication, contrary to its intended usage
Microbeads	manufactured microplastics less than 1 mm in length used in exfoliating products and toothpastes
Microplastics	polluting fragments of plastic less than 5 mm in length
Ozone layer	layer within stratosphere with high concentration of ozone (O_3) that absorbs UV radiation
Solvent	substance that dissolves another substance
Stratosphere	layer of earth's atmosphere 20–50 km above sea level

Everyday organic compounds & benefits and hazards 1.2.8.1 & 1.2.8.2

From foods to fabrics, from car parts to artificial hearts; organic compounds are used in nearly every aspect of our day-to-day lives. While there is no doubt about their usefulness to society, the potential risks they pose to health and to the environment must always be considered.

How are organic compounds both useful and harmful to society?

Organic solvents

Organic compounds are used as **solvents** in a multitude of products, including markers, paints, glues, aerosol sprays, and cleaning products. For example, 1,1,1-trichloroethane (CH_3CCl_3, figure 1) is an organic solvent once used in household cleaning products and as a machine degreaser. Although 1,1,1-trichloroethane proved very useful to manufacturing and was once considered 'the least toxic haloalkane solvent', like many solvents in paints and glues, 1,1,1-trichloroethane was often abused as an **inhalant** to produce dizziness. However, regular inhalation of the solvent can have a **carcinogenic** effect, while even small amounts can induce irregular heartbeat and potential respiratory arrest.

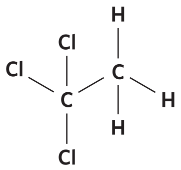

Figure 1 Structure of 1,1,1-trichloroethane

Furthermore, when 1,1,1-trichloroethane enters the **stratosphere**, it releases chlorine atoms, which catalyse the breakdown of ozone (O_3) and oxygen (O_2). It was hence found to be one of numerous chemicals responsible for depleting the **ozone layer**, removing a vital safeguard from skin cancer-causing UV rays (figure 2). In response to this, 1,1,1-trichloroethane was phased out of use in 1996.

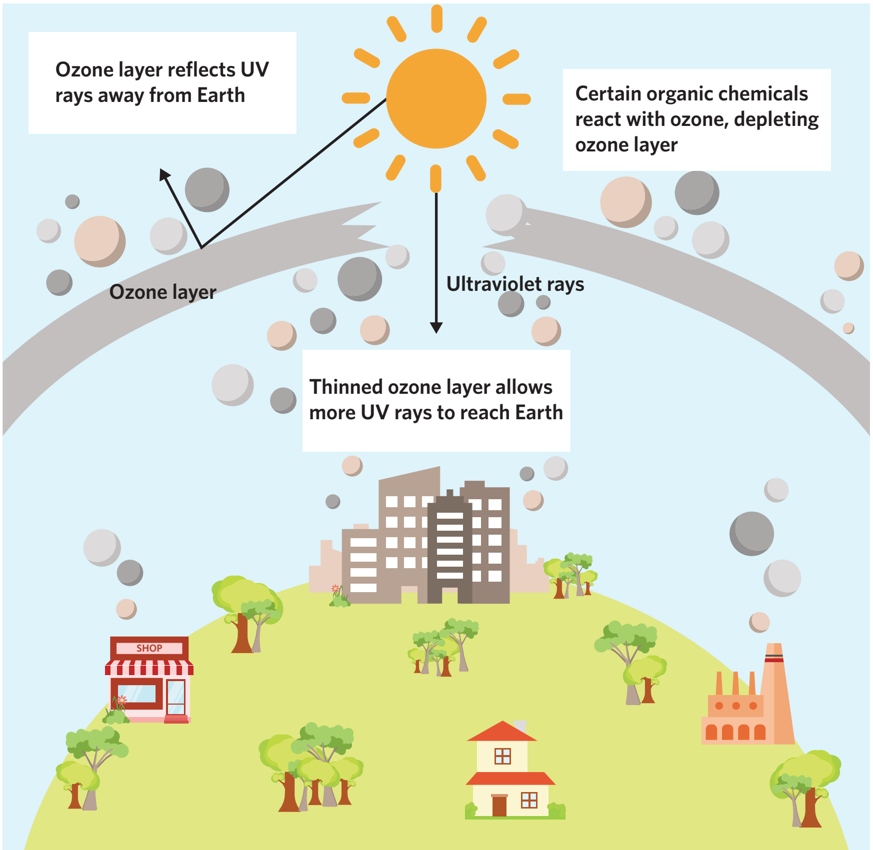

Figure 2 Diagram explaining cause and effect of ozone depletion

MISCONCEPTION

'There are holes in the ozone layer due to organic chemicals.'

Although there is significant thinning of the ozone layer in certain places, there are technically no 'holes' all the way through.

Foods

Many organic compounds are in fact edible; sugars, fats, and proteins are all organic compounds, as are artificial sweeteners, flavours, and colours. One such sweetener is Xylitol ($C_5H_{12}O_5$, figure 3), used in lollies, toothpaste, mints, and chewing gum. As sweet as sugar but with 40% less energy, chewing Xylitol gum has been thought to reduce tooth cavities, though evidence is unclear.

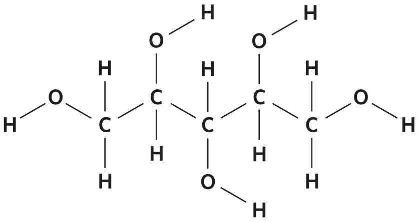

Figure 3 Structure of xylitol

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USEFUL TIP

In large quantities, Xylitol has a laxative effect. Don't eat too many mints at once!

Pesticides

Pesticides are organic chemicals used primarily in the agricultural industry to kill insects and hence reduce damage to crops, enabling a higher rate of food production. A famous example is DDT (an abbreviation of **d**ichloro**d**iphenyl**t**richloroethane, figure 4), which is a pesticide widely used on farms and in households in the mid-20th century. It was also sprayed on soldiers in tropical regions during World War II to kill disease-carrying insects; in fact, DDT-spraying programs in the 1940s temporarily eradicated malaria (a disease spread by mosquitoes) in Europe and North America, resulting in Paul Hermann Muller, who discovered its insect-killing properties, receiving a Nobel Prize for Medicine.

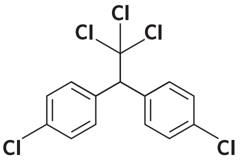

Figure 4 Skeletal formula of DDT[1]

KEEN TO INVESTIGATE?

[1] **What is that C_6H_6 hexagonal ring present in so many organic compounds?**
Search YouTube: Benzene Aromatics & Cyclic Compounds

However, although not conclusively harmful to humans, DDT posed significant harms for ecosystems, primarily due to phenomena known as **bioaccumulation** and **biomagnification**, explained in figure 5. Accordingly, the most significant victims were those highest on the **food chain**, like birds of prey. Scientists discovered that DDT poisoning caused animals like bald eagles and peregrine falcons to lay eggs with thinner shells, dramatically reducing their survival rate and devastating populations. This had a 'domino effect' on ecosystems, and so in 2004, DDT was banned in agriculture worldwide.

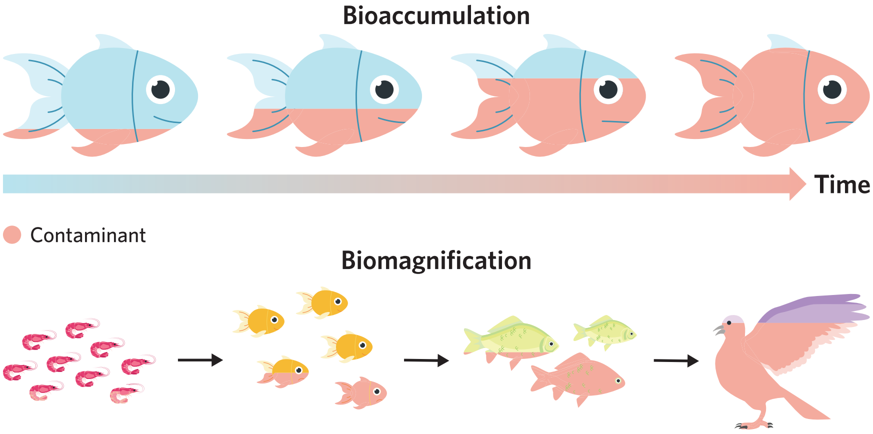

Figure 5 Biomagnification and bioaccumulation of contaminants like DDT

Medicines

As was explored in lesson 7D, although most medicines are produced from crude oil, many significant medicines are derived from plants. These include morphine, a high-strength pain relief medication with the empirical formula $C_{17}H_{19}NO_3$ (figure 6) that is found naturally in the resin of opium poppies, shown in figure 7. Morphine is prescribed to those suffering severe acute or chronic (long-lasting) pain, and it is often used to improve end-of-life quality for the terminally ill.

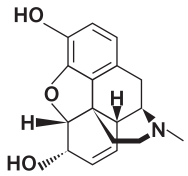

Figure 6 3D skeletal structure of morphine

KEEN TO INVESTIGATE?

[2] **How can organic compounds prevent an opioid overdose?**
Search YouTube: Naloxone

USEFUL TIP

In three-dimensional molecule structures like morphine, dotted lines are used to represent bonds going 'into the page', while thick, triangle-shaped lines are used to represent bonds coming 'out of the page'.

However, like other drugs derived from opium, it is highly addictive and often abused in society as a recreational drug. The health impacts of morphine abuse are severe and numerous, while overdose of the substance is potentially fatal.[2] As a result, the organic compound morphine, as well as compounds that can be used to synthesise it, are controlled by governments and are illegal for non-authorised persons to possess or distribute.

Cosmetics

Cosmetic products, including make-up, hair dye, moisturiser, and shampoos, are filled with organic chemicals (as figure 8 shows), and the global cosmetic industry is one of the biggest in the world, turning over an estimated $250 billion AUD per year.

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Ingredients
Water, Sodium Lauryl Sulfate, Sodium Laureth Sulfate, Cocamidopropyl Betaine, Sodium Chloride, Sodium Citrate, Citric Acid, Sodium Xylenesulfonate, Parfum, Cocamide MEA, Sodium Benzoate, Tetrasodium EDTA, PEG-60 Almond Glycerides, Butylphenyl Methylpropional, Panthenyl Ethyl Ether, Panthenol, Guar Hydroxypropyltrimonium Chloride, Benzyl Alcohol, Linalool Hexyl Cinnamal, Limonene, Benzyl Salicylate, Magnesium Nitrate, Hydroxyisohexyl 3-Cyclohexene Carboxaldehyde, Fumaric acid, Methylchloroisothiazolinone Magnesium Chloride, Methylisothiazolinone

Figure 8 Typical ingredients in shampoo, many of which are organic compounds

Image: Lukas Puchrik/Shutterstock.com

Figure 7 Poppies oozing resin, used to produce morphine

Salicylic acid (HOC_6H_4COOH, figure 9) is a common ingredient of skincare products due to its exfoliating properties (removing the outer layer of skin). As well as the aesthetic benefit of improving skin quality, it can be used to treat skin conditions like warts, acne, psoriasis, and dandruff.

However, many exfoliating products like soaps, facial scrubs and toothpaste also contain **microbeads**, tiny plastics under 1 mm in length which are able to pass through sewage systems and enter oceans and waterways. Although not directly harmful to humans, these tiny yet dangerous organic compounds tend to absorb and act as a 'vessel' for more harmful organic chemicals like DDT. Their impact on ecosystems is shown in figure 10.

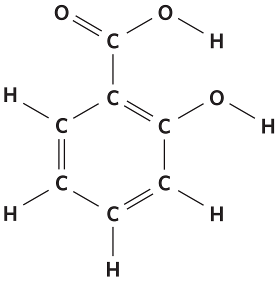

Figure 9 Structure of salicylic acid

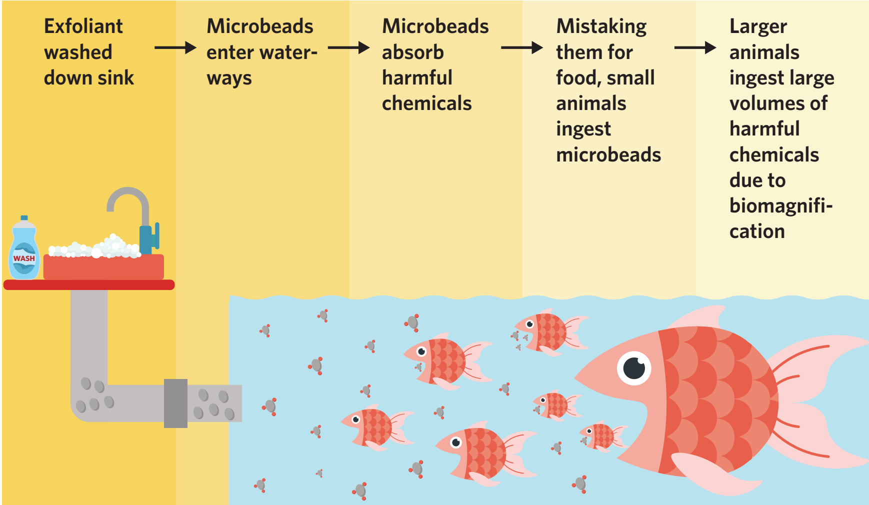

Figure 10 Flow chart of microbeads, from cosmetics to food chain destruction

Many countries have now banned the use of microbeads in exfoliating products. Although they are not banned in Australia, the federal government ensured microbeads were voluntarily phased out in Australia by 2018.

Synthetic fabrics

Synthetic fabrics, like polyester and nylon, are also made of organic compounds. For example, nylon is a class of polymer, which is itself a type of organic compound composed of many repeating small units. Polymers are explored further in lesson 8A.

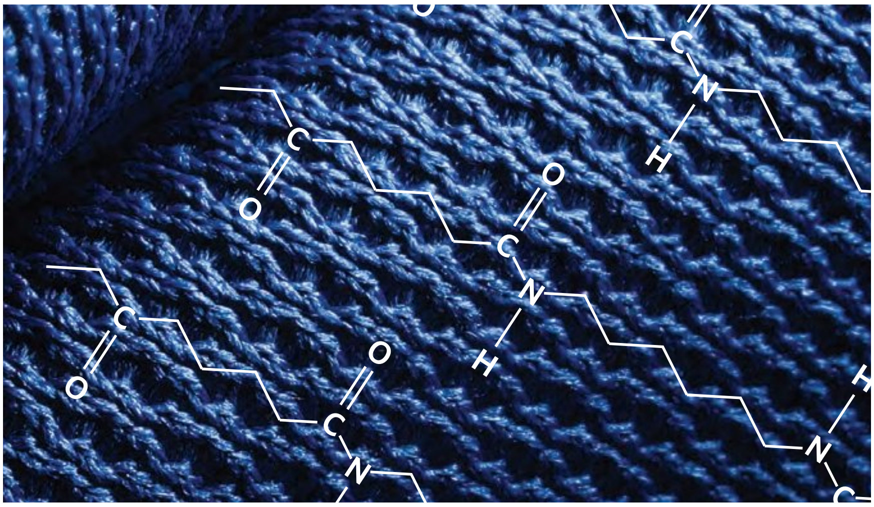

Image: Lukas Jonaitis/Shutterstock.com

Figure 11 Nylon fabric overlaid with structure of Nylon 6,6

Nylon (shown in figure 11) is used to make clothing, fishing nets, towing cables, ropes, pulleys, and more. It has a number of advantages over traditional fibres, including:

- High tensile strength
- Non-flammability
- High durability
- Light weight
- Water resistance

These properties have meant that clothes last longer and need replacing less often, ropes are less likely to break, and pulleys are able to lift heavier loads. However, nylon produces **microplastics** with the same risks as microbeads, and it takes years to decompose.

MISCONCEPTION

'Nylon is an organic compound.'

Nylon is a generic term for a family of similar compounds: polymers linked by amide groups.

Car parts

Not only are the fuels that petrol cars run on composed of organic compounds (discussed in detail in Unit 3 of VCE chemistry), but many components of the cars themselves are organic chemicals, and the gases they give off combine to produce the distinctive 'new car smell'. For example, the polymer **p**oly**v**inyl **c**hloride, also known as vinyl or its abbreviation PVC, is used in car dashboards and elsewhere in car interiors. Its flame-retardant properties reduce risk of fire spreading inside cars in crash situations, while its softness makes dashboards effective shock absorbers and thus helps reduce injury from car accidents.

However, although it helps save lives in car crashes, PVC is one of the most environmentally harmful and unrecyclable chemicals, and the vinyl chloride (shown in figure 12) used to produce it is a notable carcinogen. Environmental impact and recycling of polymers are further explored in lesson 8B.

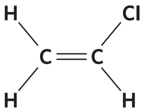

Figure 12 'Vinyl chloride', which makes up PVC

Artificial hearts

Another life-saving application of organic chemicals is in artificial hearts. The 'total artificial heart' (shown in figure 13), able to replace the function of the heart by pumping blood throughout the body, was first approved for use in 2004 and is made of segmented polyurethane solution, a mixture of organic polymers. The factors to consider when selecting an organic compound for this purpose include:

- durability: must last for years without damage
- **inertness**: must not conduct electricity or react with fluids in the body
- flexibility: must be able to slightly change shape in response to movement
- **adhesivity**: must be textured to enable red blood cells to stick to it

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Figure 13 Total artificial heart made of segmented polyurethane

The total artificial heart has been a breakthrough medical innovation in enabling better blood flow in patients with damaged hearts, hence improving quality of life and extending lifespan, particularly for those waiting for (non-artificial) heart transplants. Prosthetics (artificial body parts) made of organic chemicals comprise an active area of scientific research, and will likely save more and more lives as technology advances.

Progress questions

Question 1

Biomagnification of organic compounds most significantly affects animals ____________ on the food chain.

A. higher

B. lower

Question 2

Organic compounds should never be eaten.

A. True

B. False

Question 3

Use of organic chemicals has led to the expansion of the ozone layer.

A. True

B. False

Question 4

One advantage of using plastics is their

A. durability.

B. fragility.

Question 5

Microplastics are toxic to some animals, but not humans.

A. True

B. False

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Theory summary

- Organic compounds are used for a multitude of purposes throughout society, such as solvents, foods, pesticides, medicines, cosmetics, fabrics, car parts, and artificial hearts.
- Organic chemicals make a variety of processes and products safer, healthier, cheaper, longer-lasting, and more accessible.
- Harms caused by organic chemicals include toxicity to humans, disruption of food chains, infiltration of waterways, potential for abuse, and difficulty of recycling.

7E Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 6–8.

Microbeads are a source of plastic water pollution and a notable case study for the process of biomagnification.

Question 6 (1 MARK) 🌶

Microbeads can be

A. indirectly toxic.

B. intentionally included in cosmetic products by manufacturers.

C. up to 1 mm in length.

D. all of the above.

Question 7 (1 MARK) 🌶

Biomagnification is the process by which

A. organisms look larger under a microscope.

B. organic solvents contribute to the depletion of the ozone layer.

C. contaminant concentration increases in animals higher up the food chain.

D. contaminant concentration increases in a given organism over time.

Question 8 (3 MARK) 🌶🌶

Explain how microbeads can harm aquatic ecosystems.

Exam-style

Question 9 (9 MARKS) 🌶🌶

Plastics are a diverse class of organic compounds used for a variety of products in society. For the following applications of plastics in society, identify one advantageous property of the plastics used in these products, and explain how this property helps make these products useful and thus beneficial to society.

a. Car dashboards. (3 MARKS)

b. Total artificial hearts. (3 MARKS)

c. Nylon pulley-ropes. (3 MARKS)

Question 10 (4 MARKS)

1,1,2-Trichloro-1,2,2-trifluoroethane is a solvent used in refrigerators and aerosol sprays, which, like 1,1,1-trichloroethane, was almost entirely phased out of usage worldwide in 1996 due to its ozone-depleting property.

a. Draw the structural formula of 1,1,2-trichloro-1,2,2-trifluoroethane. (1 MARK)

b. Explain how the use of certain organic solvents like 1,1,2-trichloro-1,2,2-trifluoroethane can cause an increase in the intensity of harmful UV rays that hit earth. (3 MARKS)

Key science skills

Question 11 (8 MARKS)

The following is an excerpt from a media release by CSIRO (Commonwealth Scientific and Industrial Research Organisation) from 2001 regarding the potential toxicity of 'new car smell'.

> **New car drivers exposed to toxic emissions**
> New car headaches may involve more than minor warranty problems…
>
> …During its two-year study using three new motor vehicles from three weeks of their delivery to purchasers, CSIRO became aware of anecdotal reports, such as:
>
> - A solicitor who was ill for several days (headache, lung irritation, swelling) after collecting a new locally built car and driving it for only 10 minutes (the solicitor eventually swapped it for an 18-month-old car, which did not have any effect on her health)
> - A government worker who felt ill when driving new government cars during the first 6 months after their delivery
> - A chemically sensitised person who felt 'spaced out' when in any new car
> - A salesman who regularly updated his locally built car and found he became lethargic on long trips (e.g. from Melbourne to Geelong) when the car was new
>
> Source: CSIRO. (2001) *"New Car Drivers Exposed to Toxic Emissions"*. Commonwealth Scientific and Industrial Research Organisation.

a. Identify two necessary safety considerations when conducting measurements of potentially toxic gas levels and their effect on drivers. (2 MARKS)

b. "A salesman…found he became lethargic on long trips when the car was new"; would 'gases in new cars cause lethargy (tiredness)' be a valid conclusion to draw from this evidence? Explain. (3 MARKS)

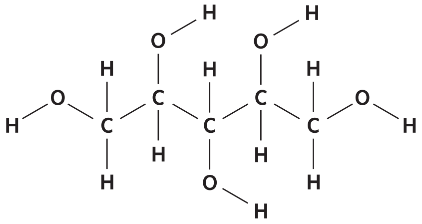
c. Explain three flaws in this component of the study, evident from this excerpt. (3 MARKS)

FROM LESSONS 16B & 16D

Questions from multiple lessons

Question 12 (6 MARKS)

Xylitol is the common name for an organic compound often added to food as an artificial sweetener.



a. Identify two other classes of organic compound commonly found in foods; one naturally present, and one additive. (2 MARKS)

b. Explain how Xylitol provides a benefit to society. (1 MARK)

c. Name the molecule pictured according to IUPAC naming conventions (note: for multiple of the same functional group, 2 = di, 3 = tri, 4 = tetra, 5 = pent). (3 MARKS)

FROM LESSON 7C

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Question 13 (9 MARKS)

The pesticide DDT ($C_{14}H_9Cl_5$) can be synthesised according to the following reaction:

$C_2H_3Cl_3O_2 + 2C_6H_5Cl \rightarrow C_{14}H_9Cl_5 + 2H_2O$

450 g of C_6H_5Cl is reacted completely with excess $C_2H_3Cl_3O_2$ to produce DDT.

a. Calculate the number of moles of C_6H_5Cl involved in this reaction. (2 MARKS)

b. Hence, calculate the number of moles of DDT produced. (2 MARKS)

c. Using your answer to part b, calculate the mass of DDT produced. (2 MARKS)

d. Explain two benefits that DDT provided to society and one environmental risk it posed. (3 MARKS)

FROM LESSON 6B

Hints

9a.	Consider the major risks posed by dashboards in car accidents.
9b.	Consider the properties a device would need in order to replace the human heart.
9c.	Consider the advantages nylon has over traditional fibres used for ropes.
10a.	The prefix 'di-' indicates that there are two instances of the given functional group in the molecule.
10b.	Consider the role and necessity of the ozone layer.
11a.	Consider how conducting the experiment may impact the driver.
11b.	Consider the exact evidence on which this conclusion is based.
11c.	Consider whether you think the evidence is reliable.
12a.	Nearly all components of food are organic compounds.
12b.	Compare Xylitol to the most common compound used for sweetening food.
12c.	First count the carbons in the longest chain, then consider the functional groups.
13a.	Equations for the number of moles of a substance can be found in the VCE Data Book.
13b.	Use the reaction equation given to determine the mole ratio of these compounds.
13c.	The molar mass of a compound can be calculated using the periodic table.
13d.	DDT was particularly useful to soldiers in tropical areas during World War 2, as well as farmers.

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Chapter 7 review

Multiple choice (10 MARKS)

Question 1 (1 MARK)

In semi-structural formulas, –COOH represents a __________ functional group.

A. hydroxyl

B. carboxylic acid

C. methyl

D. carboxyl

Question 2 (1 MARK)

Kay determines that a molecule has formula C_4H_8. Therefore, they can be absolutely certain that this molecule is

A. an alkane.

B. an alkene.

C. saturated.

D. branched.

Question 3 (1 MARK)

Crude oil is

A. a renewable source of hydrocarbons.

B. a mixture of hydrocarbons obtained from the fractional distillation of plant oils.

C. a mixture of hydrocarbons obtained from the earth using an oil drill.

D. a non-hydrocarbon form of oil.

Question 4 (1 MARK)

Which of the following correctly lists a set of organic molecules in order of increasing boiling point?

A. butanoic acid, butanol, butane, propane

B. butane, butanol, butanoic acid, propane

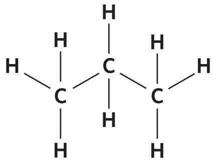
C. propane, butane, butanoic acid, butanol

D. propane, butane, butanol, butanoic acid

Question 5 (1 MARK)

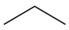
Which of the following is not a correct representation of propane?

A.



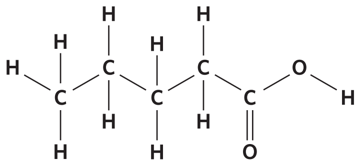
B. $CH_3CH_2CH_3$

C.



D. C_3H_{10}

Question 6 (1 MARK)



The name of the molecule shown is

A. pentanol.

B. pentanoic acid.

C. pentan-1-ol.

D. pentan-1-oic acid.

Question 7 (1 MARK)

The number of structural isomers of C_4H_9Cl is

A. 2.

B. 3.

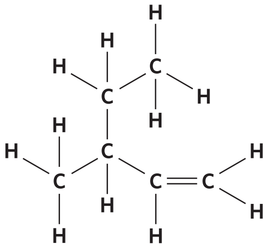
C. 4.

D. 5.

VCAA 2012 Exam 1 Multiple choice Q2

Question 8 (1 MARK)

The name of the following molecule is



A. 3-methylpent-1-ene.

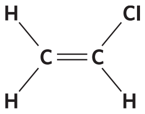
B. 3-ethylbut-1-ene.

C. 2-ethylbut-3-ene.

D. 2-ethylbutane.

Question 9 (1 MARK)

The molecule shown is used to make PVC, a plastic used in car interiors, pipes, and many more products. Its name, according to IUPAC nomenclature, is



A. chloroethene.

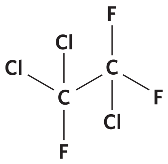
B. ethene chloride.

C. 2-chloroethene.

D. 1-chloroethene.

Question 10 (1 MARK)

The molecule pictured, known as CFC-113, was once used as a cleaning agent in the electronics industry, but has since been phased out. According to IUPAC conventions, its name is



A. chlorofluoroethane.

B. 1,1-dichloro-1-fluoro-2-chloro-2,2-difluoroethane.

C. 1,1,2-trichloro-1,2,2-trifluoroethane.

D. trichlorotrifluoroethane.

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Short answer (30 MARKS)

Question 11 (7 MARKS)

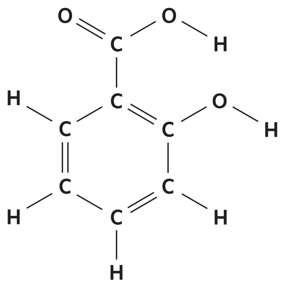
Crystal is a year 11 chemistry student who wants to study the interactions and properties of hydrocarbons with different functional groups. To begin with, she collects samples of ethanol and propanoic acid.

a. Draw the structural formulas of ethanol and propanoic acid. (2 MARKS)

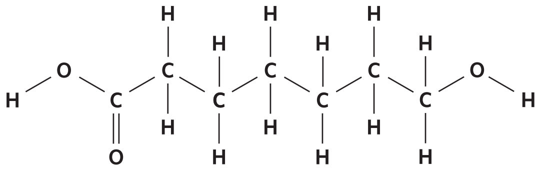
b. Crystal produces samples of pure liquid ethanol and propanoic acid and places them in test tubes. Following this, she increases the temperature of each sample, and records the temperature at which each begins boiling.

- **i.** Which sample would be expected to have a higher boiling point? Explain your answer. (3 MARKS)
- **ii.** Describe how the boiling points of ethanol and propanoic acid compare to ethane. (2 MARKS)

Question 12 (4 MARKS)



Salicylic acid



7-hydroxyheptanoic acid

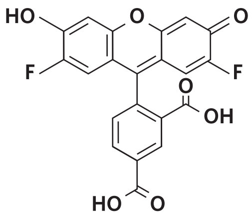
Pictured are the molecules 7-hydroxyheptanoic acid and salicylic acid, an important organic compound that can be derived from willow trees.

a. Determine the molecular formula of salicylic acid. (1 MARK)

b. Would the two molecules shown be considered isomers? Justify your answer. (2 MARKS)

c. Identify one use for salicylic acid in society. (1 MARK)

Question 13 (9 MARKS)



Dyes are often complex organic molecules. The structural formula of the synthetic dye 'Oregon Green' is shown.

a. Identify three functional groups that are present in 'Oregon Green'. (3 MARKS)

b. Explain the process by which most synthetic organic chemicals like 'Oregon Green' are produced. (3 MARKS)

c. Explain why 'Oregon Green' is classed as a non-renewable resource, and discuss why it is necessary to produce dyes and other organic compounds from biomass. (3 MARKS)

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Question 14 (10 MARKS)

Alandra is a chemist tasked with the job of labelling a series of unknown substances. The method used in this procedure is to determine the molecular formula of the substance through chemical analysis, and to determine the IUPAC systematic name of the molecule from this.

a. Categorise the following molecules as alkanes, alkenes or haloalkanes.

- **i.** CH_2CH_2 (1 MARK)
- **ii.** C_7H_{16} (1 MARK)
- **iii.** $CH_3CH_2CH_2CH_2Cl$ (1 MARK)
- **iv.** C_5H_{10} (1 MARK)

b. It is known that some of the molecules contain alkyl groups, and therefore structural isomers must be accounted for. Draw and name all the structural isomers of the molecules with the molecular formula C_5H_{12}. (6 MARKS)

Key science skills (10 MARKS)

Question 15 (4 MARKS)

Shivana has been given samples of three different carboxylic acids. She knows that one is methanoic acid, one is ethanoic acid, and one is propanoic acid, but she does not know which is which.

a. Devise an experimental method which Shivana could use to determine which carboxylic acid is which. (3 MARKS)

b. Name one safety precaution Shivana should take when handling the carboxylic acids. (1 MARK)

FROM LESSON 16B

Question 16 (6 MARKS)

A major children's stationery company is planning to release a new line of colourful markers, and they are investigating various organic solvents to use in these markers. Discuss considerations the company must make using the following themes:

a. Legal (2 MARKS)

b. Ethical (2 MARKS)

c. Environmental (2 MARKS)

FROM LESSON 16B

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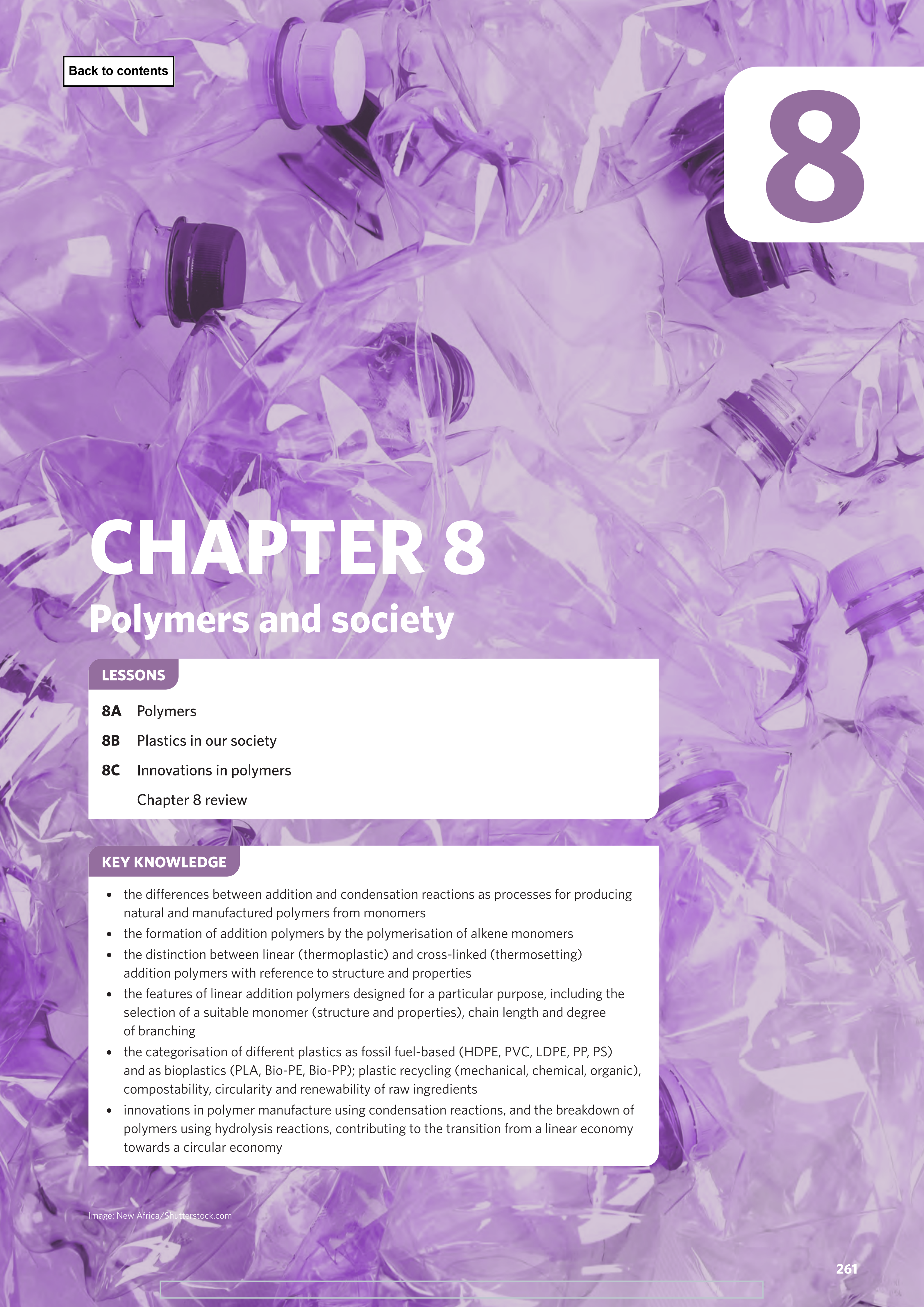

8

CHAPTER 8

Polymers and society

LESSONS

KEY KNOWLEDGE

- the differences between addition and condensation reactions as processes for producing natural and manufactured polymers from monomers
- the formation of addition polymers by the polymerisation of alkene monomers
- the distinction between linear (thermoplastic) and cross-linked (thermosetting) addition polymers with reference to structure and properties
- the features of linear addition polymers designed for a particular purpose, including the selection of a suitable monomer (structure and properties), chain length and degree of branching
- the categorisation of different plastics as fossil fuel-based (HDPE, PVC, LDPE, PP, PS) and as bioplastics (PLA, Bio-PE, Bio-PP); plastic recycling (mechanical, chemical, organic), compostability, circularity and renewability of raw ingredients
- innovations in polymer manufacture using condensation reactions, and the breakdown of polymers using hydrolysis reactions, contributing to the transition from a linear economy towards a circular economy

Image: New Africa/Shutterstock.com

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8A Polymers

STUDY DESIGN DOT POINTS

- the differences between addition and condensation reactions as processes for producing natural and manufactured polymers from monomers
- the formation of addition polymers by the polymerisation of alkene monomers
- the features of linear addition polymers designed for a particular purpose, including the selection of a suitable monomer (structure and properties), chain length and degree of branching
- the distinction between linear (thermoplastic) and cross-linked (thermosetting) addition polymers with reference to structure and properties

8A 8B 8C

1.2.9.1 Types of polymerisation reactions

1.2.10.1 & 1.2.11.1 Formation of addition polymers & properties of addition polymers

1.2.12.1 Thermoplastic vs. thermosetting polymers

ESSENTIAL PRIOR KNOWLEDGE

2A	Structural formulas
2B	Intramolecular bonds & intermolecular forces
2B	Properties of covalent compounds
7C	Functional groups

See questions 46–49.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: XXLPhoto/Shutterstock.com

Why are all plastics polymers, but not all polymers are plastics?

Many thousands of different **polymers** are produced by two basic types of chemical reactions. In this lesson, we will be discussing the formation and properties of polymers, as well as the applications of some of these molecules.

KEY TERMS AND DEFINITIONS

Addition polymerisation chemical reaction between monomers resulting in the formation of a polymer

Alkene hydrocarbon with at least one double carbon to carbon bond

Alkyne hydrocarbon with at least one triple carbon to carbon bond

Condensation polymerisation chemical reaction between monomers resulting in the formation of a polymer and water

Cross-link covalent bond between polymer chains

Cross-linked (thermosetting) polymer polymer containing cross-links that degrades or chars when heated

Elastomer polymer that forms occasional cross-links and is elastic

Functional group specific groups of atoms within a compound that affect the properties of the compound

High density polyethene (HDPE) tightly packed polymer produced at low pressures

Linear (thermoplastic) polymer polymer with no cross-links that can be remoulded when heated

Low density polyethene (LDPE) branched polymer produced at high pressures

Monomer molecule that can react with other molecules to form larger molecules (polymers)

Plastic molecule produced synthetically from monomers bonded together

Polymer molecule produced naturally or synthetically from monomers bonded together

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Types of polymerisation reactions 1.2.9.1

Polymers are covalent molecular substances made of many repeating units called **monomers** (figure 1).

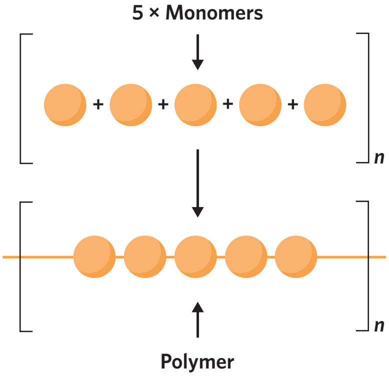


Figure 1 Polymerisation

What are the different ways polymers can be produced?

There are two main ways in which polymers can form – **addition polymerisation** (figure 2) and **condensation polymerisation** (figure 3).

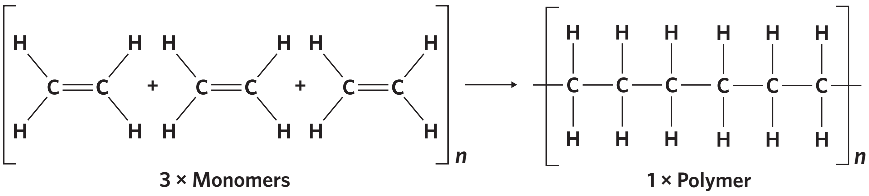
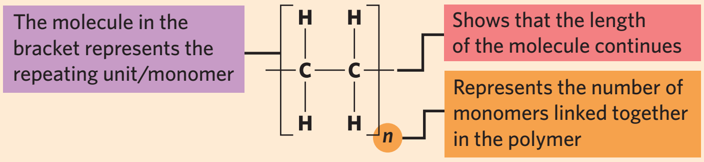


Figure 2 General formula for addition polymerisation

USEFUL TIP

Polymers can be represented using the following conventions:



The value of *n* varies between polymers molecules but is usually quite a large number, somewhere in the thousands.

$$R{-}A{-}H + R{-}B{-}OH \longrightarrow R{-}A{-}B{-}R + H_2O$$

Figure 3 General formula for condensation polymerisation

USEFUL TIP

R is an abbreviation that is used in organic chemistry for any group in which a carbon or hydrogen atom is attached to the rest of the molecule and this helps simplify what is occurring in reactions.

These two types of polymerisation reactions have many differences, and these are summarised in table 1.

Table 1 Comparison of addition and condensation reactions

Addition polymerisation	Condensation polymerisation
Alkene or **Alkyne** monomers Carbon to carbon double or triple bond is broken, allowing carbon atoms to form bonds with adjacent monomers Results in formation of a polymer	Monomers have different functional groups. One is a carboxylic acid, the other is an alcohol or an amine. Two **functional groups** react to form a bond between monomers Results in formation of a polymer and water

Condensation polymerisation occurs both naturally and artificially. In fact, it occurs in our own bodies, through the formation of proteins, as shown in figure 4. In this case, amino acid monomers form a polymer called a polypeptide.

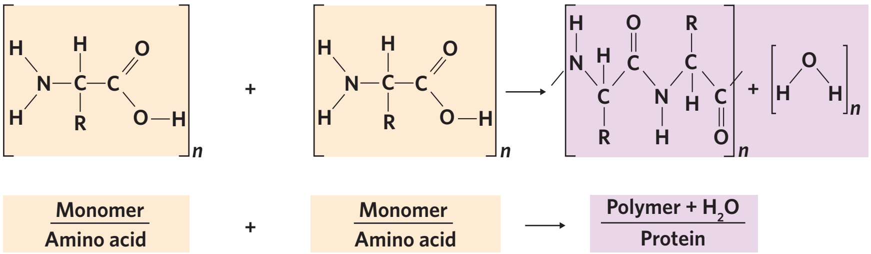


Figure 4 Condensation polymerisation of two amino acids

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Polyesters, a class of synthetic polymers with a huge variety of uses, are manufactured via condensation polymerisation (figure 5).

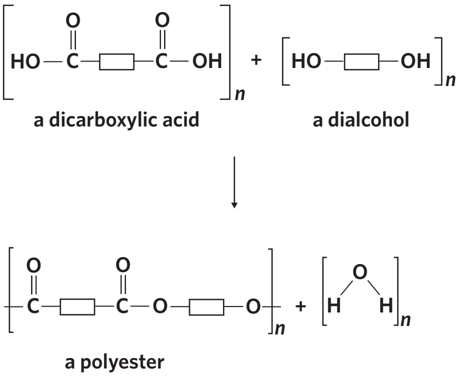


Figure 5 General formula for the formation of polyesters

USEFUL TIP

When there are two hydroxyl or carboxyl groups in a compound, the suffix 'di' is put before the functional groups name e.g. ethandiol or ethandioic acid.

USEFUL TIP

The symbol —▭— is also used to represent chains of any elements bonded together which allows the polymerisation bonds to be seen more clearly.

MISCONCEPTION

'All polymers are **plastics**.'

The key difference between a polymer and a plastic is that the polymer can be either natural or synthetic, whereas a plastic is always a synthetic polymer.

Progress questions

Question 1

__________ are covalently bonded to form __________.

A. Monomers, polymers

B. Polymers, monomers

Question 2

Addition polymerisation produces water as a by-product.

A. True

B. False

Question 3

Polyester is an example of a manufactured polymer, while polypeptide chains are an example of

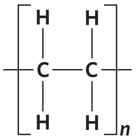
A. a natural polymer.

B. an artificial polymer.

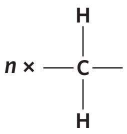
Question 4

Which of the following is an appropriate representation of a repeating unit of a polymer?

A.



B.



Formation of addition polymers & properties of addition polymers 1.2.10.1 & 1.2.11.1

There are thousands of different addition polymers that have been designed for a wide range of uses.

How can polymers be designed for a purpose?

Polymers are highly modifiable and many different polymers possess their own unique properties.

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ymers are

- light and strong,
- durable,
- highly versatile and modifiable (can possess many colours and characteristics),
- flammable,
- chemically resistant (particularly against acids),
- and, effective thermal and electrical insulators.

Table 2 gives some examples of different monomers which can undergo addition reactions to form useful polymers.

When thinking of the desired properties of a polymer, the choice of monomer is very important. Polymers made up of non-polar monomers are held together only by dispersion forces. However, when a polymer is formed from polar monomers, strong intermolecular forces such as dipole-dipole interactions or hydrogen bonds hold polymer chains together (figure 6). As a result, polar polymers are harder and more rigid than non-polar polymers.

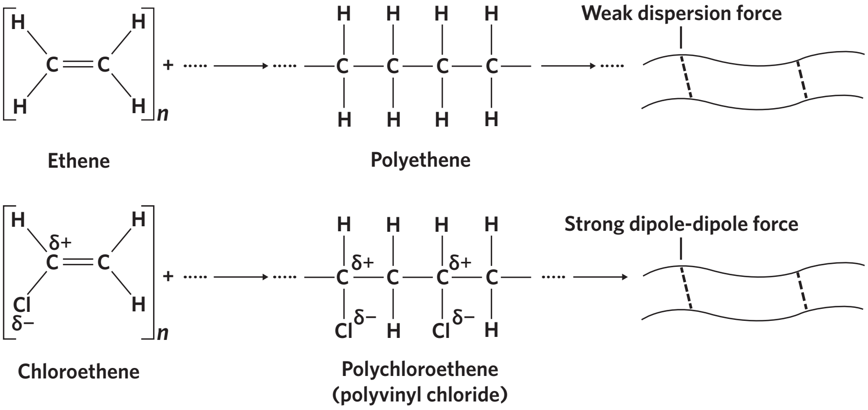


Figure 6 Polar vs non-polar polymers

An example of a non-polar monomer is tetrafluoroethene. The polymer formed from tetrafluoroethene is polytetrafluoroethene, also known as Teflon. Due to its structure, there are strong dispersion forces between polytetrafluoroethene chains, however, there are almost no dispersion forces between these chains and other substances. This is what gives Teflon its non-stick quality.[1]

Polyacrylonitrile is an example of a polar polymer. This polymer is used to make materials such as carbon fibre which is found in objects such as tennis racquets and fishing rods. The carbon-nitrogen bond in acrylonitrile – the monomer from which polyacrylonitrile is formed – creates permanent dipole-dipole interactions between chains, giving materials like carbon fibre high tensile strength and stiffness.

It is not only the monomer that will determine the properties of a polymer but also the size and branching of polymer chains. The greater the chain length of a polymer, the stronger the intermolecular forces between chains will be. This is because there is an increased surface area over which chains can interact. Some materials are made up of very large polymer chains and can be used to make Kevlar® bulletproof vests (figure 7). The strong intermolecular forces make the material very tough.

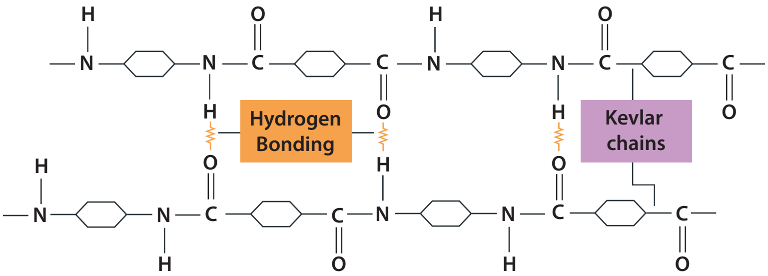
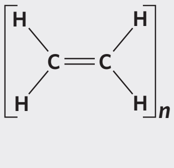
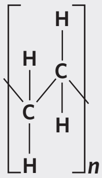
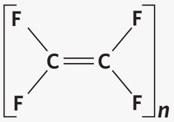
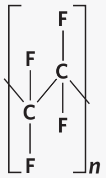
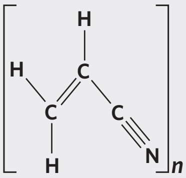
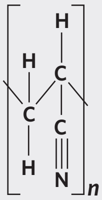


Figure 7 Polymer chains in bulletproof vests

Table 2 Addition polymerisation of monomers

Monomer	Polymer
 Ethene	 Polyethene
 Tetrafluoroethene	 Polytetrafluoroethene / Teflon
 Acrylonitrile	 Polyacrylonitrile

KEEN TO INVESTIGATE?

[1] **What are the real world applications of Teflon?**
Search YouTube: Why doesn't anything stick to Teflon?

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When polymerisation takes place at high temperatures and high pressures, it prevents polymers from forming linear chains. Polyethene chains produced under these conditions are known as **low density polyethene** (LDPE) and possess many branches extending from the parent chain. Figure 8 contains the atomic and molecular structures of LDPE, as well as a real world example.

MISCONCEPTION

'High pressures are used to produce HDPE.'

High pressure conditions create branching, which means the polymer will have a lower density.

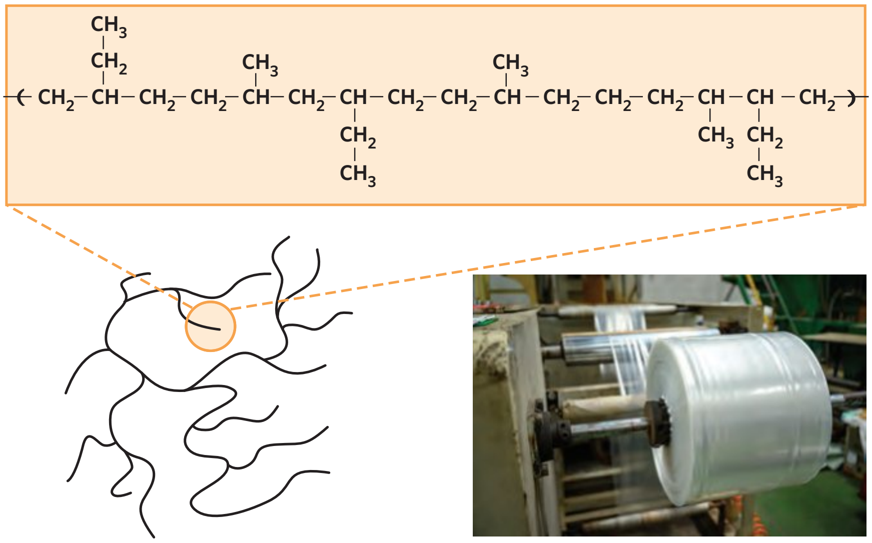

Image: Juan Enrique del Barrio/Shutterstock.com

Figure 8 Structure and example of low density polyethene

The presence of these branches prevents low density polyethene from packing together closely (figure 9), leading to weakened dispersion forces. As a result, low density polyethene possesses the following properties:

- Lower density (relative to HDPE)
- Relatively soft
- Lower melting point (105–115 °C) (relative to HDPE)
- Good chemical resistance
- Opaque, but transparent in thin forms
- Insulator of electricity

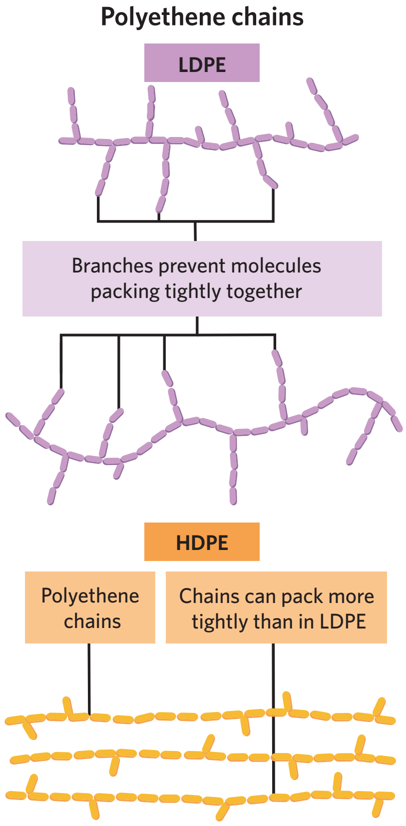

Figure 9 Branching affecting the density of polymers

In contrast to low density polyethene, **high density polyethene** (HDPE) is produced under lower temperature and pressure conditions. Thus, there is little branching and polymer chains pack together tightly (figure 9).

The properties of HDPE are as follows:

- Higher density (relative to LDPE)
- Hard
- Higher melting point (120–180 °C) (relative to LDPE)
- Weatherproof and cold-resistant
- Good chemical resistance
- Allows light to pass through
- Insulator of electricity

Figure 10 illustrates the structure of HDPE and one of its uses (container for cleaning products).

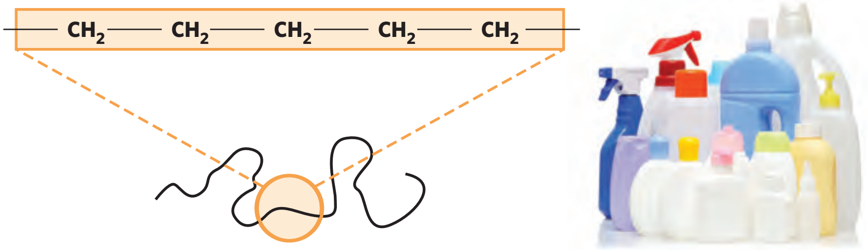

Image: MrJPEG/Shutterstock.com

Figure 10 High density polyethene

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ome examples of polymers that have properties that are designed for a particular purpose.

Table 3 The properties and applications of different polymers

Polymer	Type of polymerisation	Properties	Applications	Reference image
Polypropene	Addition	Durable, cheap	Artificial grass, ice-cream containers and rope	 Image: red mango/Shutterstock.com
Polyvinylidene chloride	Addition	Sticks to itself	Food wrap	 Image: New Africa/Shutterstock.com
Nylon (polyamide)	Condensation	Strong, resilient, elastic	Carpets, clothing, fishing line and toothbrush bristles	 Image: Pornprapa Korprasert/Shutterstock.com
Polystyrene	Addition	Can be solid or foamed	Insulation foam, cups, protective packaging, and filling for beanbags	 Image: Anna Hoychuk/Shutterstock.com

Progress questions

Question 5

Polar monomers tend to form harder, more rigid polymers than non-polar monomers.

A. True

B. False

Question 6

Increasing the chain length of polymers

A. increases strength due to increased intermolecular forces.

B. decreases strength due to increased packing.

Question 7

Increasing branching in polymers

A. allows polymer chains to pack tightly together.

B. decreases the strength of intermolecular forces.

Question 8

High density polyethene is produced under high pressure conditions.

A. True

B. False

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Image: Mr_Mrs_Marcha/Shutterstock.com

Figure 11 Thermoplastic pellets

Thermoplastic vs. thermosetting polymers 1.2.12.1

Based on their response to heat, polymers can be separated into two major categories – **linear (thermoplastic) polymers** and **cross-linked (thermosetting) polymers**.

How do the structure and properties of linear and cross-linked polymers differ?

When heated, linear polymers will soften, which enables them to be reshaped or remoulded. In fact, linear polymers are commonly sold as pellets, which are heated and then moulded to make a variety of plastic products (figure 11).

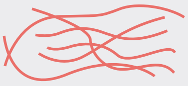
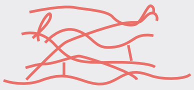
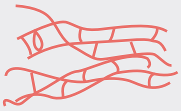
The intramolecular bonds of linear polymers are much stronger than the intermolecular forces holding the chains together. Upon exposure to sufficient heat, individual chains become mobile and the plastic can be reshaped.

Covalent bonds between polymer chains are referred to as **cross-links**. The presence of cross-links in a molecule greatly restricts the movement of a molecule, making it more rigid and heat resistant. Due to these properties, cross-setting polymers are used to make tough plastics, such as bowling balls. **Elastomers** also have cross-links but to a lesser extent than thermosetting polymers. The cross-links in elastomers are elastic, meaning the polymer can be stretched and will return to its original shape. Unsurprisingly, rubber bands are an example of elastomers.

Cross-linked polymers decompose when exposed to heat due to covalent bonding between their chains. Their chains cannot be mobilised as there is no difference in the strength of the bonds within chains or the bonds between chains. Breaks can take place at any point in the polymer and will cause degradation and decomposition.

The properties of different types of polymers are summarised in table 4.

Table 4 Summary of thermoplastics, elastomers and thermosets

Property	Thermoplastics	Elastomers	Thermosets
Cross-links	None	A small number	A much larger number
Forces between chains	Weak intermolecular forces	Weak intermolecular forces and occasional strong covalent bonds	Strong covalent bonds
Response to heat	Softens / melts	Varies	Degradation / decomposition
Hardness	Varies	Varies	Generally hard
Diagram	 Thermoplastic	 Elastomer	 Thermoset

Progress questions

Question 9

Linear polymers

A. are able to be remoulded when heated.

B. are resistant to remoulding.

Continues →

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Question 10

The bonds formed in thermosetting polymers are

A. cross-links.

B. joining bonds.

Theory summary

- Polymers are formed from both the addition and condensation polymerisation of monomers.
- The polarity of monomers affects the properties of a polymer.
- Chain length and branching affect the properties of a polymer.
- LDPE is produced at higher pressure than HDPE and consequently has increased branching.
- Linear polymers have weak intermolecular forces and can be remoulded when heated.
- Cross-linked polymers decompose (char) when exposed to intense heat.
- Elastomers have occasional cross-links, giving them elasticity.

The content in this lesson is considered fundamental prior knowledge to reactions of organic compounds (Unit 4 AOS 1).

8A Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 11–13.

A year 11 chemistry student who has recently learned about different types of polymers has set up a series of experiments to determine the difference between thermoplastic and thermosetting polymers.

Question 11 (1 MARK) 🌶

Which of the following best describes thermoplastic polymers?

A. Polymers with only weak intermolecular forces between chains, enabling remoulding upon heating

B. Polymers with covalent bonding between chains, meaning they do not degrade when being reshaped

C. Polymers with covalent intermolecular bonds which, when broken, cause the molecules to decompose

D. Polymers with highly mobile chains at all temperatures

Question 12 (1 MARK) 🌶

Which of the following best describes thermosetting polymers?

A. Polymers with weak intermolecular forces but strong intramolecular bonds, enabling remoulding upon heating

B. Polymers with covalent bonds between chains, preventing decomposition upon exposure to heat

C. Polymers with covalent intermolecular bonds which prevent chain mobilisation

D. Polymers with occasional covalent bonds between chains, giving them the property of elasticity

Question 13 (4 MARKS) 🌶🌶

Would a thermoplastic or thermoset polymer be the most suitable material for the following applications? Justify your answer using a property of these polymers.

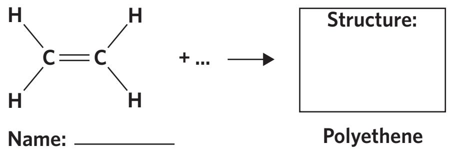
a. Squeezable shampoo bottle (2 MARKS)

b. Knob of a saucepan lid (2 MARKS)

Question 14 (6 MARKS)

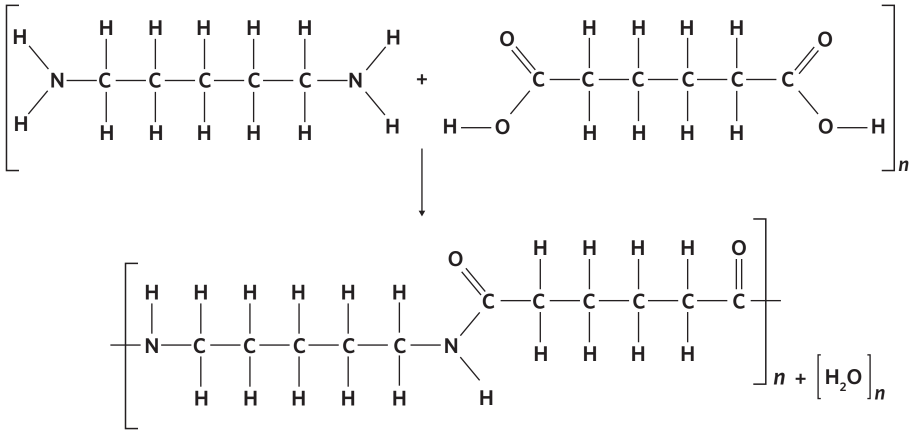
Searching through the ruins of an exploded research lab, Mei, a forensic investigator, finds the remains of a lab report.

a. She sees a template for a chemical reaction but there are some key elements missing. Give the missing name and structure. (2 MARKS)



b. Name the type of reaction that has occurred. (1 MARK)

c. Mei finds another fragment, this time with a complete reaction. What type of reaction is this? (1 MARK)



d. Compare these two reactions, identifying a similarity and a difference. (2 MARKS)

Question 15 (7 MARKS)

Leo is employed by a ballistics company to manufacture a polymer for use as a bullet-proof material. He is told he must use the molecule, ethene (C_2H_4) in the production process.

a. What type of reaction will Leo use to make this polymer? (1 MARK)

b. Leo is warned by his supervisor to carefully monitor the temperature and pressure at which this reaction is occurring. Explain the effect of these variables on polyethene's structure. (2 MARKS)

c. Under what conditions should Leo create this polymer? Justify your answer. (3 MARKS)

d. What is the name of the polymer these conditions will produce? (1 MARK)

Key science skills

Question 16 (5 MARKS)

Elisa wanted to test the properties of different types of plastics. As part of her experiment, she collected different items and tested the 'bendability' of the plastic. The results are provided.

Sample A	Sample B	Sample C
Was able to continually bend the plastic and it didn't rip or break.	Was very hard to bend the plastic. The plastic eventually snapped when I applied too much pressure.	Was able to bend the plastic backwards and forwards without it breaking.

a. Identify the dependent variable. (1 MARK)

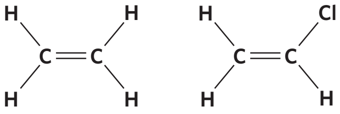
b. What type of data is being collected in the experiment? (1 MARK)

c. Comment on the reproducibility of the experiment. (3 MARKS)

FROM LESSONS 16B & 16D

Question 17 (9 MARKS)

Maya is given the choice of two monomers to design a polymer. Their respective structural formulas are shown.



Ethene Chloroethene

a. Identify the polarity of these two monomers. Label any permanent dipoles. (3 MARKS)

b. Maya is indecisive and makes two polymers – polyethene and polychloroethene. She tests the heat resistance of both polymers. Based on the intermolecular forces of these polymers, which would you expect to have the highest melting point? (3 MARKS)

c. Unexpectedly, Maya finds that polyethene has a higher melting point than polychloroethene. How might this be possible? (3 MARKS)

FROM LESSONS 2A & 2B

Hints

14a.	This is a polymerisation reaction.
14b.	Think of the two different types of polymerisation reactions.
14c.	Think of the two different types of polymerisation reactions.
14d.	Think of the reactants and products of both reactions.
15a.	Think of the two different types of polymerisation reactions.
15b.	Think in terms of branching.
15c.	Branching affects intermolecular forces.
15d.	This polymer is named according to its density.
16a.	The dependent variable is affected by the independent variable.
16b.	The type of data is affected by the nature of the data collected.
16c.	Reproducibility measures the closeness of results between experiments in certain conditions.
17a.	Polar molecules have permanent dipoles which do not cancel each other out.
17b.	Stronger intermolecular forces result in a higher melting point.
17c.	Consider variables other than the choice of monomer that affect a polymer's properties.

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8B Plastics in our society

STUDY DESIGN DOT POINT

- the categorisation of different plastics as fossil fuel-based (HDPE, PVC, LDPE, PP, PS) and as bioplastics (PLA, Bio-PE, Bio-PP); plastic recycling (mechanical, chemical, organic), compostability, circularity and renewability of raw ingredients

8A 8B 8C

1.2.13.1 Categories of plastics
1.2.13.2 Plastic recycling
1.2.13.3 Compostability
1.2.13.4 Circularity and renewability of raw ingredients

ESSENTIAL PRIOR KNOWLEDGE

7D Fossil fuels versus biomass
7E Microplastics
8A Thermoplastics and thermosets
See questions 50–52.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

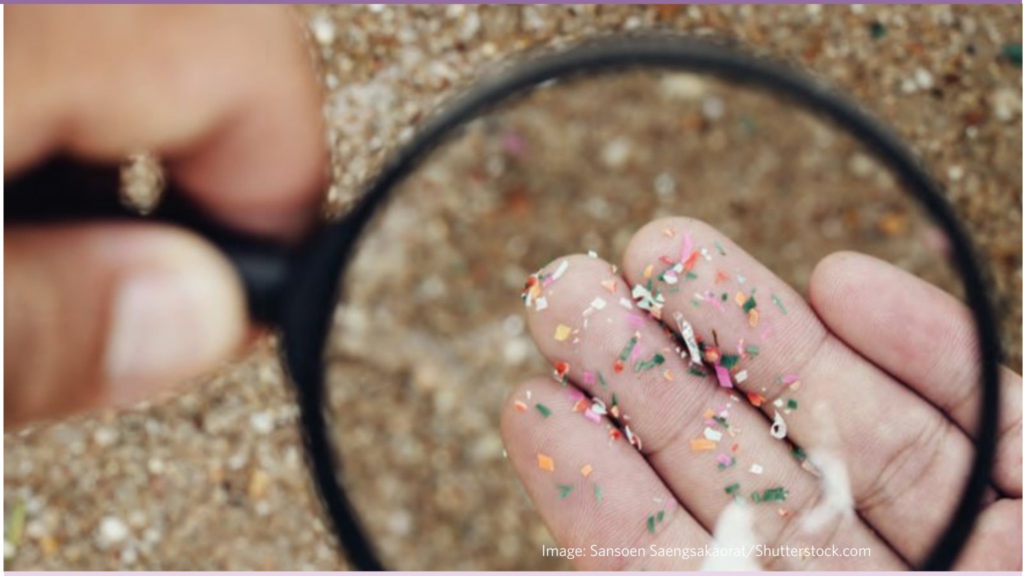

Image: Sansoen Saengsakaorat/Shutterstock.com

Why might 'biodegradable' plastics actually harm the environment?

An increasing number of plastic products are marketed as **biodegradable.** However, this label may be misleading as many of these plastics take years to decompose and often leave even more harmful residue: microplastics. In this lesson, the role, advantages, and drawbacks of various plastics and recycling methods in society will be explored.

KEY TERMS AND DEFINITIONS

Biodegradable ability to decompose in the environment

Bioplastics plastics produced from biomass

Carbon negative describes a process that absorbs more carbon dioxide than it produces

Chemolysis the use of solely chemical substances to decompose organic substances into simpler ones

Circular economy a continuous cycle that focuses on the optimal use and re-use of resources from the extraction of raw materials through to production of new materials, followed by the consumption and re-purposing of unused and waste materials

Compostable describes a polymer that can be at least 90% decomposed after 180 days in a composting environment

Dehydrate to remove H_2O from a molecule or ion

Dehydrogenate to remove a hydrogen atom or atoms from a molecule or ion

Feedstock raw material used for producing another product

Greenhouse gases molecules that absorb infrared radiation from the sun

Humus dark, nutrient-rich, organic material produced from composting

Linear economy operates on a 'take-make-dispose' model, making use of resources to produce products that will be discarded after use

Plastic dissolution dissolving plastic to extract polymers and separate them from their additives

Pyrolysis decomposition brought about by high temperatures

Steam cracking the breaking down of larger saturated hydrocarbons into smaller, often unsaturated ones

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Categories of plastics 1.2.13.1

Similar to the other organic compounds explored in chapter 7, plastics can be sorted into those sourced from fossil fuels, and those sourced from plant-based biomass.

What are the major types of fossil fuel-based plastics?

The vast majority of widely-used plastics in society are produced from fossil fuels, particularly crude oil. As discussed in lesson 7D, although crude oil is currently the most widely-applicable and accessible source of plastics, long-term reliance on fossil fuel-based plastics is not sustainable.

HDPE & LDPE

High-density polyethene (HDPE) and low-density polyethene (LDPE) are thermoplastics formed from the polymerisation of ethene monomers. Ethene is produced by a process called **steam cracking**, in which larger hydrocarbons, extracted from crude oil, are mixed with steam and heated in the absence of oxygen to produce small, unsaturated hydrocarbons. HDPE is used in milk jugs, detergent bottles, outdoor furniture, and crates, whilst LDPE is primarily used for plastic bags and wraps; the differences between HDPE and LDPE are explained in detail in lesson 8A.

PVC

Polyvinyl chloride (PVC), commonly known as vinyl, is a widely-used plastic made up of chloroethene (‘vinyl chloride’) monomers. These chloroethene monomers are synthesised by reacting chlorine gas with ethene, which is produced from crude oil in the cracking process described above. PVC is used mainly for pipes, electrical cable insulation, vinyl records, and as an alternative to leather. Its use in car interiors is explored in lesson 7E.

PP

Polypropene (PP) is a plastic composed of propene monomers, which, like ethene, are produced by the steam cracking of larger hydrocarbons from crude oil. PP’s main applications are in flexible and rigid packaging, rugs, bags, cars, and other textiles.

PS

Polystyrene (PS) is a distinctive plastic that is made up of styrene monomers. To produce it, ethylbenzene (a compound naturally present in fossil fuels like coal tar and crude oil) is **dehydrogenated** and then polymerised to form polystyrene: a process summarised in figure 1. Its soft, lightweight properties make it suitable for widespread usage in packaging, polystyrene cups, and insulated boxes.

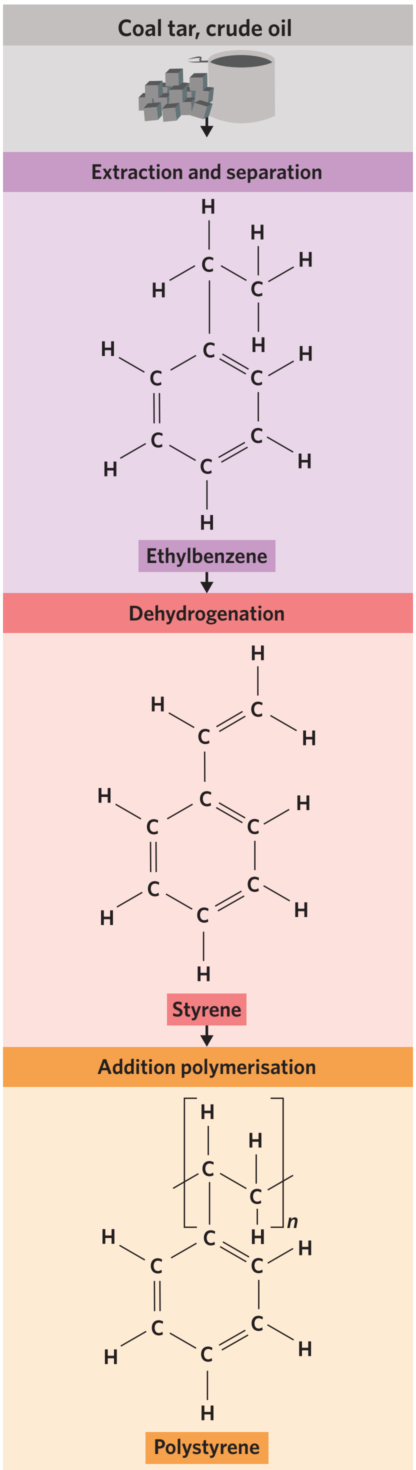

Figure 1 Production of polystyrene from fossil fuels

What are the major types of bioplastics?

Bioplastics are plastics produced from plant-based biomass. Many varieties of bioplastics are becoming increasingly popular as an alternative to their fossil fuel-based counterparts.

PLA

Polylactic acid (PLA) is a condensation polymer produced from fermented plant starch from crops, such as sugarcane or corn, whose synthesis and applications are discussed further in lesson 8C. Some of its properties include:

- low melting point,
- high strength,
- and good adhesivity between layers.

These properties have led to PLA being the most widely used filament plastic in 3D printing (figure 2), an innovation with the potential to replace traditional manufacturing methods. It is also used for disposable tableware, cutlery, upholstery, and clothing.

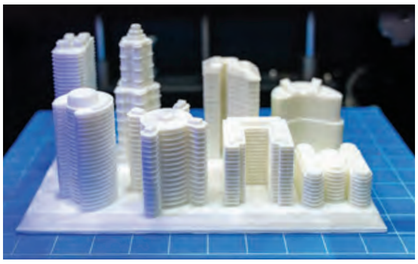

Image: asharkyu/Shutterstock.com

Figure 2 Model of a city 3D-printed using PLA

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Bio-PE

Bio-polyethene (Bio-PE), sometimes referred to as renewable polyethene, has the same formula as HDPE and LDPE, but is derived from plants. Crops, such as wheat and sugar cane, are fermented to produce ethanol, as discussed in lesson 7D, which is then **dehydrated** to form ethene and polymerised, as per figure 3. Bio-PE can be produced as HDPE or LDPE and is indistinguishable from its fossil fuel-based counterparts. The crops used to produce Bio-PE absorb carbon dioxide, meaning Bio-PE production can be **carbon negative**, hence reducing **greenhouse gas** emissions. However, this process requires a large area, and rapid harvesting can cause land degradation.

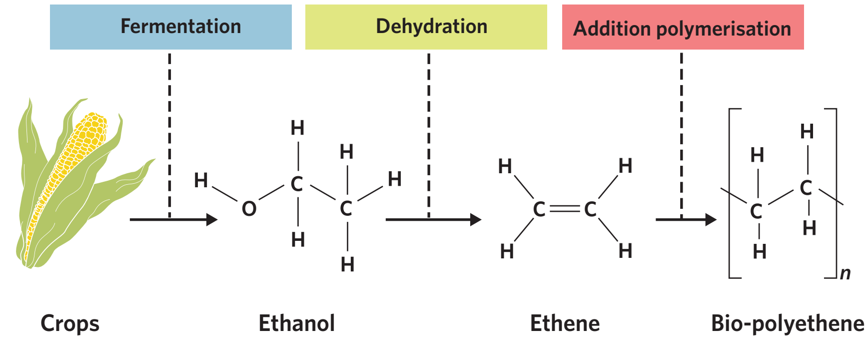

Figure 3 Production of Bio-PE from crops

Bio-PP

Bio-polypropene (Bio-PP) is polypropene primarily produced from crops, as with Bio-PE, or from vegetable oils. This production process is shown in figure 4. Also, similarly to Bio-PE, Bio-PP is a drop-in bioplastic, meaning it is chemically identical to fossil fuel-based polypropene. Therefore, for Bio-PP products, the same manufacturing equipment can be used as for their PP alternatives.

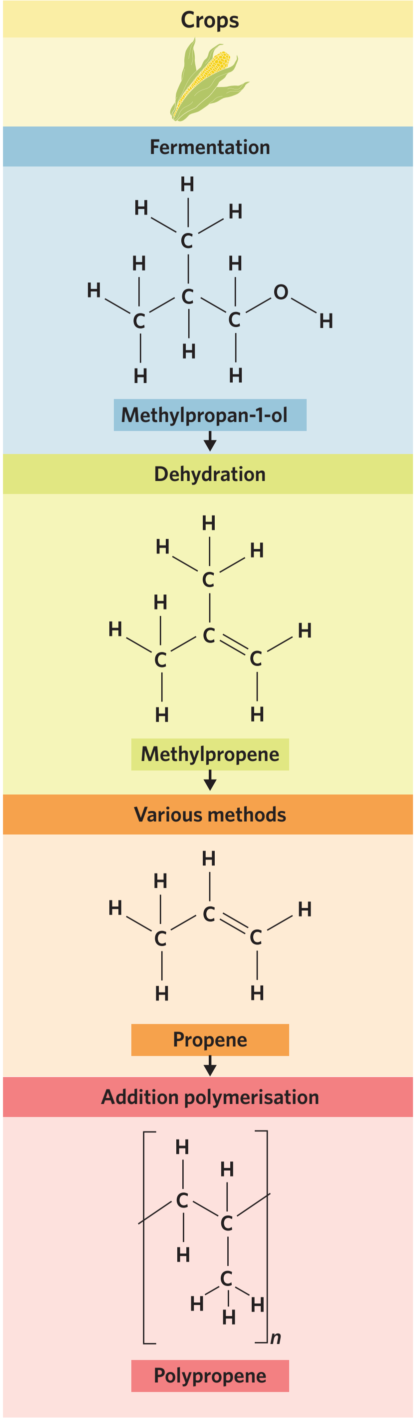

Figure 4 Production of Bio-PP from crops

Progress questions

Question 1

The majority of plastics currently used in society are

A. fossil fuel-based.

B. bioplastics.

Question 2

Bio-PP and fossil fuel-based polypropene function differently from each other.

A. True

B. False

Question 3

PLA is a major type of

A. fossil fuel-based plastic.

B. bioplastic.

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recycling 1.2.13.2

To reduce waste and demand for raw materials, plastics are regularly recycled in society, which involves collecting plastic items that would otherwise be discarded, processing them, and then repurposing them into new items. In theory, most plastics are recyclable, but some are much more recyclable than others.

How can plastics be recycled?

On plastic packaging, it is common to see a small triangle with a number in it; this is a recycling code and it indicates the type of plastic resin used in the item (figure 5). For example, PETE (polyethylene terephthalate) is most often used in plastic bottles, and it is given recycling code 1. Since plastics vary significantly in their composition and properties, these recycling codes help enable plastics to be separated before being recycled.

Figure 5 Recycling codes

USEFUL TIP

In Australia, plastics of all recycling codes can be placed in recycling bins, as plastics are separated in large recycling facilities after collection.

Following collection, plastics are sorted both automatically and manually, before being washed and dried to remove impurities. Then, plastics are recycled according to three main methods: mechanical, chemical, and organic (a.k.a. composting).

Mechanical recycling is the most common type of plastic recycling in society, involving little alteration of the chemical structure of plastics. In this process, plastics are shredded into small flakes, which are melted and crushed together to form pellets. These pellets are then sent to manufacturers as raw materials to make new products from the same plastic (figure 6).

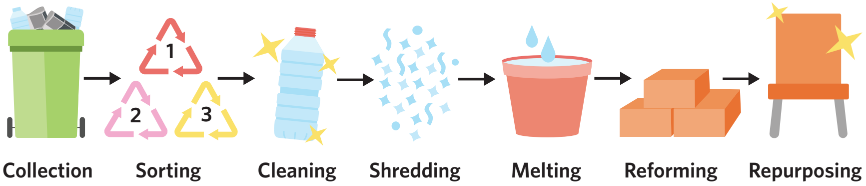

Figure 6 Mechanical recycling process

Generally, thermosets cannot be recycled, while all thermoplastics can be mechanically recycled. However, the products of mechanical recycling can usually only be used for lower-grade plastic products. For example, due to the risk of toxic contamination in the recycling process, plastic food containers can usually only be recycled into products like outdoor furniture or plant holders. Additionally, some plastics, like PVC, are more difficult and resource-intensive to recycle than others, and so not all recycling facilities consider recycling them to be worthwhile.

Chemical recycling, also referred to as advanced or **feedstock** recycling, involves the conversion of plastics back into either their constituent polymers, monomers, or organic chemicals, rather than simply the plastics themselves. One such method is **plastic dissolution**, involving the heating of plastics with solvents to produce a solution of polymers and additives, which can then be separated. Alternatively, **chemolysis** is a method of depolymerising plastics, in which chemical reactions are used to break a plastic up into its constituent monomers. However, this method can only be used for plastics from condensation polymers, like nylon and polyester.

High-temperature techniques, such as **pyrolysis** (heating to 300–900 °C in the absence of oxygen), can also be used to convert plastics to raw organic chemicals, which can then be used as fuels to produce energy. Like many other chemical recycling methods, pyrolysis is, at present, prohibitively expensive and hence not in mainstream usage, but research is continually undertaken to improve these methods. Figure 7 summarises chemical recycling methods.

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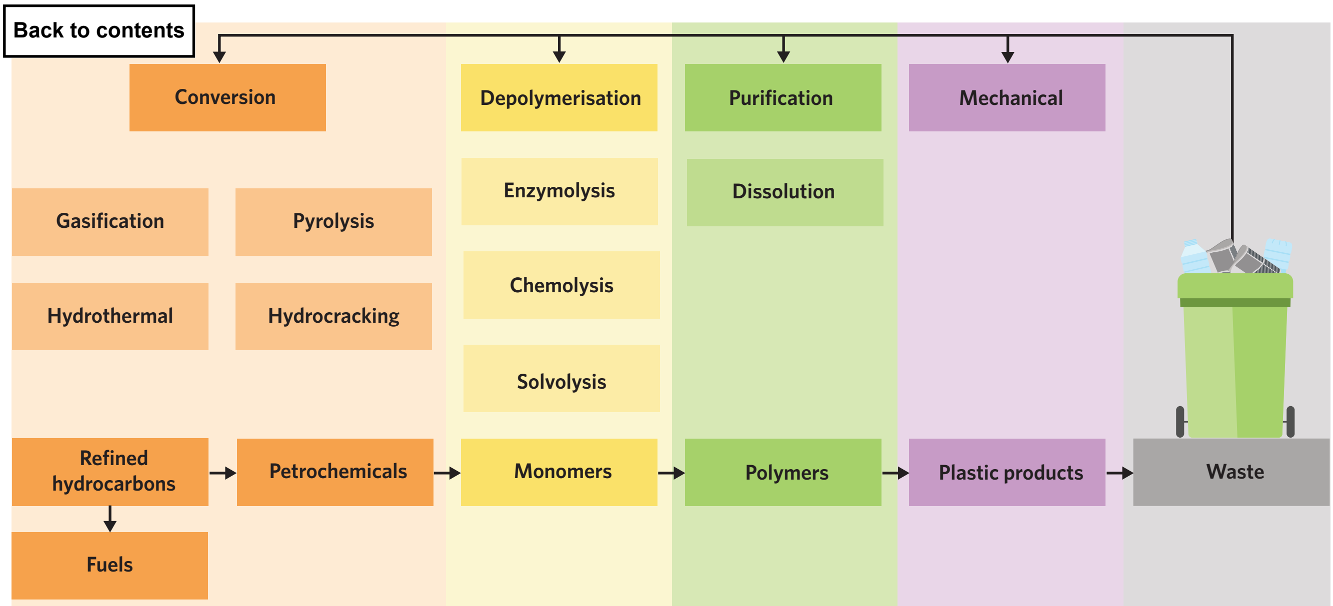

Figure 7 Chemical recycling methods and products

Progress questions

Question 4

All thermoplastics are recyclable.

A. True

B. False

Question 5

Plastics can only be recycled into the same type of plastic.

A. True

B. False

Compostability 1.2.13.3

Plastic composting, sometimes referred to as organic recycling, involves plastics being broken down by microorganisms. However, only some plastics are **compostable**, and there is some doubt as to the environmental benefits (or otherwise) of composting plastics.

How compostable are different plastics?

Many commercially available plastic products are marketed as biodegradable. This means that these plastics will naturally break down in the earth. However, this could be a process that takes millions of years, and may leave toxic residue or microplastics in the soil (figure 8), thereby harming (rather than helping) the environment. As explored in lesson 7E, tiny plastics are a huge problem; microplastics can enter aquatic environments, build up in the food chain, and devastate ecosystems.

Image: SIVStockStudio/Shutterstock.com

Figure 8 Biodegradable plastics often leave microplastics in soil.

MISCONCEPTION

‘Bioplastics are the same as compostable plastics.’

Some, but not all, bioplastics are compostable, and some, but not all, fossil fuel-based plastics are compostable. Bioplastics are defined by the source of their materials, while compostable plastics are defined by how easily they can be broken down.

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and, compostable plastics are ones that biodegrade in a short time period without leaving any toxic residue or microplastics, only nutrients for soil. Similar to the way organic waste, like fruit and vegetable scraps, can be broken down by worms or other small organisms, compostable plastics are able to be decomposed by microorganisms under certain conditions. If a plastic is labelled 'compostable', it does not mean it will break down in your garden compost heap. Compostable plastics require the high-temperature environment of an industrial composter (figure 9) to break down rapidly. In these composting environments, moisture, oxygen, and heat levels are carefully controlled to ensure a maximal breakdown rate.

The bacteria in a composter struggle to break down the long polymer chains in plastic. To account for this, the compostable plastic is first shredded, often before being reacted with water at high temperatures to break it down into its constituent monomers. Then, these bacteria decompose the monomers to produce water vapour and carbon dioxide, which are released into the atmosphere, as well as **humus** (figure 10), which is used in agriculture or home gardens to promote plant growth.

Compostable bioplastics include PLA and PHB (polyhydroxybutyrate), a polymer that is produced by microorganisms under stress. Although not widely employed as of now, PHB is used in the medical industry for internal stitches as it is non-toxic and breaks down easily in the body. Few fossil fuel-based plastics are compostable, though one exception is PBAT (polybutylene adipate terephthalate, figure 11), a tough and flexible plastic that can be used to make cling wrap for food packaging and compostable plastic bags.

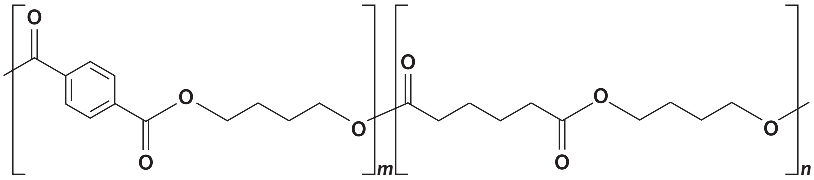

Figure 11 Structure of PBAT, a compostable plastic

Progress questions

Question 6

Plastic composting is sometimes referred to as

A. organic recycling.

B. chemical recycling.

Question 7

All compostable plastics are

A. bioplastics.

B. biodegradable.

MISCONCEPTION

'Biodegradable plastics are the same as compostable plastics.'

In theory, all plastics can technically be labelled biodegradable, since they eventually break down in the environment by natural processes; however, this could take millions of years and produce toxic leftovers. On the other hand, to be labelled compostable, a plastic must be at least 90% broken down after 180 days in an industrial composter, without leaving any harmful residue.

Image: Vitaliy Kyrychuk/Shutterstock.com

Figure 9 Industrial composter

Image: OoddySmile Studio/Shutterstock.com

Figure 10 Humus

USEFUL TIP

Humus is not to be confused with the dip hummus. Unlike hummus, humus would not taste nice on pita bread!

Circularity and renewability of raw ingredients 1.2.13.4

As was explored in lesson 3C, 'the **circular economy**' refers to the cycle of using and reusing resources in an optimal way. Like metals, plastics can also contribute to this through renewable production and recycling, yet true circularity may not be feasible.

How can plastics contribute to the circular economy?

Of all plastic generated by humans, only 9% has been recycled. As a result, the current plastic economy is considered a **linear economy**: the vast majority of plastic is created from raw materials, used once, then left in the ground for millions of years thereafter. Currently, producing new fossil fuel-based plastics is cheaper than using recycled or renewable alternatives, meaning manufacturers are not inclined to shift away from linear production methods. Other barriers to recycling more plastics are illustrated in figure 12.

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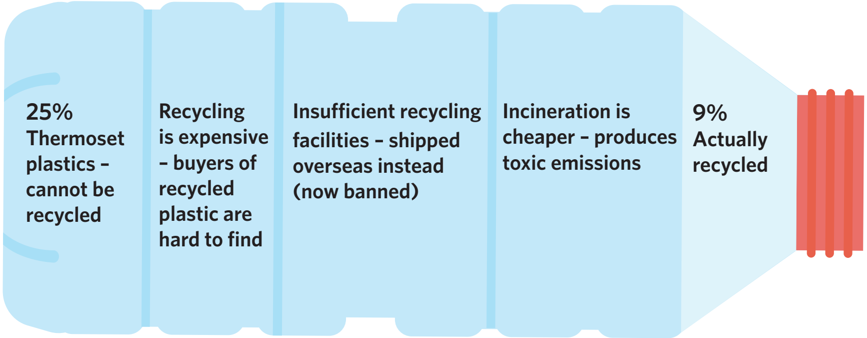

Figure 12 Current barriers to plastic recycling

Since the products of mechanical recycling are invariably of lower grade than the original plastics, it is unlikely that mechanical recycling can ever be truly ‘circular’. Contrastingly, the products of chemical recycling, which can be monomers or raw chemical materials, can be reformed into high-grade plastics, introducing the potential for circularity (figure 13). However, wastage and efficiency must be considered. For example, if 10 grams of plastic can be recycled to produce only 1 gram of new plastic, the circularity of the process is called into question.

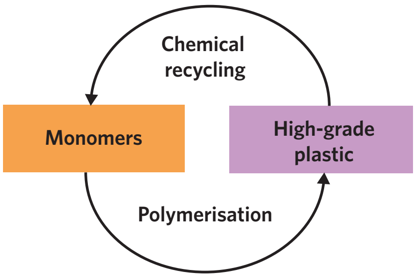

Figure 13 Potential circularity of chemical recycling

Currently, plastics produced from renewable biomass make up less than 1% of the global plastic market. Yet along with organic recycling, this could hold the key to a circular plastic economy, so long as the composting process does not leave plastic residue. For example, nutrient-rich humus produced from the composting of certain bioplastics can be used to help grow feedstock crops, like corn, which can in turn be used to produce more bioplastics, thus completing the circular economy. This concept is explored further in lesson 8C.

Progress questions

Question 8

Most plastic is recycled.

A. True

B. False

Question 9

Features of a circular plastic economy include

A. renewables and recycling.

B. microplastics and incineration.

Theory summary

- Most plastics, including HDPE, LDPE, PVC, PP, and PS, are produced from fossil fuels.
- Bioplastics are made from renewable biomass and include PLA, Bio-PE, and Bio-PP.
- All thermoplastics can be mechanically recycled to produce lower-grade plastic products.
- Some plastics can be chemically recycled back into polymers, monomers, or raw chemical materials.
- Some plastics rapidly compost in soil to produce nutrients, while others decompose slowly and produce microplastic residues.
- Using renewable plastics and changing recycling practices could help circularise the plastic economy.

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8B Questions

Deconstructed

Use the following information to answer questions 10–12.

Bio-PE is a drop-in bioplastic that has been promoted as a potential source for manufacturing HDPE and LDPE, used for products like detergent bottles and plastic bags.

Question 10 (1 MARK)

Biomass is usually defined as ___________ and is thus ___________.

A. remains of marine organisms from millions of years ago, renewable

B. remains of marine organisms from millions of years ago, non-renewable

C. plant material, renewable

D. plant material, non-renewable

Question 11 (1 MARK)

A drop-in bioplastic

A. can be dropped on the ground, and it will decompose naturally.

B. is functionally identical to its fossil fuel-based alternative.

C. can be dropped in the recycling bin.

D. is always compostable.

Question 12 (3 MARKS)

Explain the advantages of using Bio-PE for manufacturing HDPE and LDPE.

Exam-style

Question 13 (2 MARKS)

Some park benches, like the one pictured, can be made from recycled plastic. Explain how household plastics can be mechanically recycled to produce other products like furniture.

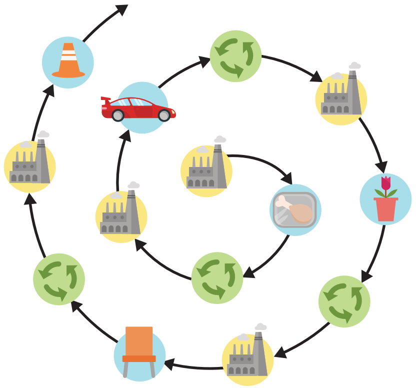
Image: Francesco Scatena/Shutterstock.com

Question 14 (4 MARKS)

It has been suggested that an achievable alternative to a circular plastic economy is a ‘spiral economy’, by which plastics are used for lower and lower-grade uses before being eventually converted into fuel, as shown in the diagram.

a. Explain why a circular plastic economy may be difficult to achieve. (2 MARKS)

b. Explain why this ‘spiral’ plastic economy model is achievable. (2 MARKS)



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kills

Question 15 (5 MARKS)

A bottled water company is planning to begin marketing an ‘eco-friendly’ plastic bottle with this label on it.

a. Explain why the label ‘biodegradable’ could potentially be misleading. (2 MARKS)

b. Devise an experiment to determine whether the bottle is compostable. (3 MARKS)

FROM LESSON 16A

Questions from multiple lessons

Question 16 (7 MARKS)

Otto develops a high-temperature chemical recycling method of PP (polypropene). He takes 35 g of polypropene and uses his method to produce 0.153 mol of propene monomers, along with waste product as the remainder.

a. What mass (in g) of propene was produced? (3 MARKS)

b. What percentage of polypropene’s mass was converted to propene? (2 MARKS)

c. Explain two likely barriers to Otto’s method being adopted on an industrial scale. (2 MARKS)

FROM LESSONS 6C & 7A

Hints

13. Mechanical recycling involves breaking plastics down and building them back up.

14a. In a true circular economy, no new products need to be made from new materials, nor are products ever left unused after being used once.

14b. Consider the products of various recycling methods.

15a. The label aims to give the impression of ‘eco-friendliness’.

15b. Consider the definition of compostable.

16a. Alkenes follow the general formula C_nH_{2n}.

16b. Use the answer to 16a to help here.

16c. Industrial processes aim to maximise efficiency while minimising cost and environmental impact.

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8C Innovations in polymers

Image: Patt Paul Studio/Shutterstock.com

STUDY DESIGN DOT POINT

- innovations in polymer manufacture using condensation reactions, and the breakdown of polymers using hydrolysis reactions, contributing to the transition from a linear economy towards a circular economy

8A 8B 8C

1.2.14.1 Innovation in condensation polymerisation

1.2.14.2 Hydrolysis of polymers

ESSENTIAL PRIOR KNOWLEDGE

3C Metal recycling

8A Condensation polymerisation

See questions 53–54.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

How are eco-friendly plastics made?

As our society becomes increasingly conscious about the environment, significant research is being undertaken into the development of eco-friendly biodegradable polymers. In this lesson, we will investigate how these polymers are made, and their influence on the environment and society.

KEY TERMS AND DEFINITIONS

Condensation polymerisation chemical reaction between monomers resulting in the formation of a polymer and water

Hydrolysis reaction involving the cleavage of bonds using water to break up polymers into monomers

Innovation introduction of new methods, procedures, or products to traditional industries

Innovations in condensation polymerisation 1.2.14.1

Condensation polymerisation is a powerful technique that allows for significant **innovation**.

How is condensation polymerisation being utilised in scientific innovation?

As mentioned in lesson 8A, the formation of polymers via condensation reactions is a useful and relatively straightforward process, used in a variety of settings from industrial manufacturing to the cellular production of proteins. Condensation reactions involve the reaction of an alcohol with another functional group, often a carboxylic acid, to form a covalent bond between two molecules, simultaneously releasing a water molecule. Polylactic acid (PLA) is a polymer formed from the condensation of lactic acid monomers, as shown in figure 1.[1]

KEEN TO INVESTIGATE?

[1] **How do condensation reactions actually occur?**
Search YouTube: Condensation Polymerisation | Organic Chemistry

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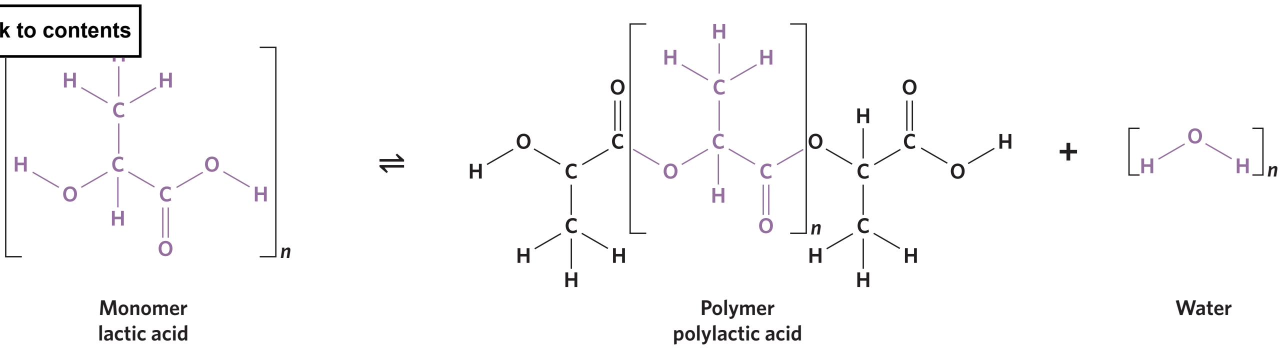

Figure 1 Lactic acid undergoes condensation polymerisation to form polylactic acid

USEFUL TIP

A detailed understanding of the mechanism of condensation polymerisation is not required in year 11 chemistry. However, if you'd like to get a headstart for year 12, see the 'Keen to investigate?' box.

PLA is characterised as a 'bioplastic', as it is synthesised from naturally-occurring lactic acid, which is usually obtained via the fermentation of sugar cane or corn starch. This allows for the 'circularity' of the raw ingredients required for the synthesis of PLA.

Due to its ability to be degraded into lactic acid (a substance present in large quantities in the body), PLA is useful in the construction of medical implants, such as screws or plates to fixate broken bones. The polymer can remain intact in the body for up to 2 years during a healing process, before being broken down and naturally removed.

Another innovative polymer formed by condensation is polybutylene succinate (PBS), which is formed from the reaction of a carboxylic acid monomer and an alcohol monomer, as shown in figure 2.

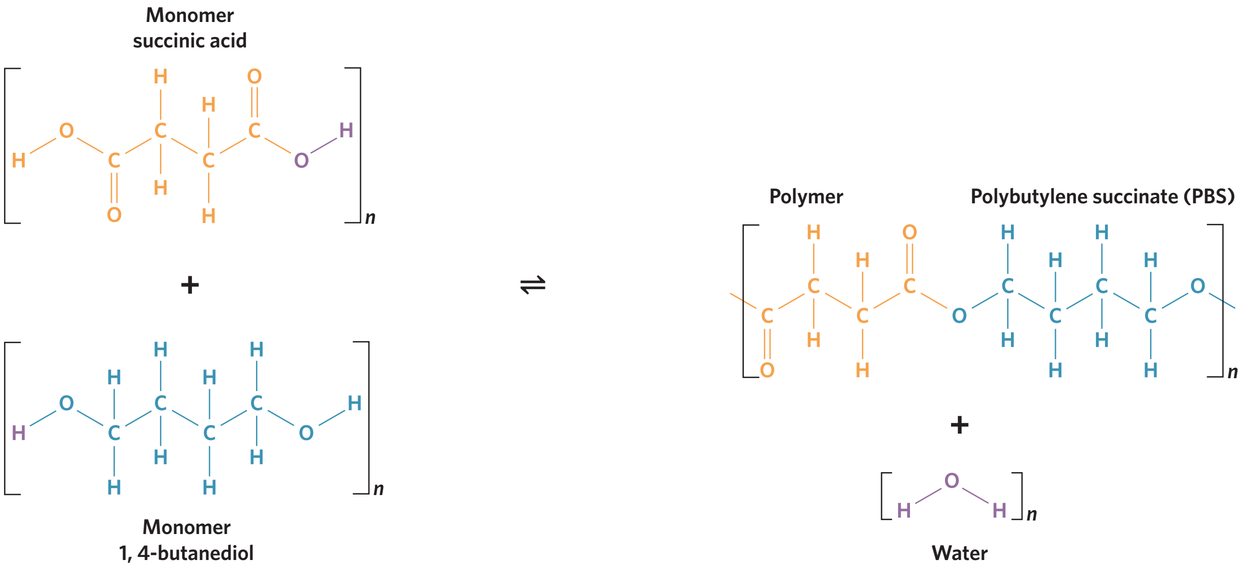

Figure 2 Succinic acid and 1,4-butanediol undergoes condensation polymerisation to form polybutylene succinate (PBS)

PBS is highly biodegradable and decomposes into biomass and water. It is also very flexible and hence, holds great potential in a variety of packaging applications. It is non-toxic and safe for use in packaging food and cosmetics, unlike many other polymers. Furthermore, the two reactants, succinic acid and 1,4-butanediol, are naturally occurring and readily obtainable via the fermentation of corn or wheat.

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Progress questions

Question 1

Condensation reactions usually produce water.

A. True

B. False

Question 2

PLA is a polymer made up of ___________ monomer subunits.

A. succinic acid and 1,4-butanediol

B. lactic acid

Question 3

The two molecules above are able to undergo condensation polymerisation.

A. True

B. False

Hydrolysis of polymers 1.2.14.2

Condensation polymers have specific structural properties that make them easy to break down, allowing for the implementation of a circular economy.

How does the structure of condensation polymers allow them to undergo hydrolysis?

In figures 1 and 2, you may have noticed that water is a product of both reactions. This is characteristic of condensation reactions, and is a result of the rearrangement of oxygen and hydrogen atoms from functional groups into water molecules. This leaves the functional groups in a reactive state, allowing polymerisation to occur.

It is also this property that makes the polymers 'biodegradable', as mentioned previously. When water is reintroduced to a polymer system, this reaction occurs gradually in reverse, through the process of **hydrolysis**, shown in figure 3.

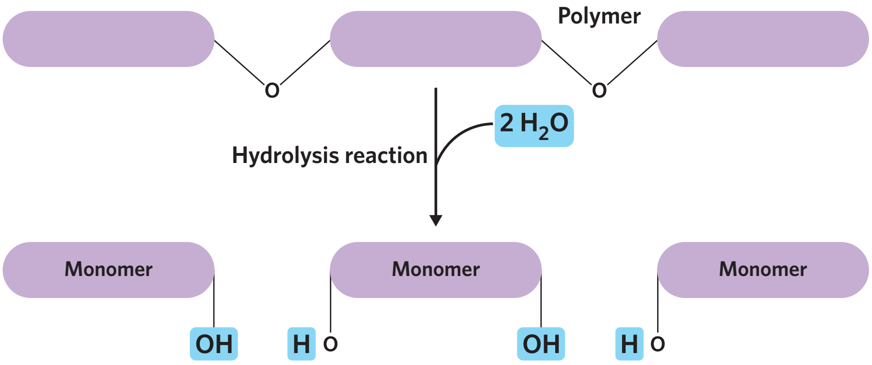
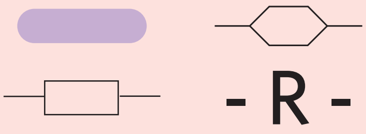


Figure 3 General representation of the hydrolysis of a polymer

USEFUL TIP

There are many different representations for a generalised individual monomer unit within a polymer, including but not limited to:



USEFUL TIP

The term 'hydrolysis' comes from the Ancient Greek 'hydro', meaning 'water', and 'lysis', meaning 'to unbind'.

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How does the hydrolysis of condensation polymers support the transition from a linear economy to a circular economy?

The ability of condensation polymers to undergo hydrolysis and therefore biodegradation, is an important property in the transition from a linear economy to a circular economy. Traditionally, polymers have been synthesised from non-renewable sources and they have degraded too slowly to be meaningfully reused. However, the rapid degrading properties of condensation polymers allow them to be used in the renewal of useful resources, as shown in figure 4.

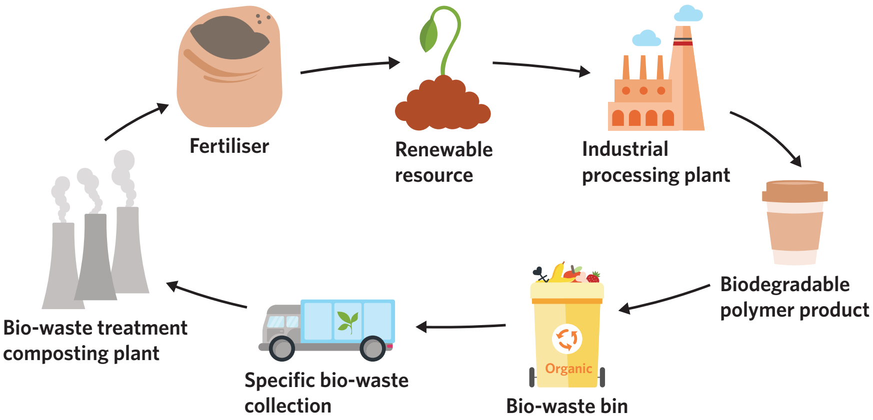
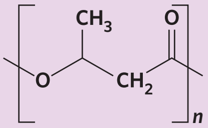


Figure 4 Polymer circular economy

WORKED EXAMPLE 1

Polyhydroxybutyrate (PHB) is a polymer formed by condensation, with the following structural formula. What chemical property allows this polymer to be biodegradable, and how does this property relate to the implementation of a circular economy?



What information is presented in the question?

The structure of the molecule, and its nature as a condensation polymer.

What is the question asking us to do?

Identify the chemical property that makes the molecule biodegradable, and relate this to the implementation of a circular economy.

What strategies do we need in order to answer the question?

1. Review properties that improve the biodegradability of molecules.
2. Consider the relationship between biodegradability and the various aspects of a circular economy.

Answer

As a condensation polymer, the molecule will be able to undergo hydrolysis (the reverse reaction). This results in the breakdown of the polymer back into its monomers, and an increased rate of biodegradation. Biodegradation is related to the implementation of a circular economy, as biodegraded condensation monomers can be used in the replenishment of renewable polymer-forming reactants.

Progress questions

Question 4

Hydrolysis always requires water to occur.

A. True

B. False

Continues →

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Question 5

With regards to what occurs to water molecules, hydrolysis reactions are opposite to condensation reactions.

A. True

B. False

Question 6

The ability of condensation polymers to undergo hydrolysis allows for the implementation of a ___________ economy.

A. circular

B. linear

Question 7

In order to conserve Earth's resources, a transition from a circular to a linear economy must occur.

A. True

B. False

Theory summary

- Condensation polymerisation is a useful technique that allows for significant innovation.
- Examples of innovative polymers formed by condensation include:
  - polylactic acid (PLA), for use in medical implants
  - polybutylene succinate (PBS), for use in food and cosmetic packaging
- Polymers formed by condensation can also be broken down by the reverse reaction, hydrolysis.
- Being able to undergo hydrolysis allows for their reuse and hence for the implementation of a circular economy.

The content in this lesson is considered fundamental prior knowledge to reactions of organic compounds (Unit 4 AOS 1).

8C Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Question 8 (1 MARK) 🌶

The synthesis of non-biodegradable polymers is an example of

A. a circular economy.

B. either a circular or linear economy.

C. a linear economy.

D. a circular economy, as long as the polymers are sourced from renewable resources.

Question 9 (1 MARK) 🌶

The 'reverse reaction' of condensation polymerisation is ___________, involving the breakdown of a ___________ into ___________.

A. hydrolysis, monomer, polymers

B. fermentation, monomer, polymers

C. hydrolysis, polymer, monomers

D. fermentation, polymer, monomers

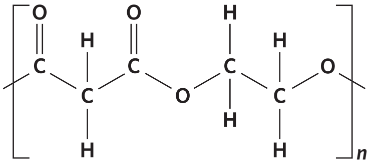
Back to contents (3 MARKS)

Explain how the hydrolysis of renewably-sourced polymers contributes to a circular economy.

Exam-style

Question 11 (5 MARKS)

Consider the following polymer X.



a. What type of polymerisation was most likely used to synthesise this polymer? (1 MARK)

b. What type of reaction can be used to break down this polymer? (1 MARK)

c. Draw the two products formed following completion of the reaction identified in part b, and the other reactant required for the reaction to occur. (3 MARK)

Key science skills

Question 12 (5 MARKS)

Imogen, an environmentalist, is conducting an experiment to determine the biodegradability of different bio-plastic coffee cups. She weighs three different types of cups, leaves them in various locations around her garden, and then weighs them again once every month for a year.

a. Identify the independent variable. (1 MARK)

b. Identify the dependent variable. (1 MARK)

c. Identify one potential source of error. (1 MARK)

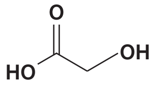
d. Upon completion of her experiment, Imogen found that her weighing scales were poorly calibrated, and had been consistently measuring the coffee cups at 6 g less than their true value. What kind of error is this, and what can Imogen do to rectify this error? (2 MARKS)

FROM LESSONS 16A & 16D

Questions from multiple lessons

Question 13 (7 MARKS)

Glycolic acid, shown below, is a naturally-occurring organic molecule, commonly used in skin-care products. Like lactic acid, it is able to undergo condensation polymerisation with itself, to form the polymer polyglycolic acid.



a. Identify all functional groups present in glycolic acid. (2 MARKS)

b. Calculate the molar mass of glycolic acid. (2 MARKS)

c. Would you expect polyglycolic acid to be biodegradable? Explain why or why not. (3 MARKS)

FROM LESSONS 6C & 7C

11a. Condensation polymers are often formed from the reaction of a carboxylic acid and an alcohol.

11b. Condensation reactions are associated with a specific 'reverse' reaction.

11c. Hydrolysis reactions involve the splitting of a polymer chain and the release of another small molecule.

12a. Independent variables are deliberately controlled.

12b. Dependent variables are carefully measured.

12c. Errors can arise from differences in environment.

12d. Calibration errors affect all values by the same amount.

13a. Alcohols and carboxylic acids are defined by specific functional groups.

13b. Molar mass can be calculated through summing all the relative atomic masses of the atoms present.

13c. Polymers formed by condensation are generally biodegradable.

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Chapter 8 review

Multiple choice (10 MARKS)

Question 1 (1 MARK) 🌶

Teflon is a non-polar polymer used

A. in pipes due to its hardness.
B. on pans due to its non-stick property.
C. on skateboards due to its adhesivity.
D. in electronics due to its conductivity.

Question 2 (1 MARK) 🌶

The main difference in producing low- and high-density polyethene is that

A. LDPE is produced at higher pressures.
B. LDPE is produced at lower temperatures.
C. HDPE is a good insulator of electricity while LDPE is not.
D. HDPE chains are longer than LDPE chains.

Question 3 (1 MARK) 🌶

Which of the following lists only fossil fuel-based plastics?

A. Bio-PP, PLA, Bio-PE
B. Polystyrene, PVC, PLA
C. PVA, PHB, LDPE
D. PBAT, PVC, HDPE

Question 4 (1 MARK) 🌶

Polar monomers

A. repel non-polar monomers.
B. increase the hardness of polymers.
C. have no effect on a polymer's properties.
D. create branches in polymer chains.

Question 5 (1 MARK) 🌶🌶

The intermolecular forces and/or bonds that may be present in thermoplastic polymers are

I. cross-links.
II. hydrogen bonds.
III. permanent dipole-dipole forces.
IV. covalent bonds.
V. dispersion forces.

A. I, II, III
B. I, IV
C. II, III, V
D. I, IV, V

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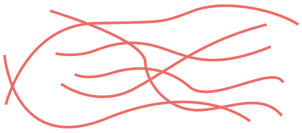
Question 6 (1 MARK)

The following diagram depicts



A. a thermoplastic.

B. a thermoset.

C. an elastomer.

D. LDPE.

Question 7 (1 MARK)

Regarding plastics, what is the definition of biodegradable?

A. Produced from renewable sources

B. Able to break down in the ground

C. Able to break down in the ground without producing toxic residues

D. Able to break down by 90% after 180 days in a composting environment

---

Question 8 (1 MARK)

In practice, the current plastic economy is primarily

A. linear, since materials used to make plastic cannot be reused.

B. linear, since most plastic is not actually recycled.

C. circular, since almost all plastics can be recycled.

D. circular, since most plastics are produced from renewable sources.

Question 9 (1 MARK)

Condensation reactions

A. always have water as a reactant.

B. always have an alcohol as a reactant.

C. always have water as a product.

D. always have an alcohol as a product.

Question 10 (1 MARK)

Polymers formed by ___________ reactions are generally ___________ due to their ability to undergo ___________.

A. addition, biodegradable, combustion

B. condensation, biodegradable, hydrolysis

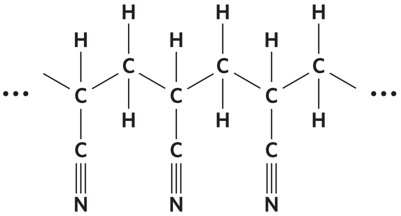
C. addition, non-biodegradable, condensation

D. condensation, non-biodegradable, hydrolysis

Short answer (30 MARKS)

Question 11 (5 MARKS)

Polyacrylonitrile is a polymer used to make carbon fibre. A segment of polyacrylonitrile is shown here.



a. Draw the structure of the alkene monomer used to make polyacrylonitrile. (1 MARK)

b. How many repeating monomer units are shown in this section of polymer? (1 MARK)

c. What are the strongest intermolecular forces operating between polyacrylonitrile chains? (1 MARK)

d. How might these forces influence the properties of the polymer? (2 MARK)

Question 12 (2 MARKS)

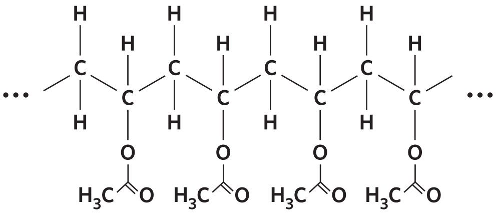
Kettle handles are made from a polymer with a specific degree of cross-linking. Explain the level of cross-linking you would expect this polymer to have, and how this level makes it appropriate for its desired use.

Question 13 (3 MARKS)

Why do thermosetting polymers decompose rather than melt when exposed to heat?

Question 14 (4 MARKS)

A segment of the polymer polyvinyl acetate is shown below.



a. Draw the structure of a monomer from this polymer. (1 MARK)

b. What intermolecular forces hold these polymer chains together? (2 MARKS)

c. Draw the structure in the condensed (bracket) form. (1 MARK)

Question 15 (6 MARKS)

Polyvinyl chloride (PVC) and high-density polyethene (HDPE) are two strong and durable polymers. They share some similarities but also have some differences.

a. The chains of PVC and HDPE have very little branching. Identify and describe two other factors that might affect the strength and hardness of these polymers. (4 MARKS)

b. In contrast to high-density polyethene, low-density polyethene is relatively soft. Explain this difference in hardness. (2 MARKS)

Question 16 (5 MARKS)

Polylactic acid (PLA) is a bioplastic.

a. Explain why PLA is considered a bioplastic. (1 MARK)

b. Outline three major ways in which PLA can be recycled. (3 MARKS)

c. Briefly explain how PLA can contribute to the circular economy. (1 MARK)

Question 17 (5 MARKS)

Polymer Z is formed by condensation.

a. Given the polymer is formed by condensation, name the reaction that would break this polymer down. (1 MARK)

b. With reference to your answer to part a, explain how condensation polymers have enabled innovation to occur in science. (4 MARKS)

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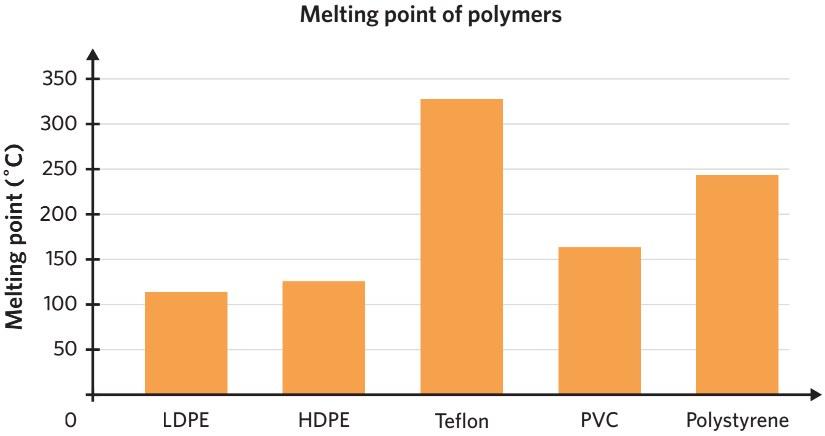
Key science skills (10 MARKS)

Question 18 (10 MARKS) 🌶🌶🌶

A polymer chemist, Savannah, investigated the melting point of different polymers. After conducting her experiment, she produced the following graph.



a. What type of graph is this, and what type(s) of data does it represent? (3 MARKS)

b. On which axis is the independent variable displayed? (1 MARK)

c. Evaluate the effectiveness of this graph as a representation of the data. (2 MARKS)

d. Could a line graph be used to represent this data? Explain. (2 MARKS)

e. Based on this graph, which of these polymers would be best suited for use in cookware? Justify your answer. (2 MARKS)

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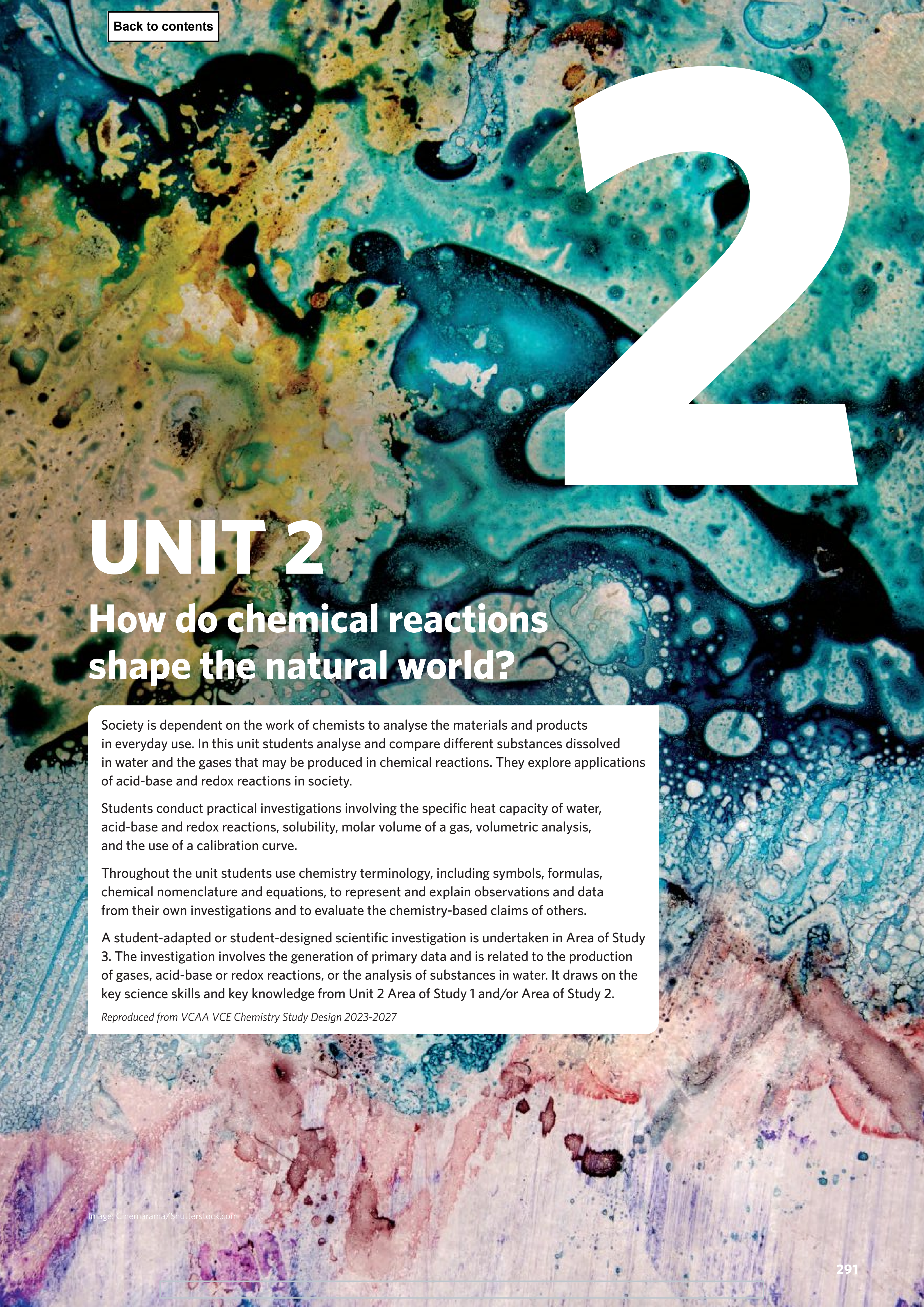

2

UNIT 2

How do chemical reactions shape the natural world?

Society is dependent on the work of chemists to analyse the materials and products in everyday use. In this unit students analyse and compare different substances dissolved in water and the gases that may be produced in chemical reactions. They explore applications of acid-base and redox reactions in society.

Students conduct practical investigations involving the specific heat capacity of water, acid-base and redox reactions, solubility, molar volume of a gas, volumetric analysis, and the use of a calibration curve.

Throughout the unit students use chemistry terminology, including symbols, formulas, chemical nomenclature and equations, to represent and explain observations and data from their own investigations and to evaluate the chemistry-based claims of others.

A student-adapted or student-designed scientific investigation is undertaken in Area of Study 3. The investigation involves the generation of primary data and is related to the production of gases, acid-base or redox reactions, or the analysis of substances in water. It draws on the key science skills and key knowledge from Unit 2 Area of Study 1 and/or Area of Study 2.

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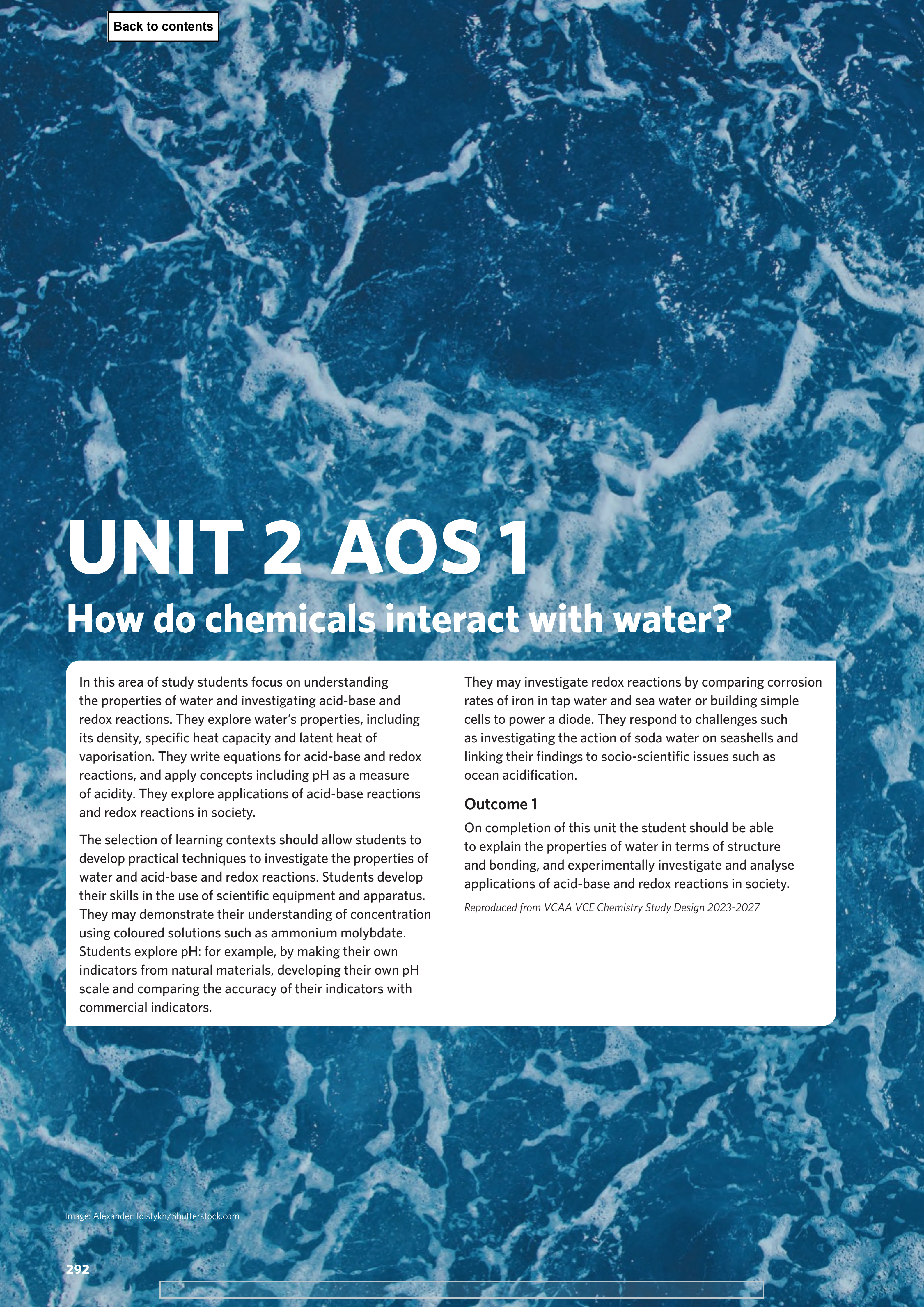

UNIT 2 AOS 1

How do chemicals interact with water?

In this area of study students focus on understanding the properties of water and investigating acid-base and redox reactions. They explore water's properties, including its density, specific heat capacity and latent heat of vaporisation. They write equations for acid-base and redox reactions, and apply concepts including pH as a measure of acidity. They explore applications of acid-base reactions and redox reactions in society.

The selection of learning contexts should allow students to develop practical techniques to investigate the properties of water and acid-base and redox reactions. Students develop their skills in the use of scientific equipment and apparatus. They may demonstrate their understanding of concentration using coloured solutions such as ammonium molybdate. Students explore pH: for example, by making their own indicators from natural materials, developing their own pH scale and comparing the accuracy of their indicators with commercial indicators. They may investigate redox reactions by comparing corrosion rates of iron in tap water and sea water or building simple cells to power a diode. They respond to challenges such as investigating the action of soda water on seashells and linking their findings to socio-scientific issues such as ocean acidification.

Outcome 1

On completion of this unit the student should be able to explain the properties of water in terms of structure and bonding, and experimentally investigate and analyse applications of acid-base and redox reactions in society.

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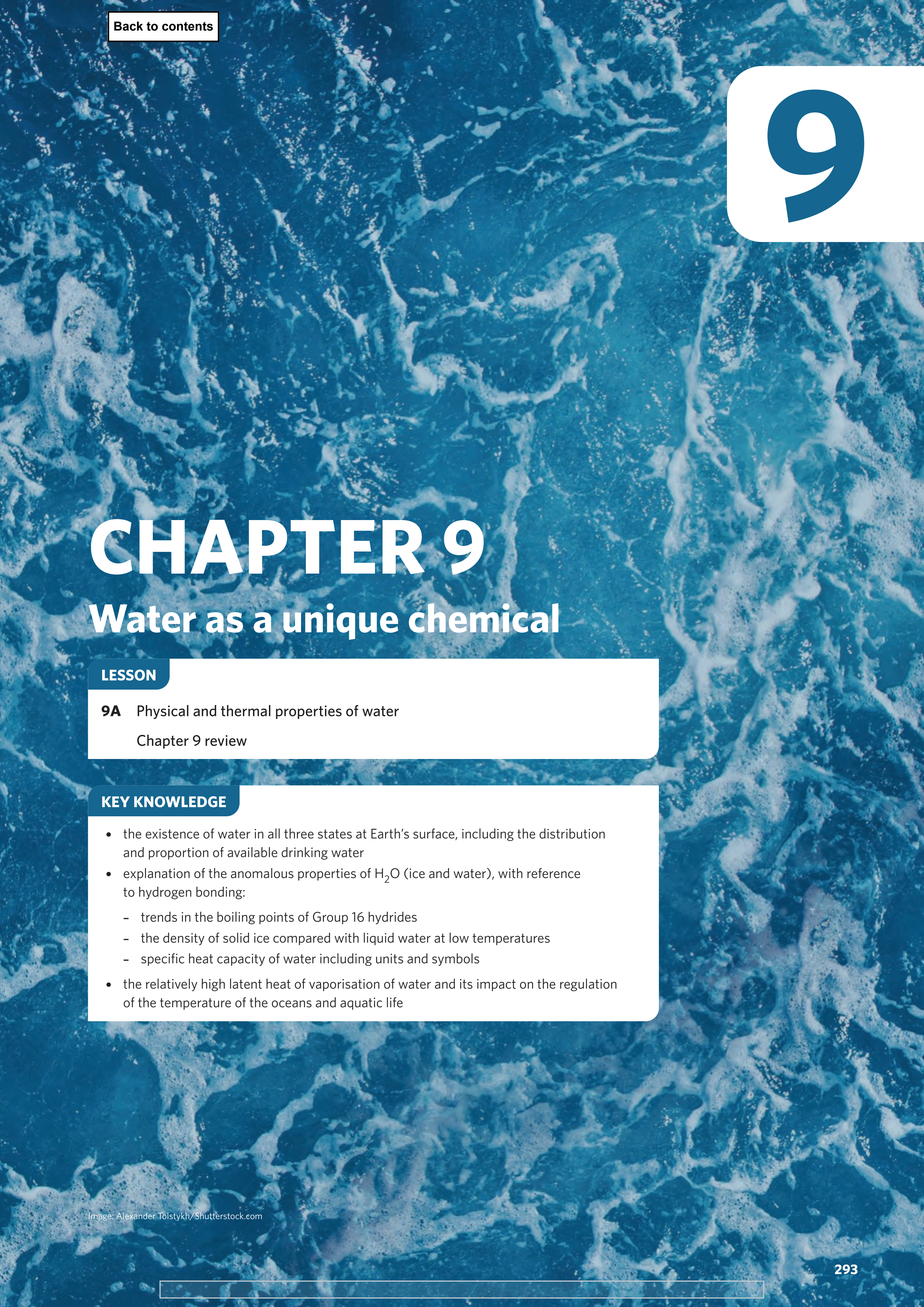

CHAPTER 9

Water as a unique chemical

LESSON

KEY KNOWLEDGE

- the existence of water in all three states at Earth's surface, including the distribution and proportion of available drinking water
- explanation of the anomalous properties of H_2O (ice and water), with reference to hydrogen bonding:
  - trends in the boiling points of Group 16 hydrides
  - the density of solid ice compared with liquid water at low temperatures
  - specific heat capacity of water including units and symbols
- the relatively high latent heat of vaporisation of water and its impact on the regulation of the temperature of the oceans and aquatic life

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9A Physical and thermal properties of water

STUDY DESIGN DOT POINTS

- the existence of water in all three states at Earth's surface, including the distribution and proportion of available drinking water
- explanation of the anomalous properties of H_2O (ice and water), with reference to hydrogen bonding:
  - trends in the boiling points of Group 16 hydrides
  - the density of solid ice compared with liquid water at low temperatures
  - specific heat capacity of water including units and symbols
- the relatively high latent heat of vaporisation of water and its impact on the regulation of the temperature of the oceans and aquatic life

9A

2.1.1.1	Precious water
2.1.2.1	Properties of group 16 hydrides
2.1.2.2	Specific heat capacity
2.1.3.1	Latent heat
2.1.3.2	Significance of the properties of water

ESSENTIAL PRIOR KNOWLEDGE

1B	Shared properties of groups in the periodic table
2A	Covalent compounds & polarity
2B	Hydrogen bonding
2B	Strength of dispersion forces

See questions 55–58.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: Michal Balada/Shutterstock.com

Why is water the only substance on Earth that naturally occurs as a solid, liquid and gas?

Water takes many different forms on Earth. In this lesson we will learn about the key properties of water that enable it to have this unique distinction of occurring as a solid, liquid and gas on our planet. We will also learn how water's properties make it such a vital chemical to life.

KEY TERMS AND DEFINITIONS

Boiling point temperature at which a liquid transitions into a gas

Condensation phase change from gas to liquid

Evaporation alternative term for vaporisation

Group 16 hydride group 16 element covalently bonded to hydrogen

Heat capacity relationship between the heat absorbed by a substance and its temperature change

Latent heat of vaporisation amount of energy required to convert 1 mol of a substance from its liquid state to its gaseous state at the boiling point of the substance

Phase change transition of a substance from one state to another

Potable alternative term for drinkable

Solar energy energy produced by the sun

Specific heat capacity energy (J) required to raise the temperature of 1 gram of a given substance by 1 °C

Vaporisation phase change from liquid to gas

Precious water 2.1.1.1

Water is abundant on Earth, yet many factors influence the availability of clean water to humans.

In what states is water found on Earth?

Water is unique in that it is the only substance that naturally occurs as a solid, liquid, and gas at the Earth's surface. Sources of these three states of water are illustrated in figure 1.

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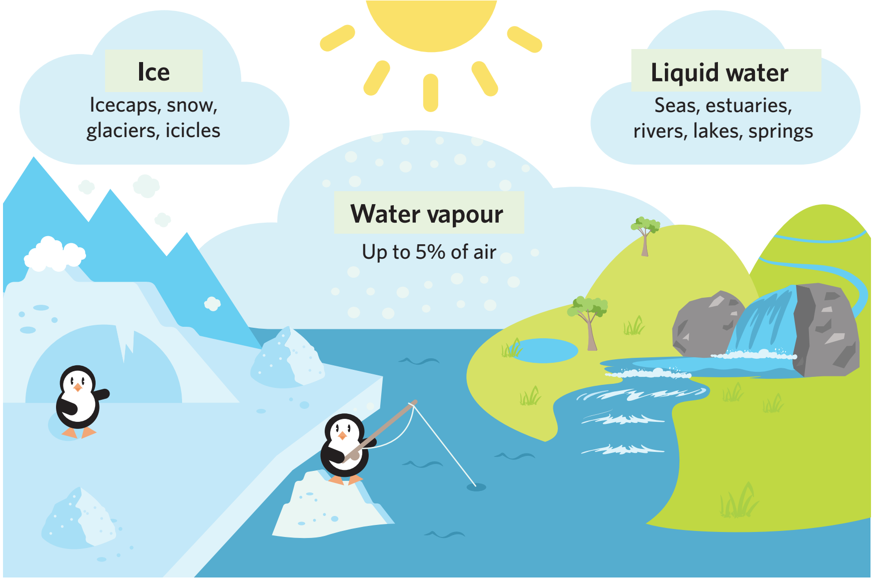

Figure 1 Sources of water as a solid, liquid, and gas

USEFUL TIP

The terms ‘phase’ and ‘state of matter’ are synonymous in this context.

MISCONCEPTION

‘Water vapour looks white and fluffy like clouds.‘

Water vapour is actually transparent, colourless, and invisible to the naked eye. The steam we see rising from boiling water, for example, is in fact visible only due to liquid water droplets suspended in the water vapour.

However, these states are not fixed; in fact, water is naturally in continuous movement between them through processes called **phase changes**: melting, freezing, **condensation**, and **evaporation** as shown in figure 2.

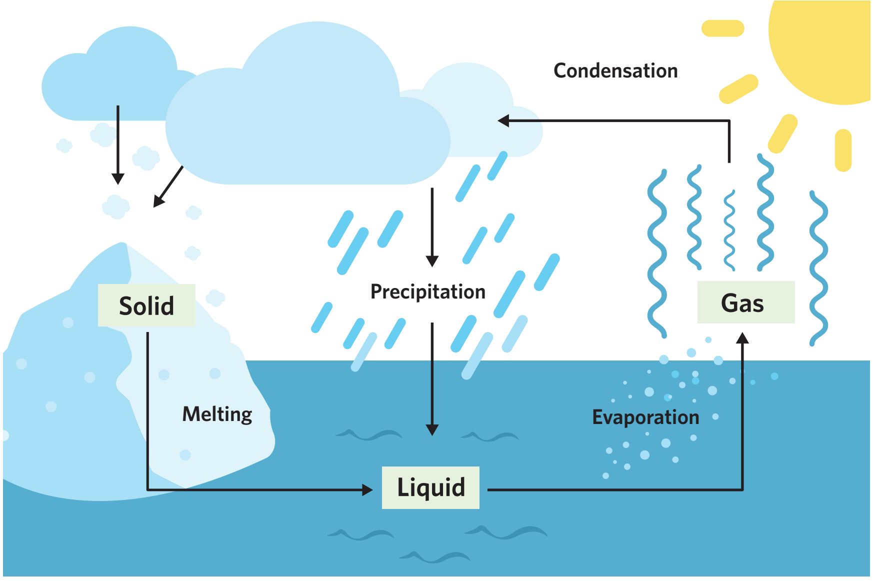

Figure 2 Natural phase changes of water

The state water takes depends on the conditions it is exposed to at a given moment; for example, solid water (or ice) is much more likely to occur where the temperature is below freezing (0 °C), so it is most common in colder regions.

How is drinking water distributed on Earth?

Water is one of the most abundant substances on Earth. It makes up more than 70% of the Earth’s surface, and it is estimated that there is approximately 1.386 billion km^3 of water on Earth which is enough to fill the Melbourne Cricket Ground 81 billion times! However, only a tiny percentage of this water is **potable** by humans, and so clean drinking water is a scarce and precious resource. A major factor behind this is salt content, and water can be separated into a number of categories based on the amount of salt it contains, as shown in figure 3.

The human body is ineffective at filtering out salt, and so in order to be potable, the water we consume needs to have a very low salt content. Yet the vast majority of water found on Earth is, unfortunately, undrinkable saltwater, as shown in figure 4.

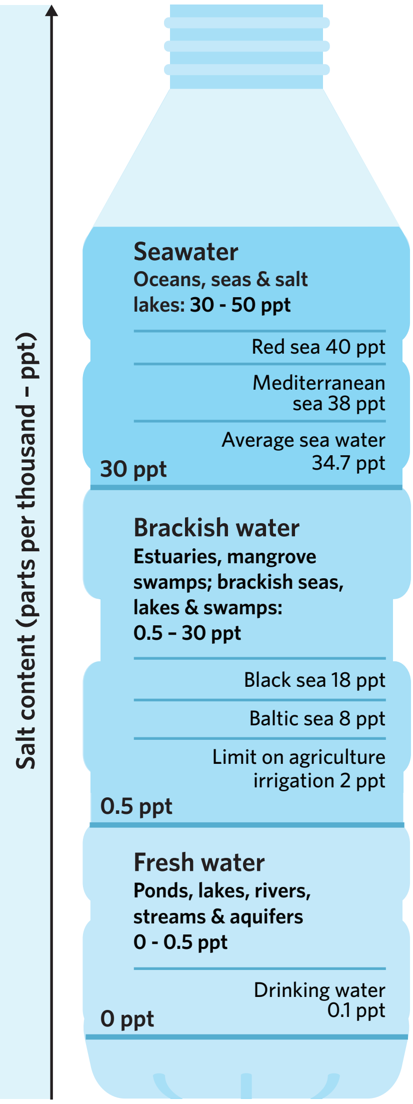

Figure 3 Types of water

MISCONCEPTION

'Freshwater is always potable.'

Although freshwater is much more likely to be potable than seawater or brackish water, only a small percentage of freshwater on Earth is sufficiently clean to be safe for human consumption.

KEEN TO INVESTIGATE?

[1] **Why is drinking water quality in Melbourne so high?**
Search: Why Melbourne's water tastes great

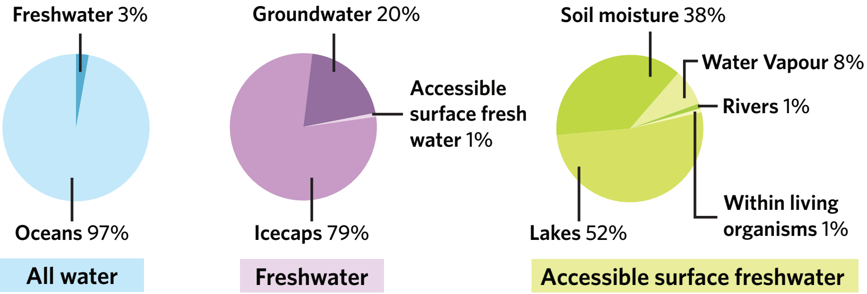

Figure 4 Distribution of water on Earth

As we can see, approximately 3% of water on Earth is freshwater, and therefore more likely to be drinkable. However, most of this freshwater is trapped in ice caps, glaciers or in groundwater, making it hard to access. In fact, just 1% of freshwater is accessible at Earth's surface; so overall, only 0.03% of water on Earth is both fresh and readily accessible.

The most readily accessible source of water on earth is surface freshwater from sources like lakes, rivers and springs, but steps need to be taken to ensure this water is safe to drink, and then distributed throughout a population.[1] Therefore, despite water's abundance, supplying clean drinking water to populated areas relies on significant infrastructure which is often less established in less economically developed regions, contributing to scarcity. Accordingly, drinking water is an unequally distributed resource.

Progress questions

Question 1

Water exists in a gaseous state at the Earth's surface.

A. True

B. False

Question 2

Seawater is unable to be consumed by humans primarily because

A. it is all contaminated with industrial waste.

B. it contains too much salt for human consumption.

Question 3

Because it comes straight from natural sources, all freshwater is safe to drink.

A. True

B. False

Question 4

The majority of the available water on Earth is

A. fresh water.

B. salt water.

Properties of group 16 hydrides 2.1.2.1

When compared with other **group 16 hydrides**, water has distinctly unique properties.

How does hydrogen bonding explain water's high boiling point?

Water exists as a liquid at the Earth's surface because its **boiling point** (100 °C) is much higher than the range of temperatures experienced on Earth, which rarely exceed 50 °C. This high boiling point is due to water's molecular structure, which was covered in lesson 2B.

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Since water contains highly polar O-H bonds and lone pairs on its oxygen atom, water molecules are able to form hydrogen bonds with other water molecules. The two lone pairs on an oxygen atom are dense in negative charge, and so each of these lone pairs is electrostatically attracted to a hydrogen atom (which carries a partial positive charge) in a neighbouring water molecule. Therefore, each water molecule can form up to four hydrogen bonds with four other water molecules, as shown in figure 5.

Additionally, water belongs to the class of group 16 hydrides, which are compounds consisting of a group 16 element covalently bonded to two hydrogen atoms. Group 16 elements are shown in figure 6; possible group 16 hydrides include water (H_2O), hydrogen sulfide (H_2S), hydrogen selenide (H_2Se), hydrogen telluride (H_2Te) and hydrogen polonide (H_2Po).

Since group 16 hydrides are all examples of molecular compounds, their respective boiling points are representative of the strength of their intermolecular forces. That is, molecular substances with stronger intermolecular forces will have higher boiling points, as explained in figure 7.

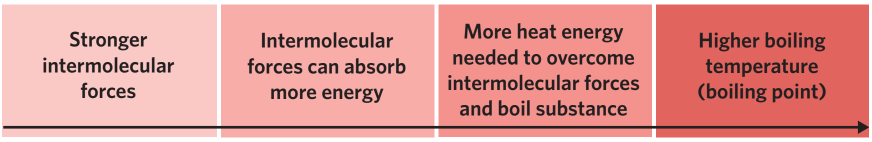

Figure 7 Flowchart linking intermolecular force strength to boiling point

As we learnt in lesson 2B, dispersion forces present between all molecules generally increase in strength as the size of these molecules increases. Since elements increase in size as we move down a group, it would be expected that the hydrides' boiling point would increase as we move down group 16. While this is partially true as shown in figure 8, water is a unique molecule in that it has the highest boiling point of all group 16 hydrides, contrary to the expected trend.

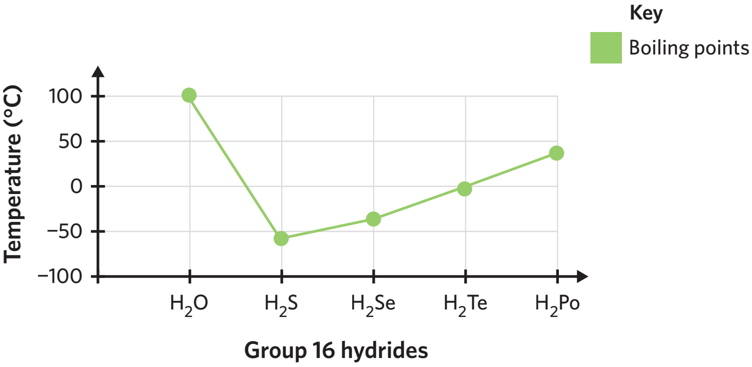

Figure 8 Graph of the boiling points of group 16 hydrides

This unexpected result is due to water's unique ability to form hydrogen bonds with four other water molecules. These hydrogen bonds are standardly much stronger than dispersion forces and thus require a significant amount of energy to disrupt. Consequently, a much higher temperature is needed to provide sufficient energy to overcome the forces between water molecules than for other group 16 hydrides, resulting in water's atypically high boiling point.

Progress questions

Question 5

Which of the following is a group 16 hydride?

A. H_2S

B. KH

Continues →

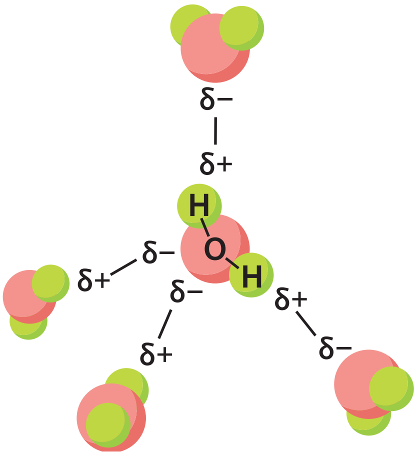

Figure 5 Water molecule with hydrogen bonds to four other water molecules

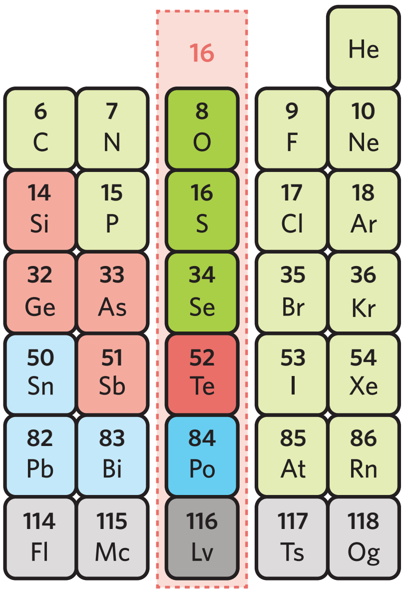

Figure 6 Elements in group 16 of the periodic table

MISCONCEPTION

'Hydrogen bonds are as strong as ionic or covalent bonds.'

Although hydrogen bonds are the strongest type of *intermolecular* force (generally stronger than dipole-dipole and dispersion forces), they are still around 100–200 times weaker than covalent and ionic bonds. For this reason, when substances boil, make sure to say that 'intermolecular forces are overcome', rather than 'bonds are broken'.

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Question 6

Hydrogen telluride is a group 16 hydride. Its boiling point is __________ than the boiling point of water.

A. greater

B. lower

Question 7

Water's high boiling point is primarily due to the presence of __________ between water molecules.

A. dispersion forces

B. hydrogen bonds

Specific heat capacity 2.1.2.2

Relative to other substances, water requires a high quantity of energy to increase in temperature. Water also has the rare property of being denser as a liquid than as a solid.

Why does water have a high specific heat capacity?

Heat capacity is a general term used to describe the amount of thermal energy a substance needs to absorb in order for its temperature to rise by a given amount. For the example in figure 9, if a glass of cooking oil and a glass of water were left in the sun for a given amount of time, the oil would feel hotter to touch than the glass of water, even though both received the same amount of thermal energy. This is due to the higher heat capacity of water compared with the lower heat capacity of oil.

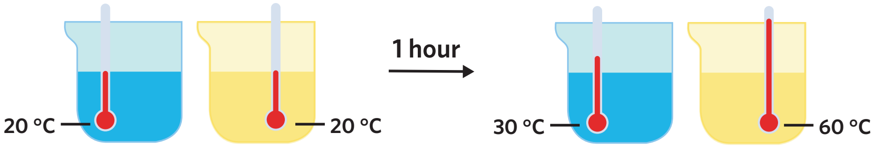

Figure 9 Effect of different heat capacities of water and cooking oil

While heat capacity can be a useful term to describe the differences in thermal properties between compounds, it is very general and has no set units. Therefore, scientists more commonly use the term **specific heat capacity** to describe the amount of energy required (measured in joules, J) to raise the temperature of 1 gram of a substance by 1 °C. Therefore, specific heat capacity, which is given the symbol *c*, is measured in units 'joules per gram per degree Celsius', written as J g^{-1} $°C^{-1}$.

USEFUL TIP

Specific heat capacity, *c*, is often given the units J g^{-1} $°C^{-1}$, but since a temperature change of 1 °C is the same as a temperature change of 1 K, the units can also be expressed as J g^{-1} K^{-1}.

Each substance has a different heat capacity, and table 1 displays the specific heat capacities of some common substances.

USEFUL TIP

The value of the specific heat capacity of water can be found in the VCE Data Book. It is given both as 4.18 J g^{-1} $°C^{-1}$ and 4.18 kJ kg^{-1} $°C^{-1}$; either can be used, as long as the units are consistent (i.e. J goes with g, kJ goes with kg).

Table 1 Specific heat capacity values of different liquids

Liquid	Specific heat capacity (J g^{-1} $°C^{-1}$)
Water	4.18
Milk	3.75
Ethanol	2.46
Coconut oil	2.10
Benzene	1.74

As can be seen, water has a relatively high specific heat capacity of 4.18 J g^{-1} $°C^{-1}$, meaning it takes 4.18 joules of energy to raise the temperature of 1 gram of water by 1 °C. For molecular substances, specific heat capacity depends on the strength of intermolecular forces within the substance. Accordingly, water's relatively high heat capacity is due to the strong hydrogen bonds between water molecules, which because of their strength are able to absorb a large quantity of thermal energy before increasing in temperature.

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How do we perform calculations using $q = mc\Delta T$?

The specific heat capacity of water is an important constant in chemistry because it enables us to calculate the amount of thermal energy required to raise the temperature of a given mass of water by a given number of degrees. This is achieved using the following equation:

$q = m \times c \times \Delta T$

q is the amount of energy that is transferred to the water (in joules), m is the mass of water (in grams), c is the specific heat capacity of water ($4.18\ \text{J g}^{-1}\ {}^\circ\text{C}^{-1}$) and ΔT is the change in temperature of the water (in °C). When calculating $q = mc\Delta T$ for a liquid other than water, substitute c with the specific heat capacity of the given liquid. As we are only measuring the change in temperature of the water, we can use either degrees Celsius or Kelvins as our unit of temperature. The change in temperature is calculated using the following equation:

$\Delta T = T_{final} - T_{initial}$

For example, if our temperature is 20 °C before heating, and 30 °C after heating,
$\Delta T = 30\ {}^\circ\text{C} - 20\ {}^\circ\text{C}$

$\Delta T = 10\ {}^\circ\text{C}$

USEFUL TIP

The equation $q = m \times c \times \Delta T$ is found in the VCE Data Book and can be remembered as $Q = MCAT$.

Since water is a liquid at room temperature, it is often measured by volume in mL, or L. This means that the volume of water needs to be converted to a mass in grams to perform calculations, using water's density value.

In the VCE Data Book, the density of water is given as $0.997\ \text{g mL}^{-1}$ at 25 °C which means that at 25 °C, 1 mL of water weighs 0.997 g. Therefore, to convert from volume (in mL) to mass (in grams), the volume of water is multiplied by 0.997. For example, 2.50 L of water, which is 2500 mL, weighs

$2500\ \text{mL} \times 0.997\ \text{g mL}^{-1} = 2492.5\ \text{g}$ which is 2.49 kg.

STRATEGY

Checking whether the units balance can be a helpful way to see if a formula is correct. For example if we have q Joules, m grams, ΔT degrees and c Joules per gram per degree Celsius, the units of the formula

$q = m \times c \times \Delta T$ are

$\text{J} = \text{g} \times \text{J g}^{-1}\ {}^\circ\text{C}^{-1} \times {}^\circ\text{C}$

J = J ✓ after cancelling, the units on each side of the equation are the same. So the formula is correct.

MISCONCEPTION

'1 mL of liquid water weighs exactly 1.000 g, so the density of liquid water is $1.000\ \text{g mL}^{-1}$.'

At 25 °C, 1 mL of liquid water actually weighs 0.997 g, so its density is technically $0.997\ \text{g mL}^{-1}$. For the most part, the approximation 1 mL = 1 g is sufficient since we usually must round answers to 2-4 significant figures. However, larger volumes of water will amplify the amount of error, and so it is best to use 1 mL = 0.997 g when working with large volumes.

WORKED EXAMPLE 1

Elizabeth is investigating how much heat energy is required to make a cup of tea. She knows that each cup of tea requires 250 mL of water and the initial temperature of the water is 25 °C. She also knows that the final temperature of water required is 100 °C and that all heat energy in the kettle is transferred to the water. How much energy (in J) is required for Elizabeth to make a cup of tea?

What information is presented in the question?

Volume of water: 250 mL

Initial temperature: 25 °C

Final temperature: 100 °C

What is the question asking us to do?

Calculate the heat energy required to raise the temperature of 250 mL of water from 25 °C to 100 °C?

What strategies do we need in order to answer the question?

1. Determine the mass of water present.
2. Calculate the temperature change of the water.
3. Using the specific heat capacity of water, substitute values into $q = m \times c \times \Delta T$ to calculate the amount of energy required.

Answer

The mass of water present is
$250\ \text{mL} \times 0.997\ \text{g mL}^{-1} = 249\ \text{g}$

The temperature change of water is calculated by:

$\Delta T = T_{final} - T_{initial}$

$\Delta T = 100\ {}^\circ\text{C} - 25\ {}^\circ\text{C}$

$\Delta T = 75\ {}^\circ\text{C}$

The specific heat capacity of water is $4.18\ \text{J g}^{-1}\ {}^\circ\text{C}^{-1}$

Substitute these values into the equation:

$q = m \times c \times \Delta T$

$q = 249\ \text{g} \times 4.18\ \text{J g}^{-1}\ {}^\circ\text{C}^{-1} \times 75\ {}^\circ\text{C}$

$q = 78\ 061.5\ \text{J}$

Hence, 7.8×10^4 J of energy is required to make the cup of tea.

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USEFUL TIP

1000 J = 1 × 10^3 J = 1 kJ; multiplying factors for SI unit conversions such as this (e.g. mega = 10^6) can be found in the VCE Data Book.

The specific heat capacity of water can also be used to calculate the temperature change that a certain volume/mass of water will undergo when a certain amount of heat energy is applied.

WORKED EXAMPLE 2

When combusted, ethanol releases 29.6 kJ per gram. Calculate the final temperature of 350 mL of water heated by the combustion of 1.0 g of ethanol. Assume the temperature of the water is initially at 25 °C and all heat energy released from the combustion of ethanol is transferred to the water.

What information is presented in the question?

Volume of water: 350 mL

Initial temperature: 25 °C

Heat energy released: 29.6 kJ

What is the question asking us to do?

Calculate the final temperature of 350 mL of water, initially at 25 °C, when 1.0 g of ethanol is combusted.

What strategies do we need in order to answer the question?

1. Determine the mass of water present.
2. Convert the amount of heat expressed in kJ to J.
3. Using the specific heat capacity of water, substitute the values into $q = m \times c \times \Delta T$ to calculate the temperature change of the water.
4. Determine the final temperature of the water by using the equation $\Delta T = T_{final} - T_{initial}$

Answer

Mass of water is 350 mL × 0.997 g mL^{-1} = 349 g.

1.0 g of ethanol releases 29.6 kJ.
29.6 kJ × 1000 = 29 600 J

The specific heat capacity of water is 4.18 J $\text{g}^{-1}\,°\text{C}^{-1}$.

Substitute these values into the equation:

$q = m \times c \times \Delta T$

$29\,600\text{ J} = 349\text{ g} \times 4.18\text{ J g}^{-1}\,°\text{C}^{-1} \times \Delta T$

$$\Delta T = \frac{29\,600\text{ J}}{349\text{ g} \times 4.18\text{ J g}^{-1}\,°\text{C}^{-1}}$$

$\Delta T = 20\ °\text{C}$

$\Delta T = T_{final} - T_{initial}$

$20\ °\text{C} = T_{final} - 25\ °\text{C}$

$T_{final} = 45\ °\text{C}$

Therefore, the final temperature of the water is 45 °C.

Why is liquid water denser than ice?

Another remarkable consequence of water's hydrogen bonding capabilities is that water is denser as a liquid than as a solid. This is because when the temperature of water is below its freezing point of 0 °C, the water molecules each bond to 4 neighbouring water molecules in a regular lattice, as per figure 10.

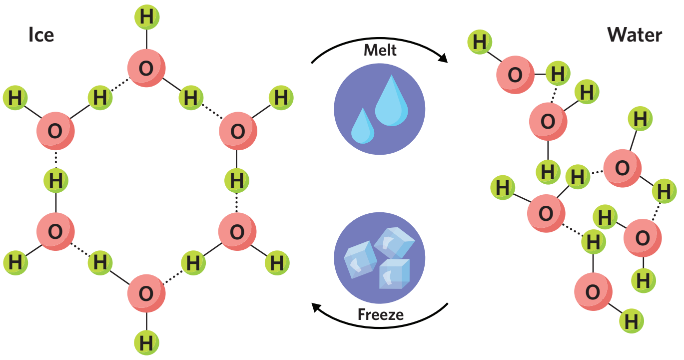

Figure 10 Arrangement of water molecules as a solid (left) and as a liquid (right)

However, this rigid, tetrahedral lattice structure prevents water molecules from being as close together as they can be. Conversely, the random arrangement of water molecules as a liquid allows the strong hydrogen bonds to pull water molecules closer together, and so there is less empty space between water molecules in a liquid state than in ice. As a result, liquid water is ~9% denser than ice. This explains why ice floats on water.

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Progress questions

Question 8

When two different substances absorb the same amount of thermal energy, they will experience different temperature changes.

A. True

B. False

Question 9

Specific heat capacity is usually measured in

A. kJ mol^{-1}.

B. J g^{-1} $°C^{-1}$.

Question 10

In order to perform calculations involving water's specific heat capacity, which of the following equations should be used?

A. $q = m \times c \times T$

B. $q = m \times c \times \Delta T$

Question 11

The units J g^{-1} $°C^{-1}$ and kJ kg^{-1} K^{-1} are equivalent for the specific heat capacity of a substance.

A. True

B. False

Question 12

Ice is __________ than liquid water.

A. denser

B. less dense

Latent heat 2.1.3.1

Relative to other molecular substances, water requires a high quantity of energy to transition from a liquid to a gas at its boiling point.

Why does water have a high latent heat of vaporisation?

As a liquid is heated under constant pressure, its temperature will increase until it reaches its boiling point. At this point, the liquid does not immediately change state into a gas; instead, energy is continually absorbed by the liquid as it boils without raising the temperature. Once a certain quantity of energy has been absorbed by the liquid at its boiling point, the state change from liquid to gas can occur. This quantity of energy is equal to the substance's **latent heat of vaporisation**, shown in figure 11.

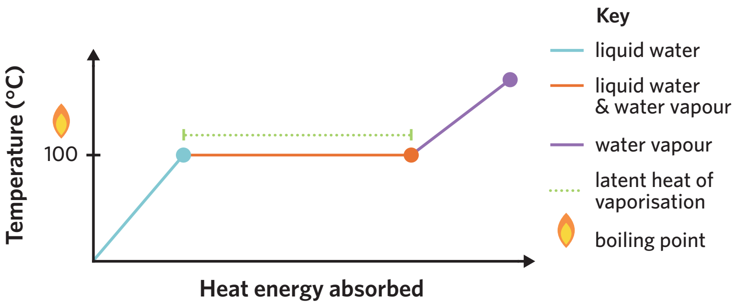

Figure 11 Graph showing water's latent heat of vaporisation

In other words, for a substance A at its boiling point, the latent heat of **vaporisation** is the amount of energy needed for the following state change to occur: $A(l) \rightarrow A(g)$

Like specific heat capacity, a substance's latent heat of vaporisation reflects the strength of intermolecular forces present. Stronger intermolecular forces are able to absorb more energy before being overcome and forming a gas; hence, water (with strong hydrogen bonds between its molecules) has a high latent heat of vaporisation relative to other molecular substances, as shown in table 2.

Table 2 Comparison of the latent heats of vaporisation of different molecular substances

Substance	Latent heat of vaporisation (kJ mol^{-1})
Water (H_2O)	44.0
Hydrogen (H_2)	0.90
Oxygen (O_2)	6.82

The latent heat of vaporisation for water is given as 44.0 kJ mol^{-1}, meaning that it requires 44.0 kJ of energy to convert 1 mol of water at 100 °C from a liquid to a gas. Just like calculations involving the specific heat capacity, we can also calculate the amount of energy required for a state change from liquid to gas to occur.
For example, for water,

$q = n \times 44.0 \text{ kJ mol}^{-1}$

where q is the quantity of heat energy (kJ) required to boil the water at 100 °C, n is the amount of water present (mol), and 44.0 kJ mol^{-1} is water's latent heat of vaporisation value.

WORKED EXAMPLE 3

Calculate the heat energy (in kJ) required to convert 450 g of liquid water to water vapour at water's boiling point (Latent heat of vaporisation of water = 44.0 kJ mol^{-1}).

What information is presented in the question?

Water is at its boiling point.
Mass of water: 450 g
Latent heat of vaporisation of water: 44.0 kJ mol^{-1}

What is the question asking us to do?

Use the latent heat of vaporisation to calculate how much heat energy is required to boil 450 g of water.

What strategies do we need in order to answer the question?

1. Calculate the amount (in mol) of water present.
2. Calculate the heat energy with the equation $q = n \times 44.0 \text{ kJ mol}^{-1}$.

Answer

$n(H_2O) = \frac{m}{M}$

$n(H_2O) = \frac{450 \text{ g}}{((2 \times 1.0) + 16.0) \text{ g mol}^{-1}}$

$n(H_2O) = 25.0 \text{ mol}$

Latent heat of vaporisation of water = 44.0 kJ mol^{-1}

$q = n \times 44.0 \text{ kJ mol}^{-1}$

$q = 25.0 \text{ mol} \times 44.0 \text{ kJ mol}^{-1}$

$q = 1.10 \times 10^3 \text{ kJ}$

Therefore, 1.10×10^3 kJ of heat energy is required to boil 450 g of water.

Progress questions

Question 13

Latent heat of vaporisation is measured in

A. kJ mol^{-1}.

B. J g^{-1} $°C^{-1}$.

Question 14

Water needs __________ energy to change from a liquid to a gas at its boiling point than most other molecular substances.

A. more

B. less

Question 15

In order to perform calculations involving water's latent heat of vaporisation, which of the following equations should be used?

A. $q = m \times c \times \Delta T$

B. $q = n \times 44.0$

Significance of the properties of water 2.1.3.2

Water's thermal properties (high boiling point, high specific heat capacity, and high latent heat of vaporisation) are part of why water is such an important chemical, particularly to life forms who make water their home.

Why are water's thermal properties vital for aquatic life?

Water's high specific heat capacity and latent heat of vaporisation values are vital for all life forms on Earth. All living organisms are predominantly made up of water, so water's high specific heat capacity ensures that our internal temperature remains relatively constant even as the external temperature fluctuates. Liquid water is also vital to a multitude of biological processes; water's high boiling point ensures it remains a liquid even in hot climates, allowing life to be found almost everywhere on Earth.

Aquatic organisms are especially reliant on water's unique thermal properties to survive. Like with internal temperatures, water's high specific heat capacity prevents **solar energy** from significantly raising the temperature of aquatic habitats like rivers and oceans. This is vital because many aquatic creatures can only thrive in water of a narrow range of temperatures, as shown in table 3.

Table 3 Ideal habitat temperature ranges for different aquatic species

Species name	Ideal habitat temperature (°C)
Sea turtle	23–27
Common octopus	22–28
Orange clownfish	25–28

The high latent heat of vaporisation of water is also necessary for ocean temperature regulation. As oceans absorb solar energy, water is constantly evaporating at their surface. Since a large amount of energy is absorbed by liquid water as it becomes a gas, the evaporation process results in a similarly large amount of energy being removed from the water, as explained in figure 12.

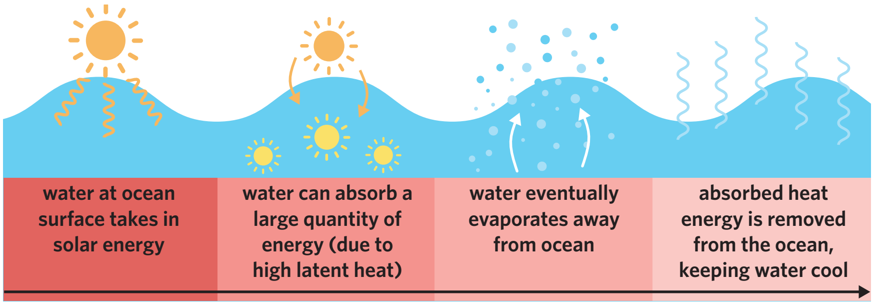

Figure 12 Flow chart explaining how water's high latent heat of vaporisation helps to regulate ocean surface temperature

This aids in keeping ocean temperatures relatively constant and thus maintaining marine life. However, rising global temperatures are beginning to counteract water's natural thermal regulation, causing an increase in average ocean surface temperatures worldwide, as per figure 13.

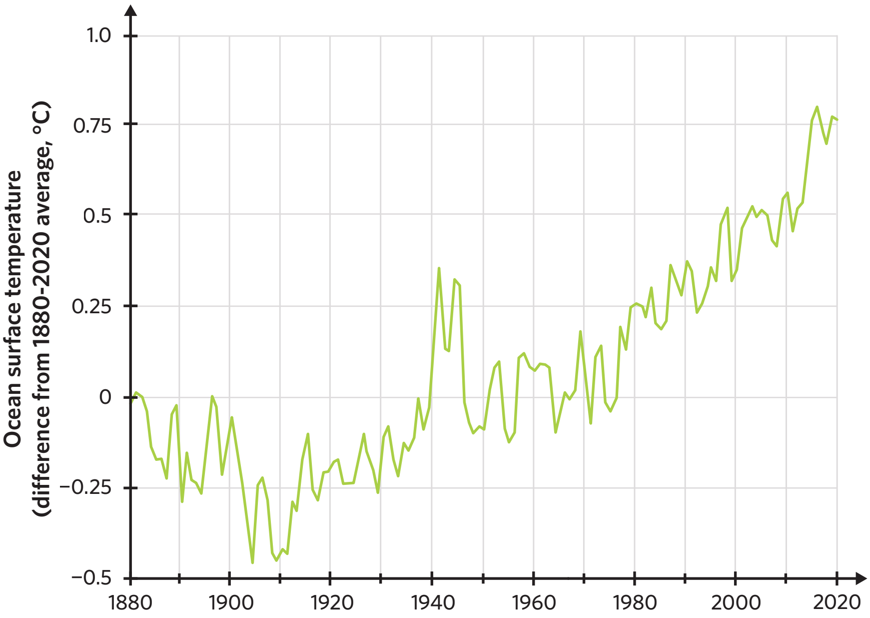

Figure 13 Graph of rise in average ocean surface temperature from 1880 to 2020

As table 3 indicates, an average ocean temperature change of even a few degrees will have devastating consequences for millions of aquatic species,[2] amplifying the need to reverse global warming.

KEEN TO INVESTIGATE?

[2] **How has global warming caused 99% of Australian green turtles to be born female?**
Search: Sea turtle sex ratio crisis

Progress questions

Question 16

Water's thermal properties ensure the ocean's temperature ___________.

A. fluctuates often

B. remains relatively constant

Question 17

Since water has a high latent heat of vaporisation, evaporation at the ocean surface causes a ___________ amount of energy to be released.

A. large

B. small

Question 18

The rise of ocean surface temperatures will likely affect aquatic life forms.

A. True

B. False

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Theory summary

- Water is the only substance that naturally occurs as a solid, liquid, and gas on Earth.
- Freshwater makes up a very small portion of the total water on Earth, and only a small fraction of this freshwater is potable and accessible to humans.
- Water has an atypically high boiling point compared with other group 16 hydrides, and is denser as a liquid than a solid.
- Water has a relatively high specific capacity (measured in J g^{-1} °C^{-1}), so it requires a large amount of energy to increase in temperature.
- Water has a relatively high latent heat of vaporisation value (measured in kJ mol^{-1}), meaning it requires a large amount of energy to transition from a liquid to a gas at its boiling point.
- The thermal properties of water are due to the strong hydrogen bonds between water molecules, and are vital to all life on Earth, especially aquatic organisms.

The content in this lesson is considered fundamental prior knowledge to measuring changes in chemical reactions (Unit 3 AOS 1).

9A Questions

Mild Medium Spicy

Deconstructed

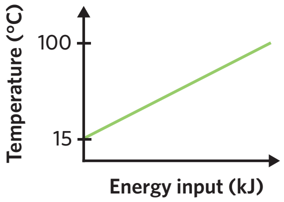
Use the following information to answer questions 19–22.

Suzie wants to determine how much energy is required to completely convert 150 g of liquid water at 15 °C to water vapour at 100 °C.

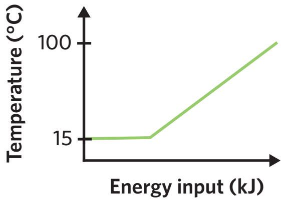
Question 19 (1 MARK)

Which of the following graphs best represents the process described above?

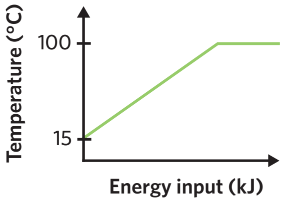
A.



B.



C.



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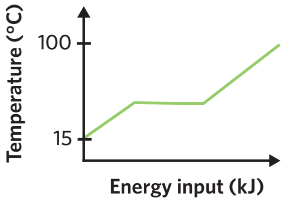

Question 20 (1 MARK)

The flat section on your chosen graph represents

A. specific heat capacity.

B. latent heat of vaporisation.

C. the heating element being turned off for a period of time.

D. there is no flat section.

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Question 21 (1 MARK)

In order to determine how much heat energy is required, Suzie needs to use water's

I. latent heat of vaporisation value.

II. specific heat capacity value.

A. II only

B. I only

C. I and II

D. Neither I nor II

Question 22 (4 MARKS)

Calculate how much heat energy Suzie requires (latent heat of vaporisation of water = 44.0 kJ mol^{-1}). Show all working.

Exam-style

Question 23 (11 MARKS)

Water is abundant on Earth, and its various forms can be categorised in numerous ways. In Samuel Taylor Coleridge's poem 'The Rime of the Ancient Mariner', a sailor despairs that there is 'water, water, everywhere, and not a drop to drink'.

a. Identify a natural source for each of the following forms of water: freshwater, brackish water, water vapour. (3 MARKS)

b. Assess the sailor's claim, with reference to the proportion of potable water on Earth. (2 MARKS)

c. Identify two factors that must be considered when determining if water from a natural source is suitable for drinking. (2 MARKS)

d. Another major concern for sailors is icebergs. The Titanic, the largest passenger ship of its time, was shipwrecked after colliding with an iceberg in 1912. Explain why icebergs float on the ocean surface, referring to chemical structure and using a diagram to support your answer. (4 MARKS)

Question 24 (1 MARK)

What would be the final temperature of 50 g of water initially at 27 °C after being heated with 12 kJ of energy?

A. 28 °C

B. 84 °C

C. 62 °C

D. 93 °C

Question 25 (6 MARKS)

The high latent heat of vaporisation and specific heat capacity of water have many practical applications.

a. Define specific heat capacity and latent heat of vaporisation. (2 MARKS)

b. With reference to the intermolecular forces present, explain why water has high values for its specific heat capacity and latent heat of vaporisation. (2 MARKS)

c. Water is used as a coolant in electronic products to absorb excess heat. What property of water makes it desirable for this purpose? Justify your response. (2 MARKS)

Question 26 (5 MARKS)

Freya is looking to compare the merits of cooking food in alcohol (ethanol) versus in water.

Liquid	Boiling point (°C)	Specific heat capacity (J g^{-1} °C^{-1})
Water	100	4.18
Ethanol	78.4	2.46

a. Referring to thermal properties, identify and explain one advantage of cooking food in ethanol rather than water. (2 MARKS)

b. Referring to thermal properties, identify and explain one advantage of cooking food in water, rather than ethanol. (2 MARKS)

c. What do the values in the table suggest about the strength of intermolecular forces in ethanol, compared with water? (1 MARK)

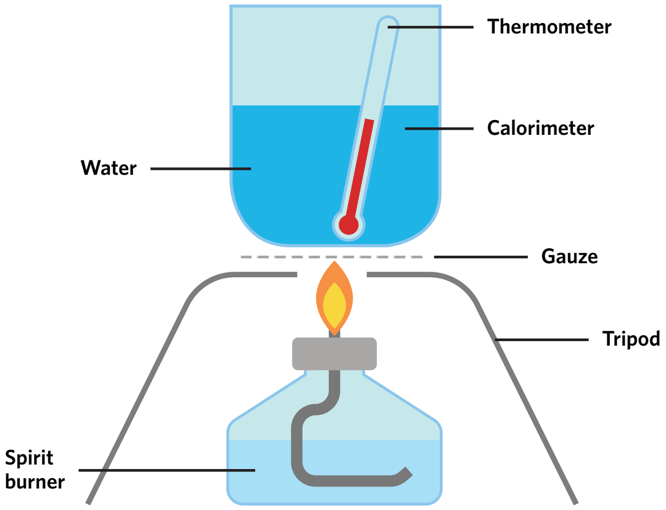
Question 27 (4 MARKS)

What is the energy (in kJ) required to completely convert 250 mL of water, initially at 25.0 °C, to water vapour at 100 °C (latent heat of vaporisation of water = 44.0 kJ mol^{-1})? Show all working.

Key science skills

Question 28 (9 MARKS)

Calorimetry is the process of measuring the amount of energy released from combustion of a substance by determining the temperature change of the water. An example of a calorimeter is shown.



Ishaan wants to determine by how much water's temperature will increase when he combusts 10.0 g of a certain substance in a calorimeter. He repeats this experiment 5 times, and his results were as follows.

	Trial 1	Trial 2	Trial 3	Trial 4	Trial 5
Initial water temp (°C)	28.3	45.1	61.6	76.6	92.1
Final water temp (°C)	46.4	62.5	78.0	92.1	100.0
ΔT (°C)	18.1	17.4	16.4	15.5	7.9

a. What type of data is collected in a calorimetry experiment? (1 MARK)

b. Identify a potential systematic error in this experiment. (1 MARK)

c. Suggest two ways to improve the experiment. (2 MARKS)

d. Why is it important to repeat experiments? (1 MARK)

e. What variable has Ishaan erroneously failed to control across the 5 trials? (1 MARK)

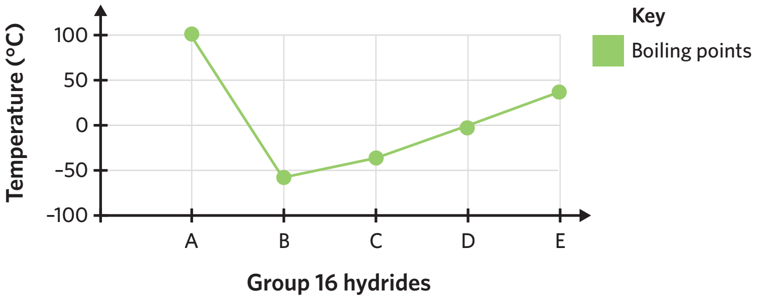
f. Trial 5 produced a vastly different value for ΔT compared with the other trials. Referring to water's thermal properties, explain why this outlying result was obtained. (3 MARKS)

FROM LESSONS 16B, 16C, 16D & 16E

Questions from multiple lessons

Question 29 (7 MARKS)

Water is one member of the group 16 hydrides, and Jessie is investigating the properties of group 16 hydrides for a scientific poster. Shown is a graph of the boiling points of group 16 hydrides from her experimental results, but unfortunately it is missing labels on the horizontal axis.



a. What is a group 16 hydride? (1 MARK)

b. What is the name and molecular formula of compound A? Explain why it has such a high boiling point compared to other group 16 hydrides. (3 MARKS)

c. Explain why the boiling points increase from compound B to compound E (excluding compound A) with reference to the forces present in these substances. (3 MARKS)

FROM LESSON 2B

Question 30 (9 MARKS)

Ethanol is a very common compound used in thermochemistry.

a. Draw the structural formula of ethanol. What is the functional group present? (2 MARKS)

b. Explain why ethanol has a relatively high specific heat capacity, but a lower specific heat capacity than water. (3 MARKS)

c. Ethanol can also be combusted to heat water. The combustion of 1.0 g of ethanol releases 29.6 kJ of energy. 300 mL of water at 20 °C is heated by the combustion of 3.0 g of ethanol.

What is the final temperature of the water? Assume all the energy released from the combustion of ethanol is transferred to the water. (4 MARKS)

FROM LESSON 7C

Hints

23a. Consider where you have encountered different forms of water.

23b. Potable water is only one type of water available on Earth.

23c. The body is inefficient at processing salt, and water contaminants can be harmful.

23d. Less dense substances will float on top of denser substances.

24. $q = m \times c \times \Delta T$ calculates the temperature change, not the final temperature.

25a. The units of latent heat and specific heat capacity reflect what is measured.

25b. Different types of intermolecular forces require different amounts of heat to overcome them.

25c. Excess heat can be transferred to another substance.

26a. Substances with higher specific heat capacity need more energy to heat up.

26b. Upon reaching their boiling point, substances begin to evaporate and rise.

26c. Stronger intermolecular forces need less energy to be overcome.

27. Water's specific heat capacity is 4.18 J g^{-1} $°C^{-1}$ and its latent heat of vaporisation is 44.0 kJ mol^{-1}.

28a. The measured variable is the amount by which the water's temperature increases.

28b. Systematic errors are errors that affect all trials similarly.

28c. As it stands, the beaker is poorly insulated.

28d. Repeating experiments gives multiple sets of results.

28e. If simply repeating trials of the same experiment, it is important to control as many variables as possible.

28f. In a graph of a substance's temperature vs. energy absorbed, there is always a flat section at the boiling point.

29a. Group 16 elements include oxygen, sulphur and tellurium.

29b. Water is the only group 16 hydride able to form hydrogen bonds.

29c. Dispersion force strength increases as molecules grow larger.

30a. Ethanol is an alcohol.

30b. Intermolecular force strength affects specific heat capacity.

30c. $q = m \times c \times \Delta T$ calculates the temperature change, not the final temperature.

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Chapter 9 review

Multiple choice (10 MARKS)

Question 1 (1 MARK)

Water in its gaseous state is present

A. above natural hot springs.

B. between clouds.

C. in polar regions.

D. All of the options correctly complete the statement.

Question 2 (1 MARK)

Which of the following characteristics is common to all group 16 hydrides?

A. High specific heat capacity

B. High boiling point

C. V-shape of the molecule

D. Lower density of solid compared with liquid

Question 3 (1 MARK)

Water's boiling point is __________ °C. This is considered relatively __________ compared to other group 16 hydrides.

A. 44, low

B. 44, high

C. 100, low

D. 100, high

Question 4 (1 MARK)

H_2O contradicts the boiling point trend of other group 16 hydrides mainly because

A. H_2O is the smallest molecule.

B. H_2O has the strongest dispersion forces between its molecules.

C. H_2O has hydrogen bonds between molecules.

D. H_2O is v-shaped.

Question 5 (1 MARK)

Which of the following statements is false?

A. There is no life form on Earth that can exist without water.

B. There is no substance on Earth that naturally exists as a solid, liquid, and gas, except water.

C. There is no liquid with a higher specific heat capacity than liquid water.

D. There are no molecular substances which lack dispersion forces between their molecules.

Use the following information to answer questions 6–8.

Liquid	Specific heat capacity ($J\ g^{-1}\ °C^{-1}$)
Skim milk	3.97
Whole milk	3.89

Source: American Association of Cereal Chemists

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Question 6 (1 MARK)

Equal masses of skim milk, whole milk, and water are each poured into three identical beakers. Starting from the same initial temperature, the three beakers are heated with equal quantities of heat energy (but not enough energy to reach the boiling point of any of the liquids). After heating, which liquid will be the warmest?

A. Skim milk

B. Whole milk

C. Water

D. There is not enough information provided to draw a conclusion.

Question 7 (1 MARK)

In a separate experiment, 100 g of skim milk, starting at 25.0 °C, was heated with 4.30 kJ of energy. What was the final temperature of the skim milk?

A. 10.9 °C

B. 25.1 °C

C. 35.3 °C

D. 35.8 °C

Question 8 (1 MARK)

To find the latent heat of vaporisation of skim milk, one student suggests calculating the energy required to completely evaporate the skim milk, then subtracting the energy required to heat the skim milk to 100 °C. What is the problem with the student's proposed method?

A. Latent heat of vaporisation takes into account the energy needed to heat the substance to its boiling point, so this should not be subtracted.

B. It is not certain that skim milk's boiling point is 100 °C.

C. Skim milk does not necessarily have a latent heat of vaporisation.

D. There is no problem with the student's method.

Question 9 (1 MARK)

Grace begins boiling water to make tea, but then decides to conduct an experiment. She calculates that once this water has reached its boiling point, 500 kJ of energy is needed to convert it completely to vapour. What mass of water did Grace boil?

A. 205 g **B.** 11.4 g **C.** 120 g **D.** 2150 g

Question 10 (1 MARK)

What is the energy required to completely evaporate 100 g of water starting from room temperature at 20 °C?

A. 33 kJ

B. 244 kJ

C. 277 kJ

D. 4420 kJ

Short answer (30 MARKS)

Question 11 (6 MARKS)

Many of water's unique properties arise because water molecules form strong hydrogen bonds with each other.

a. Why can a water molecule form hydrogen bonds with 4 other water molecules? (3 MARKS)

b. Order the following compounds from lowest to highest boiling point: H_2Te, H_2O, H_2Po, H_2Se, H_2S. Why is this order not entirely as expected? (3 MARKS)

Question 12 (9 MARKS) 🌶🌶

Perspiration, or sweating, is one of the most important biological processes for keeping humans cool during hot weather or exercise. This involves droplets of liquid water forming on the surface of the skin, which cool the body down.

a. What property of water enables sweat to absorb a large quantity of heat? (1 MARK)

b. Ashari is an athlete who carefully tracks her training sessions. She finds that during a session of intense physical activity, her muscles have generated 100 kJ of heat energy. If all this heat energy was used to evaporate sweat, how much sweat, in mL, did Ashari produce (density of water = 0.997 g mL^{-1}, latent heat of vaporisation of water = 44.0 kJ mol^{-1})? (3 MARKS)

c. Following her training session, Ashari cools off further by drinking a cup of water. She observes that the ice cubes float on top of the water. Why do these ice cubes float? Explain with reference to the intermolecular forces present. (3 MARKS)

d. After winning her event, Ashari rewards herself with a small glass of whiskey with ice. She notices that the ice cubes have sunk to the bottom of the whiskey. Which is denser – whiskey or liquid water? Justify your answer using the information provided. (2 MARKS)

Question 13 (7 MARKS) 🌶🌶

Water is not only a molecule with many unique properties, but also one that is vital to all life on Earth.

a. One such rare property is that water is less dense as a solid than as a liquid. Ice also has the ability to reflect solar energy back into the lower atmosphere. Explain why in regions where icebergs are common, these properties also help to regulate ocean surface temperature. (2 MARKS)

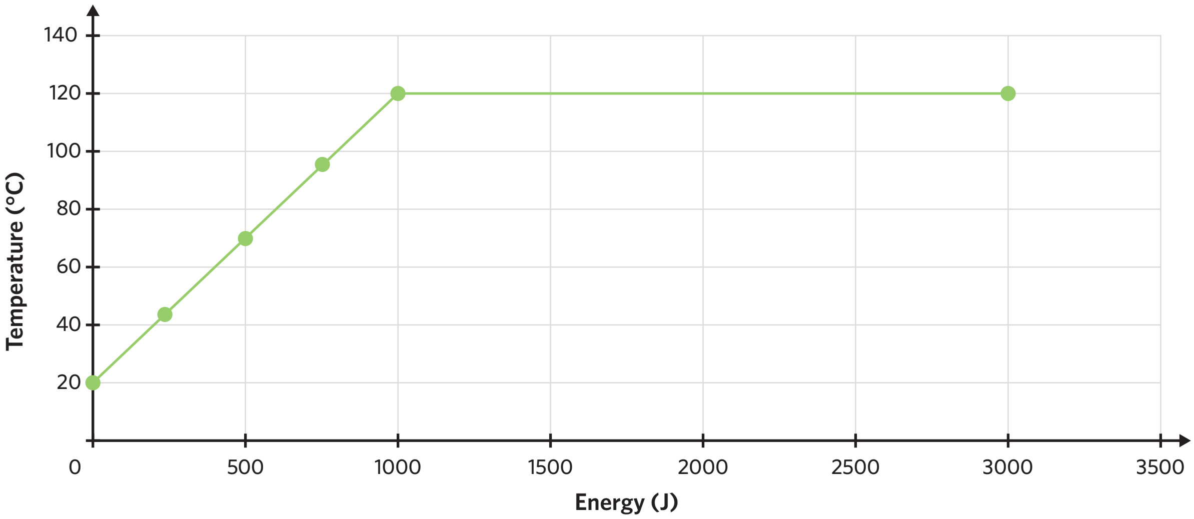
b. Another of these properties is that water has a high latent heat of vaporisation. Why does water have a high latent heat of vaporisation? Explain. (2 MARKS)

c. The angelfish is particularly sensitive to water temperature, only inhabiting water at temperatures between 26 °C and 29 °C. Explain why water's high latent heat of vaporisation is vital to the angelfish's survival. (3 MARKS)

Question 14 (8 MARKS) 🌶🌶🌶

Liquid	Molar mass (g mol^{-1})	Specific heat capacity (J g^{-1} °C^{-1})
Ethanol	46.0	2.46
Ethanoic acid	60.0	2.04
Chloroform	119.5	0.96

Dilruk is determining the identity of an unknown liquid. He knows the liquid is either ethanol, ethanoic acid, or chloroform, and he has access to the information in the given table. He heats 4.90 g of the liquid, starting at 20 °C, until it all vaporises, and using the data obtained he draws the following graph.



a. Using the graph, calculate the specific heat capacity of Dilruk's unknown substance, and hence determine its identity from the table. (4 MARKS)

b. Calculate the latent heat of vaporisation (in kJ mol^{-1}) of the substance whose identity you determined in part a. (4 MARKS)

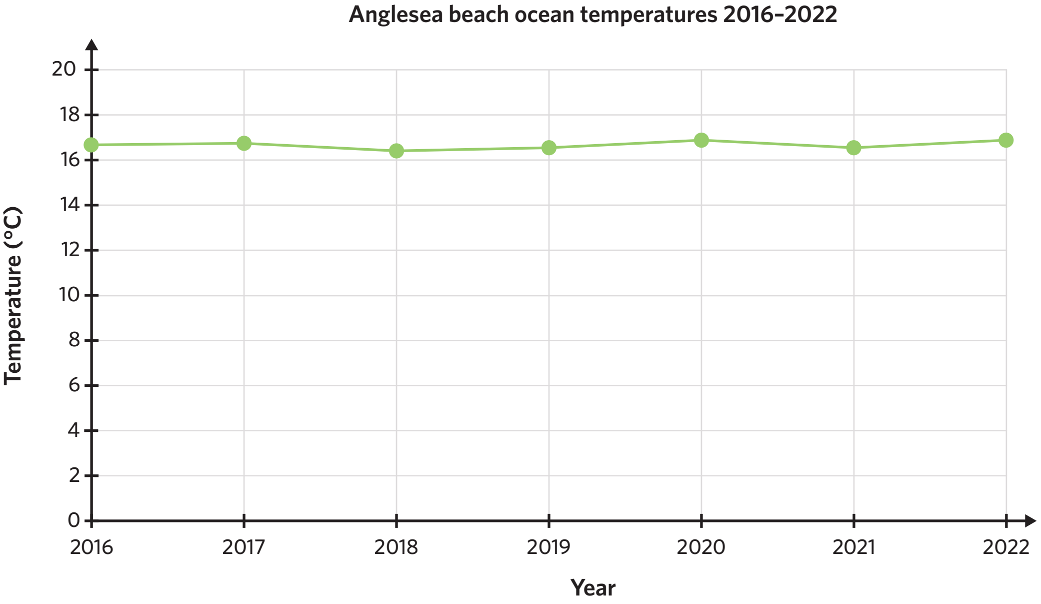
Key science skills (10 MARKS)

Question 15 (10 MARKS)

A student aims to investigate whether ocean temperatures are rising globally. To achieve this, each year on their family holiday from 2016 to 2022, they measured the temperature of water at Anglesea beach, where they were staying. The following is an extract from their scientific report.

Discussion

From the measurements taken between 2016 and 2022, the ocean temperature increased from 16.6 °C to 17.0 °C. However, as is clear in figure 1, the graph is essentially flat, so this change can be considered negligible, and any variation can be attributed to random error. Furthermore, anyone will tell you that the water does not feel any warmer than it used to.



Conclusion

It was found that the temperature of water at Anglesea Beach did not increase significantly from 2016 to 2022 and it was therefore concluded that global ocean temperatures are not rising.

a. Evaluate the validity of the student's conclusion. Justify your response. (2 MARKS)

b. Suggest two improvements that could be made to the experiment design to better achieve the outlined aim. (2 MARKS)

c. "the ocean temperature increased from 16.6 °C to 17.0 °C. However...the graph is essentially flat". Identify the major issue with the graph, and suggest a change to improve the student's data representation. (2 MARKS)

d. "...this change is negligible..." Evaluate this statement with reference to aquatic ecosystems. (3 MARKS)

e. Give an example of a non-scientific idea presented in this scientific report extract. (1 MARK)

FROM LESSONS 16A, 16C, 16D & 16E

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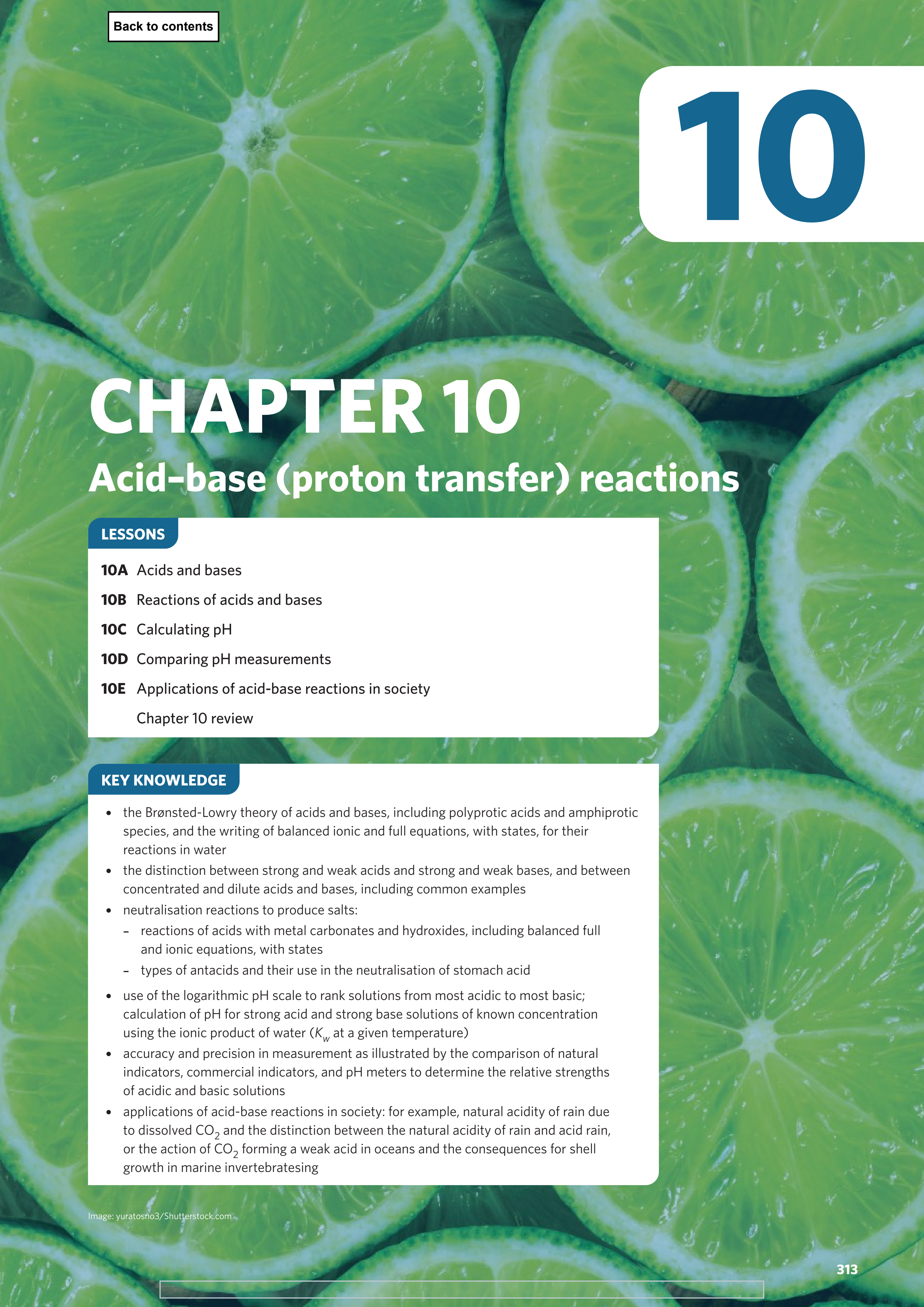

10

CHAPTER 10

Acid–base (proton transfer) reactions

LESSONS

10A Acids and bases

10B Reactions of acids and bases

10C Calculating pH

10D Comparing pH measurements

10E Applications of acid-base reactions in society

Chapter 10 review

KEY KNOWLEDGE

- the Brønsted-Lowry theory of acids and bases, including polyprotic acids and amphiprotic species, and the writing of balanced ionic and full equations, with states, for their reactions in water
- the distinction between strong and weak acids and strong and weak bases, and between concentrated and dilute acids and bases, including common examples
- neutralisation reactions to produce salts:
  - reactions of acids with metal carbonates and hydroxides, including balanced full and ionic equations, with states
  - types of antacids and their use in the neutralisation of stomach acid
- use of the logarithmic pH scale to rank solutions from most acidic to most basic; calculation of pH for strong acid and strong base solutions of known concentration using the ionic product of water (K_w at a given temperature)
- accuracy and precision in measurement as illustrated by the comparison of natural indicators, commercial indicators, and pH meters to determine the relative strengths of acidic and basic solutions
- applications of acid-base reactions in society: for example, natural acidity of rain due to dissolved CO_2 and the distinction between the natural acidity of rain and acid rain, or the action of CO_2 forming a weak acid in oceans and the consequences for shell growth in marine invertebratesing

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10A Acids and bases

STUDY DESIGN DOT POINTS

- the Brønsted-Lowry theory of acids and bases, including polyprotic acids and amphiprotic species, and the writing of balanced ionic and full equations, with states, for their reactions in water
- the distinction between strong and weak acids and strong and weak bases, and between concentrated and dilute acids and bases, including common examples

10A 10B 10C 10D 10E

2.1.4.1 The Brønsted-Lowry theory

2.1.4.2 Types of acids and bases

2.1.5.1 Strength of acids and bases

ESSENTIAL PRIOR KNOWLEDGE

1A Proton

See question 59.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

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What makes an acid a 'super acid'?

Some **acids** are so dangerous that they can dissolve plastic and glass. In this lesson, we will be learning about the definition, nature and **strengths** of acids and **bases.**

KEY TERMS AND DEFINITIONS

Acid species that donates a proton, H^+, during an acid-base reaction

Acid-base reaction chemical reaction where a proton, H^+, is transferred between two chemical species

Alkali a base that is soluble in water and neutralises acids

Amphiprotic species that can either donate or accept a proton, H^+

Base species that accepts a proton, H^+, during an acid-base reaction and are a group of substances that neutralise acids

Brønsted-Lowry theory definition of an acid as a chemical species which donates a proton, H^+, to a base

Corrosive a highly reactive substance that causes obvious damage to tissue

Deprotonation loss of a proton, H^+, from an acid

Diprotic species that can donate up to two protons, H^+, during ionisation

Hydronium ion positively charged chemical species with molecular formula H_3O^+

Hydroxide ion negatively charged chemical species with molecular formula OH^-

Ionisation process by which a chemical species gains or loses an electron to gain an electric charge

Monoprotic species that can only donate one proton, H^+, during ionisation

Neutralise react with an acid or a base to produce water

Polyprotic species that donates multiple protons, H^+, during ionisation

Protonated water alternate name for the hydronium ion, H_3O^+

Strength how readily an acid or base will donate or accept a proton, H^+

Triprotic species that can donate up to three protons, H^+, during ionisation

The Brønsted-Lowry theory 2.1.4.1

The **Brønsted-Lowry theory** of acids and bases defines an **acid-base reaction** as the donation of a proton, H^+, from an acid to a base (figure 1).

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How do hydrogen ions affect the nature of solutions?

Acids and bases are some of the most recognisable chemicals in both our everyday lives and the wider world of the chemical industry. Since these substances are so widespread, it is important that we understand and are able to define how they function in chemical reactions. The names and formulas of a number of common acids and bases are given in table 1.

Table 1 The names and formulas of common acids and bases

Acids		Bases	
Name	**Formula**	**Name**	**Formula**
Sulfuric acid	H_2SO_4	Sodium hydroxide	$NaOH$
Hydrochloric acid	HCl	Ammonia	NH_3
Nitric acid	HNO_3	Magnesium hydroxide	$Mg(OH)_2$
Phosphoric acid	H_3PO_4	Potassium hydroxide	KOH
Ethanoic acid	CH_3COOH	Sodium carbonate	Na_2CO_3
Carbonic acid	H_2CO_3	Oxide	O^{2-}
Ammonium	NH_4^+	Calcium hydroxide	$Ca(OH)_2$

We can understand the nature of acids and bases using the Brønsted-Lowry theory. According to this theory:

- An acid is a substance that donates a proton, H^+
- A base is a substance that accepts a proton, H^+
- An acid-base reaction involves the transfer of a proton from an acid to a base
- Acids[1] and bases are **corrosive** substances.

An example of an acid-base reaction is the reaction between nitric acid, HNO_3, and water:

$$HNO_3(aq) + H_2O(l) \rightarrow H_3O^+(aq) + NO_3^-(aq)$$

$$HNO_3(aq) + H_2O(l) \rightarrow H_3O^+(aq) + NO_3^-(aq)$$

We can see that in this reaction, nitric acid has donated a H^+ to the water molecule, forming a **hydronium ion**, H_3O^+ (also known as **protonated water**) and a nitrate ion, NO_3^-. Using the Brønsted-Lowry theory, we can confirm that the HNO_3 acted as a proton donor, or acid, and that H_2O acted as a proton acceptor, or base. Even though we usually think of H_2O as being a 'neutral' molecule in terms of acid-base chemistry, it is important to note that it can act as either an acid or base.

The **ionisation** reaction of HNO_3 in water can also be written without the presence of H_2O as shown.

$$HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$$

In this scenario, the $H_3O^+(aq)$ hydronium ion is instead represented as a $H^+(aq)$ ion.

We can also use the example of the base ammonia, NH_3, reacting with water:

$$NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$$

$$NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$$

In this reaction, ammonia has accepted a H^+ from the water molecule to form ammonium, NH_4^+ and the **hydroxide ion**, OH^-. Again, using the Brønsted-Lowry theory – since we know that NH_3 acted as a proton acceptor it must have been a base, and H_2O, since it acted as a proton donor, was the acid.

What is the difference between a hydronium ion and hydroxide ion?

Whenever there is a transfer of a proton, H^+, we know that an acid-base reaction must have taken place. If these reactions are happening between an acid or base and water, then either a hydronium, $H_3O^+(aq)$, or hydroxide, $OH^-(aq)$, ion is produced.

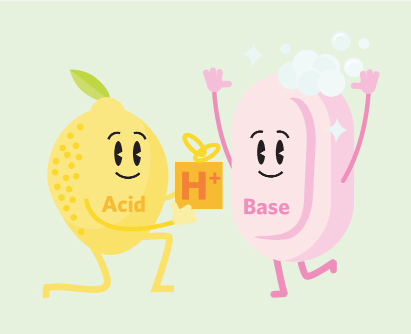

Figure 1 Acids are proton donors and bases are proton acceptors.

KEEN TO INVESTIGATE?

[1] **What is the strongest acid in the world?**
Search YouTube: This SUPERACID Will Dissolve Anything!

STRATEGY

Throughout this chapter, $H^+(aq)$ and $H_3O^+(aq)$ are essentially identical and will only affect whether or not $H_2O(l)$ is included explicitly as a reactant.

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MISCONCEPTION

'All bases are alkalis.'

All alkalis are bases, but not all bases are alkalis. Bases and alkalis both **neutralise** acids. The main difference is that alkalis are a group of bases that are soluble in water (e.g. metal hydroxides), whereas all bases are not necessarily soluble in water (e.g. calcium carbonate).

As shown in figure 2, the $H_3O^+(aq)$ and $OH^-(aq)$ ions are formed whenever $H_2O(l)$ either gains or loses a proton. **Alkalis** form hydroxide ions as they are soluble in water.

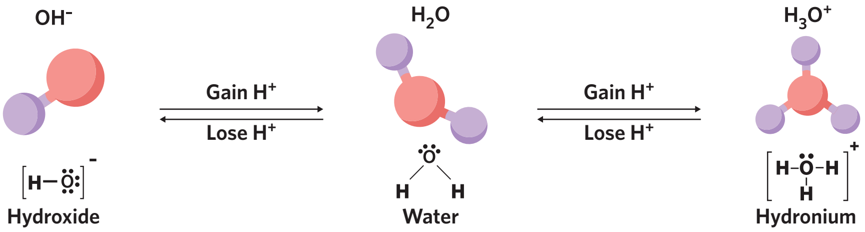

Figure 2 Hydroxide and hydronium ions are formed from either the gain or loss of a H^+ ion.

Not all acid-base reactions involve water specified as a reactant in the equation, but they do require water to occur. This is shown through other reactants being in an aqueous solution with water. In fact, we could just as easily have a reaction between $HNO_3(aq)$ and $NH_3(aq)$. In this reaction, nitric acid will act as an acid and ammonia as a base through the transfer of a proton:

$$HNO_3(aq) + NH_3(aq) \rightarrow NH_4{}^+(aq) + NO_3{}^-(aq)$$

$$HNO_3(aq) + NH_3(aq) \rightarrow NH_4{}^+(aq) + NO_3{}^-(aq)$$

WORKED EXAMPLE 1

Write the balanced equation for the acid-base reaction of hydrochloric acid, HCl, in aqueous solution (with water).

What information is presented in the question?

The molecular formula of hydrochloric acid, HCl.

The reaction is occurring in aqueous solution.

What is the question asking us to do?

Write a balanced equation for the acidic behaviour of HCl in water.

What strategies do we need in order to answer the question?

1. Remember that aqueous solution means that the reaction is occurring in water.
2. Understand that an acid will donate a H^+ ion during an acid-base reaction.

Answer

$$\underset{\text{Acid}}{HCl(aq)} + \underset{\text{Base}}{H_2O(l)} \rightarrow H_3O^+(aq) + Cl^-(aq)$$

As an acid, HCl(aq) will donate a H^+ ion to the base in an acid-base reaction. Since this reaction is occurring in aqueous solution, $H_2O(l)$ is present and can act as a base. The $H_2O(l)$ will accept the H^+ ion from HCl, resulting in the formation of the hydronium ion, $H_3O^+(aq)$. The HCl(aq), since it ionises to donate its H^+ ion, will become an aqueous chloride ion, $Cl^-(aq)$.

This gives the full overall balanced equation as:

$$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

Progress questions

Question 1

Acid-base reactions involve the

A. transfer of neutrons.

B. transfer of H^+ ions.

Question 2

Bases ________ H^+ ions, acids ________ H^+ ions.

A. accept, donate

B. donate, accept

Continues →

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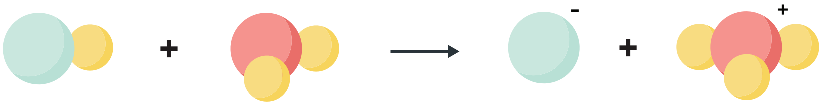
Question 3



$HCl(aq) + H_2O(l) \rightarrow Cl^-(aq) + H_3O^+(aq)$

In this reaction,

A. $HCl(aq)$ is acting as a base, $H_2O(l)$ is acting as an acid.

B. $HCl(aq)$ is acting as an acid, $H_2O(l)$ is acting as a base.

Types of acids and bases 2.1.4.2

Acids and bases can be classified according to their capacity to donate or accept a number of protons.

How can a compound be both an acid and a base?

We have established that for a chemical species to be classified as an acid, it must be able to donate a proton, H^+, and to be a base, it must be able to accept a proton. However, if we investigate the following two reactions, we can see that some species can do both:

$$\underset{\text{Acid}}{HNO_3\,(aq)} + \underset{\text{Base}}{H_2O(l)} \rightarrow H_3O^+(aq) + NO_3^-(aq)$$

$$\underset{\text{Base}}{NH_3(aq)} + \underset{\text{Acid}}{H_2O(l)} \rightarrow NH_4^+(aq) + OH^-(aq)$$

In these two acid-base reactions, we can see that $H_2O(l)$ can act either as a base or as an acid depending on which species it is reacting with. Any substance that can act as either an acid or a base is known as an **amphiprotic** substance, meaning that it can both donate and accept a proton in an acid-base reaction.

A list of amphiprotic substances and the ions they form when acting as either an acid or a base is given in figure 3.

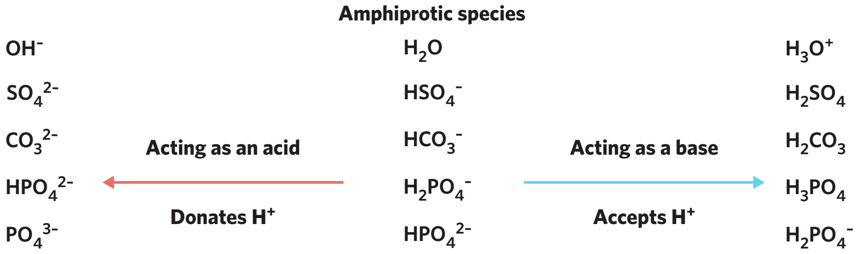


Figure 3 Amphiprotic species can donate or accept a H^+ ion, acting as either an acid or a base.

How do we classify different types of acids?

We know that any species that can donate a H^+ ion is classified as an acid. We can define this process of giving up a proton as **deprotonation**. Any acid that will only donate one proton is classified as **monoprotic**. Once these acids donate their only H^+ ion, they can no longer react as acids in acid-base reactions. We have already investigated the ionisation reaction of nitric acid, HNO_3, and, removing $H_2O(l)$ from the equation, we can follow nitric acid's behaviour more closely.

$HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$

Once $HNO_3(aq)$ gives up its only H^+ ion, the nitrate ion that forms has no more H^+ ions to donate. Therefore, we would classify HNO_3 as a monoprotic acid.

There are a number of acids that are capable of undergoing deprotonation multiple times consecutively. This means that they can donate a H^+ and then the compound formed from this deprotonation can again donate a H^+ ion.

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These species are defined as being **polyprotic**, meaning they can donate a H^+ ion more than once. If an acid can donate two H^+ ions, it is classified as **diprotic**. An example of a diprotic acid is sulfuric acid, H_2SO_4. Sulfuric acid ionises in water according to the following equations.

(1) $H_2SO_4(aq) \rightarrow HSO_4^-(aq) + H^+(aq)$

(2) $HSO_4^-(aq) \rightleftharpoons SO_4^{2-}(aq) + H^+(aq)$

First, one proton is lost to form HSO_4^-, and then another is donated to form the sulfate ion, SO_4^{2-}.

Note that when H_2SO_4 fully ionises in water, this second deprotonation step only occurs to about 10% of the extent of the first stage resulting in partial ionisation. However, when H_2SO_4 reacts with a base, such as NH_3, this second proton is lost to a much greater (almost complete) extent.

If an acid can donate three H^+ ions, then it is classified as **triprotic**. A common example of a triprotic acid is phosphoric acid, H_3PO_4. Phosphoric acid deprotonates in three stages, since it has the potential to lose three H^+ ions. These reactions are given below.

(1) $H_3PO_4(aq) \rightarrow H_2PO_4^-(aq) + H^+(aq)$

(2) $H_2PO_4^-(aq) \rightleftharpoons HPO_4^{2-}(aq) + H^+(aq)$

(3) $HPO_4^{2-}(aq) \rightleftharpoons PO_4^{3-}(aq) + H^+(aq)$

However, since phosphoric acid is relatively weak, the second and third deprotonation steps here only occur to a low extent. We can use 'bidirectional harpoon' arrows ($\rightleftharpoons$) to show that this reaction will have a large proportion of reactants that haven't ionised. A simple way to explain this is that once a polyprotic acid loses an H^+ ion, it becomes negatively charged. Since opposite charges attract, it then becomes much harder for the positive H^+ ion to escape the pull of the full compound's negative charge. As a result, the second and third deprotonations occur to an increasingly lower extent than the first.

If we were to have a look at a solution of phosphoric acid in water, rather than seeing only PO_4^{3-} and H^+ ions, we would instead see a mixture of H_3PO_4, $H_2PO_4^-$, HPO_4^{2-}, PO_4^{3-} and H^+ ions, as shown in figure 4.

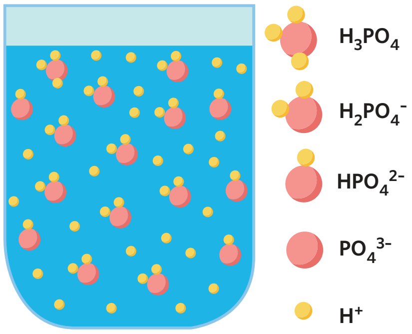

Figure 4 A solution of phosphoric acid will contain a mixture of many different ions.

Table 2 contains a list of common acids and whether they are classified as monoprotic, diprotic or triprotic. Note that a number of amphiprotic species are formed from the deprotonation of a polyprotic acid, since any proton that is lost can be re-accepted if the polyprotic acid acts as a base in another reaction.

Table 2 List of some monoprotic, diprotic and triprotic acids

Monoprotic		Diprotic		Triprotic	
Name	**Formula**	**Name**	**Formula**	**Name**	**Formula**
Nitric acid	HNO_3	Sulfuric acid	H_2SO_4	Phosphoric acid	H_3PO_4
Hydrochloric acid	HCl	Carbonic acid	H_2CO_3	Citric acid	$C_6H_8O_7$
Ethanoic acid	CH_3COOH	Hydrogen sulfide	H_2S	Boric acid	H_3BO_3
Ammonium	NH_4^+	Selenous acid	H_2SeO_3	Arsenic acid	H_3AsO_4

Progress questions

Question 4

Hydrofluoric acid, HF, is a/an

A. amphiprotic acid.

B. monoprotic acid.

Continues →

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Question 5

Dihydrogen phosphate, $H_2PO_4^-$ can be described as

A. diprotic only.

B. diprotic and amphiprotic.

Question 6

All triprotic acids are polyprotic.

A. True

B. False

Strength of acids and bases 2.1.5.1

The strength of an acid or base is related to how readily they donate or accept a H^+ ion.

Why are some acids classified as either strong or weak?

So far, we have classified acids and bases according to whether they donate or accept H^+ ions and how many H^+ ions they are capable of donating. We will now investigate the strength of acids and bases, which depends on how readily an acid donates a H^+ ion or how readily a base accepts a H^+ ion.

A strong acid is one that will undergo almost complete ionisation when added to water. This means that essentially all of the acid molecules in the solution will donate their protons to form aqueous H^+ ions. Three common strong acids are shown below, all of which will lose their first proton very readily in aqueous solution.

Sulfuric acid: $H_2SO_4(aq) \rightarrow HSO_4^-(aq) + H^+(aq)$

Hydrochloric acid: $HCl(aq) \rightarrow Cl^-(aq) + H^+(aq)$

Nitric acid: $HNO_3(aq) \rightarrow NO_3^-(aq) + H^+(aq)$

For example, in a 1.00 M solution of HCl, well over 99.9% of HCl molecules will have ionised completely into $Cl^-(aq)$ and $H^+(aq)$ ions. For most purposes, zero molecules of un-ionised HCl remain in solution and so we can use a complete, unidirectional arrow (→) to show that the reaction goes to completion. The complete ionisation of HCl is depicted in figure 5.

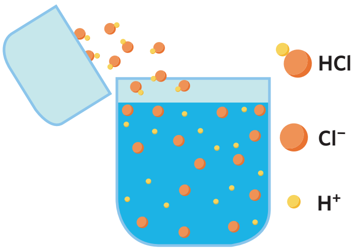

Figure 5 Solutions of HCl(aq) will be almost completely composed of just $H^+(aq)$ and $Cl^-(aq)$ ions.

A weak acid, on the other hand, is one that does not readily give up its H^+ ion in solution and only partially ionises. The ionisation equations for three common weak acids are given.

Ethanoic acid: $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$

Carbonic acid: $H_2CO_3(aq) \rightleftharpoons HCO_3^-(aq) + H^+(aq)$

Ammonium: $NH_4^+(aq) \rightleftharpoons NH_3(aq) + H^+(aq)$

None of these acids will ionise in water to a significant extent and so they are classified as weak acids. We can use the bidirectional arrows (⇌) to show that these reactions do not go to completion. For example, only about 0.5% of the CH_3COOH molecules in a 1.0 M solution will ionise into CH_3COO^- and H^+ ions (1 out of every 200). The rest will simply remain as aqueous CH_3COOH molecules. The partial ionisation of $CH_3COOH(aq)$ is depicted in figure 6.

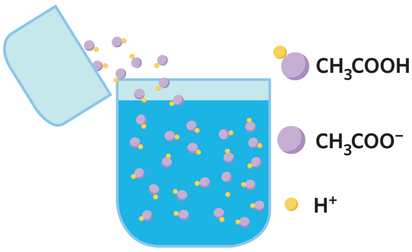

Figure 6 Ethanoic acid in aqueous solution will contain mostly $CH_3COOH(aq)$ molecules, with very few $CH_3COO^-(aq)$ and $H^+(aq)$ ions.

Why are some bases classified as either strong or weak?

The rules for classifying strong and weak bases are very similar to those used for acids, except that a base's strength depends on how readily it will accept a H^+ ion from an acid.

Examples of strong bases include the oxide ion, O^{2-}, the hydroxide ion, OH^-, and hydrogen phosphate, HPO_4^{2-}. All of these species will readily accept a H^+ ion during an acid-base reaction. The reaction of O^{2-} in water is given in figure 7. This reaction will reach completion with a majority of O^{2-} ions having accepted a H^+ ion to form a OH^- ion. A second OH^- ion is formed as $H_2O(l)$ donates its proton.

$O^{2-}(aq) + H_2O(l) \rightarrow 2OH^-(aq)$

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USEFUL TIP

Organic acids contain carbon and are generally weak acids, whilst strong acids usually do not contain carbon and are called inorganic or mineral acids.

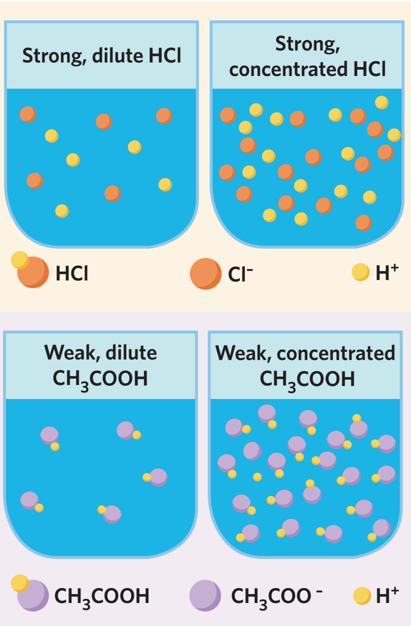

Figure 8 The differences between strong, weak, concentrated and dilute acids.

USEFUL TIP

Adjectives such as 'concentrated', 'dilute', 'strong' and 'weak' are all qualitative descriptions of an acid. Precise measurements of concentration with units such as mol L^{-1} (M) or g L^{-1} would be quantitative descriptions of concentration. In VCE chemistry, quantitative measurements of an acid's strength are not required.

Since the hydroxide ion, OH^-, is such a strong base, any compound that more or less completely **dissociates** (separates into smaller particles) to form OH^- ions will create a basic solution. Ionic compounds such as NaOH, KOH and $Ca(OH)_2$ all function as strong bases since they readily dissociate to form OH^- ions. However, according to Brønsted-Lowry theory, these ionic compounds are not technically bases themselves, since they do not accept any H^+ ions directly. They instead act as a source of the strong base OH^-.

Weak bases, on the other hand, do not readily accept H^+ ions. An example of the reaction between the weak base ammonia, NH_3, and water is given below. Again, we use bidirectional arrows to show that this reaction is only partially ionised.

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

We classify ammonia as a weak base as only a small proportion of the NH_3 molecules present in the aqueous solution will accept a H^+ ion to form NH_4^+. This means that only a small fraction of the molecules present in the solution will be hydroxide ions, OH^-, which is an indicator of how basic the resulting solution is. Most of the molecules in solution will remain as ammonia, NH_3.

Figure 7 shows the issues with distributions of ions that would be present in solutions of the strong base $O^{2-}(aq)$ and the weak base $NH_3(aq)$.

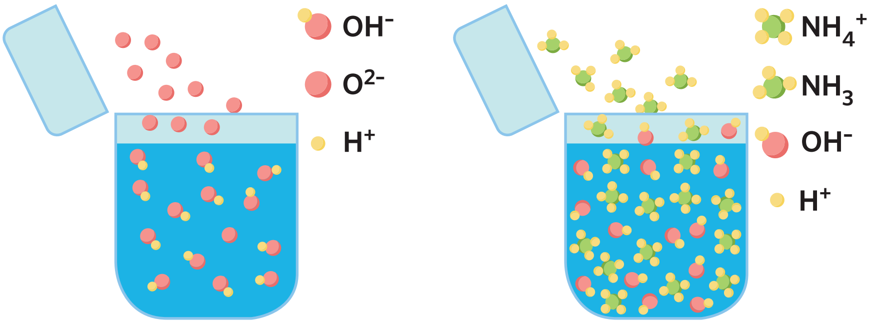

Figure 7 Solutions of the strong base $O^{2-}(aq)$ and the weak base $NH_3(aq)$ will have different proportions of $OH^-(aq)$ ions present.

What is the difference between strength and concentration?

It is important to note the difference between the concentration of an acid or base in solution and its strength. The strength of an acid relates to what proportion of its molecules will donate a H^+ ion. Its concentration depends on how many of those acid molecules there were to begin with.

In the case of hydrochloric acid, if we use a small amount of HCl, we will end up with a dilute solution. If we use a large amount, then we will have a concentrated solution. However, both of these solutions will be solutions of the strong acid HCl.

In the case of ethanoic acid, if we use a small amount of CH_3COOH, we will still end up with a dilute solution. If we use a large amount, then we will still have a concentrated solution. The only difference is that in both these cases, we will end up with only a small percentage of $CH_3COO^-(aq)$ and $H^+(aq)$ ions. However, both of these solutions will be solutions of the weak acid CH_3COOH.

Adding or removing acid molecules from solution will only change its concentration, not its strength, since an acid or base's strength depends only on the identity of the acid or base itself. Figure 8 shows common acids and solutions that would be classified as some combination of strong or weak and concentrated or dilute.

Progress questions

Question 7

Which of the following is a weak acid?

A. Nitric acid, HNO_3

B. Ammonium, NH_4^+

Continues →

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Question 8

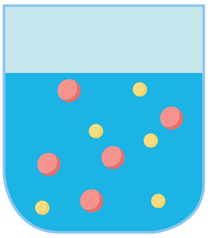
As a strong acid, a 1.0 M solution of HCl will contain

A. almost zero HCl molecules.

B. almost zero H^+ and Cl^- ions.

Question 9

The given diagram most likely depicts a



A. weak, dilute acid.

B. strong, concentrated acid.

C. weak, concentrated acid.

D. strong, dilute acid.

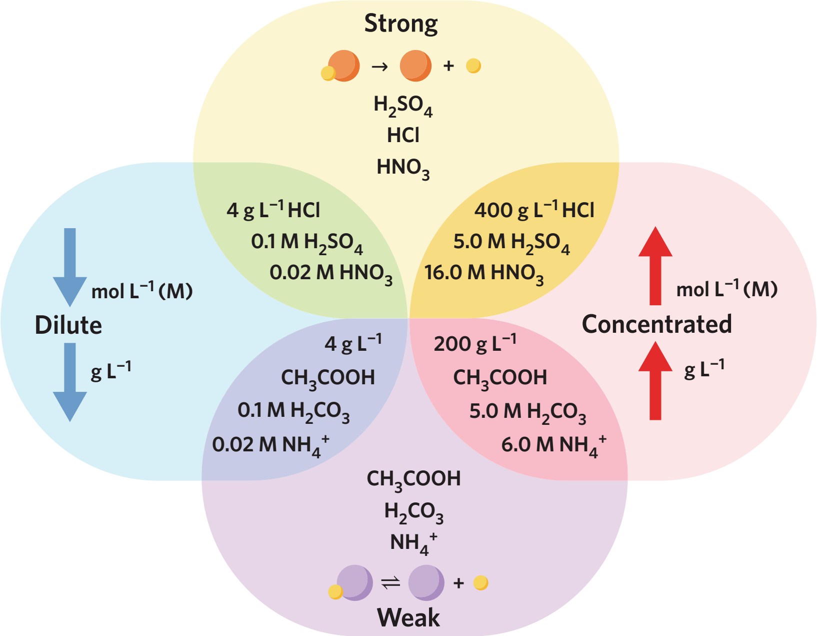
MISCONCEPTION

'Strength and concentration mean the same thing.'

Strength depends on what type of acid is present. Concentration depends on how much of that acid is present in a given volume. It is possible to have a dilute solution of a strong acid or a concentrated solution of a weak acid.

Theory summary

- According to the Brønsted-Lowry theory of acids and bases:
  - Acids donate H^+ ions.
  - Bases accept H^+ ions.
  - Acids and bases are corrosive.
- The hydronium ion, $H_3O^+(aq)$, and $H^+(aq)$ can be used interchangeably for acid-base reactions to represent an aqueous proton.
- Classification of acids and bases:
  - Amphiprotic species are capable of accepting or donating a H^+ ion.
  - Monoprotic acids can donate 1 proton.
  - Polyprotic acids can donate multiple protons.
  - Diprotic acids can donate 2 protons.
  - Triprotic acids can donate 3 protons.
- Each subsequent deprotonation step becomes harder and harder for polyprotic acids.
- Strong acids and bases ionise completely in water.
- Weak acids and bases ionise only partially in water.



The content in this lesson is considered fundamental prior knowledge to the laboratory analysis of organic compounds (Unit 4 AOS 2).

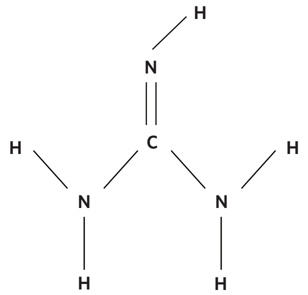
10A Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 10-12.

The structural formula of guanidine is given below. It is a strong base often used in the manufacturing of plastic and explosives. Guanidine is capable of accepting one H^+ ion during acid-base reactions.



Question 10 (1 MARK) 🌶

Which of the following gives the molecular formula of guanidine?

A. $HNC(NH_2)_2$

B. $C(NH_2)_4$

C. $HNC(NH)_2$

D. $HC(NH_2)_2$

Question 11 (1 MARK) 🌶🌶

Consider an aqueous solution of guanidine. Which of the following species would most likely be present in high concentrations in this solution?

I. OH^-

II. H_3O^+

III. $NC(NH_2)_2^-$

IV. $C(NH_2)_3^+$

V. $HNC(NH_2)_2$

A. I, III, V

B. I, IV

C. II, III, V

D. I, IV, V

Question 12 (5 MARKS) 🌶🌶🌶

Guanidine is highly soluble in water because of its multiple highly electronegative nitrogen atoms.

a. Give the balanced equation for the acid-base reaction between guanidine and water. (1 MARK)

b. Which chemical species in this reaction is acting as the acid? Justify your answer with reference to the Brønsted-Lowry theory of acids and bases. (2 MARKS)

c. Could the species from your answer to part b be considered as amphiprotic? Explain your choice. (2 MARKS)

Exam-style

Question 13 (1 MARK) 🌶

0.03 M H_2SO_4 is an example of a

A. weak, dilute acid.

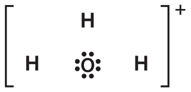
B. strong, concentrated acid.

C. weak, concentrated acid.

D. strong, dilute acid.

Use the following information to answer questions 14–15.

‘Perhaps it should be mentioned that H^+ cations are unique in having no electrons. Their small size makes their charge density very high. The H^+ ion is a proton, and free protons do not really exist in water solution. Although often written as H^+ for the sake of simplicity, the hydrogen ions in acid solutions are actually attached to molecules of the solvent. One way to indicate this in equations is by writing $H^+(aq)$ to signify the ‘aqueous’ solvation of the ion. Another way is by using the hydronium ion, H_3O^+, which is a proton attached to a molecule of water.



Since liquid water molecules exist in hydrogen bonded clusters, with about four H_2O units in an average group at room temperature, it may be that the actual formula for the aquated hydrogen ion is closer to $H_9O_4^+$ (although it is rarely written this way). In any case, that fact does not lessen the usefulness of the hydronium ion, which remains the simplest and most practical way to indicate the covalent attachment of H^+ ions to water.’

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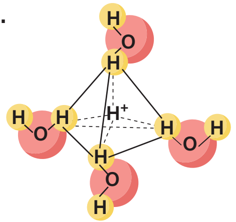
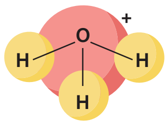
Question 14 (1 MARK)

Using the information presented in the article and your knowledge of hydrogen bonding and electronegativities, which of the following best represents the true structure of a solvated proton in aqueous solution?

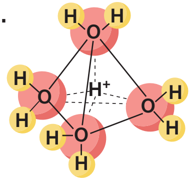
A.



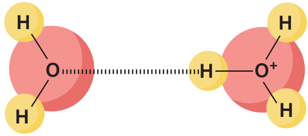
B.



C.



D.



Question 15 (1 MARK)

Which of the representations of an aquated proton, $H^+(aq)$, does the article suggest is the most suitable?

A. H^+ because it is the simplest and most efficient way of showing protons in solution.

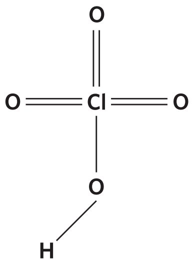
B. $H_9O_4^+$ because it represents the true and most accurate structure of an aquated proton.

C. H_3O^+ because it offers a structure that represents the dispersion of the proton’s charge.

D. H_3O^+ because it conveniently displays an aquated proton’s covalent bond with the solvent molecules.

Question 16 (2 MARKS)

Perchloric acid, $HClO_4$, is a considerably stronger acid than both sulfuric acid and nitric acid. For these reasons, it is actually classified as a ‘super acid’. In high concentrations, it is dangerously corrosive and is used in the synthesis of rocket fuel. Its structure is shown.



a. Is perchloric acid a monoprotic, polyprotic or amphiprotic acid? (1 MARK)

b. Give the balanced equation for the acid-base reaction between $HClO_4$ and the alkali NH_3 in aqueous solution. (1 MARK)

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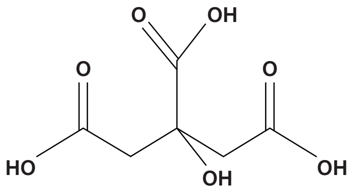
Use the following information to answer questions 17–18.

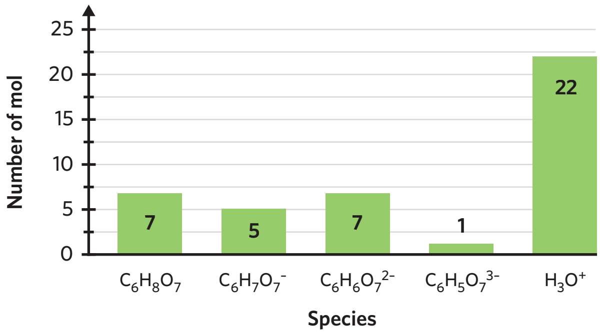
Citric acid, $C_6H_8O_7$ is a relatively weak triprotic acid that is found naturally in citrus fruits. The structure of citric acid is given. 20 mol of citric acid is added to a vessel of water and left for a few minutes until no further activity is observed.



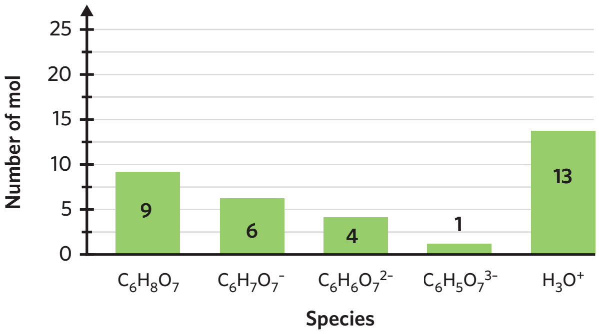
Question 17 (1 MARK)

Which of the following charts would represent the most likely distribution of aqueous species present in the resulting solution?

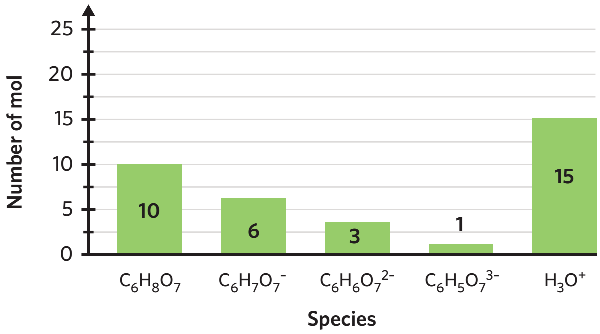
A.



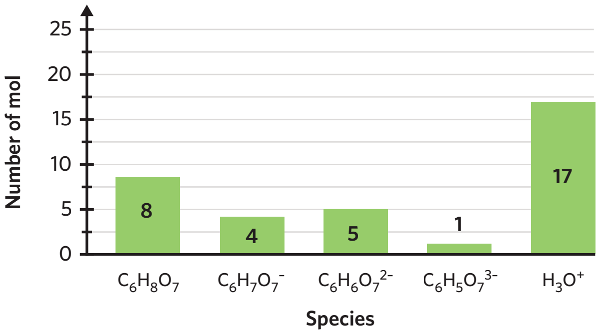
B.



C.



D.



Question 18 (1 MARK)

Which of the following best explains the distribution of the ions present in the citric acid solution?

A. Polyprotic acids deprotonate randomly and spontaneously, leading to a random distribution of ions.

B. With each deprotonation, it becomes much harder for the positive H^+ ion to escape the pull of the remaining compound's negative charge, which leads to a progressively decreasing distribution of ions.

C. As each proton is lost to form a H_3O^+ ion, these positive charges electrostatically resist and repel any new positive H^+ ions from entering the solution, leading to a decreasing distribution of fully deprotonated species.

D. Deprotonation results in the destabilisation of polyprotic acids and so they resist deprotonation to maintain this stability, leading to the species in highest concentration being the complete molecule.

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Key science skills

Question 19 (16 MARKS)

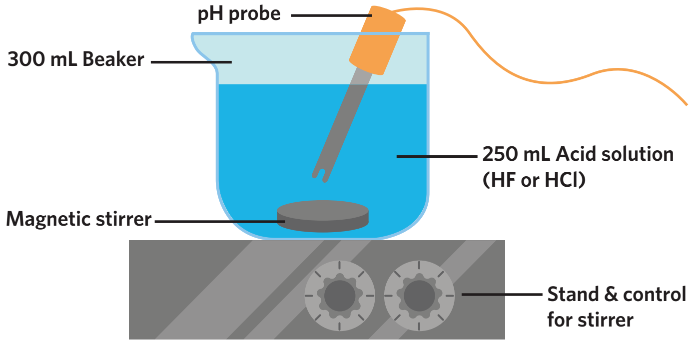
Caoimhe is a chemistry student interested in investigating the relative strengths of hydrofluoric acid, HF, and hydrochloric acid, HCl. Various sections of her experimental report are shown.

Aim

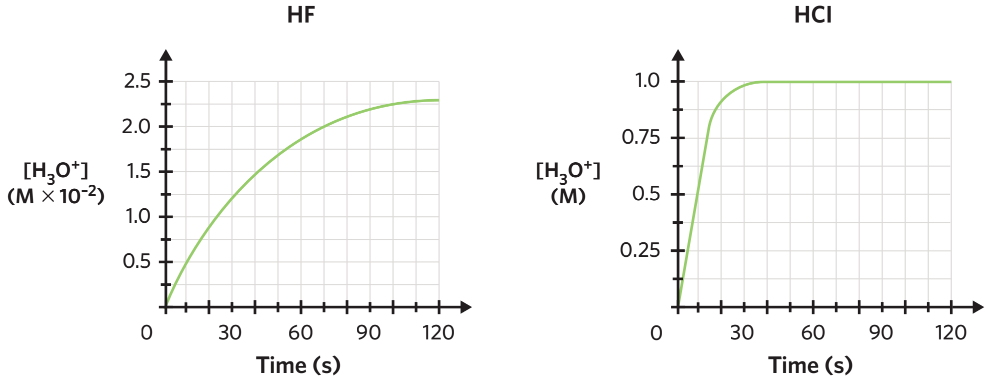
To determine the relative strengths of HF and HCl in aqueous solution by measuring their extent of ionisation over time.

Method

1. Two 300 mL beakers were rinsed and 250 mL of deionised water was added to each.
2. A pH probe was added to each beaker to measure the concentration of $H_3O^+(aq)$ ions in the solution.
3. 5.00 g of HF(s) was weighed in a pressurised container and then added to one of the beakers.
4. 9.125 g of HCl(s) was weighed in a pressurised container and then added to the second beaker.
5. The solutions were stirred using an automatic magnetic stirring bead until all solid had dissolved.
6. The results from the pH probe were recorded over 2 minutes.
7. Steps 1–6 were repeated 4 times and all results were averaged.



Averaged Results



a. Caoimhe wishes to ensure that she is following correct experimental procedures.

- **i.** Identify the independent and dependent variables in this experiment. (2 MARKS)
- **ii.** Identify three variables that should be controlled in this experiment. (1 MARK)

b. One of Caoimhe's classmates analysing her method is concerned that she has made a number of errors.

- **i.** Why were the masses of HF and HCl used in the experiment different? Justify your answer with the use of a calculation. (2 MARKS)
- **ii.** What is the purpose of step 7 in the method? What impact, if any, does this have on the accuracy of the results obtained? (2 MARKS)

c. How can Caoimhe use the data she has collected to determine each acid's strength? Is hydrofluoric acid or hydrochloric acid stronger? Justify your answer with reference to the data collected. (3 MARKS)

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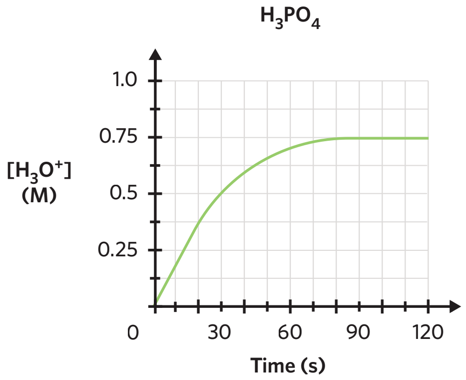
d. While doing research about the risks associated with her experiment, Caoimhe reads the following statement:

Hydrofluoric acid is corrosive to tissue due to its proton and is corrosive to bones because of its fluoride. HF also interferes with nerve function, meaning that low concentration burns may not be noticed until a significant portion of tissue and bone has been dissolved. This makes it one of the most dangerous inorganic acids. HF corrodes glass, whereas HCl does not. This is because silicon tetrafluoride exists as a more stable complex than the unstable silicon tetrachloride."

Identify two safety measures that Caoimhe should put in place to ensure safe laboratory practice. (2 MARKS)

e. Caoimhe now repeats the experiment but this time she uses an appropriate amount of phosphoric acid. Her results are given.



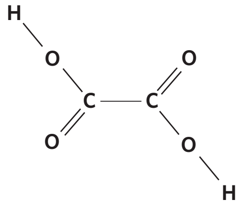
Is this a valid addition to the experimental results if she intends to compare the strengths of HCl, HF and H_3PO_4? Explain. (4 MARKS)

FROM LESSONS 16A, 16C & 16D

Questions from multiple lessons

Question 20 (11 MARKS)

Oxalic acid is the simplest possible dicarboxylic acid, which means that it contains two carboxyl functional groups. It has the semi-structural formula HOOCCOOH or $(COOH)_2$. Oxalic acid is a naturally occurring compound found in the flowers of the wood-sorrel or *Oxalis* genus. The structural formula of oxalic acid is shown.



a. Sketch a diagram of the hydrogen bonding interaction that would occur between two molecules of oxalic acid. (2 MARKS)

b. Consider a 5.0 M solution of $(COOH)_2$(aq). The percentage composition by mass of the solution is shown in the given table.

Species	% composition by mass
$(COOH)_2$	25.7
$(COOH)(COO)^-$	4.30
$(COO)_2^{2-}$	0.912
H_3O^+	1.29

i. Why does the '% composition by mass' column not sum to a total of 100%? (1 MARK)

ii. Would oxalic best be classified as a strong or weak acid? Justify your answer with reference to the data presented in the table. (3 MARKS)

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c. Oxalic acid will react with the strong base $OH^-(aq)$ in aqueous solution according to the following balanced reaction.

$$(COOH)_2(aq) + 2OH^-(aq) \rightarrow (COO)_2^{2-}(aq) + 2H_2O(l)$$

If 11.5 mol of oxalic acid reacts completely with $OH^-(aq)$ ions, calculate

i. the number of molecules of $(COOH)_2(aq)$ that are reacting. (1 MARK)

ii. the mass, in grams, of $OH^-(aq)$ ions required for this reaction. (2 MARKS)

d. What kind of acid would oxalic acid be classified as according to its ability to donate a number of $H^+(aq)$ ions? (1 MARK)

e. The reaction given in part c actually occurs in two distinct stages. The equations for these reactions are given.

$$(COOH)_2(aq) + OH^-(aq) \rightarrow (COOH)(COO)^-(aq) + H_2O(l)$$

$$(COOH)(COO)^-(aq) + OH^-(aq) \rightarrow (COO)_2^{2-}(aq) + H_2O(l)$$

There are three amphiprotic chemical species in these reactions. Give the molecular formulas of these amphiprotic species. (1 MARK)

FROM LESSONS 2B & 6B

Hints

13. Sulfuric acid, H_2SO_4, is a common strong acid.

14. The highly electronegative oxygen in H_2O will be attracted towards the highly positive H^+ cation in the $H_9O_4^+$ representation.

15. "...the hydronium ion, which remains the simplest and most practical way to indicate the covalent attachment of H^+ ions to water."

16a. Acids can be classified according to how many protons they can donate.

16b. During acid-base reactions, bases accept a proton donated by an acid.

17. It becomes more difficult for polyprotic acids to lose a proton after each deprotonation.

18. When a polyprotic acid donates a proton, it forms a negatively charged species.

19ai. Changing the independent variable affects the dependent variable.

19aii. Variables are controlled so that the experimenter knows that only their independent variable is responsible for influencing the dependent variable.

19bi. The concentration of an acidic solution will depend on the mol L^{-1}.

19bii. Random errors occur in some experiment trials, but not in all of them.

19c. An acid's strength is related to its extent of ionisation.

19d. Safety measures are intended to reduce the risk to whomever is performing the experiment.

19e. The experiment is currently set up to measure only the total amount of $H_3O^+(aq)$ ions in the solution.

20a. Hydrogen bonding involves a H atom 'flanked' by two highly electronegative atoms.

20bi. % composition by mass includes all species present in the solution.

20bii. An acid's strength depends on how readily it will donate a proton.

20ci. The number of molecules in a mole is related to Avogadro's constant.

20cii. The mass of a reactant depends on how many moles of that species are reacting.

20d. Acids are classified according to the number of protons they can donate.

20e. Amphiprotic species can both donate and accept protons.

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10B Reactions of acids and bases

STUDY DESIGN DOT POINT

- neutralisation reactions to produce salts:
  - reactions of acids with metal carbonates and hydroxides, including balanced full and ionic equations, with states
  - types of antacids and their use in the neutralisation of stomach acid

10A 10B 10C 10D 10E

2.1.6.1 Reactions of acids

2.1.6.2 Reactions of antacids

ESSENTIAL PRIOR KNOWLEDGE

10A	Base
10A	Acid
10A	Neutralise
4C	Soluble/insoluble

See questions 60–63.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: Emily Frost/Shutterstock.com

How does an antacid actually work?

Acidic and basic compounds are found in all facets of life: in food, the environment and in the body. In this lesson we will investigate how acid-base reactions are used for medicinal purposes.

KEY TERMS AND DEFINITIONS

Antacid	a weak base that is used to neutralise stomach acid
Heartburn	a burning sensation in the chest, caused by acid regurgitation into the oesophagus
Indigestion	pain or discomfort in the stomach
Metal hydroxide	compound that neutralises acids to form a salt and water
Metal carbonate	compound that neutralises acids to form a salt, water and carbon dioxide
Metal hydrogencarbonate	compound that neutralises acids to form a salt, water and carbon dioxide
Neutralisation reaction	a reaction between an acid and an alkali or base that forms a salt and water (and carbon dioxide when a metal carbonate is a reactant)
Ulcer	a break in the lining of an organ like the stomach

Reactions of acids 2.1.6.1

Acids react with **metal carbonates** to produce an ionic ‘salt’, water and carbon dioxide gas, whereas **metal hydroxides** react with acids to produce an ionic ‘salt’ and water.

MISCONCEPTION

‘All bases can give off carbon dioxide gas when they react with acids.’

Only bases that actually contain carbonates can produce carbon dioxide gas.

How do acids react with metal carbonates?

Metal carbonates are bases formed from a metal cation and carbonate ions, CO_3^{2-}. Metal carbonates can be either soluble or insoluble in water.

Examples of soluble metal carbonates include the group 1 metal carbonates: sodium carbonate, Na_2CO_3, and potassium carbonate, K_2CO_3. When added to water, a soluble metal carbonate will dissociate to form its cation and carbonate ions, as shown for Na_2CO_3.

$$Na_2CO_3(aq) \rightarrow 2Na^+(aq) + CO_3^{2-}(aq)$$

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When a metal carbonate is added to an acidic solution, the products of the reaction are an ionic salt, carbon dioxide gas, $CO_2(g)$ and liquid water, $H_2O(l)$. The general equation for the reaction between an acid and a metal carbonate is as follows.

acid + metal carbonate → ionic salt + $CO_2(g) + H_2O(l)$

For example, consider the reaction between sodium carbonate, $Na_2CO_3(aq)$, and nitric acid, HNO_3.

$2HNO_3(aq) + Na_2CO_3(aq) \rightarrow 2NaNO_3(aq) + CO_2(g) + H_2O(l)$

This reaction is represented in figure 1.

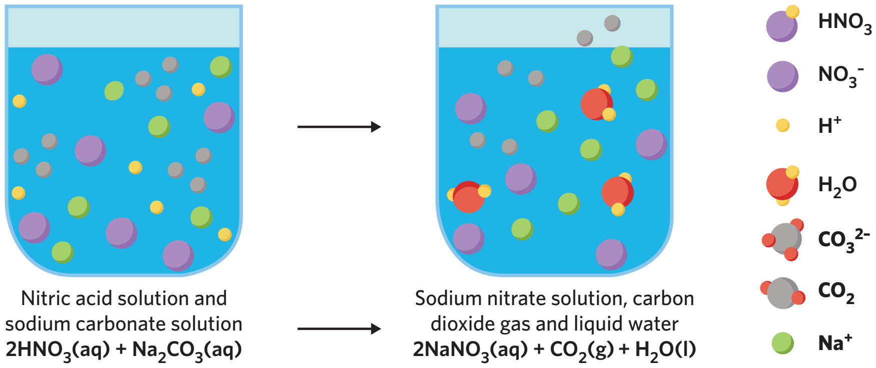


Figure 1 Nitric acid reacts with sodium carbonate to form sodium nitrate, water and carbon dioxide gas.

In order to determine the ionic equation for this reaction, we can rewrite the full equation as shown.

Full equation: $2HNO_3(aq) + Na_2CO_3(aq) \rightarrow 2NaNO_3(aq) + CO_2(g) + H_2O(l)$

Ions: $2H^+(aq) + 2NO_3^-(aq) + 2Na^+(aq) + CO_3^{2-}(aq) \rightarrow 2Na^+(aq) + 2NO_3^-(aq) + CO_2(g) + H_2O(l)$

Ionic equation: $2H^+(aq) + CO_3^{2-}(aq) \rightarrow CO_2(g) + H_2O(l)$

MISCONCEPTION

'The reactions of reactive metals and acids are a **neutralisation reaction.**'

When metals react with an acid, an ionic salt and hydrogen gas are formed; this is called a single displacement reaction. Neutralisation reactions must produce an ionic salt and water as products.

Insoluble metal carbonates also follow the same overall reaction as soluble metal carbonates. These include magnesium carbonate, $MgCO_3$, calcium carbonate, $CaCO_3$, and zinc carbonate, $ZnCO_3$. These compounds will remain in solid form when added to water. This means that the ionic equations for their reactions with acids will be different. An example of the reaction between nitric acid and solid calcium carbonate is given below.

Full equation: $2HNO_3(aq) + CaCO_3(s) \rightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$

Ions: $2H^+(aq) + 2NO_3^-(aq) + CaCO_3(s) \rightarrow Ca^{2+}(aq) + 2NO_3^-(aq) + CO_2(g) + H_2O(l)$

Ionic equation: $2H^+(aq) + CaCO_3(s) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(l)$

Metal hydrogencarbonates, or bicarbonates, behave with acids in a similar fashion to metal carbonates. Metal hydrogencarbonates are compounds containing a metal cation and the bicarbonate ion, HCO_3^-. Examples include sodium hydrogencarbonate, $NaHCO_3$, potassium hydrogencarbonate, $KHCO_3$, magnesium hydrogencarbonate, $Mg(HCO_3)_2$ and calcium hydrogencarbonate, $Ca(HCO_3)_2$. Metal hydrogencarbonates follow the same general reaction with an acid as metal carbonates and have similar ionic equations as well. This is shown in the following example for the reaction between nitric acid and potassium hydrogencarbonate.

General equation: acid + metal hydrogencarbonate → ionic salt + $CO_2(g) + H_2O(l)$

Full equation: $HNO_3(aq) + KHCO_3(s) \rightarrow KNO_3(aq) + CO_2(g) + H_2O(l)$

Ions: $H^+(aq) + NO_3^-(aq) + KHCO_3(s) \rightarrow K^+(aq) + NO_3^-(aq) + CO_2(g) + H_2O(l)$

Ionic equation: $H^+(aq) + KHCO_3(s) \rightarrow K^+(aq) + CO_2(g) + H_2O(l)$

WORKED EXAMPLE 1

Write the ionic equation for the reaction between hydrochloric acid, $HCl(aq)$ and solid magnesium carbonate, $MgCO_3(s)$, in aqueous solution.

What information is presented in the question?

The acid $HCl(aq)$ is reacting with the insoluble metal carbonate $MgCO_3(s)$ in aqueous solution (water).

What is the question asking us to do?

Write the ionic equation for this reaction.

What strategies do we need in order to answer the question?

1. Identify the general equation for an acid reacting with a metal carbonate.
2. Determine which ions in the equation are spectator ions.
3. Write the ionic equation for the reaction, remembering to omit any spectator ions.

Answer

The general equation for an acid reacting with a metal carbonate is:

acid + metal carbonate → ionic salt + $CO_2(g) + H_2O(l)$

The ionic salt $MgCl_2(aq)$ will be formed from the acid's anion ($Cl^-(aq)$) and the metal's cation ($Mg^{2+}(aq)$). Therefore, the overall balanced equation is:

$2HCl(aq) + MgCO_3(s) \rightarrow MgCl_2(aq) + CO_2(g) + H_2O(l)$

Note that $MgCO_3$ is insoluble in water and so will be represented as solid $MgCO_3(s)$. This equation can be rewritten to show all individual ions:

$2H^+(aq) + 2Cl^-(aq) + MgCO_3(s) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq) + CO_2(g) + H_2O(l)$

Chloride, $Cl^-(aq)$, is the only species that appears as an ion on both sides of the equation, making it a spectator ion. As a spectator ion, it can be omitted from the balanced ionic equation:

$2H^+(aq) + MgCO_3(s) \rightarrow Mg^{2+}(aq) + CO_2(g) + H_2O(l)$

USEFUL TIP

Remember that, in solution, ionic compounds such as $KNO_3(aq)$ exist as aqueous forms of anions and cations. If we were to evaporate the solution's water, however, we would be left with a solid $KNO_3(s)$ residue.

How do acids react with metal hydroxides?

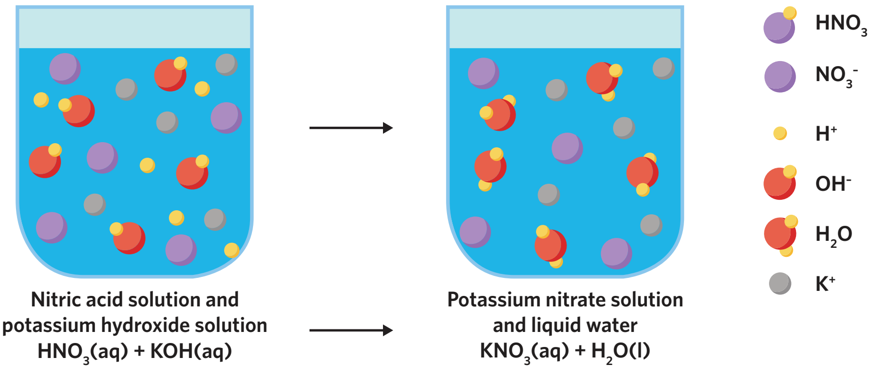
The second reaction we will consider is the reaction between a metal hydroxide and an acid. Metal hydroxides are ionic compounds formed from a metal cation and hydroxide ions, OH^-. Examples of metal hydroxides include sodium hydroxide, $NaOH$, barium hydroxide, $Ba(OH)_2$, and potassium hydroxide, KOH. When added to aqueous solution, a soluble metal hydroxide will dissociate into its cation and hydroxide ions, as shown for KOH.

$KOH(aq) \rightarrow K^+(aq) + OH^-(aq)$

When an acidic solution is added to a metal hydroxide, the products of the reaction are an ionic salt and liquid water, $H_2O(l)$, giving the following general reaction:
acid + metal hydroxide → ionic salt + $H_2O(l)$

For example, consider the full reaction between potassium hydroxide, KOH, and nitric acid, HNO_3.
$HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$

Once again, the ionic salt will be composed of the metal hydroxide's cation ($K^+(aq)$) and the conjugate base/anion from the acid ($NO_3^-(aq)$), resulting in potassium nitrate, $KNO_3(aq)$. This reaction is shown in figure 2.



Nitric acid solution and potassium hydroxide solution
$HNO_3(aq) + KOH(aq)$

Potassium nitrate solution and liquid water
$KNO_3(aq) + H_2O(l)$

Figure 2 Nitric acid reacts with potassium hydroxide to form soluble potassium nitrate and liquid water.

In order to determine the ionic equation for this reaction, we can rewrite the full equation as shown. Nitric acid is completely ionised into $H^+(aq)$ and $NO_3^-(aq)$ and potassium nitrate exists as $K^+(aq)$ and $NO_3^-(aq)$ ions in solution.

Full equation: $HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$

Ions: $H^+(aq) + NO_3^-(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + NO_3^-(aq) + H_2O(l)$

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Since $K^+(aq)$ and $NO_3^-(aq)$ exist as spectator ions on both sides of the equation, we can eliminate them both. Water is a molecular compound and so must be present in the ionic equation. This leaves us with the following balanced ionic equation:
Ionic equation: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

When a reaction between an acid and a soluble metal hydroxide is shown as an ionic equation, we can see that $H^+(aq)$ ions from the acid and $OH^-(aq)$ ions from the base are reacting together to form water, $H_2O(l)$. This is known as a neutralisation reaction, as the acid and base are effectively removing each other from the solution and forming pure water, which has a neutral pH. When all of the basic $OH^-(aq)$ ions have reacted with the acidic $H^+(aq)$ ions to form water, the solution is said to be 'neutralised'.

USEFUL TIP

If the acidic hydronium ions in the given neutralisation reaction are expressed as $H_3O^+(aq)$ instead of $H^+(aq)$, the balanced ionic equation would be written as $H_3O^+(aq) + OH^-(aq) \rightarrow 2H_2O(l)$.

WORKED EXAMPLE 2

Write the full equation for the reaction between sulfuric acid, $H_2SO_4(aq)$, and sodium hydroxide, $NaOH(aq)$, in aqueous solution.

What information is presented in the question?

The acid $H_2SO_4(aq)$ is reacting with the metal hydroxide $NaOH(aq)$ in aqueous solution (water).

What is the question asking us to do?

Write the full equation for this reaction.

What strategies do we need in order to answer the question?

1. Identify the general equation for an acid reacting with a metal hydroxide.
2. Determine which ions will contribute to form the ionic salt.
3. Write the full (overall) equation for the reaction.
4. Balance the equation with respect to atoms and charges.

Answer

The general equation for an acid reacting with a metal hydroxide is:

acid + metal hydroxide → ionic salt + $H_2O(l)$

The ionic salt in this reaction will be formed from the conjugate base/anion of the acid (SO_4^{2-}) and the cation from the metal hydroxide (Na^+). Therefore, the overall (unbalanced) equation is:

$H_2SO_4(aq) + NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$

The Na and H must then be balanced, producing the overall balanced equation:

$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

Progress questions

Question 1

An ionic equation does not include species present only as aqueous ions in a chemical reaction. These species are known as ____________.

A. background ions

B. spectator ions

Question 2

Water molecules, $H_2O(l)$, are never included in ionic equations.

A. True

B. False

Question 3

Which of the following is an ionic salt?

A. $NaNO_3$

B. CH_4

Continues →

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Question 4

When an acid reacts with a metal hydroxide, what are the products?

A. $H^+(aq)$, $OH^-(aq)$

B. ionic salt, $H_2O(l)$

Question 5

Which of the following is a reaction between an acid and a metal carbonate?

A. $2HCl(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$

B. $H_3PO_4(aq) + 3KOH(aq) \rightarrow K_3PO_4(aq) + 3H_2O(l)$

Question 6

What is the ionic equation for the reaction between an acid and a metal hydroxide?

A. $2HNO_3(aq) + Ba(OH)_2(aq) \rightarrow Ba(NO_3)_2(aq) + 2H_2O(l)$

B. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

Reactions of antacids 2.1.6.2

Antacids[1] are a class of medicines that are commonly used to relieve the discomfort of excess hydrochloric acid, HCl, in the stomach that can cause **ulcers**, **heartburn** and **indigestion**.

KEEN TO INVESTIGATE?

[1] **How do indigestion medications work?**
Search YouTube: How Do Reflux Medications Work?

How do different types of antacids neutralise stomach acid?

Antacids are weak bases that are used to treat the symptoms of excess concentrated hydrochloric acid produced in the stomach. Strong bases cannot be used as they will corrode the internal linings within the body and cause severe permanent damage. The different types of compounds are commonly used as a remedy for neutralising excess stomach acid are shown in table 1.

Table 1 Neutralisation reactions of common antacids

Name of active ingredient	Reaction equation	Action and side effect
Magnesium hydroxide	$Mg(OH)_2(l) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$	Commonly taken as a liquid suspension gel and is very fast acting. Acts as a laxative.
Aluminium hydroxide	$Al(OH)_3(l) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$	Commonly taken as a liquid suspension gel and is slower acting. Can cause constipation.
Sodium hydrogencarbonate (also called sodium bicarbonate)	$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$	Commonly taken as a soluble tablet and is very fast acting. Bloating, burping and flatulence.
Calcium carbonate	$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$	Commonly taken as a chewable tablet and is very fast acting. Bloating, burping and flatulence.

Progress questions

Question 7

What type of reaction occurs in the stomach when an antacid is taken?

A. Displacement

B. Neutralisation

Continues →

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Question 8

Most antacids have possible side-effects.

A. True

B. False

Theory summary

- The general equations for acids with different compounds:
  - acid + metal hydroxide → ionic salt + $H_2O(l)$
  - acid + metal carbonate → ionic salt + $CO_2(g)$ + $H_2O(l)$
  - acid + metal hydrogencarbonate → ionic salt + $CO_2(g)$ + $H_2O(l)$
- These reactions can all be represented as either overall/full equations or ionic equations.
- A range of different weak bases can be used to neutralise stomach acid (HCl) and are a class of medicines known as antacids.

The content in this lesson is considered fundamental prior knowledge to the laboratory analysis of organic compounds (Unit 4 AOS 2).

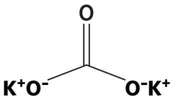
10B Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 9–11.

Potassium carbonate, K_2CO_3, is a highly soluble white salt. Its solubility is so high that over 100 g of potassium carbonate will dissolve completely in only 100 mL of water. The structure of K_2CO_3 is shown. Consider what will happen if potassium carbonate reacts with the strong acid hydroiodic acid, HI.



Question 9 (1 MARK) 🌶

What is the general equation for the reaction between hydroiodic acid and a metal carbonate?

A. $HI(aq)$ + metal carbonate → ionic salt + $CO_2(g)$ + $H_3O^+(aq)$

B. $2HI(aq)$ + metal carbonate → ionic salt + $CO_2(g)$ + $H_2O(l)$

C. $2HI(aq)$ + metal carbonate → cation + $CO_2(g)$ + $H_2O(l)$

D. $HI(aq)$ + metal carbonate → anionic salt + $NO_2(g)$ + $OH^-(l)$

Question 10 (1 MARK) 🌶

Which of the following is the correct balanced equation for the overall reaction between hydroiodic acid and potassium carbonate?

A. $2HI(aq) + K_2CO_3(aq) \rightarrow 2HK(aq) + IO_3(s) + C(s)$

B. $HI(aq) + K_2CO_3(aq) \rightarrow KI(aq) + CO_2(g) + H_2O(l)$

C. $HI(aq) + K_2CO_3(s) \rightarrow 2KI(aq) + HCO_3^-(aq)$

D. $2HI(aq) + K_2CO_3(aq) \rightarrow 2KI(aq) + CO_2(g) + H_2O(l)$

Question 11 (2 MARKS) 🌶🌶

The reaction between an acid and a metal carbonate can also be represented as an ionic equation.

a. Identify the spectator ion(s) from the overall reaction between $HI(aq)$ and $K_2CO_3(aq)$. (1 MARK)

b. Write the balanced ionic equation for the reaction between $HI(aq)$ and $K_2CO_3(aq)$. (1 MARK)

Exam-style

Question 12 (1 MARK)

Which of the following reactions would have an ionic equation that is identical to its overall equation?

A. $3H^+(aq) + 2PO_4^{3-}(aq) + Ca(s) \rightarrow Ca_3(PO_4)_2(aq) + H_2(g)$

B. $CdSO_4(aq) + K_2S(aq) \rightarrow CdS(s) + 2K^+(aq) + SO_4^{2-}(aq)$

C. $HCl(aq) + Na^+(aq) + HCO_3^-(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$

D. $Fe^{3+}(l) + 2Al(s) \rightarrow 2Fe(s) + Al^{3+}(l)$

Question 13 (1 MARK)

Which of the following is the balanced ionic equation for the reaction between solid zinc and chloric acid, $HClO_3(aq)$?

A. $HClO_3(aq) + Zn(s) \rightarrow Zn(ClO_3)_2(aq) + H_2(g)$

B. $H^+(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + H_2(g)$

C. $2H^+(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + H_2(g)$

D. Solid zinc would not react with chloric acid.

Question 14 (5 MARKS)

Ethanoic acid, CH_3COOH, also known as acetic acid, is the acid found in household vinegar. A solution of ethanoic acid is added to a 40.0 L reaction vessel containing 2.0 M sodium hydroxide.

a. Write the balanced overall equation for the reaction that would occur. (1 MARK)

b. Write the balanced ionic equation for this reaction. (1 MARK)

c. Explain your reasoning for the omission or inclusion of chemical species in the ionic equation in part b. (2 MARKS)

d. What is the name given to this kind of reaction? (1 MARK)

Question 15 (6 MARKS)

Milk of magnesia is a gel-like liquid that is insoluble in water and is used to treat indigestion. It contains magnesium hydroxide which reacts with excess hydrochloric acid in the stomach.

a. Write the following types of equations to represent the reaction of milk of magnesia and stomach acid – no state symbols are required: (2 MARKS)

- Word
- Symbol

b. A second indigestion remedy contains limestone, another name for calcium carbonate. Name the gas produced when calcium carbonate reacts with hydrochloric acid and suggest a possible side-effect of using this type of antacid. (2 MARKS)

c. Why are compounds like calcium hydroxide and sodium hydroxide not used as a treatment for indigestion? (2 MARKS)

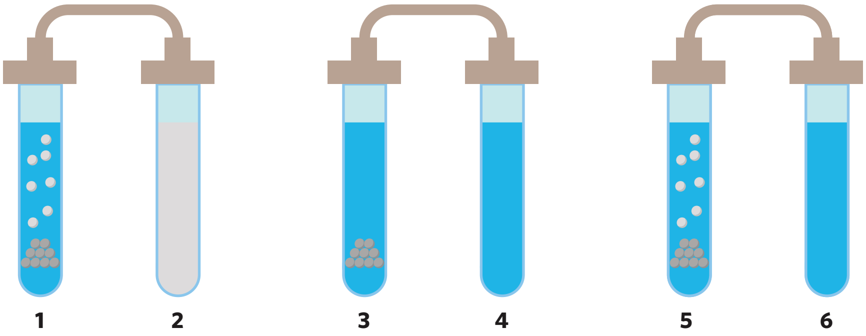
Key science skills

Question 16 (5 MARKS)

The limewater test is a test used by chemists to determine the presence of carbon dioxide gas, $CO_2(g)$. Limewater is a saturated solution of calcium hydroxide, $Ca(OH)_2(aq)$, that will transition to a white 'milky' or 'cloudy' solution if $CO_2(g)$ is present. This is due to the precipitation of insoluble calcium carbonate, $CaCO_3(s)$, that becomes suspended in the solution.

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

A chemist has mixed up the labels of three of her metal compounds - $Pb(s)$, $MgCO_3(s)$ and $Zn(OH)_2(s)$ – and so she sets up an experiment to determine their identities. To accomplish this, she prepared the three sets of test tubes as shown.



The contents of each test tube are shown in the given table.

Test tube	Contents	
1	Unknown compound 1	$HNO_3(aq)$
2	$Ca(OH)_2(aq)$	
3	Unknown compound 2	$HNO_3(aq)$
4	$Ca(OH)_2(aq)$	
5	Unknown compound 3	$HNO_3(aq)$
6	$Ca(OH)_2(aq)$	

a. Determine which test tube contains $MgCO_3(s)$. Justify your answer with reference to the data collected. (3 MARKS)

b. Is the data collected best classified as quantitative or qualitative? Explain. (2 MARKS)

FROM LESSON 16B

Question 17 (2 MARKS)

Fatima decided to do some research on what the most effective antacid would be to take for her indigestion. She found the data shown in the table from a scientific journal.

Product	Standard dose to neutralise a sample of stomach acid
Alka-Seltzer (tablet)	15 tablets
Amphogel (liquid)	25 tsp
Amphogel (tablet)	17 tablets
Maalox (liquid)	11 tsp
Maalox Therapeutic Concentrate (liquid)	5 tsp
Mylanta (liquid)	12 tsp
Mylanta II (liquid)	6 tsp
Mylanta II (tablet)	7 tablets
Titralac (liquid)	8 tsp
Tums (tablet)	16 tablets

What conclusion can she make about which is the most effective antacid in terms of the amount of product required to neutralise a fixed amount of stomach acid, and is this conclusion valid? Explain your reasoning.

FROM LESSON 16D

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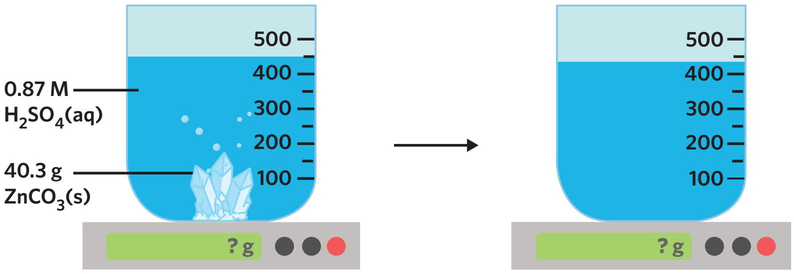
Questions from multiple lessons

Question 18 (6 MARKS)

Smithsonite is a mineral ore of zinc, with the molecular formula $ZnCO_3$.

Although insoluble in water, smithsonite can be broken down by the addition of concentrated sulfuric acid in aqueous solution. Consider the following 500 mL reaction vessel:

Image: Reload Design/Shutterstock.com



a. Write the balanced equation for the overall reaction occurring in the beaker. (1 MARK)

b. What is the balanced ionic equation for this reaction? (1 MARK)

c. The chemist investigating this reaction noticed that the water level in the beaker changed after the reaction had finished. Explain this observation. (2 MARKS)

d. The chemist weighed the reaction vessel before and after the reaction had finished. Calculate the change in mass that would be observed. (2 MARKS)

FROM LESSON 4B

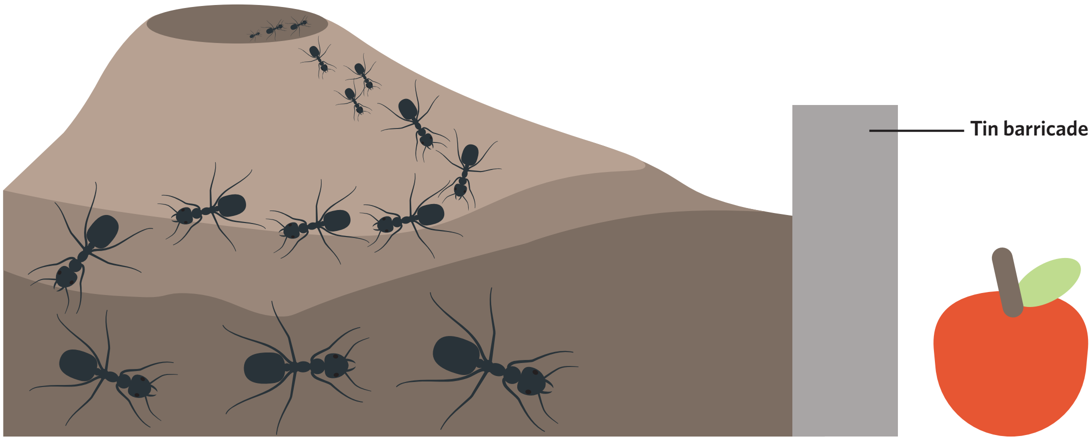
Question 19 (10 MARKS)

Methanoic acid, also known as formic acid, is used in the defence mechanism of many different ant species to protect their colonies and subdue their prey. It is a colourless monoprotic organic acid that exists in a highly soluble, aqueous form in the ant's abdominal glands.

a. Draw the structural formula of methanoic acid. (1 MARK)

b. What is the name of the functional group present in methanoic acid? (1 MARK)

c. Consider the following scenario: a number of ants are on a scouting mission from their colony. On their journey, they encounter a barricade made out of the metal tin (Sn). These intrepid explorer ants identify the tin barricade as an enemy that is preventing them from getting to a nearby apple.



In order to defend themselves and acquire the apple, the ants spray methanoic acid onto the tin barricade to break it down.

i. Write the balanced overall equation for the reaction that would occur between the methanoic acid from the ants and the tin barricade. (1 MARK)

ii. Write the balanced ionic equation for this reaction. (1 MARK)

d. Each ant can produce an average of 0.600 μg (1 μg = 1×10^{-6} g) of formic acid. If the tin barricade whas a mass of 40.0 g, how many ants would be needed to completely break it down and reach the apple? (4 MARKS)

e. How many molecules of $H_2(g)$ would be produced in the reaction from part d? (2 MARKS)

FROM LESSONS 4B, 6B & 7D

Question 20 (6 MARKS)

The human stomach contains cells that produce hydrochloric acid, HCl. The average human stomach has a volume of 1.20 L and a concentration of hydrochloric acid equal to 3.2×10^{-4} M. Although stomach tissue has evolved to handle these highly acidic conditions, the oesophagus has not. If acid leaves the stomach and comes into contact with this unprotected tissue, the result is acid reflux (heartburn).

a. Is hydrochloric acid an amphiprotic, monoprotic or polyprotic acid? (1 MARK)

b. Antacids are tablets taken to reduce the effects of acid reflux. A common antacid is aluminium hydroxide, $Al(OH)_3$, which comes in 250 mg tablets. Aluminium hydroxide is insoluble in water but will react with the hydrochloric acid in the stomach.

- **i.** Give the balanced overall equation for the reaction between aluminium hydroxide and hydrochloric acid. (1 MARK)
- **ii.** Give the balanced ionic equation for this reaction. (1 MARK)

c. If a patient suffering from a severe case of acid reflux requires 0.05 mol of $Al(OH)_3$ to return their stomach to average acidity, how many antacid tablets should they take? (3 MARKS)

FROM LESSONS 3B, 6B & 10A

Hints

12. Ionic equations do not include spectator ions.

13. Zinc is a reactive metal that will react with chloric acid to produce an ionic salt from its cation (Zn^{2+}) and chlorate ion ($ClO_3^-(aq)$) and hydrogen gas, $H_2(g)$.

14a. acid + metal hydroxide → ionic salt + water

14b. Ionic equations do not include spectator ions.

14c. Spectator ions are chemical species that are present as ions on both sides of a chemical equation.

14d. Water has a neutral pH.

15a. Balance formula first before balancing the equation.

15b. All antacids can have potential side-effects.

15c. Antacids are always weak bases.

16a. acid + metal carbonate → ionic salt + carbon dioxide + water

16b. Data can be classified according to whether it gives numerical information or describes trends.

17. Validity determines whether the method followed a scientific process to produce meaningful results.

18a. acid + metal carbonate → ionic salt + carbon dioxide + water

18b. Ionic equations do not include spectator ions.

18c. Since there is no lid on the beaker, any evolved gases will escape.

18d. The amount of $CO_2(g)$ produced will depend on the amount of $ZnCO_3(s)$.

19a. A structural formula shows all atoms and bonds in a molecule.

19b. Methanoic acid contains a functional group that can donate a proton, $H^+(aq)$.

19ci. acid + reactive metal → ionic salt + hydrogen gas

19cii. Ionic equations do not include spectator ions.

19d. In reactions between acids and reactive metals, the metal becomes a dissolved cation.

19e. The amount of $H_2(g)$ produced will be proportional to the amount of tin consumed.

20a. Acids can be classified according to how many protons they can accept or receive.

20bi. acid + metal hydroxide → ionic salt + water

20bii. Ionic equations do not include spectator ions.

20c. Start by calculating the amount, in mol, of $Al(OH)_3$ in one antacid tablet.

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10C Calculating pH

STUDY DESIGN DOT POINT

- use of the logarithmic pH scale to rank solutions from most acidic to most basic; calculation of pH for strong acid and strong base solutions of known concentration using the ionic product of water (K_w at a given temperature)

10A 10B 10C 10D 10E

2.1.7.1 The ionic product of water

2.1.7.2 The pH scale

ESSENTIAL PRIOR KNOWLEDGE

10A Water as an amphiprotic substance

10A Strong acids and strong bases

See questions 64–65.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: H_Ko/Shutterstock.com

How can we use pH to quantify the suitability of soil for plant growth?

The **pH** scale has a number of applications in the real world, such as in determining whether soil samples are optimal for growing vegetation. In this lesson, we will learn how to quantify the acidity and basicity of substances using the **ionic product of water**, and how to subsequently rank these substances on a **pH scale.**

KEY TERMS AND DEFINITIONS

Acidic solution	solution containing a greater concentration of hydronium ions than hydroxide ions
Autoionisation	process by which an atom or molecule spontaneously forms ions
Basic solution	solution containing a greater concentration of hydroxide ions than hydronium ions
Ionic product of water	K_w, expression of the concentration of ions present in water at 25 °C
Neutral solution	solution containing an equal concentration of hydronium and hydroxide ions
pH	measure of the concentration of hydronium ions in a solution
pH scale	pH values arranged on a logarithmic scale

The ionic product of water 2.1.7.1

Solutions of pure water contain both hydronium and hydroxide ions, but only in very low concentrations.

What happens when water is ionised?

In lesson 10A, we discovered that water is an amphiprotic substance, meaning it is capable of acting as either an acid or a base during an acid-base reaction. Therefore, although water acts as both a very weak acid and a very weak base, it can nonetheless accept and donate protons. In fact, two water molecules are actually capable of reacting with each other, one as the acid and the other as the base, as shown by the equation:

$$\underset{\text{acid}}{H_2O(l)} + \underset{\text{base}}{H_2O(l)} \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

Figure 1 shows the donation of one proton (hydrogen ion) from one water molecule to another represented in the equation.

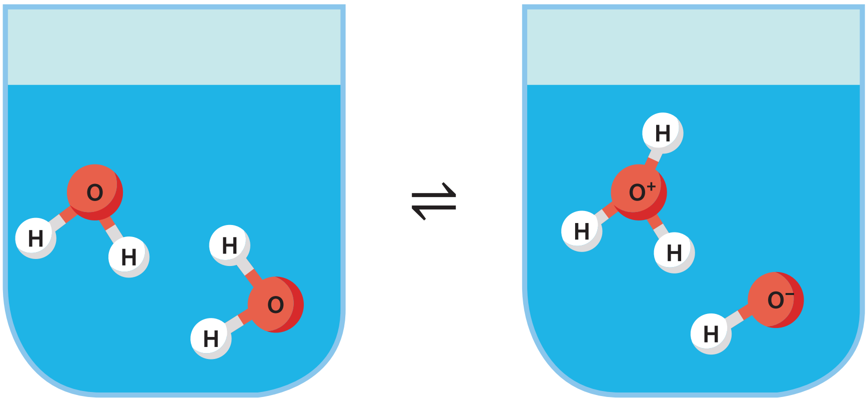

Figure 1 Autoionisation of water

The process by which water molecules spontaneously gain an electric charge is known as autoionisation. Of course, since water acts as a relatively weak acid or base, this reaction only occurs to a very small extent. This means that in pure water, there will only be very low concentrations of H_3O^+ and OH^-. In fact, for pure water at a temperature of 25 °C, the concentrations of both ions are equal to 1.00×10^{-7} M:

$$[H_3O^+] = 1.00 \times 10^{-7} \text{ M}$$

$$[OH^-] = 1.00 \times 10^{-7} \text{ M}$$

USEFUL TIP

mol L^{-1} and M and [X] all represent the same units. Square brackets are often used when doing pH and ion concentration calculations.

To try and imagine how low this concentration is, consider the Olympic—sized swimming pool shown in figure 2. If the pool were filled with 2 x 2 500 000 L of pure water, then the volumes of H_3O^+ and OH^- would each be approximately 5 mL.

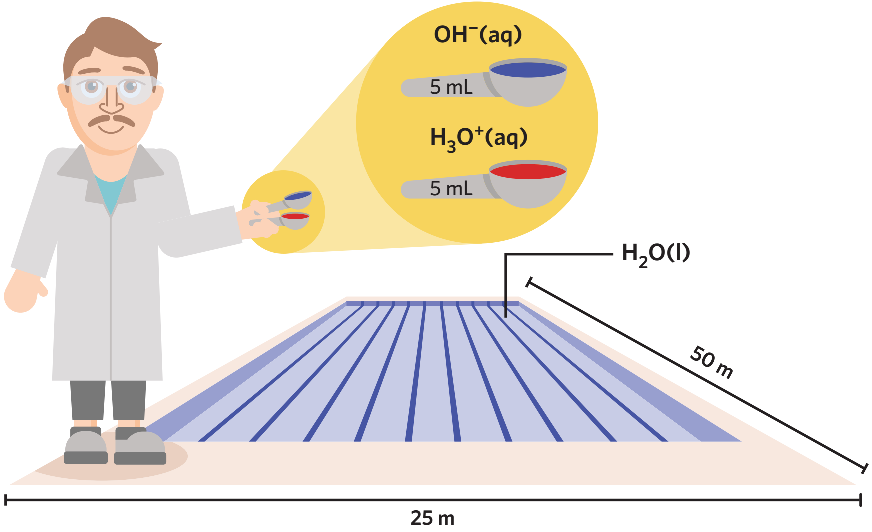

Figure 2 The concentrations of H_3O^+(aq) and OH^-(aq) in an olympic swimming pool filled with pure water

When a sample of pure water is ionised at 25 °C, we know that, regardless of the individual concentrations of hydroxide and hydronium ions, the product of their concentrations is always equal to 1.00×10^{-14} M^2. This value is known as the ionic product of water and is represented by the symbol K_w:

$$K_w = [H_3O^+] \times [OH^-]$$

$$K_w = (1.00 \times 10^{-7} \text{ M}) \times (1.00 \times 10^{-7} \text{ M})$$

$$K_w = 1.00 \times 10^{-14} \text{ M}^2$$

With the ionic product of water in mind, we can say that if the concentration of hydronium ions in a sample at 25 °C increases, the concentration of hydroxide ions must decrease to maintain a constant K_w value. Similarly, if the concentration of hydroxide increases, the concentration of hydronium ions must decrease to again maintain a constant value. This balancing act is shown in figure 3.

USEFUL TIP

When the unit of M is multiplied by itself (M × M), the unit becomes M^2. Although it is difficult to make any real physical sense of a unit like M^2, this unit is still included by convention.

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MISCONCEPTION

'Pure water is 50% hydroxide ions and 50% hydronium ions.'

Since autoionisation only occurs to a very small degree in pure water, the vast majority of a water sample is composed of H_2O molecules. Ions make up only a very small amount of pure water.

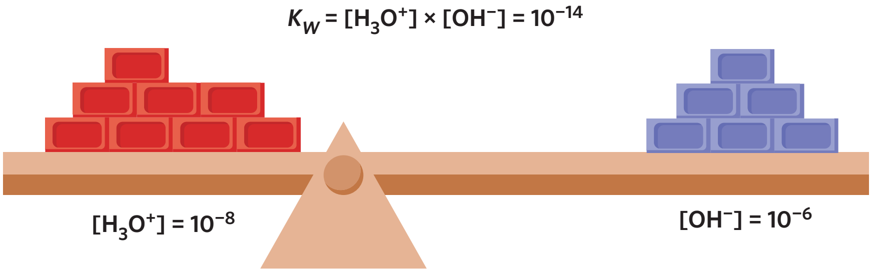

Figure 3 The balancing act between the concentration of hydronium and hydroxide ions in pure water

USEFUL TIP

The relationship $K_W = [H_3O^+] \times [OH^-] = 1.00 \times 10^{-14}$ M^2 is only true if the solution is at a temperature of 25 °C.

WORKED EXAMPLE 1

The concentration of hydronium ions in an aqueous solution at 25 °C is equal to 5.32×10^3 M. Calculate the concentration of hydroxide ions present in this solution.

What information is presented in the question?

The concentration of hydronium ions in an aqueous solution at 25 °C.

What is the question asking us to do?

Calculate the concentration of hydroxide ions present in the solution.

What strategies do we need in order to answer the question?

1. Substitute the concentration of hydronium ions into the relationship $K_w = [H_3O^+] \times [OH^-] = 1.00 \times 10^{-14}$ M^2 at 25 °C.
2. Solve for the concentration of hydroxide ions.

Answer

Since the water is at 25 °C, we know that the relationship $K_w = [H_3O^+] \times [OH^-] = 1.00 \times 10^{-14}$ M^2 is applicable. Substituting in the known concentration of hydronium ions, we get:

$$K_w = [H_3O^+] \times [OH^-] = 1.00 \times 10^{-14}\ M^2$$

$$K_w = 5.32 \times 10^3\ M \times [OH^-] = 1.00 \times 10^{-14}\ M^2$$

Rearranging this, we can determine the concentration of hydroxide ions in the solution:

$$[OH^-] = \frac{10^{-14}\ M^2}{5.32 \times 10^3\ M}$$

$$[OH^-] = 1.88 \times 10^{-18}\ M$$

USEFUL TIP

The higher the concentration of H_3O^+ in a solution, the more acidic the solution is. Likewise, the higher the concentration of OH^- in a solution, the more basic the solution is.

The values of $[H_3O^+]$ and $[OH^-]$ in pure water provide a baseline for a **neutral solution**, a solution in which the concentrations of $[H_3O^+]$ and $[OH^-]$ are equal. If we add an acid to a solution of pure water, then it will form new H_3O^+ ions by donating protons to H_2O molecules. Since the ionic product of water remains constant at 1.00×10^{-14} M^2, this increase in $[H_3O^+]$ above the baseline of 10^{-7} M results in an **acidic solution**. Similarly, adding a base to pure water results in the formation of OH^- ions. This increase in $[OH^-]$ will result in a **basic solution**. This information is summarised in figure 4.

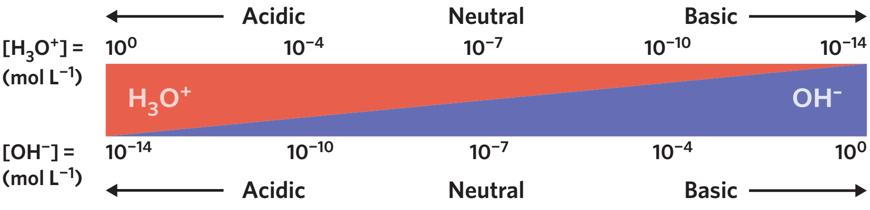

Figure 4 The acidity of a solution depends on the relative concentrations of H_3O^+ and OH^- ions.

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Progress questions

Question 1

The ionic product of water, K_w, shows that a solution of pure water

A. has no ions at all present.

B. has hydronium and hydroxide ions, but only in very low concentrations.

Question 2

At 25 °C, the ionic product of water is

A. constant.

B. dependent on whether an acid or base has been added to the solution.

Question 3

An acidic solution has a ________ concentration of H_3O^+ ions and a ________ concentration of OH^- ions than a basic solution.

A. higher, lower

B. lower, higher

The pH scale 2.1.7.2

The pH scale is used to quantitatively measure how acidic or basic a solution is.

How can we measure and rank the pH of solutions?

In lesson 10A, we classified acids and bases qualitatively by describing them as either strong or weak. The pH scale gives us a way to quantitatively classify acidic and basic solutions. By measuring how many H_3O^+ ions are present in an acidic or basic solution, we can label it as either strong or weak.

We can use pH values to define what is considered an acidic, basic or neutral solution at 25 °C:

- A solution with a pH of 7 is a neutral solution e.g. pure water.
- A solution with a pH less than 7 is an acidic solution.
- A solution with a pH greater than 7 is a basic solution.

The pH's of several everyday substances are shown in figure 5.

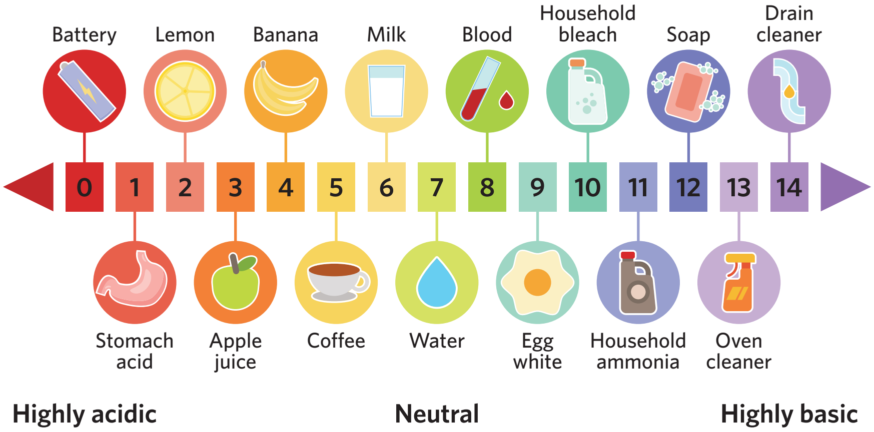

Figure 5 The pH of common acidic and basic solutions

USEFUL TIP

pH is usually measured on a scale from 0–14, with 0 representing a highly acidic solution and 14 representing a highly alkaline solution. However, if either H_3O^+ or OH^- are present in high enough concentrations, pH values below 0 or above 14 are possible.

How can we calculate the pH of a solution?

As we learned in lesson 10A, when an acid is added to water, H_3O^+ ions are formed, and when a base is added to water, OH^- ions are formed. Therefore, we can use the concentrations of H_3O^+ and OH^- to determine just how acidic or basic the solution we are investigating is.

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USEFUL TIP

Some scientific calculators will require the base number '10' to be manually entered into the '$\log_{10}$' expression, whilst others will only have the expression 'log' with a default base of 10. It's important to familiarise yourself with how to use your scientific calculator for VCE chemistry.

USEFUL TIP

Since pH values depend on the negative logarithm of $[H_3O^+]$, the pH of a solution will decrease as the concentration of H_3O^+ ions increases.

Using the relationship for the ionic product of water, K_w, we only need to know the concentration of one of these ions to determine the concentration of the other. In chemistry, we primarily use $[H_3O^+]$ to measure the acidity or basicity of a solution. This measure of acidity is known as the pH of the solution and can be calculated from the following formula:

$$pH = -\log_{10}[H_3O^+]$$

or, rearranged: $[H_3O^+] = 10^{-pH}$

Since the expression for pH involves a logarithm ($\log_{10}$), every increase or decrease in pH by 1 means that there was a change in the concentration of H_3O^+ by a factor of 10. Similarly, if there is a change in pH of 4, then the concentration of H_3O^+ has been changed by a factor of 10 000! This allows a very wide range of H_3O^+ concentrations to be represented on a relatively small scale.

For the purposes of VCE chemistry, we will only investigate how to calculate the pH of a strong acid or base solution. This is because their complete ionisation in water gives much simpler calculations. For example, a solution of 1.0 M hydrochloric acid, a strong acid, will ionise to a much greater degree than a solution of 1.0 M ethanoic acid, a weak acid. As a result, a solution of 1.0 M hydrochloric acid will have a greater concentration of H_3O^+ ions, and will therefore have a lower pH.

WORKED EXAMPLE 2

Calculate the pH of a 0.30 M solution of hydrochloric acid at 25 °C.

What information is presented in the question?

The concentration of a hydrochloric acid solution at 25 °C.

What is the question asking us to do?

Determine the solution's pH.

What strategies do we need in order to answer the question?

1. Determine $[H_3O^+]$ for the solution.
2. Substitute $[H_3O^+]$ into the formula for the pH of a solution.

Answer

Given HCl is a strong acid, it will completely ionise. Therefore, the concentration of $H_3O^+(aq)$ in the solution will be equal to the concentration of HCl(aq):

$$[HCl] = [H_3O^+] = 0.30\ M$$

Substituting $[H_3O^+]$ into the formula for the pH of a solution, we get:

$$pH = -\log_{10}[H_3O^+]$$

$$pH = -\log_{10}(0.30\ M)$$

$$pH = 0.52$$

Progress questions

Question 4

An acidic solution has a pH ________, whereas a basic solution has a pH ________.

A. above 7, below 7

B. below 7, above 7

Question 5

If the pH value of a solution decreases by 2, then the concentration of H_3O^+ ions in the solution has ________ by a factor of ________.

A. decreased, 2

B. increased, 100

Question 6

A solution of 1.0 M ethanoic acid (vinegar) has a pH of 4.0. A solution of 5.0 M ethanoic acid would have a ________ pH than this solution.

A. lower

B. higher

Theory summary

- The ionic product of water, K_w, is constant at 25 °C and is a result of the spontaneous autoionisation of H_2O into hydronium and hydroxide ions.
  - $H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$
  - $K_w = [OH^-][H_3O^+] = 1.0 \times 10^{-14}\ M^2$
- At 25 °C:
  - Neutral solution: $pH = 7$, $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ M
  - Acidic solution: $pH < 7$, $[H_3O^+] > 1.0 \times 10^{-7}$ M and $[OH^-] < 1.0 \times 10^{-7}$ M
  - Basic solution: $pH > 7$, $[H_3O^+] < 1.0 \times 10^{-7}$ M and $[OH^-] > 1.0 \times 10^{-7}$ M
- On the pH scale:
  - 0 represents a highly acidic solution.
  - 7 represents a neutral solution (pure water).
  - 14 represents a highly basic solution.
- The pH of strong acids and bases is calculated using the formulas:
  - $pH = -\log_{10}[H_3O^+]$ and $K_w = [OH^-][H_3O^+]$

10C Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 7–9.

Sulfuric acid, H_2SO_4, is a corrosive substance used in the manufacture of products such as explosives and detergents. Sulfuric acid is considered to be a strong acid.

Question 7 (1 MARK) 🌶

Sulfuric acid would ________ in pure water at 25 °C.

A. partially ionise

B. explode

C. partially dissociate

D. fully ionise

Question 8 (1 MARK) 🌶

Which of the following is not true of sulfuric acid?

A. Sulfuric acid has a pH below 7.

B. Sulfuric acid would form an acidic solution with pure water.

C. 1.0 M sulfuric acid is a basic substance.

D. Sulfuric acid would have a pH similar to stomach acid.

Question 9 (2 MARKS) 🌶🌶

Calculate the concentration of hydronium, H_3O^+, ions in a solution of sulfuric acid with a pH of 1.3.

Exam-style

Question 10 (1 MARK)

Which of the following correctly lists the three solutions in order of increasing pH?

A. 0.01 M $CH_3COOH(aq)$, 2.0 M $H_2SO_4(aq)$, 0.025 M $NaOH(aq)$

B. 2.0 M $H_2SO_4(aq)$, 0.01 M $CH_3COOH(aq)$, 0.025 M $NaOH(aq)$

C. 0.01 M $CH_3COOH(aq)$, 2.0 M $H_2SO_4(aq)$, 0.025 M $NaOH(aq)$

D. 0.025 M $NaOH(aq)$, 0.01 M $CH_3COOH(aq)$, 2.0 M $H_2SO_4(aq)$

Question 11 (6 MARKS)

The human body requires a blood pH in the range 7.35 – 7.45 in order to survive. Any pH outside this range can have rapid and devastating effects on the body's ability to function.

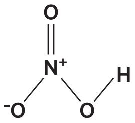
a. Is human blood acidic, basic or neutral? (1 MARK)

b. Based on your answer to part a, would human blood contain a higher concentration of OH^- or H_3O^+ ions? (1 MARK)

c. Calculate the highest and lowest concentrations of H_3O^+ ions that could be present in the blood when the human body is functioning normally. (4 MARKS)

Question 12 (9 MARKS)

Nitric acid, also known as 'aqua fortis' (Latin for 'strong water'), is a strong and highly corrosive acid. Nitric acid completely ionises in water. The structure of nitric acid is given.



a. Calculate the pH of a solution of 0.50 M nitric acid at 25 °C. (3 MARKS)

So far, we have only investigated the value of K_w at 25 °C. As the temperature of a solution changes, the value of K_w also changes.

b. At 100 °C, the value of K_w is actually 51.3×10^{-14} M^2. How might this affect the pH value of the solution of nitric acid calculated in part a? (2 MARKS)

c. If neutral pH is still defined as when $[H_3O^+] = [OH^-]$, what would be considered neutral pH at 100 °C, assuming K_w is equal to 51.3×10^{-14} M^2? (4 MARKS)

Question 13 (6 MARKS)

Sodium hydroxide, NaOH, is a strong base commonly used as a household cleaning product. Jamie has been tasked with scrubbing the bathroom tiles and wants to use sodium hydroxide, but he recently learned that it's only safe to use products at or below a pH of 11.

a. Calculate the highest concentration of sodium hydroxide that Jamie can safely use. (4 MARKS)

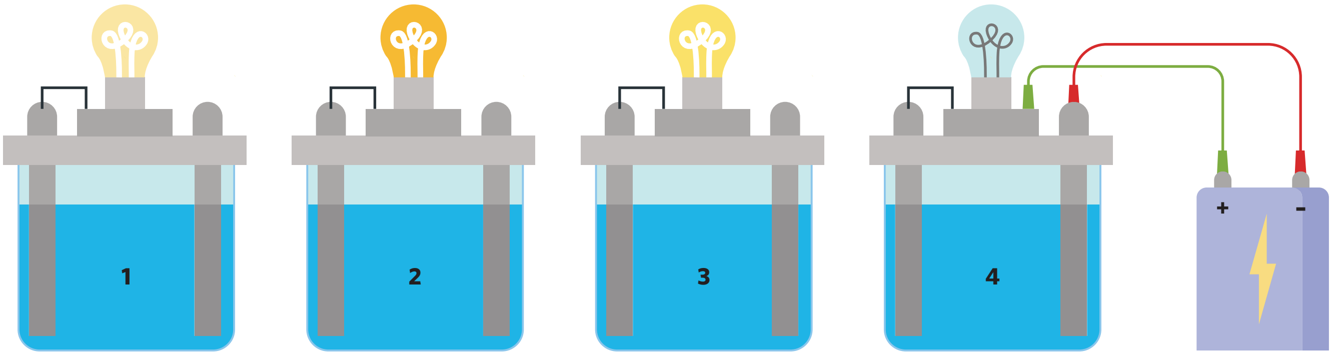
b. Jamie only has a 0.1 M sodium hydroxide solution at his disposal. Based on your answer to part a, is this safe for him to use? (2 MARKS)

Key science skills

Question 14 (9 MARKS)

When acids or bases are added to a solution, they form positively and negatively charged ions (H_3O^+, OH^-, etc.). These ions are capable of conducting electricity if a current from an external power supply is run through the solution.

Carlos wants to investigate the electrical conductivity of acid-base solutions and so sets up the experiment as shown. He connected a battery to four different acidic solutions and measured how brightly each lamp shone. These beakers are labelled 1–4. Only one such circuit is shown in the diagram, but all lamps are shown with how brightly they shone when connected to the power supply.



The results of Carlos's experiment are shown in the table.

Beaker	Solution	Brightness (lumens)
1	1.0 M $CH_3COOH(aq)$	130
2	1.0 M $HNO_3(aq)$	780
3	1.0 M HF(aq)	290
4	Deionised water, $H_2O(l)$	0.00

Carlos knows that solutions containing high concentrations of ions are strong conductors of electricity and that solutions with low concentrations of ions are poor conductors. Assume that the temperature of all solutions remains constant at 25 °C.

a. What is the purpose of beaker 4 in this experiment? (1 MARK)

b. Identify the independent and dependent variables in this experiment. (2 MARKS)

c. List the solutions in beakers 1–4 in order of decreasing pH. Justify your answer. (4 MARKS)

d. Carlos conducts a second experiment using the solutions in the given table.

Beaker	Solution	Brightness (lumens)
1	2.0 M $CH_3COOH(aq)$	
2	2.0 M $HNO_3(aq)$	
3	2.0 M HF(aq)	
4	Deionised water, $H_2O(l)$	

i. Propose a hypothesis for the brightness of the lamps connected to these solutions relative to those of the first experiment. (1 MARK)

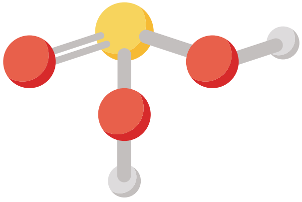
ii. Would the data Carlos collects in this second experiment be classified as qualitative or quantitative? (1 MARK)

FROM LESSONS 16A & 16B

Questions from multiple lessons

Question 15 (8 MARKS)

Selenous acid, H_2SeO_3, is an acid used in the protection and alteration of the colour of steel. The 3D structure of selenous acid is shown.



Since selenous acid is not typically classified as a 'strong' acid, it does not deprotonate fully. It deprotonates according to the two given equations.

$H_2SeO_3(aq) \rightleftharpoons H^+(aq) + HSeO_3^-(aq)$

$HSeO_3^-(aq) \rightleftharpoons H^+(aq) + SeO_3^{2-}(aq)$

a. Would selenous acid be classified as monoprotic or diprotic? (1 MARK)

b. If selenous acid were a strong acid, calculate the pH of a 0.080 M solution of H_2SeO_3 at 25 °C. (2 MARKS)

c. Why does the second deprotonation of selenous acid happen to a much lower extent than the first deprotonation? (2 MARKS)

d. If selenous acid were actually an even stronger acid, would the pH of the solution be higher or lower than the pH calculated in part b? Explain. (3 MARKS)

FROM LESSON 10A

Hints

10. Acids in higher concentrations have greater $[H_3O^+]$, and therefore lower pH. Bases have higher pH.

11a. Neutral pH is 7 at 25 °C.

11b. The acidity or basicity of a solution can be used to determine the relative concentrations of hydronium and hydroxide ions.

11c. The higher the pH, the lower the $[H_3O^+]$. The lower the pH, the higher the $[H_3O^+]$.

12a. The concentration of a solution of a strong acid is equal to $[H_3O^+]$.

12b. The higher the value of K_w, the higher the value of $[H_3O^+][OH^-]$.

12c. Neutral pH is the pH at which $[H_3O^+]$ is equal to $[OH^-]$, regardless of the temperature.

13a. The concentration of a sodium hydroxide solution is equal to $[OH^-]$.

13b. The concentration of a sodium hydroxide solution is equal to $[OH^-]$.

14a. A sample in which the independent variable is unchanged is a key part of every experiment.

14b. Changing the independent variable affects the dependent variable.

14c. pH is dependent on the concentration of H_3O^+ ions present in the solution.

14di. The acids are significantly more concentrated in the second experiment.

14dii. Qualitative data is used to describe characteristics, whilst quantitative data is countable or measurable.

15a. Selenous acid can deprotonate twice.

15b. The pH of a solution is related to the concentration of H_3O^+ ions.

15c. Negatively and positively charged species have a partial attractive force between them.

15d. The strength of an acid depends on how readily it donates a proton.

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10D Comparing pH measurements

Image: Rabbitmindphoto/Shutterstock.com

How can the pH of a solution be estimated from its colour?

Indicators are chemical substances that change colour in acidic and basic solutions. In this lesson, we will learn about how to determine the approximate pH of solutions using these colour changes. We will then learn how pH readings from indicators compare to those of **pH meters**.

STUDY DESIGN DOT POINT

- accuracy and precision in measurement as illustrated by the comparison of natural indicators, commercial indicators, and pH meters to determine the relative strengths of acidic and basic solutions

10A 10B 10C **10D** 10E

2.1.8.1 Indicators and pH meters

2.1.8.2 Accuracy and precision

ESSENTIAL PRIOR KNOWLEDGE

10A Strength of acids and bases
16D Accuracy and precision
See questions 66–67.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

KEY TERMS AND DEFINITIONS

Commercial indicator	an artificially prepared indicator
Indicator	a substance that changes colour to indicate a pH range
Litmus paper	paper used to indicate the acidity or basicity of a substance
Natural indicator	an indicator made from naturally sourced substances
pH meter	an instrument used to measure the pH of a substance

Indicators and pH meters 2.1.8.1

Indicators and pH meters can be used to measure the pH of various substances.

How can indicators be used to approximate the pH of a solution?

In lesson 10C, we learned the formula for calculating the pH of a strong acid or base of known concentration. However, in the case where we have a solution of unknown concentration, the pH can be approximated using indicators, chemical compounds that change colour depending on the pH they are immersed in. One of the simplest and most common indicators is **litmus paper**, which uses a dye derived from lichens.[1] As shown in figure 1, blue litmus paper will turn red in an acidic solution, and red litmus paper will turn blue in a basic solution.

Whilst litmus paper is an example of a **natural indicator** – an indicator made from naturally sourced substances – there are also a number of **commercial indicators**. That is, indicators that are prepared artificially, such as in a laboratory. Table 1 contains a number of natural and commercial indicators, as well as their colour changes.

KEEN TO INVESTIGATE?

[1] **How is litmus paper made from lichens?**
Search: Lichens and Litmus and pH! Oh My!

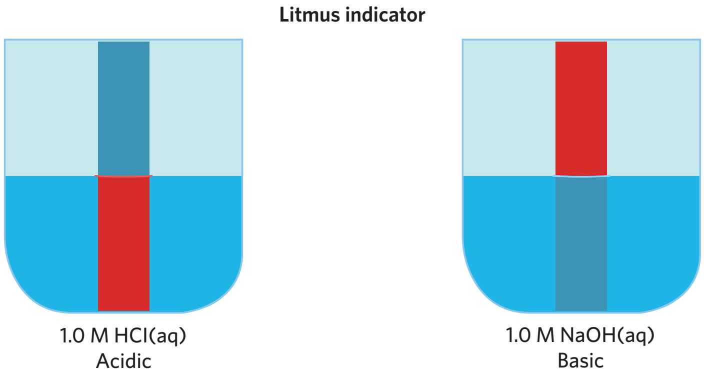

Figure 1 Litmus paper in acidic and basic solutions

Table 1 The colours of natural and commercial indicators in acidic and basic solutions

Natural			**Commercial**		
Indicator	**Colour in acid**	**Colour in base**	**Indicator**	**Colour in acid**	**Colour in base**
Red litmus	Red	Blue	Phenolphthalein	Colourless	Pink
Blue litmus	Blue	Red	Methyl violet	Yellow	Blue
Turmeric	Yellow	Reddish brown	Methyl orange	Red	Yellow
Red cabbage juice	Reddish	Greenish yellow	Thymolphthalein	Colourless	Blue

How can more specific pH measurements be obtained?

So far, we have only learned about pH indicators, which can give us an approximation of the pH of an acid or base and its relative strength. However, pH meters can be used to provide a specific, quantitative measurement of pH. pH meters work by measuring the difference in electrical potential between two probes due to the concentration of charged particles – or ions – in the solution. As shown in figure 2, the meter then gives a reading on a digital display.

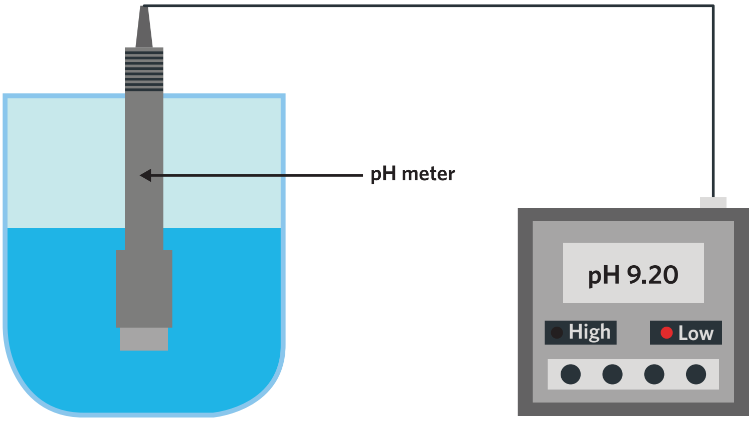

Figure 2 pH meter in a basic solution

How can the relative strengths of acids and bases be determined?

As we learned in lesson 10A, acids and bases can be described qualitatively by classifying them as either weak or strong. Using the chart shown in figure 3, a number of indicators can be used to determine the strengths of acids and bases.

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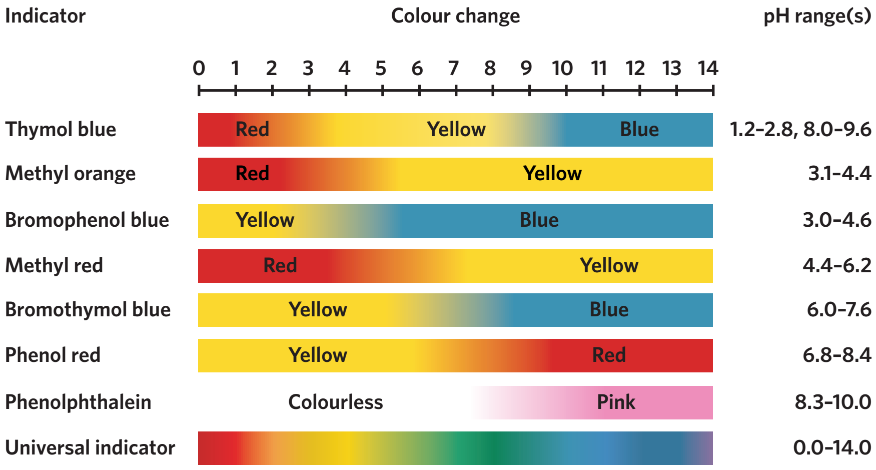


Figure 3 Common indicators, their colour changes, and pH ranges

USEFUL TIP

Aside from the universal indicator, the colour changes and pH ranges of the indicators given can all be found in the VCE Data Book.

In general, a solution with a pH of 3 or less is considered to be a strong acid. Therefore, if methyl orange was added to a solution and a red colour was produced, the solution would be a strong acid. At the other end of the spectrum, a solution with a pH of 10 or higher is considered to be a strong base. Therefore, if phenolphthalein was added to a solution and a pink colour was produced, the solution would be a strong base.

Universal indicator can be used to indicate the pH of a solution all the way from 0 to 14. If universal indicator was added to a solution of a weak acid (pH of 3–7), the solution would likely appear green or yellow. On the other hand, if the universal indicator was added to a solution of a weak base (pH of 7–10), the solution would likely appear green or blue.

WORKED EXAMPLE 1

A chemist added an acid of unknown strength to water. He has thymol blue, bromothymol blue and phenolphthalein at his disposal. Using figure 3, which of these indicators would be most suitable to determine whether the solution is a strong or weak acid?

What information is presented in the question?

An acid is added to water.

What is the question asking us to do?

Decide which of the three indicators is best for distinguishing between strong and weak acids.

What strategies do we need in order to answer the question?

1. Identify the pH ranges of each indicator.
2. Identify the pH range for strong and weak acids.
3. Match the pH range of the indicator to the pH range for strong and weak acids.

Answer

Thymol blue changes colour between a pH of 1.2 and 2.8, as well as between a pH of 8.0 and 9.6. Bromothymol blue changes colour between a pH of 6.0 and 7.6. Phenolphthalein changes colour between a pH of 8.3 and 10.0.

A solution with a pH below 3 is typically considered to be a strong acid, and a solution with a pH between 3 and 7 is considered to be a weak acid.

Therefore, thymol blue would be the most suitable indicator to determine the strength of the acid, as it would turn red in a solution of a strong acid, and yellow in a solution of a weak acid.

Progress questions

Question 1

What colour is likely to be observed if thymolphthalein is added to a basic solution?

A. Colourless

B. Blue

Question 2

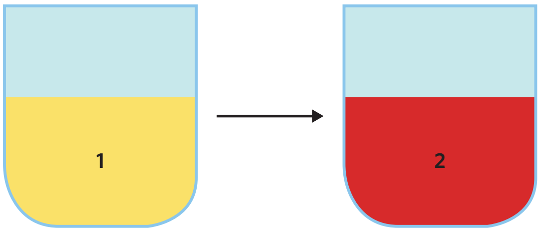
Red litmus paper is immersed in a solution and the paper remains red. The solution is likely to be

A. acidic.

B. basic.

Question 3

Beaker 1 contains a solution of unknown pH. Beaker 2 contains the same solution with several drops of methyl orange added.



The solution is likely to be ________.

A. acidic

B. basic

Question 4

A pH meter provides ________ data.

A. qualitative

B. quantitative

Question 5

Several drops of thymol blue are added to a solution and the solution turns yellow. Which of the following is the likely pH range of the solution?

A. 1.2–9.6

B. 2.8–8.0

Question 6

What would be observed if phenolphthalein was added to a colourless, acidic solution?

A. No colour change would be observed.

B. The solution would turn pink.

Question 7

A solution of a weak base would turn ________ when several drops of universal indicator are added.

A. red

B. blue

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Accuracy and precision 2.1.8.2

Why are some pH measurements more accurate or precise than others?

In lesson 16D, we learned about the accuracy and precision of results obtained from experimental methods. Whilst both indicators and pH meters are highly useful pH measurements in industries like aquaculture and wastewater treatment, one of the two provides typically more accurate and precise measurements. Unlike pH meters, indicators only provide a pH range for a given solution. Therefore, since a specific pH value can only be approximated from a range, indicators do not tend to give accurate measurements of the pH of a solution. pH meters, on the other hand, typically provide a high level of accuracy. The accuracy of most pH meters is around ± 0.01 pH units, and is built into the calibration of the meter.

Given pH meters can give reproducible measurements of pH, they can provide precise results, provided there is no change to the calibration of the instrument between measurements. Indicators, however, can only be used to approximate a pH range, and hence, if the true pH of a solution is on the border of the indicator's colour change range, the estimated pH of a given solution may differ greatly between multiple trials. Therefore, indicators cannot be said to provide very precise measurements of pH. A comparison of the accuracy and precision of pH measurements obtained from techniques covered in this lesson is summarised in table 2.

Table 2 Accuracy and precision of pH measurements obtained from indicators and pH meters

pH Measuring Technique	Accuracy	Precision
Indicator	Low	Low
Calibrated pH meter	High	High

Progress questions

Question 8

Indicators provide relatively inaccurate but highly precise pH measurements.

A. True

B. False

Question 9

pH meters can provide both accurate and precise pH measurements.

A. True

B. False

Theory summary

- Indicators are used to determine the pH range of a solution through colour changes.
- pH meters are used to give a specific pH value for a solution.
- pH meters give more accurate and precise measurements of pH than indicators.

The content in this lesson is considered fundamental prior knowledge to analysis of organic compounds by volumetric analysis (Unit 4 AOS 1).

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10D Questions

Mild Medium Spicy

Deconstructed

Use the following information to answer questions 10-12.

In agriculture, knowing the pH of soil is incredibly important for plant growth, as the pH level affects the amount of nutrients available to plants. pH meters are used by farmers to measure the acidity or basicity of the soil before using it. The optimal soil pH for most plants to grow is around 5.8–6.8.

Question 10 (1 MARK)

Which of the following options would be best for measuring the pH of soil accurately?

A. Bromothymol blue

B. Phenolphthalein

C. A pH meter

D. Red litmus paper

Question 11 (1 MARK)

Consider a solution with a pH somewhere between 5.8–6.8. Which of the following indicators would be most suitable for determining if the solution was too basic?

A. Thymol blue

B. Phenolphthalein

C. Phenol red

D. Methyl orange

Question 12 (3 MARKS)

Wanting to test the pH of her land, a farmer inserts 4 pieces of red litmus paper into a solution of her soil. In reference to the accuracy of the results obtained, suggest an alternative method that the farmer can use to measure the pH of the soil.

Exam-style

Question 13 (1 MARK)

To each of three samples of a solution, a different acid-base indicator is added. The following colours are observed.

Indicator	Colour
thymol blue	yellow
methyl red	yellow
phenolphthalein	colourless

The pH of the solution is between

A. pH = 2.8 and pH = 4.4

B. pH = 4.4 and pH = 6.2

C. pH = 6.2 and pH = 8.0

D. pH = 8.0 and pH = 10

Adapted from VCAA 2011 Exam 1 Multiple choice Q12

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Key science skills

Question 14 (7 MARKS)

Jean-luc is testing the pH of his pool. He has a pool cleaning kit which tells him that the safe pH range for pool water is 7.0–7.6. At his disposal, he has several strips of blue litmus paper, a vial of universal indicator, and a pH meter. He scoops up a cup of water to use as a sample, estimating that the water's current pH would be approximately 8.0.

a. Which of the three measuring techniques should Jean-luc use to determine a specific value for the pH of the pool water? Explain your answer. (3 MARKS)

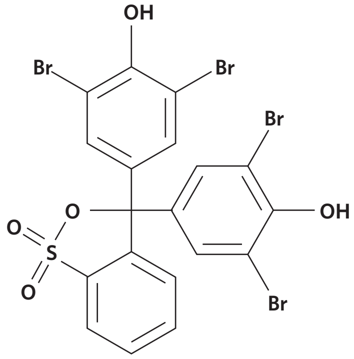
b. Jean-luc uses his pH meter to take one reading of the pH of the water in his cup. Comment on the precision of the results obtained from this technique. (2 MARKS)

c. Suggest one way he can improve the precision of his final measurement. Explain your answer. (2 MARKS)

FROM LESSONS 16C & 16D

Questions from multiple lessons

Question 15 (3 MARKS)



The structure of the indicator bromophenol blue is shown. Its molecular formula is $C_{19}H_{10}Br_4O_5S$. Bromophenol blue changes colour from yellow to blue as the pH of a solution transitions from 3.0 (yellow) to 4.6 (blue).

Bromophenol blue can react with ammonia, NH_3, according to the given equation.

$C_{19}H_{10}Br_4O_5S(aq) + NH_3(aq) \rightarrow C_{19}H_9Br_4O_5S^-(aq) + NH_4^+(aq)$

a. What would the colour of bromophenol blue be in a neutral solution? (1 MARK)

b. Is bromophenol blue acting as an acid or a base in this reaction? Explain. (2 MARKS)

FROM LESSON 10A

Hints

13. The pH range of each indicator can be used to narrow down the pH of the solution.

14a. Some indicators can help to narrow down the pH range more than others.

14b. There is a correlation between the number of trials in an experiment and the precision of the results obtained.

14c. There is a correlation between the number of trials in an experiment and the precision of the results obtained.

15a. A neutral solution has a pH of 7.

15b. During an acid-base reaction, a proton is transferred from an acid to a base.

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10E Applications of acid-base reactions in society

STUDY DESIGN DOT POINT

- applications of acid-base reactions in society: for example, natural acidity of rain due to dissolved CO_2 and the distinction between the natural acidity of rain and acid rain, or the action of CO_2 forming a weak acid in oceans and the consequences for shell growth in marine invertebrates

10A 10B 10C 10D **10E**

2.1.9.1 Acid rain

2.1.9.2 Ocean acidification

ESSENTIAL PRIOR KNOWLEDGE

3B Oxidation
10A Weak acid
10A Acidic and Basic
10A Neutralisation
10A Hydrogen ions
See questions 68–72.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.


Image: Richard Whitcombe/Shutterstock.com

Is it possible that humans are changing the chemistry of the oceans?

Acid-base reactions can have a strong impact on society and the environment. An increase in the emissions of acid-forming gases over the past few hundred years is accelerating the decline of shelled marine life, destroying forests and decreasing the pH of natural waters. In this lesson we will learn about the negative impact of human-enhanced acid-base reactions.

KEY TERMS AND DEFINITIONS

Combustion when a fuel reacts with oxygen (also called burning)

Calcification a chemical process that forms calcium carbonate

Decalcification a chemical process that breaks down calcium carbonate

Carbonic acid formed by the reaction of carbon dioxide and water

Exoskeleton the outer shell of a marine creature that is made of calcium carbonate

Marine invertebrates ocean and sea creatures that do not have a backbone

Nitric acid can be formed from a series of reactions in the atmosphere between nitrogen gas, oxygen gas and water

Sulfuric acid can be formed from a series of reactions in the atmosphere between sulfur, oxygen gas and water

Acid rain 2.1.9.1

Acid rain is an environmental problem and has significant consequences for all living organisms. Understanding acid rain requires an understanding of acid-base reactions in aqueous solution.

Why is rain naturally acidic?

The atmosphere contains the soluble gas carbon dioxide, which dissolves in the rainwater creating a weak acidic solution with a pH of between 5 and 6. The equations below show how **carbonic acid** (H_2CO_3) is formed, and then ionises to form a proton and a bicarbonate ion.

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$$

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How do human activities create more acidic rain?

The **combustion** of fossil fuels is the major cause of the increased acidity (decreasing the pH) of rainwater. A mixture of acids make up what is known as acid rain.

The formation of dilute **sulfuric acid** occurs due to most fossil fuels containing compounds of sulfur. When these compounds are given energy through combustion sulfur undergoes a series of chemical reactions resulting in the formation of the sulfuric acid. The creation of dilute sulfuric acid in the atmosphere is outlined by the following equations and is shown in figure 1.

1. Sulfur is released during the combustion of fossil fuels and this reacts with oxygen gas in the atmosphere to form sulfur dioxide gas.
   $S(s) + O_2(g) \rightarrow SO_2(g)$
2. The sulfur dioxide gas then reacts with water in the atmosphere to form a dilute solution of sulfurous acid.
   $SO_2(g) + H_2O(l) \rightleftharpoons H_2SO_3(aq)$
3. The dilute sulfurous acid then reacts with oxygen gas in the atmosphere to form a dilute solution of sulfuric acid.
   $2H_2SO_3(aq) + O_2(g) \rightleftharpoons 2H_2SO_4(aq)$
4. The overall equation can be simplified.
   $2SO_2(g) + O_2(g) + 2H_2O(l) \rightleftharpoons 2H_2SO_4(aq)$

MISCONCEPTION

‘Rainwater is acidic due to human activity.’

Pure rainwater collected in areas free from human activity is slightly acidic (pH of less than 6) due to dissolved carbon dioxide, which reacts with water to give carbonic acid, a weak acid.

USEFUL TIP

Nitrogen gas makes up around 78% of air and is a very unreactive gas as it has a triple covalent bond and exists as a diatomic molecule, N_2. It requires a lot of energy to break this bond.

USEFUL TIP

Sulfur dioxide gas is also naturally produced in large amounts from active volcanoes. The natural cycling of sulfur from the atmosphere into oceans and rocks has kept the acidity of rain in balance – it is the human production and combustion of fossil fuels that has made it imbalanced.

The formation of dilute **nitric acid** also originates from the combustion of fossil fuels, especially petrol and diesel from transportation vehicles. When fuels in vehicles undergo combustion, the high temperatures produced allow nitrogen gas from the atmosphere to react with oxygen gas from the atmosphere to form a variety of nitrogen oxide compounds. These nitrogen oxides are collectively symbolised by the formula, NO_x, and consist mainly of nitrogen monoxide (NO) and nitrogen dioxide gas (NO_2). The creation of dilute nitric acid in the atmosphere is outlined by the following equations and is shown in figure 1.

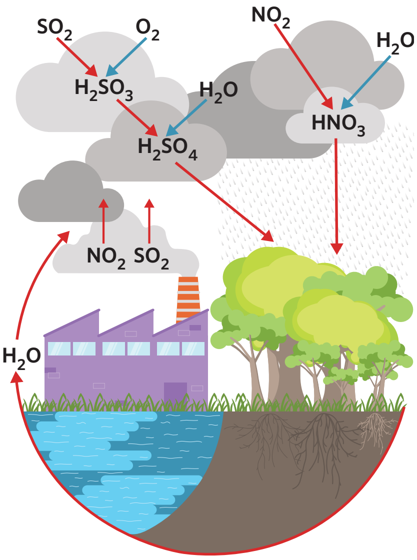

Figure 1 Simplified acid rain production pathway

USEFUL TIP

The single arrow in equations (→) means that the extent of the reaction forming products is complete. The double arrow (⇌) means that the extent of the reaction is not complete and reactants will always be present.

1. Nitrogen gas from the atmosphere reacts to form nitrogen dioxide gas in a two step process.
   $N_2(g) + O_2(g) \rightarrow 2NO(g)$
   $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$
2. Nitrogen dioxide gas reacts with water in the atmosphere to form a dilute mixture of nitrous acid (HNO_2) and nitric acid (HNO_3).
   $H_2O(l)+ 2NO_2(g) \rightleftharpoons HNO_2(aq) + HNO_3(aq)$
3. Oxygen in the atmosphere oxidises the nitrous acid to additional nitric acid.
   $2HNO_2(g) + O_2(g) \rightarrow 2HNO_3(aq)$
4. The overall equation can be simplified.
   $2N_2(g) + 5O_2(g) + 2H_2O(l) \rightleftharpoons 4HNO_3(aq)$

WORKED EXAMPLE 1

Give the complete balanced equation for the reaction that occurs when nitrogen dioxide reacts with water to form a mixture of acids.

What information is presented in the question?

One of the chemicals that reacts and the product that forms.

What is the question asking us to do?

Write an equation.

What strategies do we need in order to answer the question?

1. The formulas for the different acids.
2. Find out which chemical causes this reaction to occur.
3. Write, balance and put state symbols into the equation.
4. Check which type of arrow to use – double or single.

Answer

$H_2O(l) + 2NO_2(g) \rightleftharpoons HNO_2(aq) + HNO_3(aq)$

What are the effects of acid rain?

Vegetation is widely affected as the acidic rainwater causes metal ions like aluminium that are normally bonded in a compound to become aqueous ions that are able to be absorbed by plants. The following equation illustrates how aluminium ions are released into the soil by an acid-base reaction.

$Al(OH)_3(s) + 3H^+(aq) \rightleftharpoons Al^{3+}(aq) + 3H_2O(l)$

When absorbed by plant roots aluminium is poisonous and results in the death of the plant. The extra hydrogen ions found in acid rain can also allow other metal ions (Ca^{2+}, K^+ and Mg^{2+}) essential for healthy plant growth to be released as ions and then become washed away.

Other possible effects of acid rain include:

- damage to the waxy layer on the leaves of plants which may eventually lead to the death of the plant (figure 2),
- acid rain runoff which can also cause the acidification of lakes, reducing the pH to the extent that it can harm or kill aquatic life,
- and reactions with metals that are exposed to the weather.

Image: reisezielinfo/Shutterstock.com

Figure 2 Trees are affected by acid rain

A large proportion of the Earth's surface and many buildings are made from limestone, or calcium carbonate ($CaCO_3$), which is a base that reacts readily with acids. Figure 3 shows the effect of acid rain on limestone. These acid-base reactions result in the formation of cave systems and the weathering of buildings and statues. The following equations show how this human enhanced acid rain mixture of sulfuric and nitric acid reacts with limestone.

$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + CO_2(g) + H_2O(l)$

$CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$

Image: corlaffra/Shutterstock.com

Figure 3 Acid rain and limestone is an example of an acid-base reaction

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Progress questions

Question 1

Acid-base reactions are classified as what type of chemical reaction?

A. Combustion

B. Neutralisation

Question 2

Rainwater is naturally acidic because of which acid?

A. Carbonic acid

B. Nitric acid

Question 3

Acid rain is a mixture of natural and human produced acids.

A. True

B. False

Question 4

Which of these is a consequence of acid rain?

A. Higher pH of the water in rivers and streams

B. Damaged trees and forests

Question 5

What type of chemical reaction occurs when both sulfur and nitrogen react with oxygen in the air?

A. Oxidation

B. Reduction

Ocean acidification 2.1.9.2

Oceans are basic, with an average pH of around 8. When carbon dioxide, CO_2, is absorbed by the oceans, reactions occur that reduce pH, carbonate ion concentration, and the solubility of the biologically important calcium carbonate which is essential in creating shells for marine life.

What happens to the oceans as they absorb more carbon dioxide?

Like rainwater that is naturally acidic due to the reaction

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$$

the surface of the ocean is constantly dissolving carbon dioxide by the same process. Historically, there has been a balance between carbon dioxide being generated, and carbon dioxide being dissolved. The problem now is that carbon dioxide is being created at a much faster rate than can be absorbed through natural processes. This excess human made carbon dioxide results in the oceans becoming more acidic due to the increased formation of hydrogen ions. Recent estimates suggest that around 30% of all human made carbon dioxide emissions have been dissolved in the oceans.

MISCONCEPTION

'Ocean acidification means that the oceans have become acidic.'

Oceans are naturally basic and have a pH of around 8. Ocean acidification means that the pH will decrease and will therefore become more acidic and less basic. However, overall they remain basic.

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How does the amount of carbon dioxide in the oceans affect shelled creatures?

Many **marine invertebrates** have shells (**exoskeletons**) that are made from calcium carbonate by a process called **calcification** (figure 4). As we have learned, calcium carbonate is a base that reacts with weak acids. Put simply, as the oceans become more acidic, these creatures are increasingly unable to grow their shells and exoskeletons, as well as the possibility of their shells slowly dissolving[1] (figure 5) by a process called **decalcification** (figure 6).

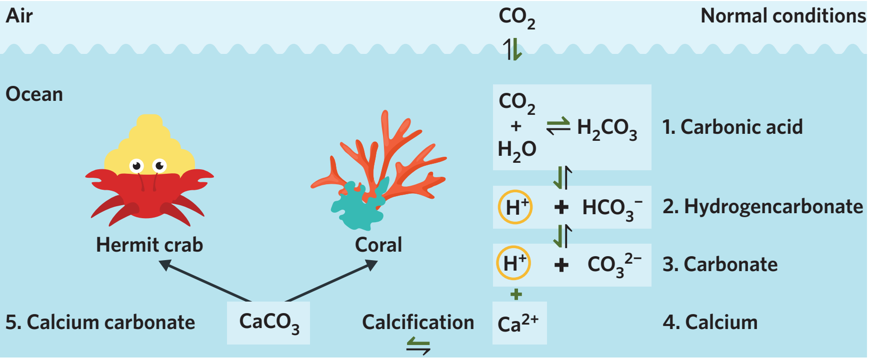

Figure 4 Natural acid-base calcification reactions

Figure 6 outlines the issue for these shelled creatures as the ocean becomes more acidic. Instead of forming calcium carbonate the acidic conditions favour the processes that form the hydrogencarbonate ion (HCO_3^-).

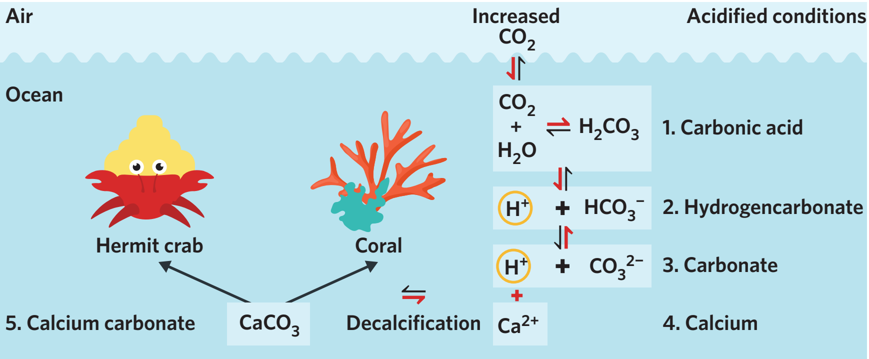

Figure 6 Changed acid-base decalcification process due to increased ocean acidification[2]

KEEN TO INVESTIGATE?

[1] **How does ocean acidification affect coral reefs?**
Search YouTube: How does ocean acidification affect coral reefs

Image: Daniel Wright98/Shutterstock.com

Figure 5 Stages of a sea shell dissolving by a neutralisation reaction

KEEN TO INVESTIGATE?

[2] **How exactly does shell building occur?**
Search: simulation Carbon Dioxide, Shell Building, and Ocean Acidification

Progress questions

Question 6

What is the main cause of the increase in ocean acidification?

A. The natural process of carbon dioxide dissolving into the oceans

B. The dissolving of excess carbon dioxide gas into the oceans generated by human activities

Question 7

Why is the formation of carbonic acid harmful to shelled creatures?

A. Because shells are made of a base

B. Because shells are made of an acid

Question 8

Decalcification is the process of

A. breaking down calcium carbonate.

B. forming calcium carbonate.

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Theory summary

- Rainwater is naturally a weak acid.
- Acid rain is a mixture of mainly carbonic, nitric and sulfuric acids.
- Combustion of fossil fuels is a major cause of acid rain.
- Ocean water is naturally basic.
- The oceans are a huge sink for a lot of the Earth's carbon dioxide.
- Further acidification of the oceans can reduce the ability of marine invertebrates to produce shells and exoskeletons.

The content in this lesson is considered fundamental prior knowledge to the extent of chemical reactions (Unit 3 AOS 2).

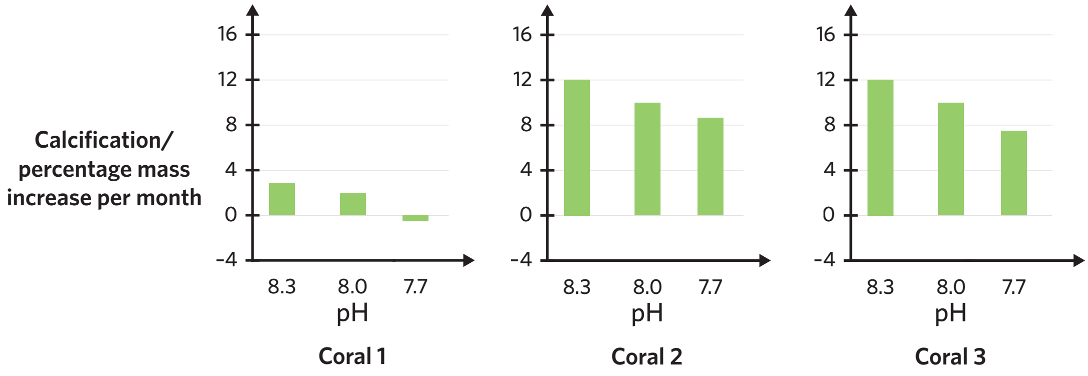
10E Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 9–11.

Coral reefs are very sensitive to changes in the pH of ocean water and these changes can reduce their ability to create calcium carbonate shells through a process called calcification. A series of experiments were conducted to see what effect changing the pH has on the amount of calcification that occurs. The pH of the ocean water that coral were growing in was altered by dissolving carbon dioxide in the water. Three different coral species were used at constant temperature and at three different pH values. The results are shown in the graphs.



Question 9 (1 MARK) 🌶

What relationship is being investigated in this experiment?

A. Temperature and percentage mass calcification change
B. pH and percentage mass calcification change
C. Temperature on different coral species
D. Global warming and coral calcification

Question 10 (1 MARK) 🌶

Decreasing the pH of the ocean water increases the

A. basicity of the ocean water.
B. alkalinity of the ocean water.
C. acidity of the ocean water.
D. neutrality of the ocean water.

Question 11 (3 MARKS) 🌶🌶

Explain what effect the different pH values have on the mass increase per month of the three coral species investigated.

Exam-style

Question 12 (1 MARK)

Increasing concentrations of atmospheric carbon dioxide are causing

A. a rise in the density of the atmosphere.

B. an increase in the ozone layer.

C. acidification of the oceans.

D. increased saltiness of the oceans.

Adapted from VCAA 2018 exam Multiple choice Q21

Question 13 (1 MARK)

Burning sulfur in dry air can produce

A. oxygen.

B. sulfuric acid.

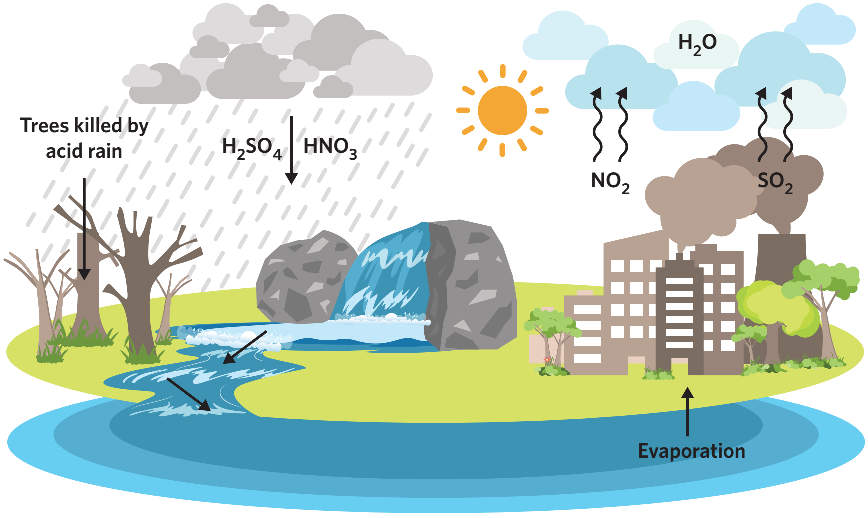
C. sulfur dioxide.

D. carbon dioxide.

VCAA 2006 Exam 2 Multiple choice Q11

Question 14 (10 MARKS)

Acid rain is both a current and future issue that affects the environment.



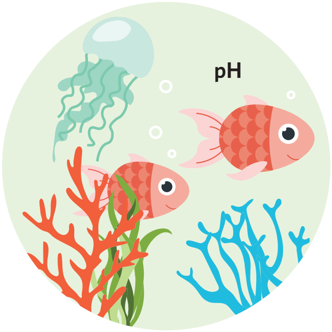
a. What are the main sources of the two acid forming gases and by what chemical process are these gases released? (2 MARKS)

b. Give the two overall balanced equations (with state symbols) that show the production of the two main acids found in acid rain. (2 MARKS)

c. Other than the negative effects shown in the image give two examples of the negative impact acid rain can have on society and explain using chemical terminology and equations. (6 MARKS)

Question 15 (5 MARKS)

Explain why the pH of the oceans is decreasing and identify one possible consequence of this decrease on marine life.



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Question 16 (3 MARKS)

The following article was published in Oceanus, 2010, and is called ‘Ocean Acidification: A Risky Shell Game.’

“Marine ecosystems—particularly those based on calcium-carbonate shell-building, such as coral or oyster reefs—could change with increasing atmospheric CO_2 (carbon dioxide),” said Justin Ries, a marine biogeochemist and lead author of the study, published online Dec. 1, 2009, in the journal *Geology*. Sensitive species could lose their protective shells and eventually die out, while other species that build stronger shells could become dominant in a future ocean that continues to absorb the buildup of CO_2 in the atmosphere caused by industrial emissions, deforestation, and other human activities.

Excess CO_2 dissolves into the ocean and is converted to corrosive carbonic acid, a process known as ocean acidification. At the same time, the CO_2 also supplies carbon that combines with calcium already dissolved in seawater to provide the main ingredient for shells—calcium carbonate ($CaCO_3$), the same material found in chalk and limestone.”

Source: Madin, K. (2010). *Ocean Acidification: A Risky Shell Game. How will climate change affect the shells and skeletons of sea life?* Oceanus.

a. What chemical characteristic of acidification reactions allows the oceans to become more acidic or basic over thousands of years? (1 MARK)

b. In the article it reads, ‘Sensitive species could lose their protective shells and eventually die out’. What is the specific chemical term given to the process that occurs when the shell of a creature is broken down in the ocean due to acidification? (1 MARK)

c. In the article it mentions that the ‘Excess CO_2 dissolves into the ocean and is converted to corrosive carbonic acid, a process known as ocean acidification.’ Give the chemical equation that outlines this process. (1 MARK)

Key science skills

Question 17 (8 MARKS)

The following data was collected from a sampling station in a large city. The researchers measured the pH of rain and the concentration of sulfuric acid in rainwater for a month. The table shows the data from each of the 12 rain events that occurred during the month.

Rain event	Sulfuric acid concentration (mg L^{-1})	pH
1	0.50	5.0
2	0.40	5.0
3	0.30	5.5
4	0.85	4.6
5	0.80	4.4
6	0.20	5.8
7	0.20	5.7
8	0.45	5.1
9	0.30	5.3
10	0.35	5.2
11	0.65	4.8
12	0.40	5.1

a. What type of data was collected in this investigation? (1 MARK)

b. What are the dependent and independent variables in this investigation? (1 MARK)

c. How many significant figures should any calculations of the above data be calculated to? (1 MARK)

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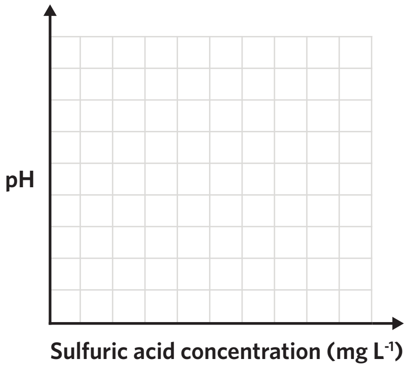
d. On a graph similar to the one shown, plot sulfuric acid concentration versus pH. Include an appropriate scale on each axis. (3 MARKS)

e. What conclusion can be drawn from your graph? (2 MARKS)

Adapted from VCAA 2006 Exam 2 Short answer Q2c,d

FROM LESSONS 16B, 16C & 16D



Questions from multiple lessons

Question 18 (6 MARKS)

When metals are left in the sun, they become very hot quickly. This means that, on a hot day, the exterior of a car can sometimes fry an egg.

a. Describe metallic bonding. (2 MARKS)

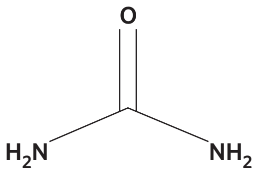
b. Explain why metals are good conductors of heat. (2 MARKS)

c. Explain why the property of heat conduction leads to a low specific heat capacity. (2 MARKS)

FROM LESSONS 3A & 9A

Question 19 (5 MARKS)

In mammals, one source of energy that is sometimes used is the oxidation of amino acids. During this process, ammonia, NH_3, is produced which is highly toxic and is immediately converted into urea, whose structure is shown.



a. Is urea soluble in water? Justify your answer with reference to a diagram. (3 MARKS)

b. Based on its solubility properties, explain how urea can be transported in the blood stream and then excreted in urine. (2 MARKS)

FROM LESSONS 2A & 2B

Hints

12. Carbon dioxide readily dissolves in water.
13. The conditions in the atmosphere are important.
14a. The main source of energy production and pollution on Earth.
14b. This includes the gas in its natural state reaction with oxygen and water.
14c. Vegetation and buildings can be chemically affected by acid rain.
15. Human activities and the release of extra gases are a major factor.
16a. Not all reactants in chemical reactions are used up.
16b. Calcium carbonate is the main chemical in shells.
16c. Carbonic acid is formed in the oceans.
17a. Numbers or words determine the type of data.
17b. Independent is the variable that is chosen.
17c. The least amount of significant figures in any calculation is important.
17d. A line graph suits this sort of data.
17e. Use the axes and gradient.
18a. Metals have a unique form of bonding.
18b. Metals have a sea of delocalised electrons.
18c. Heat conduction and specific heat capacity are related.
19a. Hydrogen bonds can occur when hydrogen is bonded to either fluorine, oxygen or nitrogen.
19b. Blood and urine consist mostly of water.

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Chapter 10 review

Multiple choice (10 MARKS)

Question 1 (1 MARK)

Which of the following solutions contains the largest concentration of $OH^-(aq)$ ions?

A. Solution with a pH of 7.3

B. Solution with a pH of 0.3

C. Solution with a pH of 13.8

D. Solution with a pH of 10.2

Question 2 (1 MARK)

Which of the following molecules contributes to acid rain formation?

A. NH_3

B. CH_4

C. O_3

D. SO_2

Question 3 (1 MARK)

Which of the following chemical equations represents a Brønsted-Lowry acid-base reaction?

A. $HPO_4^{2-}(aq) + HBr(aq) \rightarrow Br^-(aq) + PO_4^{3-}(aq)$

B. $HClO_4(aq) + NH_3(aq) \rightarrow ClO_4^-(aq) + NH_4^+(aq)$

C. $C_2H_6(l) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$

D. $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$

Question 4 (1 MARK)

Identify the species containing phosphorus that would be present in the reaction vessel if 12 mol of $HPO_4^{2-}(aq)$ were added to an aqueous solution.

A. $HPO_4^{2-}(aq)$, $PO_4^{3-}(aq)$

B. $HPO_4^{2-}(aq)$, $H_2PO_4^-(aq)$, $H_3PO_4(aq)$

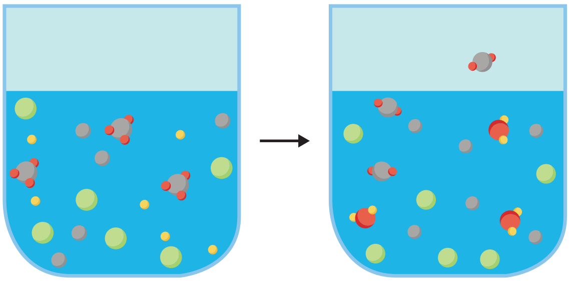
C. $PO_4^{3-}(aq)$, $HPO_4^{2-}(aq)$, $H_2PO_4^-(aq)$, $H_3PO_4(aq)$

D. $PO_4^{3-}(aq)$, $HPO_4^{2-}(aq)$, $H_2PO_4^-(aq)$, $H_3PO_4(aq)$, $H_4PO_4^+(aq)$

Question 5 (1 MARK)

A representation of a reaction occurring between an acid and a metal carbonate is given.

Which of the following represents the balanced ionic equation for this reaction?



A. 

B. 

C. 

D. 

Question 6 (1 MARK)

Which of the following species would be classified as both amphiprotic and polyprotic?

A. $H_2PO_4^-(aq)$

B. $H_2SO_4(aq)$

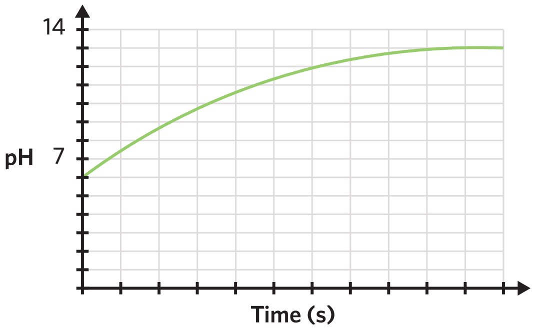
C. $NO_3^-(aq)$

D. $H_3PO_4(aq)$

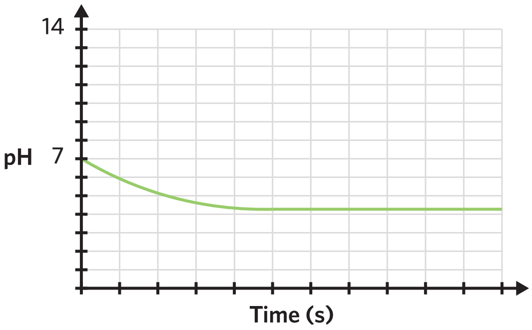
Question 7 (1 MARK)

320 g of ethanoic acid, $CH_3COOH(aq)$, is added to 90.0 L of pure water. Which of the following graphs would be the most likely representation of how the solution's pH would change over time?

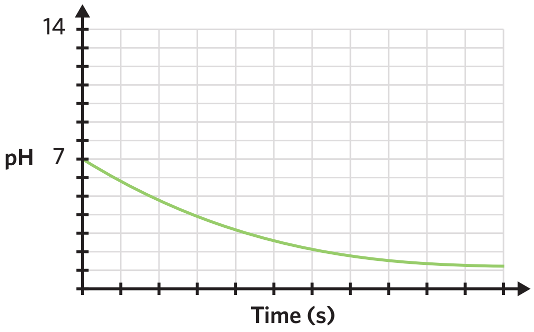
A.



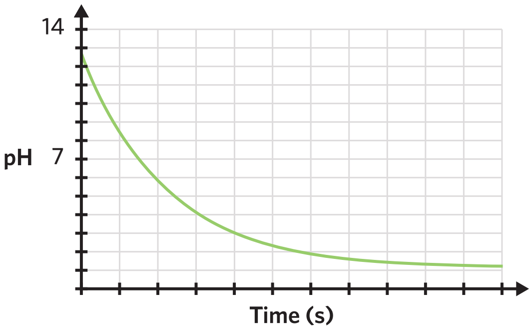
B.



C.



D.



Question 8 (1 MARK)

Consider the following reaction:

$H_2SeO_3(aq) + H_2O(l) \rightarrow HSeO_3^-(aq) + H_3O^+(aq)$

In this reaction $HSeO_3^-(aq)$ is

A. the conjugate base of $H_2O(l)$.

B. the conjugate base of $H_2SeO_3(aq)$.

C. the conjugate acid of $H_3O^+(aq)$.

D. the conjugate acid of $H_2SeO_3(aq)$.

Question 9 (1 MARK)

Which of the following is the balanced overall equation for the reaction between solid magnesium and iodic acid, $HIO_3(aq)$?

A. $HIO_3(aq) + Mg(s) \rightarrow Mg(IO_3)_2(aq) + H_2(g)$

B. $2HIO_3(aq) + Mg(s) \rightarrow Mg(IO_3)_2(aq) + H_2(g)$

C. $2H^+(aq) + Mg(s) \rightarrow Mg^{2+}(aq) + H_2(g)$

D. Solid magnesium would not react with iodic acid.

Question 10 (1 MARK)

The human stomach has an acidic pH of 3.40 due to the presence of hydrochloric acid, HCl. This destroys 99% of bacteria that are ingested, but people can still become ill and may die if they eat rotten or dangerously contaminated food. Turkey vultures rarely have this problem, as their stomach pH is slightly above zero at 0.40, which is lower than car battery acid.

Which of the following statements are **not** correct?

I. Humans and turkey vultures must have different acids present in their stomachs.
II. Turkey vulture stomachs are over 10 000 times more acidic than human stomachs.
III. Human stomachs have a greater concentration of $OH^-(aq)$ ions than turkey vulture stomachs.
IV. Drinking a glass of water will briefly increase the pH of a human's stomach.
V. Turkey vulture stomachs have a lower concentration of $H^+(aq)$ ions than human stomachs.

A. I, II, IV
B. III, IV
C. II, III, V
D. I, II, V

Short answer (30 MARKS)

Question 11 (3 MARKS)

The following table shows the colours observed when 3 unknown indicators are added to 3 different solutions.

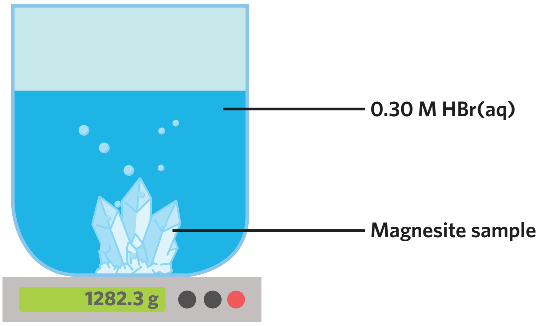
	Solution A – pH = 1.0	Solution B – pH = 9.6	Solution C – pH = 4.7
Colour Observed	Colourless	Red	Blue

Identify the indicators that have been added to each solution.

Question 12 (6 MARKS)

Hydrobromic acid, HBr, is an even stronger acid than hydrochloric acid, HCl. Since it is such a strong acid, it is useful in quickly extracting certain metal ores during industrial mining processes.

A chemist has recently uncovered a small deposit of magnesite in her backyard, which is a mineral with the chemical formula $MgCO_3(s)$. However, when she analyses her sample, she discovers that there is actually an enormous golden nugget buried within the magnesite! To remove the magnesite exterior, the chemist places the sample in a large beaker of 0.30 M hydrobromic acid.



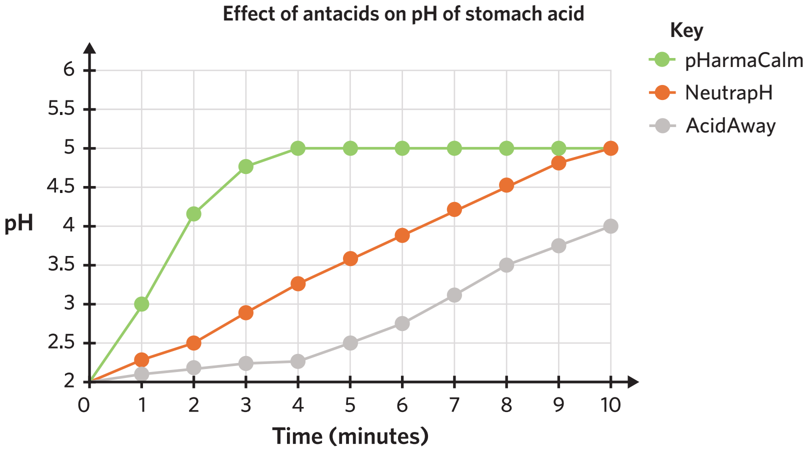
a. Give the balanced overall equation for the reaction occurring in this experiment. (1 MARK)

b. Give the balanced ionic equation for the reaction occurring in this experiment. (1 MARK)

c. Identify the conjugate base of hydrobromic acid. If a solution of this conjugate base were added to a vessel of pure water, would the concentration of $OH^-(aq)$ ions increase by any noticeable extent? Explain. (4 MARKS)

Question 13 (8 MARKS)

Antacid tablets treat indigestion by neutralising stomach acid. A study was done to test the effectiveness of 3 brands of antacid tablet – pHarmaCalm, NeutrapH and AcidAway – and the graph shown was obtained.



a. According to the graph, which of the antacid tablets was least effective at neutralising the pH of stomach acid? Justify your answer. (2 MARKS)

b. According to the graph, which of the antacid tablets would be most effective for treating indigestion? Explain your answer. (3 MARKS)

c. Milk of magnesia is a common antacid composed of magnesium hydroxide, $Mg(OH)_2(aq)$. Give the balanced overall equation for the neutralisation reaction between magnesium hydroxide and hydrochloric acid. (1 MARK)

d. Based on your understanding of neutralisation reactions, would magnesium hydroxide be classified as an acid or base? Explain your answer. (2 MARKS)

Question 14 (7 MARKS)

Acid rain and ocean acidification are predicted to have severe environmental impacts in the future. Sulfur dioxide, $SO_2(g)$, and nitrogen dioxide, $NO_2(g)$, are the main gases that result from human activities that contribute to the formation of increasingly acidic rain and ocean water.

a. Which other gas is responsible for the formation of acidic water? (1 MARK)

b. Give the overall balanced equation for the formation of sulfuric acid in rainwater. (2 MARKS)

c. What is the chemical process of making calcium carbonate shells called? (1 MARK)

d. As ocean water becomes more acidic, which ion is produced in larger quantities? (1 MARK)

e. Identify and explain using a chemical equation one negative impact an acid-base reaction can have on plant growth. (2 MARKS)

Question 15 (6 MARKS)

Brass is an alloy of zinc and copper, though the exact proportions of each metal vary depending on the kind of brass. Prince's metal is the name given to a version of brass that is 75% copper and 25% zinc by weight. Due to its vibrant yellow colour, Prince's metal is often used as a cheaper substitute for gold.

Unfortunately, a jeweller has mixed up an ingot of gold and an ingot of Prince's metal and, due to his poor eyesight, cannot visually tell the difference between the two. Fortunately, he remembers his year 11 chemistry and prepares an experiment to solve his problem. The experimental set-up is shown.

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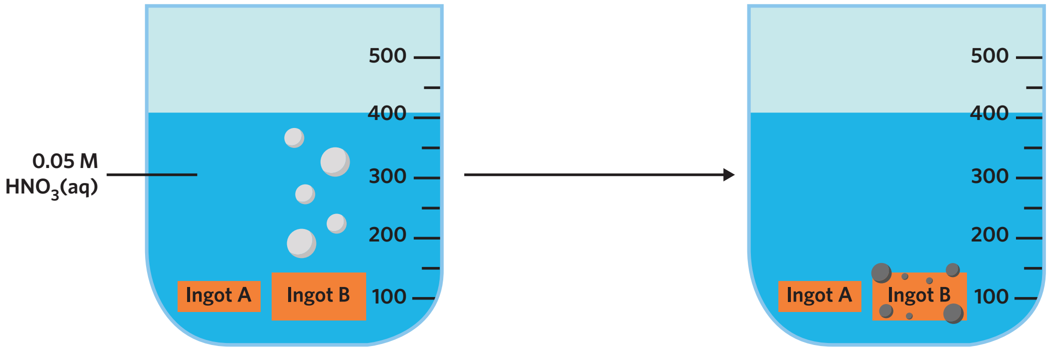


a. Give the balanced overall equation for the reaction(s) occurring in this experiment. (1 MARK)

b. Give the balanced ionic equation for the reaction(s) occurring in this experiment. (1 MARK)

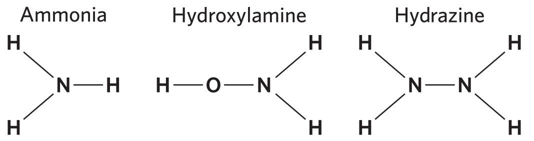
c. Determine the identity of each of the two ingots. Justify your answer with reference to the experimental observations. (2 MARKS)

d. When the reaction had finished, would the pH of the solution have increased, decreased or stayed the same from the beginning of the experiment? Explain. (2 MARKS)

Key science skills (10 MARKS)

Question 16 (10 MARKS)

Vignesh is in the process of designing an experiment to measure the strength of a number of different basic solutions. These bases are ammonia, NH_3, hydroxylamine, $HONH_2$, and hydrazine, N_2H_4. The structures of these molecules are shown.



Vignesh's experimental report is due tomorrow and so he has quickly assembled the following materials. However, not all of the materials will need to be used in the experiment.

Materials

Solutions and beakers:

- 750 mL glass beakers (3)
- 500 mL of 1.0 M $NH_3(aq)$
- 90 mL of 1.2 M HCl(aq)
- 500 mL of 0.70 M $NH_3(aq)$
- 250 mL of 2.0 M $N_2H_4(aq)$
- 500 mL of 1.0 M $HONH_2(aq)$
- 500 mL measuring cylinder
- Deionised water

Measurement techniques

- Litmus paper (5 strips)
- pH probe and recording device
- Phenolphthalein indicator

a. Write the method of a valid experiment Vignesh could conduct to determine the relative strengths of each of the bases under investigation. Ensure that as many variables are controlled for as possible. (2 MARKS)

b. Explain your reasoning for using/not using the three measurement techniques that Vignesh assembled. (3 MARKS)

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c. Explain how the data collected from the selected measurement techniques could be used to determine the relative strengths of each of the basic solutions. (3 MARKS)

d. Unfortunately, Vignesh was not able to design a valid experiment and had to search online to determine each base's strength. During his research, he discovered the following list of weak bases and their strengths.

Increasing strength ↑

Butylamine	$C_4H_9NH_2$
Methylamine	CH_3NH_2
Ammonia	NH_3
Hydrazine	N_2H_4
Strychnine	$C_{21}H_{22}N_2O_2$
Morphine	$C_{17}H_{19}NO_3$
Hydroxylamine	$HONH_2$
Pyridine	C_5H_5N
Aniline	$C_6H_5NH_2$

Complete the following table using the words 'highest', 'lowest' and 'middle' to classify how the concentrations of OH^-(aq) and H_3O^+(aq) ions would appear in each solution. (2 MARKS)

	Ammonia	**Hydroxylamine**	**Hydrazine**
$[OH^-]$			
$[H_3O^+]$			

FROM LESSONS 16B & 16E

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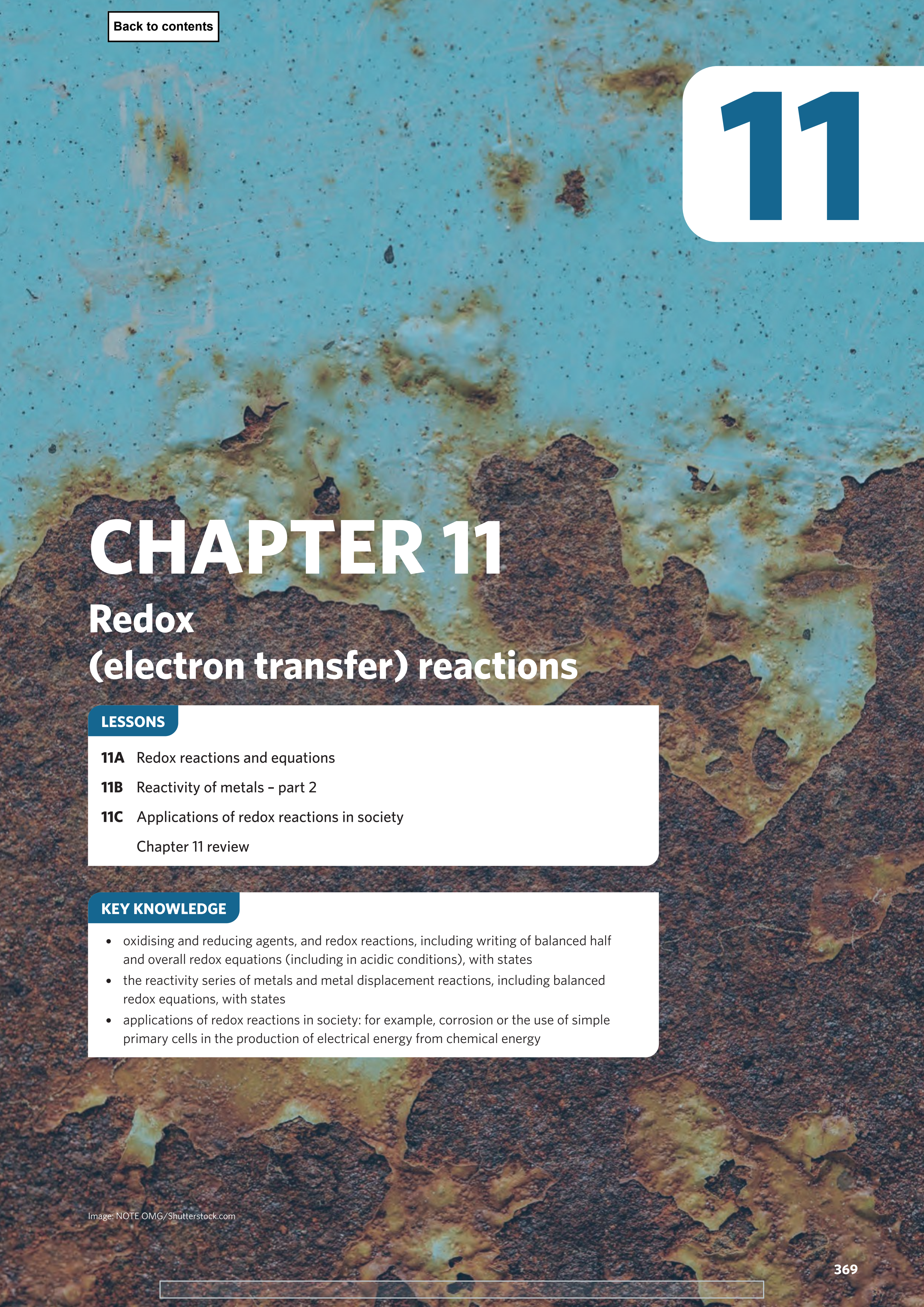

11

CHAPTER 11

Redox (electron transfer) reactions

LESSONS

KEY KNOWLEDGE

- oxidising and reducing agents, and redox reactions, including writing of balanced half and overall redox equations (including in acidic conditions), with states
- the reactivity series of metals and metal displacement reactions, including balanced redox equations, with states
- applications of redox reactions in society: for example, corrosion or the use of simple primary cells in the production of electrical energy from chemical energy

Image: NOTE OMG/Shutterstock.com

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11A Redox reactions and equations

STUDY DESIGN DOT POINT

- oxidising and reducing agents, and redox reactions, including writing of balanced half and overall redox equations (including in acidic conditions), with states

11A 11B 11C

2.1.10.1 & 2.1.10.2 Defining redox reactions & oxidising and reducing agents

2.1.10.3 & 2.1.10.4 Identifying redox reactions & writing redox half-equations

2.1.10.5 Writing overall redox equations

ESSENTIAL PRIOR KNOWLEDGE

1A Electrons as subatomic particles

4A Electrons can be transferred between atoms

See questions 73–74.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: Engineer studio/Shutterstock.com

Why are bones the only parts that remain of prehistoric creatures?

The organic matter in prehistoric remains decomposes over time due to **redox reactions** with oxygen in the air, leaving only bones behind. In this lesson, we will learn about redox reactions like decomposition, and how we can write the half and full equations for redox reactions.

KEY TERMS AND DEFINITIONS

Half-equation reduction or oxidation equation of a complete redox reaction

Oxidation chemical reaction in which a chemical species loses electrons

Oxidation state number assigned to an atom that can be used to determine the movement of electrons in redox reactions

Oxidise cause a chemical species to lose electrons

Oxidising agent chemical species that oxidises another substance by accepting electrons

Polyatomic ion ions made up of a group of atoms covalently bonded together

Redox reaction chemical reaction involving the transfer of electrons between chemical species

Reduce cause a chemical species to gain electrons

Reducing agent chemical species that reduces another substance by giving away electrons

Reduction chemical reaction in which a chemical species gains electrons

Defining redox reactions & oxidising and reducing agents 2.1.10.1 & 2.1.10.2

Redox reactions involve the balanced transfer of electrons between chemical species.

How are redox reactions defined?

Redox reactions are a very common type of chemical reaction that involve the transfer of electrons from one chemical species to another. As shown in figure 1, the term redox is a combination of two chemical processes: **reduction** and **oxidation**.

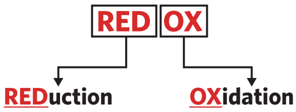

Figure 1 Formation of the term redox

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As such, a redox reaction is actually the combination of a reduction and oxidation reaction occurring simultaneously.

Consider the reaction occurring between atoms A and B shown in figure 2.

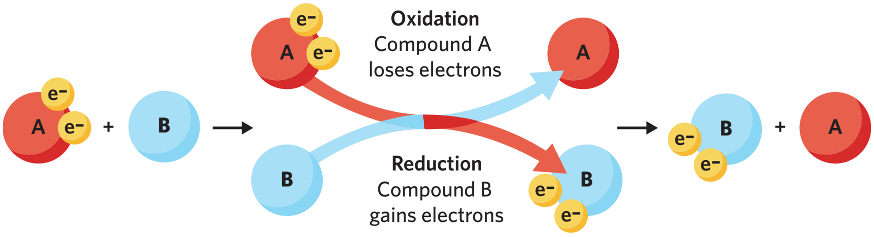

Figure 2 Model for redox reactions

As we can see, compound A loses 2 electrons, and compound B gains 2 electrons. As such, compound A is said to undergo oxidation, and compound B is said to undergo reduction. In general, oxidation is defined as the loss of electrons, and reduction is defined as the gain of electrons.

USEFUL TIP

Oxidation can also be defined as the gain of oxygen or loss of hydrogen. Similarly, reduction can also be defined as the loss of oxygen or gain of hydrogen.

STRATEGY

OIL RIG is a useful mnemonic for remembering the two parts to any redox reaction:

Oxidation **I**s **L**oss (of electrons)

Reduction **I**s **G**ain (of electrons)

Image: Thaiview/Shutterstock.com

What are oxidising and reducing agents?

So far, we have learned that redox reactions always involve two types of reactions: oxidation and reduction, both occurring simultaneously. As a result, the chemical species in each half-reaction and in the whole chemical equation can be classified based on whether they have gained or lost electrons in the reaction.

Oxidising agents (or oxidants) are chemical species that **oxidise** another chemical species. In other words, the oxidising agent causes a species to lose electrons.

In contrast, **reducing agents** (or reductants) are chemical species that **reduce** another chemical species. In other words, the reducing agent causes a species to gain electrons.

MISCONCEPTION

'Oxidising agents are oxidised.'

Since redox reactions involve the balanced transfer of electrons between species, oxidising agents must gain electrons to balance the electrons they cause other species to lose (during oxidation). Therefore, oxidising agents are reduced – not oxidised.

Progress questions

Question 1

A redox reaction is a reaction involving either reduction or oxidation.

A. True

B. False

Question 2

Reduction is the __________ of electrons.

A. gain

B. loss

Continues →

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Question 3

Oxidising agents

A. undergo oxidation.

B. gain electrons.

Identifying redox reactions & writing redox half-equations 2.1.10.3 & 2.1.10.4

Oxidation states are useful tools that we can use to help identify redox reactions.

How can oxidation states be used to identify redox reactions?

One technique that can be used to identify redox reactions is the use of oxidation states. Oxidation states indicate the charge that an atom may have as a result of the movement of electrons during a redox reaction. A change in oxidation state of a chemical species indicates that electrons have either been lost or gained, and therefore that a redox reaction has occurred. Hence, oxidation states also allow us to identify which species has undergone oxidation and which species has undergone reduction. Table 1 shows some common rules for determining the oxidation state of different forms of elements.

Table 1 Common rules to determine the oxidation state of elements

Rule	Example
Oxidation state of a free element (an element by itself or bonded to itself) is always zero	O_2, C, Cl_2
Oxidation state of a simple ion is equal to the charge of the ion	Na^+, Cl^-, O^{2-}, Al^{3+}
The sum of oxidation states in a neutral compound is 0	KCl, CO_2, $MgSO_4$
The sum of oxidation states in a **polyatomic ion** is equal to the charge of the ion	SO_4^{2-}, NH_4^+, MnO_4^-
The normal oxidation state for hydrogen in a compound is +1*	HCl, HF, NH_3
The normal oxidation state for oxygen in a compound is −2*	H_2O, CO, NO_3^-

*There are exceptions to these rules which will be introduced in VCE chemistry units 3 & 4.

USEFUL TIP

Oxidation states can also be referred to as oxidation numbers.

MISCONCEPTION

'−6 is a higher oxidation state than −2.'

The more negative an oxidation state, the lower the oxidation state of the given atom is. Similarly, the more positive an oxidation state, the higher it is.

Some of the trends in the oxidation states of individual elements of the periodic table are shown in figure 3.

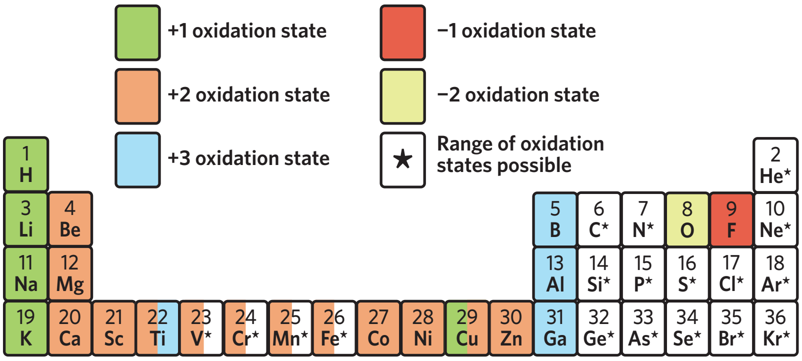

Figure 3 Periodic table showing the oxidation states of some common elements

Let's take a look at an example of how to determine these oxidation states. Carbon dioxide, CO_2, is a compound that can take part in redox reactions. Based on the rules outlined in table 1, there are a couple of ideas that we need to consider:

- Carbon dioxide is a neutral compound (it has no overall charge). Therefore, the sum of oxidation states of the one carbon atom and the two oxygen atoms need to add up to 0.
- Each of the two oxygen atoms normally has an oxidation state of −2.

Considering we know the oxidation state of each oxygen atom, we can use this information to find the oxidation state of the carbon atom in CO_2, as shown in figure 4.

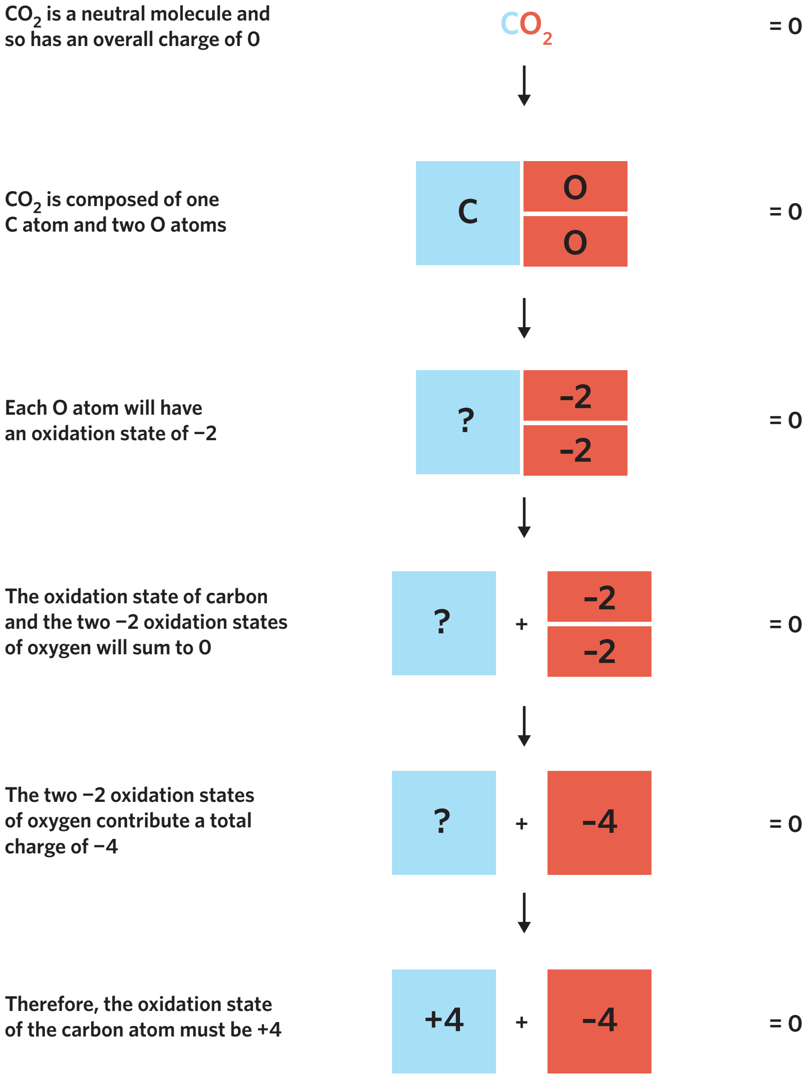

Figure 4 Calculating the oxidation state of the carbon atom in CO_2

MISCONCEPTION

'In CO_2, the oxidation state of oxygen is −4.'

In a molecule or compound, oxidation states are assigned to each atom of the element – not the element as a whole. Therefore, in the case of CO_2, the oxidation state for each oxygen atom is −2, as the two oxygen atoms together cancel out the +4 oxidation state of the carbon atom.

USEFUL TIP

When attempting to determine the oxidation state of an atom, it's important to note the most electronegative atom in a compound or molecule will have a negative oxidation state. For example, in NO_2, the oxygen atoms will both have negative oxidation states, whereas in NH_3, the nitrogen will have a negative oxidation state.

WORKED EXAMPLE 1

Sulfur trioxide, SO_3, is a gas that contributes to the greenhouse effect. Determine the oxidation state of the sulfur atom in a molecule of SO_3.

What information is presented in the question?

The molecular formula of sulfur trioxide, SO_3.

What is the question asking us to do?

Find the oxidation state of the sulfur atom in SO_3.

What strategies do we need in order to answer the question?

1. Recall the oxidation state rule for oxygen.
2. Calculate the oxidation state of sulfur.

Answer

According to the rules in table 1, oxygen has an oxidation state of −2. However, sulfur can have a range of possible oxidation states. Sulfur trioxide is a neutral compound with an overall charge equal to 0. Since the compound has 3 oxygen atoms, the oxidation state of sulfur in SO_3 must be +6 to give an overall charge of 0.

Using oxidation states, we can determine

- whether a redox reaction has taken place, shown by a change in oxidation states.
- which species has undergone oxidation, shown by an increase in oxidation state.
- which species has undergone reduction, shown by a decrease in oxidation state.

MISCONCEPTION

'Overall oxidation states in a compound are always whole numbers.'

Overall oxidation states in a compound do not always have to be whole numbers. The tetrathionate ion, $S_4O_6^{2-}$, has four sulfur ions with a combined charge of +10, which means that the average is +2.5 per sulfur ion. This occurs because, in tetrathionate, sulfur exists as 2 ions of +2 and 2 ions of +3.

How can we use oxidation states to write redox half-equations?

So far, we have investigated the processes of oxidation (the loss of electrons) and reduction (the gain of electrons). These two processes occur simultaneously, with oxidation reactions coupled to reduction reactions that allow for the transfer of electrons between chemical species. A redox **half-equation** is a chemical equation that describes either the individual oxidation or reduction reaction that is occurring within the overall redox reaction.

As we learned in the previous section of this lesson, changes in oxidation states are useful indicators of whether a redox reaction has taken place, and can be used to determine which species have undergone oxidation and reduction. Consider the following oxidation reaction.

$$\overset{0}{K}(s) \rightarrow \overset{+1}{K^+}(s)$$

Oxidation is taking place as the oxidation state of potassium increases from 0 to +1. Since a loss of one electron corresponds to an increase in oxidation state by one, a total of one electron must have been lost for each potassium atom in this oxidation reaction.

$$K(s) \rightarrow K^+(s) + e^-$$

USEFUL TIP

Redox half-equations must be balanced both in the number of atoms of each species, but also in the amount of charge present on each side of the equation.

We can use a similar line of reasoning to determine how many electrons are gained during a reduction reaction. Consider the following reduction reaction.

$$\overset{0}{Cl_2}(g) \rightarrow \overset{-1}{2Cl^-}(s)$$

Reduction is taking place as the oxidation state of each chlorine atom decreases from 0 to −1. Since a gain of one electron corresponds to a decrease in oxidation state by one, a total of two electrons must have gained, one by each chlorine atom, in this reduction reaction.

$$Cl_2(g) + 2e^- \rightarrow 2Cl^-(s)$$

USEFUL TIP

Recall from lesson 1B that electrons are not molecules or atoms, and so they do not need to be written with a state (solid, liquid, gas or aqueous) in chemical reactions.

WORKED EXAMPLE 2

Write the balanced redox half-equation for the conversion of $O_2(g)$ to $O^{2-}(aq)$.

What information is presented in the question?

The molecular formulas of $O_2(g)$ and $O^{2-}(aq)$.

What is the question asking us to do?

Write the balanced redox half-equation for the production of $O^{2-}(aq)$ from $O_2(g)$.

What strategies do we need in order to answer the question?

1. Determine the oxidation states of $O_2(g)$ and $O^{2-}(aq)$.
2. Identify whether an oxidation or reduction reaction is taking place.
3. Balance the half-equation with respect to the number of atoms and charge.

Answer

$O_2(g)$ is a free element and so has an oxidation state of 0. $O^{2-}(aq)$ is a simple ion, so its oxidation state is equal to its charge, −2.

$$\overset{0}{O_2}(g) \rightarrow \overset{-2}{O^{2-}}(aq)$$

There is a decrease in oxidation state, from 0 to −2, indicating the gain of negatively charged electrons. Therefore, this is a reduction reaction.

First, we balance the number of atoms of each element:

$$O_2(g) \rightarrow 2O^{2-}(aq)$$

On the reactants' side of the equation, there is a charge of 0.

On the products' side of the equation, there is a charge of $2 \times (-2) = -4$.

Therefore, four electrons must be placed on the reactants' side, producing the balanced reduction half-equation.

$$O_2(g) + 4e^- \rightarrow 2O^{2-}(aq)$$

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Polyatomic ions - ions that are composed of more than one atom – can also accept or donate electrons to take part in redox reactions. Some of the most common polyatomic ions that take part in redox reactions are strong oxidising agents used in laboratory settings. These oxidising agents include the permanganate (VII) ion, $MnO_4^-(aq)$, and the dichromate ion (VI), $Cr_2O_7^{2-}(aq)$.

As these ions act as strong oxidising agents, they themselves must undergo reduction when taking part in a redox reaction and have observable changes in colour in acidic conditions.

$$\underset{\text{purple}}{\overset{+7}{MnO_4^-(aq)}} \rightarrow \underset{\text{colourless}}{\overset{+2}{Mn^{2+}(aq)}}$$

Manganese readily accepts 5 electrons and is a strong oxidising agent.

$$\underset{\text{orange}}{\overset{+6}{Cr_2O_7^{2-}(aq)}} \rightarrow \underset{\text{green}}{\overset{+3}{Cr^{3+}(aq)}}$$

Chromium readily accepts 3 electrons and is a strong oxidising agent.

We can instantly see that changing just the number of manganese or chromium ions and electrons will not result in a balanced half-equation due to the presence of the multiple oxygen atoms. In fact, balancing these half-equations requires us to follow a specific sequence of balancing steps involving $H^+(aq)$ ions and $H_2O(l)$ molecules as well. To help remember the order in which these ions and molecules are added, we can use the following acronym.

USEFUL TIP

In Units 1 & 2 VCE chemistry, we will only be writing equations for redox reactions occurring in acidic solution. This allows us to assume an excess of water molecules and hydrogen ions are present to help us balance redox equations.

Key elements need to be balanced.

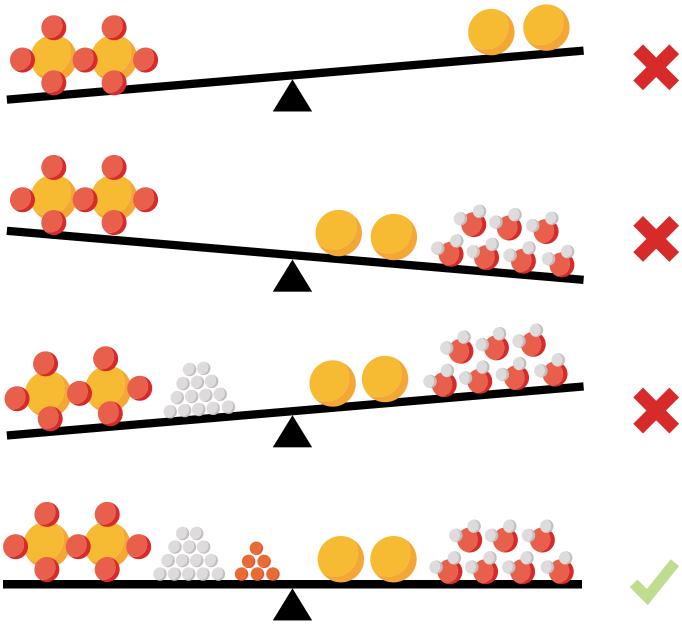
Oxygen atoms need to be balanced with water molecules, $H_2O(l)$.

Hydrogen atoms need to be balanced with hydrogen ions when in acidic conditions, $H^+(aq)$.

Electrons need to be added to balance charge on each side of the equation, e^-.

States need to be included.

Figure 5 shows how KOHES can be applied to the reduction equation for the dichromate ion.



K $Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$

O $Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$

H $Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$

E
S $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$

Figure 5 Balancing the half-equation for the reduction of $Cr_2O_7^{2-}(aq)$ to $2Cr^{3+}(aq)$

USEFUL TIP

When balancing the amount of charge in half-equations, remember that the amount of charge contributed by a chemical species in the equation depends on both the coefficient of the compound and its charge. For example, $2Cr^{3+}(aq)$ will contribute a $2 \times 3 = +6$ total charge.

WORKED EXAMPLE 3

Hydrogen sulfite, HSO_3^-, is oxidised to form the polyatomic sulfate ion, SO_4^{2-}, in an acidic aqueous solution. Write the balanced oxidation half-equation for this reaction.

What information is presented in the question?

The molecular formulas of hydrogen sulfite, HSO_3^-, and sulfate, SO_4^{2-}.

The oxidation reaction is occurring in an acidic aqueous solution.

What is the question asking us to do?

Write the balanced half-equation for the oxidation of HSO_3^- to SO_4^{2-}.

What strategies do we need in order to answer the question?

1. Balance key elements.
2. Balance oxygen atoms using H_2O.
3. Balance hydrogen atoms using H^+.
4. Balance charge using electrons.
5. Add states.

Answer

The key element sulfur is already balanced:

$HSO_3^- \rightarrow SO_4^{2-}$

H_2O is added to balance the four oxygen atoms on the products' side:

$HSO_3^- + H_2O \rightarrow SO_4^{2-}$

$3H^+$ is added to balance the three hydrogen atoms on the reactants' side:

$HSO_3^- + H_2O \rightarrow SO_4^{2-} + 3H^+$

2 electrons are added to balance the −1 charge on the reactants' side:

$HSO_3^- + H_2O \rightarrow SO_4^{2-} + 3H^+ + 2e^-$

Appropriate states are included for all ions and covalent compounds:

$HSO_3^-(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + 3H^+(aq) + 2e^-$

STRATEGY

The location of electrons in a half-equation can be used to determine whether the equation represents a reduction or an oxidation reaction. If the electrons are on the left hand side of the equation, reduction is occurring. If the electrons are on the right hand side of the equation, oxidation is occurring.

Progress questions

Question 4

What is the oxidation state of nitrogen in ammonia, $NH_3(g)$?

A. −3

B. +3

Question 5

An increase in the oxidation state of a chemical species means that it has undergone

A. oxidation.

B. reduction.

Question 6

For a reduction half-equation, electrons will appear on the ___________ of the chemical equation because electrons are ___________ during a reduction reaction.

A. left-hand side, gained

B. right-hand side, lost

Question 7

Which of the following represents a balanced half-equation?

A. $SO_4^{2-}(aq) + 2H^+(aq) \rightarrow SO_2(g) + H_2O(l)$

B. $SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow SO_2(g) + 2H_2O(l)$

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Writing overall redox equations 2.1.10.5

Overall redox equations are formed from the combination of the individual oxidation and reduction half-equations during a redox reaction.

How can we combine half equations to create overall equations for redox reactions?

Since oxidation and reduction half-equations are always coupled together, an overall redox equation that combines these two half-equations can provide an overview of the reactants and products of a redox reaction.[1]

Consider the oxidation and reduction half-equations that we looked at previously.

Oxidation: $K(s) \rightarrow K^+(s) + e^-$

Reduction: $Cl_2(g) + 2e^- \rightarrow 2Cl^-(s)$

Since redox reactions involve the balanced transfer of electrons, we can alter the equation for the oxidation of K(s) so that there are the same number of electrons on both sides of the equation.

Oxidation: $2 \times K(s) \rightarrow K^+(s) + e^-$

$2K(s) \rightarrow 2K^+(s) + 2e^-$

Now that there is an equal number of electrons on each side, we can add the two half-equations together, producing the overall equation.

Overall: $2K(s) + Cl_2(g) + \cancel{2e}^- \rightarrow 2K^+(s) + 2Cl^-(aq) + \cancel{2e}^-$

Since there are two electrons present on both sides of the equation, they can be removed, leaving us with the final overall balanced redox equation.

Overall: $2K(s) + Cl_2(g) \rightarrow 2K^+(s) + 2Cl^-(aq)$

MISCONCEPTION

'Electrons will be present in an overall redox equation.'

No additional electrons are consumed or produced during a redox reaction. Therefore, overall redox equations will not have electrons present on either the reactants' or products' side of the equation.

KEEN TO INVESTIGATE?

[1] How are redox reactions used in rocket fuel?
Search: What Makes a Rocket Soar?

USEFUL TIP

Electrons, as well as any other chemical species (such as $H^+(aq)$ or $H_2O(l)$), that appear on both sides of a chemical equation, can be removed during the balancing process.

WORKED EXAMPLE 4

A piece of solid cadmium is added to a solution of acidified potassium dichromate, resulting in the following two half-reactions occurring.

(1) $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$

(2) $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$

What is the balanced overall redox equation for the reaction taking place?

What information is presented in the question?

The balanced half-equations for oxidation and reduction.

What is the question asking us to do?

Determine the balanced overall redox equation for this reaction.

What strategies do we need in order to answer the question?

1. Ensure that there is an equal number of electrons in each half-equation.
2. Add the two reactions together.
3. Remove any chemical species that are present on both sides of the overall redox equation.

Answer

The reduction half-equation of $Cr_2O_7^{2-}(aq)$ shows that $6e^-$ are consumed. The oxidation half-equation for Cd(s) shows that only $2e^-$ are being produced. We require an equal number of electrons to be consumed and produced. Therefore, half-equation (2) must be multiplied by a factor of three:

$3 \times (Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-)$

$3Cd(s) \rightarrow 3Cd^{2+}(aq) + 6e^-$

We can now add this half-equation to half-equation (1).

$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 3Cd(s) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 3Cd^{2+}(aq) + 6e^-$

As there are six electrons present on both sides of the equation, they can be removed, leaving the final balanced overall redox equation.

$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 3Cd(s) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 3Cd^{2+}(aq)$

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WORKED EXAMPLE 5

Using the half-equations provided, give the overall equation for the redox reaction between permanganate ions, $MnO_4^-(aq)$, and iron, $Fe(s)$.

Oxidation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Reduction: $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

What information is presented in the question?

The equation for the oxidation of iron.

The equation for the reduction of permanganate ions.

What is the question asking us to do?

Give the overall equation for the redox reaction between permanganate ions and iron.

What strategies do we need in order to answer the question?

1. Balance the number of electrons in each equation.
2. Combine the two half-equations.
3. Remove common terms on each side of the equation.

Answer

Oxidation $5 \times (Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-)$

$5Fe(s) \rightarrow 5Fe^{2+}(aq) + 10e^-$

Reduction: $2 \times (MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l))$

$2MnO_4^-(aq) + 16H^+(aq) + 10e^- \rightarrow 2Mn^{2+}(aq) + 8H_2O(l)$

Combined: $5Fe(s) + 2MnO_4^-(aq) + 16H^+(aq) + \cancel{10e^-} \rightarrow 5Fe^{2+}(aq) + 2Mn^{2+}(aq) + 8H_2O(l) + \cancel{10e^-}$

Overall: $5Fe(s) + 2MnO_4^-(aq) + 16H^+(aq) \rightarrow 5Fe^{2+}(aq) + 2Mn^{2+}(aq) + 8H_2O(l)$

USEFUL TIP

When multiplying an equation by a given factor, remember that it is not only the number of electrons that is multiplied. All other chemical species will also be multiplied by the same number.

Progress questions

Question 8

A balanced overall redox equation will not contain electrons.

A. True

B. False

Question 9

Which of the following represents an overall redox equation?

A. $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$

B. $2H_2O(l) \rightarrow O_2(g) + 2H_2(g)$

Theory summary

- Redox reactions are a common type of chemical reaction that involve the transfer of electrons from one chemical species to another.
- Oxidation is the loss of electrons. Reduction is the gain of electrons.
  - Remember OIL RIG.
- Oxidising agents cause the oxidation of another chemical species, and are themselves reduced.
- Reducing agents cause the reduction of another chemical species, and are themselves oxidised.
- Oxidation states are values assigned to the atoms in molecular compounds to track the movement of electrons between chemical species inredox reactions.

Continues →

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- Changes in oxidation states can be used to determine whether a redox reaction has taken place.
    - An increase in the oxidation state means that species has undergone oxidation e.g. −2 to −1.
    - A decrease in the oxidation state means that species has undergone reduction e.g. +7 to +2.
- Redox half-equations describe the individual oxidation and reduction reactions occurring within a redox reaction.
    - Redox half-equations will show the movement (loss or gain) of electrons.
- Redox half-equations and overall redox equations must be balanced with respect to both the number of each atom and the amount of charge present on each side of the equation.
    - The balancing of redox half-equations in acidic solution follows a specific sequence of steps described by the acronym KOHES.
    - The permanganate ion, MnO_4^-, and the dichromate ion, $Cr_2O_7^{2-}$, are polyatomic ions that are strong oxidising agents.
- Overall redox equations are formed from the combination of the oxidation and reduction half-equations.
    - Overall redox equations will not explicitly show the movement (loss or gain) of electrons.

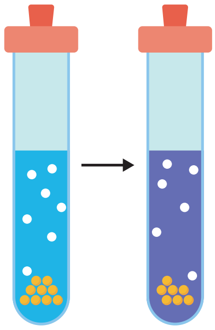
The content in this lesson is considered fundamental prior knowledge to primary galvanic cells and fuel cells as sources of energy (Unit 3 AOS 1).

11A Questions

Deconstructed

Use the following information to answer questions 10–12.

If solid copper is added to a solution of nitric acid, $HNO_3(aq)$, the colourless gas nitric oxide, $NO(g)$, can be formed. The solution will also transition from colourless to vibrant blue due to the formation of $Cu^{2+}(aq)$ ions as shown.



The unbalanced overall redox equation for this reaction is given.

$Cu(s) + NO_3^-(aq) \rightarrow NO(g) + Cu^{2+}(aq)$

Question 10 (1 MARK)

Which of the following correctly identifies the oxidation states of all chemical species present in these unbalanced equations?

A. 0 +6 −2 +2 −2 +2
$Cu(s) + NO_3^-(aq) \rightarrow NO(g) + Cu^{2+}(aq)$

B. 0 +5 −2 +2 −2 +2
$Cu(s) + NO_3^-(aq) \rightarrow NO(g) + Cu^{2+}(aq)$

C. 0 +2 −1 +1 −1 +2
$Cu(s) + NO_3^-(aq) \rightarrow NO(g) + Cu^{2+}(aq)$

D. −2 +5 −2 +2 −2 +2
$Cu(s) + NO_3^-(aq) \rightarrow NO(g) + Cu^{2+}(aq)$

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Question 11 (1 MARK)

Which of the following correctly classifies the species undergoing oxidation and reduction in this reaction?

	Oxidation	Reduction
A.	$Cu(s)$	$NO_3^-(aq)$
B.	$Cu^{2+}(aq)$	$NO(g)$
C.	$NO_3^-(aq)$	$Cu(s)$
D.	$NO(g)$	$Cu^{2+}(aq)$

Question 12 (3 MARKS)

By first writing the balanced oxidation and reduction half-equations for this reaction, give the balanced overall redox equation for the reaction between nitric acid and solid copper.

Exam-style

Question 13 (1 MARK)

Which of the following represents a balanced reduction half-reaction?

A. $VO_2^+ + H^+ + 2e^- \rightarrow VO_2^+ + H_2O$

B. $VO_2^+ + H_2 \rightarrow VO^{2+} + H_2O + e^-$

C. $VO_2^+ + 2H^+ + e^- \rightarrow VO^{2+} + H_2O$

D. $VO_2^+ + 4H^+ + 3e^- \rightarrow VO^{2+} + 2H_2O$

VCAA 2010 Exam 2 Multiple choice Q16

Question 14 (1 MARK)

The oxidation state of phosphorus in the pyrophosphate ion $P_2O_7^{4-}$ is

A. +3.5

B. +5

C. +7

D. +10

VCAA 2012 Exam 1 Multiple choice Q19

Question 15 (1 MARK)

Which one of the following reactions is a redox reaction?

A. $2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$

B. $Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

C. $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

D. $CH_3OH(l) + HCOOH(l) \rightarrow HCOOCH_3(l) + H_2O(l)$

VCAA 2013 exam Multiple choice Q6

Question 16 (1 MARK)

Based on the following chemical equation and half-equations, what is the oxidation half-equation in its simplest form?

Overall: $2Au^{3+}(aq) + 6I^-(aq) \rightarrow 2Au(s) + 3I_2(s)$

(1) $2Au^{3+}(aq) + 6e^- \rightarrow 2Au(s)$

(2) $6I^-(aq) \rightarrow 3I_2(s) + 6e^-$

A. $Au^{3+}(aq) \rightarrow Au(s) + 3e^-$

B. $Au^{3+}(aq) + 3e^- \rightarrow Au(s)$

C. $2I^-(aq) \rightarrow I_2(s) + 2e^-$

D. $I_2(s) + 2e^- \rightarrow 2I^-(aq)$

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Question 17 (3 MARKS)

Fertilisers are mixtures of different compounds that provide the nutrients that plants need to be able to grow. Nitrogen is very important for plants because it is an atom that is found in chlorophyll, the pigment that gives plants their distinctive green colour. Therefore, a common ingredient in fertiliser is nitric acid, which can be produced by the following redox reaction:

$NH_3(g) + 2O_2(g) \rightarrow HNO_3(g) + H_2O(g)$

a. The term oxidising agent is derived from the fact that one of the first and strongest oxidising agents to be investigated was oxygen. Given this information, identify the chemical species undergoing reduction. (2 MARKS)

b. The next step in the production of fertiliser is the following chemical reaction where nitric acid is neutralised by ammonia: $HNO_3(g) + NH_3(g) \rightarrow NH_4NO_3(s)$. This reaction is not a redox reaction. What does this mean in terms of the transfer of electrons between chemical species? (1 MARK)

Question 18 (12 MARKS)

a. Identify the oxidising and reducing agent in the following redox reactions. Justify your answers with reference to oxidation states.

i. $2FeCl_3(s) + H_2S(g) \rightarrow 2FeCl_2(s) + S(s) + 2HCl(g)$ (4 MARKS)

ii. $3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$ (4 MARKS)

b. Write the balanced oxidation and reduction half-equations for these two redox reactions. For ionic compounds, include only the ion undergoing a change in oxidation state. (4 MARKS)

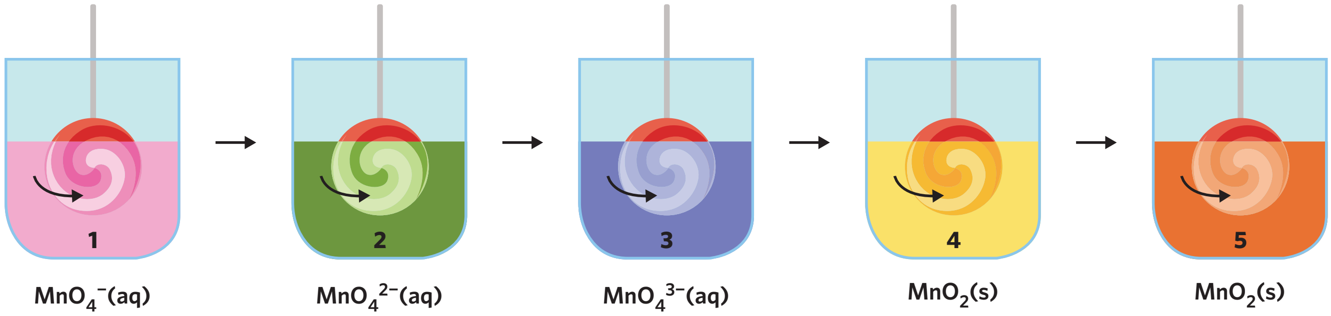
Key science skills

Question 19 (9 MARKS)

Manganese is a transition metal that can take a number of oxidation states. Depending on manganese's oxidation state in a compound, it can form a number of differently coloured solutions.

Pallas is a chemistry student interested in investigating these colour changes through the use of a potassium permanganate, $KMnO_4(aq)$, solution, which he knows is a strong oxidising agent. He also knows that the sugars present in confectionery such as lollipops can act as reducing agents.

Pallas designs an experiment in which he stirs a lollipop in an alkaline solution of $KMnO_4(aq)$ and records any colour changes he observes throughout the experiment. A diagram of his experiment and results is shown.



a. Identify the dependent variable in this experiment. (1 MARK)

b. Complete Pallas's table of results shown. (5 MARKS)

	1	2	3	4	5
Colour →	Pink	Green	Blue	Yellow	Orange
Main species present	$MnO_4^-(aq)$	$MnO_4^{2-}(aq)$	$MnO_4^{3-}(aq)$	$MnO_2(s)$	$MnO_2(s)$
Mn oxidation state					

c. Give the balanced redox half-equation for the reaction responsible for the solution's colour transition from green to blue. (1 MARK)

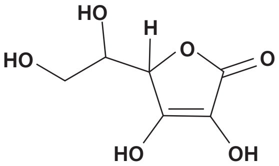
d. Suggest a reason for the second colour observed in the beaker at point 5, even though there was the same chemical species present. (2 MARKS)

FROM LESSONS 16A, 16C & 16E

Questions from multiple lessons

Question 20 (7 MARKS)

Antioxidants are compounds that are commonly used in the food industry. As the name suggests, antioxidants are chemical compounds that stop another chemical compound from being oxidised. They function as preservatives in foods because oxidation leads to a decline in food quality.



a. If antioxidants are oxidised themselves to protect food from being oxidised, what property must they be good at? (2 MARKS)

b. A student thought that since antioxidants are oxidised, they are oxidising agents. Is the student correct? Explain your answer. (2 MARKS)

c. Vitamin C is a naturally occurring antioxidant and its structure is shown.

What is the name of the most common functional group in this compound? (1 MARK)

d. What property of this functional group might enable vitamin C to be such a good antioxidant? (2 MARKS)

FROM LESSONS 2A & 7C

Question 21 (7 MARKS)

Chlorate, $ClO_3^-(aq)$, is an incredibly strong oxidising agent that was once commonly used in pyrotechnic (e.g. firework) displays due to its explosive reactivity with organic compounds. When taking part in a redox reaction in an acidic solution, the chlorate ion will be reduced to form chlorine gas, $Cl_2(g)$.

a. Write the balanced half-equation for the reduction of $ClO_3^-(aq)$ to $Cl_2(g)$. Identify the oxidation states for all atoms present in the half-equation. (2 MARKS)

b. Chlorate is capable of oxidising nitric oxide, $NO(g)$, into nitrate ions, $NO_3^-(aq)$. Give the balanced oxidation half-equation for this reaction. Identify the oxidation states for all atoms present in the half-equation. (2 MARKS)

c. Using these two half-equations, write the balanced overall redox equation for the reaction between chlorate and nitric oxide. (1 MARK)

d. Would you expect the pH of the solution in which this reaction was occurring to increase, decrease or remain constant as a result of this reaction? Explain. (2 MARKS)

FROM LESSON 10C

Hints

13. A reduction half-reaction involves the gain of electrons (electrons on the reactant side of the equation).

14. The overall oxidation state of the pyrophosphate ion is −4.

15. In a redox reaction, at least one species should be reduced and at least one species should be oxidised (neutralisation reactions are never a redox reaction).

16. Oxidation is the loss of electrons (electrons on the product side of the equation).

17a. If one species is oxidised, another must be reduced.

17b. Redox reactions are defined by the transfer of electrons from one chemical species to another.

18ai. An increase in oxidation state is oxidation, whilst a decrease in oxidation state is reduction.

18aii. An increase in oxidation state is oxidation, whilst a decrease in oxidation state is reduction.

18bi. The electrons will be on the reactant side of the equation in reduction, and on the product side of the equation in oxidation.

18bii. The electrons will be on the reactant side of the equation in reduction, and on the product side of the equation in oxidation.

19a. The dependent variable is measured in an experiment.

19b. Use oxidation state rules for oxygen to find the oxidation state that will balance the charge on each compound.

19c. A greater negative charge means a gain of electrons has occurred.

19d. The concentration of a compound in a solution can affect its colour.

20a. Reducing agents are oxidised.

20b. Additional electrons are not consumed or lost from a redox reaction.

20c. Functional groups are sets of atoms which give a chemical compound unique properties.

20d. Oxygen atoms in hydroxyl groups have 2 lone pairs of electrons to lose.

21a. Use KOHES.

21b. Use KOHES.

21c. Overall redox reactions are formed from their individual half-equations.

21d. pH is proportional to the number of H^+ ions present in a solution.

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11B Reactivity of metals – part 2

Why does solid silver spontaneously form on the surface of copper wire?

The difference in reactivity between copper and silver causes a silver coating to form on a copper wire when it is immersed in a solution of silver ions. In this lesson, we will investigate the oxidation and reduction processes enabling reactions like this, and how metals can be organised according to their reactivity.

STUDY DESIGN DOT POINT

- the reactivity series of metals and metal displacement reactions, including balanced redox equations, with states

11A 11B 11C

2.1.11.1 Reactivity series of metals

2.1.11.2 Metal displacement reactions

ESSENTIAL PRIOR KNOWLEDGE

3B Reactivity of metals

10B Reactions of acids and metals

See questions 75–76.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

KEY TERMS AND DEFINITIONS

Metal displacement reaction redox reaction in which a more reactive metal displaces a less reactive metal's cation from solution

Reactivity series of metals organised scale of metals and their cations, ranked according to their strength as reducing and oxidising agents

Spontaneous redox reaction feasible redox reaction that occurs without the addition of any extra heat or energy

Reactivity series of metals 2.1.11.1

Metals and their cations can be organised according to their tendency to give away electrons (reducing agents) or accept (oxidising agents) electrons.

How can the reactivity of metals be predicted?

As we learned in lesson 3B, metals can be classified according to their reactivity with acids, water, and oxygen. All of these reactions are examples of redox reactions. When a metal in its solid form reacts with oxygen, water or an acid, it acts as a reducing agent, losing electrons as it forms its corresponding metal cation. We can rank these metals from strongest to weakest in terms of reducing strength.[1] This is known as the **reactivity series of metals**. The reactivity series of a number of common metals is given in figure 1, showing the forward reaction as a reduction half-equation for each metal cation.

Metal cations are shown on the left-hand side of the reactivity series, with the weakest oxidising agent (Li^+) in the bottom-left position. The strength of the metal cation as an oxidising agent increases from bottom to top, with the strongest oxidising agent (Pt^{2+}) present in the top-left position.

USEFUL TIP

Strong oxidising agents are good at stealing electrons from other chemical species, whilst strong reducing agents are good at giving away electrons to other chemical species.

KEEN TO INVESTIGATE?

[1] **How are redox reactions used in photochromic sunglasses?**
Search: The Ultimate Guide on How Photochromic Lenses Work

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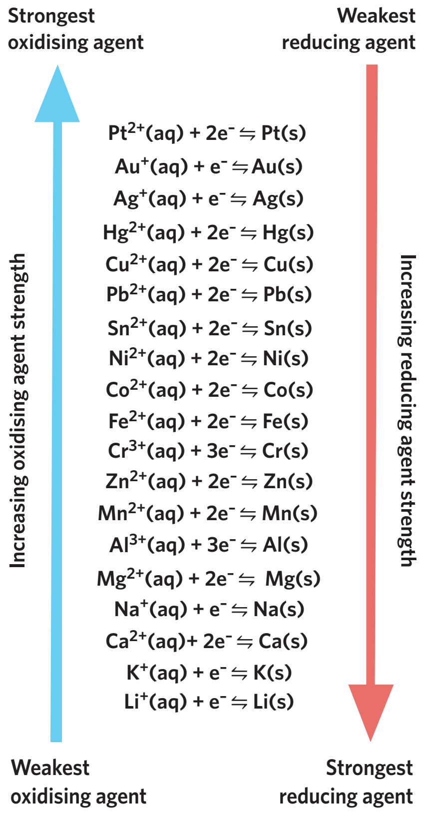

$Pt^{2+}(aq) + 2e^- \rightleftharpoons Pt(s)$
$Au^{+}(aq) + e^- \rightleftharpoons Au(s)$
$Ag^{+}(aq) + e^- \rightleftharpoons Ag(s)$
$Hg^{2+}(aq) + 2e^- \rightleftharpoons Hg(s)$
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$
$Cr^{3+}(aq) + 3e^- \rightleftharpoons Cr(s)$
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$
$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s)$
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$
$Na^{+}(aq) + e^- \rightleftharpoons Na(s)$
$Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$
$K^{+}(aq) + e^- \rightleftharpoons K(s)$
$Li^{+}(aq) + e^- \rightleftharpoons Li(s)$

Figure 1 Reactivity series of metals

USEFUL TIP

Highly reactive metals are strong reducing agents. Therefore, they will readily give away electrons and undergo oxidation.

Metals in their solid form are shown on the right-hand side of the reactivity series, with the weakest reducing agent (Pt) in the top-right position. The strength of the metal as a reducing agent increases from top to bottom, with the strongest reducing agent (Li) present in the bottom-right position.

USEFUL TIP

In Units 3 & 4 chemistry, a more comprehensive reactivity series is given in the VCE Data Book. This version, known as the electrochemical series, also incorporates the relative reactivity of non-metal compounds such as $Cl_2(g)$ and $H_2O(l)$.

As shown in figure 1, the stronger reducing agent a metal in solid form is, the weaker oxidising strength its cation will have. The inverse also applies – the stronger oxidising agent a metal cation is, the weaker reducing strength it will have in solid form. The more reactive a metal is, the more readily it undergoes oxidation, reacting vigorously and readily with oxygen, water and dilute acids. A metal like gold is valued highly for both its colour, and because it is very unreactive. It will not react and become degraded by oxygen or water.

A highly reactive metal such as potassium, on the opposite end of the reactivity series, is highly reactive with water and oxygen. In fact, potassium is so reactive that it is not found in pure metal form in nature. Instead, potassium is always present as K^+ ions. When it is refined to form K(s) metal, it must be carefully stored in oil to prevent it from reacting violently with oxygen or water in the atmosphere.

In a solution of dilute 0.01 M hydrochloric acid, gold, nickel, zinc, and magnesium all react to different degrees; in fact, gold doesn't react at all. As we learned in lesson 3B, the reaction between a reactive metal and an acid will produce hydrogen gas, $H_2(g)$. However, how readily and vigorously the reaction occurs depends on the reactivity of the specific metal, as shown in figure 2.

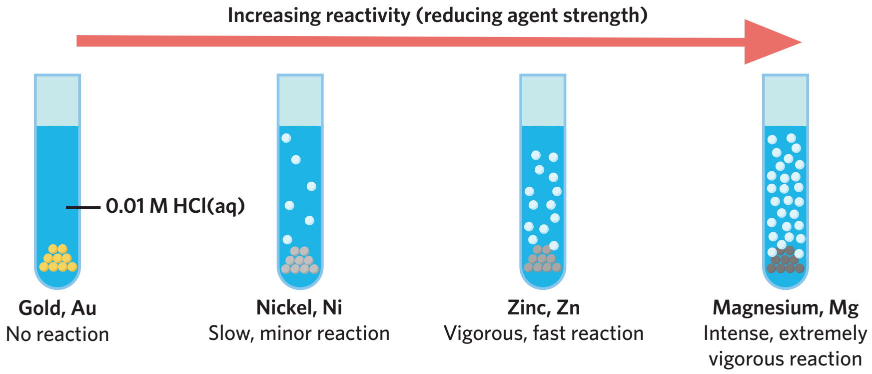

Figure 2 Metals of greater reactivity will react more vigorously with dilute acid to produce $H_2(g)$.

WORKED EXAMPLE 1

A cube of sodium (Na) metal and a cube of iron (Fe) metal are each added to a separate 500 mL beaker of deionised water. In which of the two beakers would a more vigorous reaction be expected to be observed? Why?

What information is presented in the question?

Na(s) and Fe(s) are added to solutions of deionised water.

What is the question asking us to do?

Establish which metal will react more vigorously with water.

What strategies do we need in order to answer the question?

1. Determine whether Na(s) or Fe(s) is a more reactive metal.
2. Link the relative reactivities of each metal to its expected reaction with water.

Answer

From the reactivity series, we can see that Na(s) is a considerably stronger reducing agent than Fe(s). Therefore, Na(s) is likely to be a more reactive metal than Fe(s).

This means that the beaker containing Na(s) will most likely react much more vigorously with the deionised water than the beaker containing Fe(s).

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Progress questions

Question 1

Metals that are highly reactive with oxygen, water and dilute acids are also strong

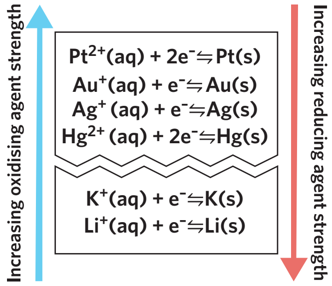
A. reducing agents.

B. oxidising agents.

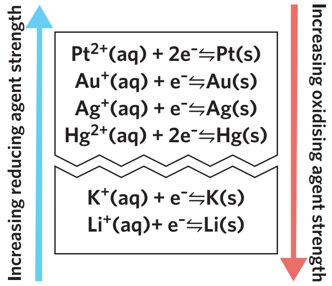
Question 2

Which of the following correctly labels the reactivity series of metals?

A.



B.



Question 3

$Au^+(aq)$ is a strong

A. reducing agent.

B. oxidising agent.

Question 4

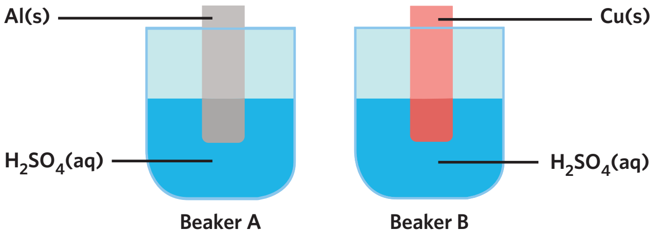
A highly reactive metal such as potassium, K(s), will react vigorously with a solution of $HNO_3(aq)$, but will not react with oxygen from the atmosphere.

A. True

B. False

Question 5

In which of the following beakers would you expect to observe a more vigorous reaction?



A. Beaker A

B. Beaker B

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Metal displacement reactions 2.1.11.2

Metal displacement reactions occur when a more reactive metal displaces a less reactive metal's cations from its solution.

How can we predict which metals will displace each other in solution?

So far we have only considered the reaction that would occur between a metal and oxygen, water or a solution of dilute acid. We will now investigate what happens when metals are added to solutions of other metal ions.

Consider the following redox reaction that occurs between a strip of iron metal, $Fe(s)$, and a solution of copper ions, $Cu^{2+}(aq)$.

Overall redox reaction: $Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$

Oxidation half-equation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Reduction half-equation: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

From the reactivity series, we can see that $Fe(s)$ is a more reactive metal than $Cu(s)$ as it has greater strength as a reducing agent. Therefore, $Fe(s)$ will readily undergo oxidation and lose its electrons, donating two electrons to the cations of the less reactive metal, $Cu^{2+}(aq)$, to form $Fe^{2+}(aq)$ and $Cu(s)$. This kind of reaction – where a more reactive metal displaces a less reactive metal's cations from solution – is known as a metal displacement reaction.

Metal displacement reactions are redox reactions that can occur spontaneously without needing to apply extra heat and energy, and without introducing any additional chemical substances. This means this reaction can be classified as a **spontaneous redox reaction**. Whilst non-spontaneous metal displacement reactions can occur, in this course we will only consider metal displacement reactions that occur spontaneously.

The reaction between $Fe(s)$ and $Cu^{2+}(aq)$ is simply a reaction between the strongest reducing agent ($Fe(s)$) and the strongest oxidising agent ($Cu^{2+}(aq)$) present in the solution. Using the reactivity series, we can predict the outcome of metal displacement reactions by determining which chemical species is highest on the left-hand side (the strongest oxidising agent) and which chemical species is lowest on the right-hand side (the strongest reducing agent).

USEFUL TIP

Metal displacement reactions occur when a more reactive metal displaces a less reactive metal from a solution of its cations.

USEFUL TIP

When just copying a half-equation from the reactivity series in the VCE Data Book use the reversible arrow, ⇌, but when a reaction is actually occurring use the single arrow → or ←.

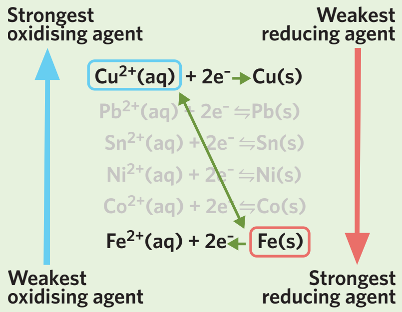
MISCONCEPTION

'A reaction will always spontaneously occur between a strong oxidising agent and a strong reducing agent.'

The reactivity series of metals predicts whether reactions are feasible, not necessarily whether they will occur spontaneously, or the rate of reaction.

STRATEGY

To quickly determine whether a reaction will occur between a solid metal and a solution of metal ions, we can draw an imaginary line to connect the two metal reactants. If the line runs diagonally down from left to right, a reaction is likely to occur spontaneously.



WORKED EXAMPLE 2

Write the balanced overall redox equation, if any, that would occur in the following scenario. Do not include spectator ions in the overall redox equation.

A strip of zinc metal, $Zn(s)$ is added to a solution of silver nitrate, $AgNO_3(aq)$.

What information is presented in the question?

$Zn(s)$ is added to a solution of $AgNO_3(aq)$.

What is the question asking us to do?

Determine the balanced equation for the reaction between $Zn(s)$ and $AgNO_3(aq)$.

Continues →

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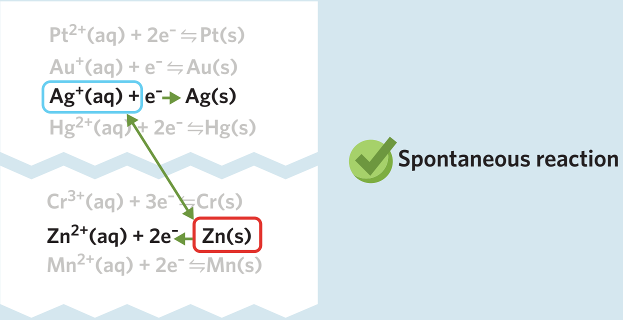
What strategies do we need in order to answer the question?

1. Identify the strongest oxidising agent and strongest reducing agent present in the beaker on the reactivity series.
2. Determine whether the reaction will be a spontaneous redox reaction.
3. Remember that the oxidising agent will undergo reduction and the reducing agent will undergo oxidation.
4. Write the individual reduction and oxidation half-equations using single arrows.
5. Using the half-equations, determine the balanced overall redox reaction.

Answer



We can see that the $Ag^+(aq)$ ions in the solution are the strongest oxidising agent and $Zn(s)$ is the strongest reducing agent. Since the line connecting the two species goes down from left to right, we know that this will be a spontaneous redox reaction. Note that the nitrate ion, $NO_3^-(aq)$, will not take part in the reaction as it is a spectator ion.

The oxidising agent, $Ag^+(aq)$ will undergo reduction according to the following half-equation:

$Ag^+(aq) + e^- \rightarrow Ag(s)$

The reducing agent (more reactive metal), $Zn(s)$ will undergo oxidation according to the following half-equation:

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

In order to write the overall balanced redox equation, the reduction half-equation must be multiplied by two so that it also contains two electrons.

$2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$

Combining the two half-equations and cancelling the two electrons leaves us with the following balanced overall redox equation:

$2Ag^+(aq) + Zn(s) \rightarrow 2Ag(s) + Zn^{2+}(aq)$

WORKED EXAMPLE 3

Write the balanced overall redox equation, if any, that would occur in the following scenario. Do not include spectator ions in the overall redox equation.

A strip of zinc metal, $Zn(s)$ is added to a solution of aluminium sulfate, $Al_2(SO_4)_3(aq)$.

What information is presented in the question?

$Zn(s)$ is being added to a solution of $Al_2(SO_4)_3(aq)$.

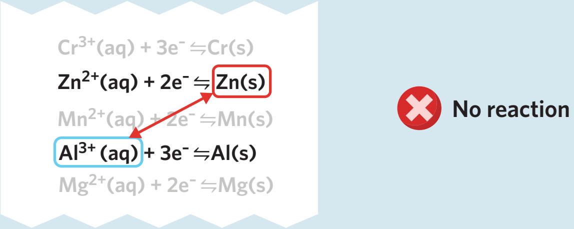
What is the question asking us to do?

Determine the balanced equation for the reaction between $Zn(s)$ and $Al_2(SO_4)_3(aq)$.

What strategies do we need in order to answer the question?

1. Identify the strongest oxidising agent and strongest reducing agent present in the beaker on the reactivity series.
2. Determine whether the reaction will be a spontaneous redox reaction.

Answer



Since zinc is a less reactive metal than aluminium (it is a weaker reducing agent than $Al(s)$), it will not be able to displace the $Al^{3+}(aq)$ ions from the $Al_2(SO_4)_3(aq)$ solution. Therefore, there will be no observed redox reaction between $Zn(s)$ and $Al_2(SO_4)_3(aq)$.

Note that the sulfate ion, $SO_4^{2-}(aq)$, will not take part in any reaction as it is a spectator ion.

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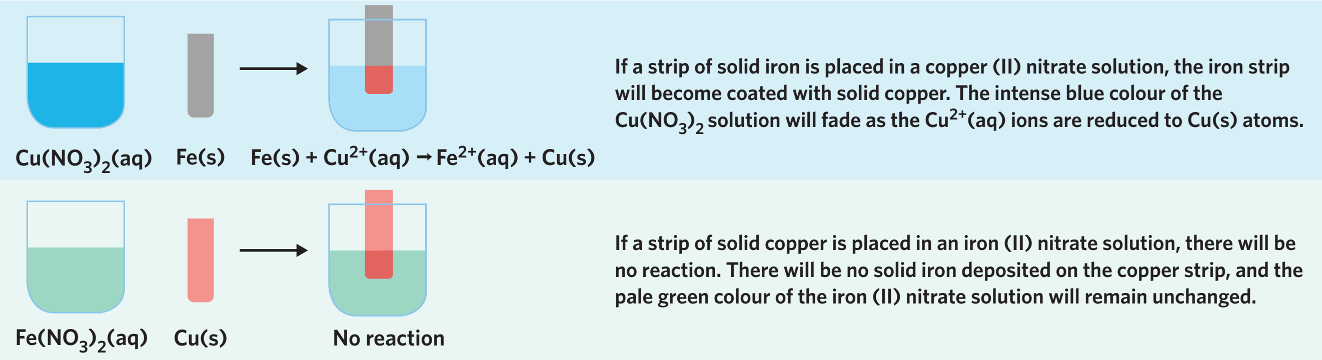
Experimentally, metal displacement reactions can be used to determine the relative reactivity of two metals. In the case of copper and iron, two beakers can be set up, one containing $Cu^{2+}(aq)$ and $Fe(s)$, the other containing $Cu(s)$ and $Fe^{2+}(aq)$ as shown in figure 3. Remember that the more reactive metal (the stronger reducing agent) will lose electrons and form ions in the solution. Therefore, the metal cation will accept electrons from any other more reactive solid metal atoms that are placed in the solution, resulting in these cations forming solid metal atoms themselves. However, if the solid metal is less reactive than the metal cation, there will be no reaction.



If a strip of solid iron is placed in a copper (II) nitrate solution, the iron strip will become coated with solid copper. The intense blue colour of the $Cu(NO_3)_2$ solution will fade as the $Cu^{2+}(aq)$ ions are reduced to $Cu(s)$ atoms.

If a strip of solid copper is placed in an iron (II) nitrate solution, there will be no reaction. There will be no solid iron deposited on the copper strip, and the pale green colour of the iron (II) nitrate solution will remain unchanged.

Figure 3 Experimental evidence of the relative reactivity of copper and iron

MISCONCEPTION

'Metal displacement reactions are precipitation reactions.'

A metal displacement reaction is a single displacement reaction where one element is replaced for another in a compound.

A precipitation reaction is a type of double displacement reaction in which either the cations or anions from two soluble compounds are swapped to form an insoluble compound.

The experimental set up in figure 3 reveals that iron is more reactive and a stronger reducing agent than copper. These results are confirmed by the position of these metals on the reactivity series with Fe(s) closer to the bottom right than Cu(s).

Progress questions

Question 6

Reactions in which a __________ reactive metal displaces a __________ reactive metal from a solution of its cations are known as metal displacement reactions.

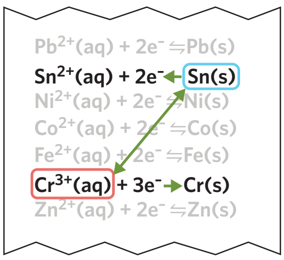
A. more, less

B. less, more

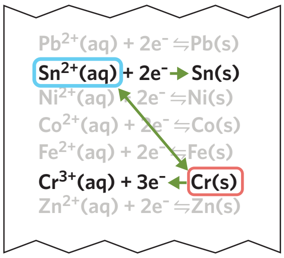
Question 7

Which of the following best represents the half-equations on the reactivity series that would occur in a metal displacement reaction between tin metal (Sn), tin ions (Sn^{2+}), chromium metal (Cr) and chromium ions (Cr^{3+})?

A.



B.



Question 8

When a strip of solid zinc, Zn(s), is added to a solution of $Pb^{2+}(aq)$ ions, __________ will displace __________ from the solution.

A. lead, zinc ions

B. zinc, lead ions

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Theory summary

- The greater the reactivity of a metal, the more readily it will undergo oxidation (give away electrons) in the presence of oxygen, water and dilute acids.
- Metals and their cations can be organised according to their tendency to give away electrons (reducing agents) or accept (oxidising agents) electrons.
  - Metal cations are found on the left-hand side of the reactivity series, increasing in oxidising strength from bottom to top.
  - Metals in their solid form are found on the right-hand side of the reactivity series, increasing in reducing strength from top to bottom.
- Metal displacement reactions occur when a more reactive metal (a stronger reducing agent) is oxidised by a less reactive metal's cations in solution.
  - The more reactive metal donates its electrons to the less reactive metal's cations, reducing them to form a solid metal deposit.
  - Having lost electrons, the more reactive metal enters the solution as cations, displacing the less reactive metal.
- To determine whether a spontaneous redox reaction will occur between a metal and a cation, the reactivity series of metals can be used.
  - First, the strongest oxidising agent and reducing agent must be identified.
  - If the two chemical species have a stronger oxidising (top left) and stronger reducing agent (bottom right) from the reactivity series, then the redox reaction can feasibly occur spontaneously.

The content in this lesson is considered fundamental prior knowledge to primary galvanic cells and fuel cells as sources of energy (Unit 3 AOS 1).

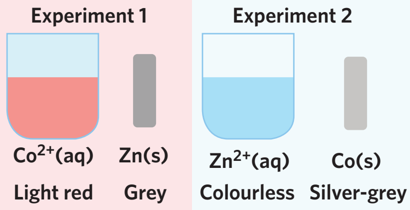
11B Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 9–11.

A student, Aeneas, prepares solutions of cobalt (II) and zinc ions. Cobalt ions form a light red solution, whilst zinc ions are colourless in solution. He places a strip of grey Zn(s) into the Co^{2+}(aq) solution (experiment 1) and a strip of silver-grey Co(s) into the Zn^{2+}(aq) solution (experiment 2).



Question 9 (1 MARK) 🌶

Which of the following is a correct statement about the reactivity of zinc and cobalt?

A. Zn(s) is a stronger oxidising agent than Co(s).

B. Co(s) is a stronger oxidising agent than Zn(s).

C. Zn(s) is a stronger reducing agent than Co(s).

D. Co(s) is a stronger reducing agent than Zn(s).

Question 10 (1 MARK) 🌶

Which of the following best describes the reaction(s) that Aeneas would observe in experiments 1 and 2?

A. There will be a reaction in both experiment 1 and experiment 2.

B. There will be a reaction in experiment 1, but not in experiment 2.

C. There will be a reaction in experiment 2, but not in experiment 1.

D. There will not be a reaction in either experiment 1 or experiment 2.

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Question 11 (3 MARKS)

Give the balanced oxidation and reduction half-equations for any reactions occurring in the two experiments. Hence, give the balanced overall equation(s) for these reaction(s).

Exam-style

Question 12 (1 MARK)

A strip of chromium metal is added to a solution of tin (II) cations. Which of the following gives the balanced equation for the reaction, if any, that would occur?

A. $Cr^{3+}(aq) + Sn^{2+}(aq) \rightarrow Cr(s) + Sn(s)$

B. $Cr(s) + Sn^{2+}(aq) \rightarrow Cr^{3+}(aq) + Sn(s)$

C. $2Cr(s) + 3Sn^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 3Sn(s)$

D. No spontaneous reaction would occur between Cr(s) and $Sn^{2+}(aq)$ ions.

Question 13 (1 MARK)

Which of the following only lists chemical species that Fe(s) will displace from solution?

A. Zn(s), Mn(s), Ca(s), Mg(s)

B. Pb(s), Ni(s), Cu(s), Co(s)

C. $Zn^{2+}(aq)$, $Mn^{2+}(aq)$, $Ca^{2+}(aq)$, $Mg^{2+}(aq)$

D. $Pb^{2+}(aq)$, $Ni^{2+}(aq)$, $Cu^{2+}(aq)$, $Co^{2+}(aq)$

Question 14 (1 MARK)

Solution I – 1.0 M NaCl

Solution II – 1.0 $CuCl_2$

Solution III – 1.0 M $MgCl_2$

Which solution or solutions above will react with Zn powder?

A. Solution I only

B. Solution II only

C. Solutions I and III only

D. Solutions I, II and III

VCAA 2015 exam Multiple choice Q25

Question 15 (6 MARKS)

Consider the following pairs of solid metals and metal cations. If a spontaneous metal displacement reaction occurs, give its balanced overall equation. If no metal displacement reaction occurs, state ‘No reaction’.

a. $Pb^{2+}(aq)$ and Co(s) (1 MARK)

b. $Ag^{+}(aq)$ and Cu(s) (1 MARK)

c. $Sn^{2+}(aq)$ and Al(s) (1 MARK)

d. $Fe^{2+}(aq)$ and Cu(s) (1 MARK)

e. Mg(s) and $Au^{+}(aq)$ (1 MARK)

f. $Na^{+}(aq)$ and Zn(s) (1 MARK)

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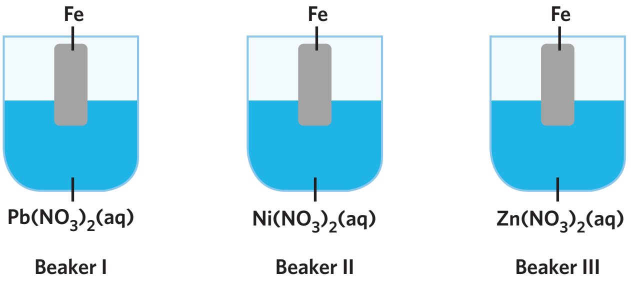
Question 16 (1 MARK)

A shiny metal coating is formed when a strip of metal X(s) is added to solution of A^{2+}(aq) cations, but not when metal X(s) is added to a solution of B^{2+}(aq) cations. Which of the following most likely represents the identities of metals X, A and B?

	X	A	B
A.	Ni	Au	Mn
B.	Co	Cu	Ag
C.	Al	Ca	Pb
D.	Zn	Sn	Mg

Question 17 (6 MARKS)

Three beakers, each containing an iron strip and a 1.0 M solution of a metal salt, were set up as shown.



a. Identify whether a reaction would occur in beaker III. Justify your answer using the reactivity series of metals. (2 MARKS)

b. Name the substance that would form on the surface of the iron strip in beaker I. (1 MARK)

c. Give the half-equations for the oxidation and reduction reactions that would occur in beaker II, and hence give the overall redox equation for the metal displacement reaction that occurs. (3 MARKS)

Adapted from VCAA 2013 exam Multiple choice Q24

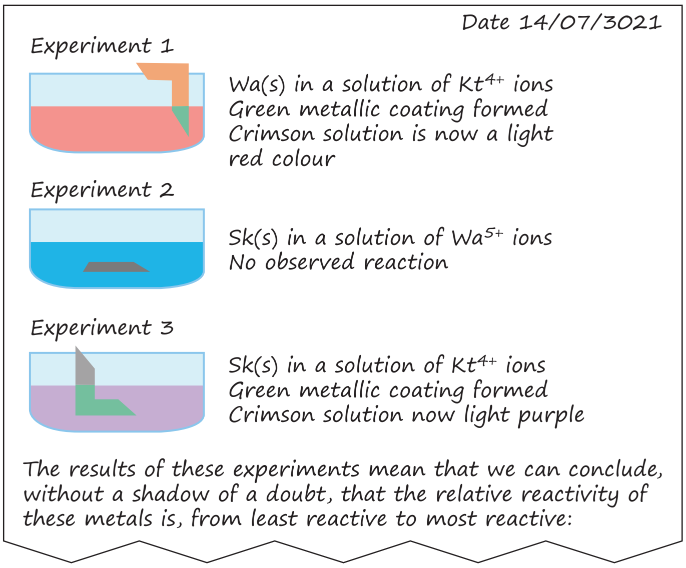
Key science skills

Question 18 (6 MARKS)

It is the year 3023 and a recent interstellar expedition to the exoplanet Sutherlandia by the chemistry syndicate Igao Industries has led to the discovery of three new metal elements. Of course, the first thing Igao Industries does is determine the relative reactivity of these three new elements – Kostrethium (Kt), Sparkonium (Sk) and Wallacium (Wa). The ions formed by each of these elements are given in the table.

Element	Ion
Kostrethium, Kt	Kt^{4+}
Sparkonium, Sk	Sk^{2+}
Wallacium, Wa	Wa^{5+}

An extract from the notebook of the expedition's chief scientist, Dr Komezyrov, is given.



Date 14/07/3021

Experiment 1

Wa(s) in a solution of Kt^{4+} ions
Green metallic coating formed
Crimson solution is now a light red colour

Experiment 2

Sk(s) in a solution of Wa^{5+} ions
No observed reaction

Experiment 3

Sk(s) in a solution of Kt^{4+} ions
Green metallic coating formed
Crimson solution now light purple

The results of these experiments mean that we can conclude, without a shadow of a doubt, that the relative reactivity of these metals is, from least reactive to most reactive:

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a. Give the balanced overall redox equation for the reaction that occurred in experiment 1. (1 MARK)

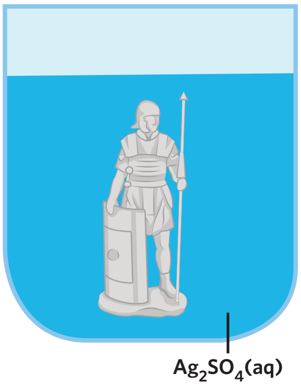
b. Tragically, a horde of space edroliphants ambushed the expedition's colony and Dr Komezyrov's notebook was damaged. Complete Dr Komezyrov's notes, making sure to justify your answer with reference to the results of experiments 1, 2 and 3. (5 MARKS)

FROM LESSONS 16C & 16E

Questions from multiple lessons

Question 19 (9 MARKS)

When preparing a solution of $Ag_2SO_4(aq)$, Dido accidentally knocked her teacher's favourite figurine into the beaker – a Roman soldier made of 102 g of pure tin metal. After spending a few minutes panicking about what to do, Dido extracted the tin figurine from the $Ag_2SO_4(aq)$ solution, hoping that her teacher wouldn't realise anything had happened. Unfortunately for Dido, her teacher noticed immediately and as punishment asked her to answer the following questions.



a. Give the balanced redox half-equations for the metal displacement reaction that occurred in the beaker. (2 MARKS)

b. Hence, give the balanced overall redox equation for this reaction. (1 MARK)

c. When the figurine was re-weighed, Dido found that it weighed an extra 37.4 g.

  i. Explain why a change in mass was observed. (2 MARKS)

  ii. Calculate the mass of Ag(s) that has been deposited on the tin figurine. (4 MARKS)

FROM LESSONS 6B & 6C

Hints

12. Metals with greater reducing strength will be oxidised.

13. More reactive metals displace less reactive metal cations from solution.

14. Metal displacement reactions can only occur if the two species have a stronger oxidising (top left) and stronger reducing agent (bottom right) from the reactivity series.

15. More reactive metals displace less reactive metal cations from solution.

16. Metal displacement reactions can only occur if the two species have a stronger oxidising (top left) and stronger reducing agent (bottom right) from the reactivity series.

17a. Metal displacement reactions can only occur if the two species have a stronger oxidising (top left) and stronger reducing agent (bottom right) from the reactivity series.

17b. Metal ions are displaced from solution.

17c. The cations in solution will be reduced to solid metal. The solid metal will dissolve into solution.

18a. More reactive metals displace less reactive metal cations from solution.

18b. Metal displacement reactions are usually accompanied by observable changes.

19a. Metal cations are oxidising agents and will undergo reduction.

19b. Overall redox equations do not include electrons or spectator ions.

19ci. A metal cation has been displaced from the solution.

19cii. Solid metal is both formed and lost during this process.

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11C Applications of redox reactions in society


Image: Maxim Trukhin/Shutterstock.com

How can redox reactions be used to protect ships from corrosion?

Ships are made of the reactive metal iron which can be coated with a layer of zinc or have a block of magnesium attached to the hull to prevent them from rusting in seawater. In this lesson, we will explore the principles of **galvanic cells**, and how they can be applied to the **corrosion** of metals.

STUDY DESIGN DOT POINT

- applications of redox reactions in society: for example, corrosion or the use of simple primary cells in the production of electrical energy from chemical energy

11A 11B 11C

2.1.12.1 Simple galvanic cells

2.1.12.2 Corrosion

ESSENTIAL PRIOR KNOWLEDGE

11A Writing equations for redox reactions

11B The reactivity series of metals

See questions 77–78.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

KEY TERMS AND DEFINITIONS

Anode	negatively charged electrode in a galvanic cell where oxidation occurs
Anion	negative ion formed when an atom gains one or more valence electrons
Cathode	positively charged electrode in a galvanic cell where reduction occurs
Cation	positive ion formed when an atom loses its valence electron(s)
Corrosion	occurs when some metals react with gases in the atmosphere (mainly oxygen)
Dry corrosion	oxidation of metals due to oxygen in the air
Electrode	electrically conductive medium
Galvanic cell	electrochemical cell in which chemical energy from spontaneous redox reactions is converted into electrical energy
Galvanised	coated with zinc
Rust	a flaky, reddish precipitate formed from the corrosion of iron with water and oxygen
Sacrificial anode	a readily oxidised anode used to coat other metals for protection against corrosion
Sacrificial protection	the use of a protective layer that is readily oxidised
Salt bridge	device used to separate the solutions in each half-cell whilst completing the circuit
Wet corrosion	oxidation of metals due to oxygen and water in the air

Simple galvanic cells 2.1.12.1

Galvanic cells are simple electrical circuits that use redox reactions to generate electrical energy from chemical energy.

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What is a galvanic cell?

Galvanic cells are used in many everyday electronic devices, such as mobile phones and calculators, as a source of energy. In a galvanic cell, the chemical energy released from spontaneous redox reactions is converted to usable electrical energy. A typical galvanic cell set-up is shown in figure 1.

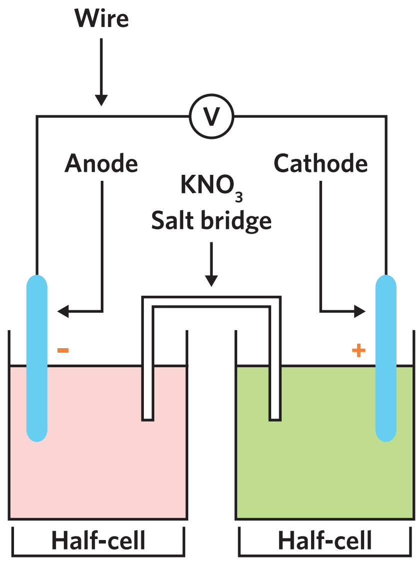

Figure 1 Design of a simple galvanic cell

In figure 1, there is a negatively charged **electrode** known as the **anode** in one half-cell, and a positively charged electrode known as the **cathode** in the other. In simple galvanic cells, each electrode – the general term for an anode or cathode – is a solid metal immersed in a solution of that same metal's cations. The two electrodes are connected by an external wire for energy output.

As shown in figure 1, in addition to the external wire connecting the electrodes, there's also a **salt bridge** infused with an aqueous salt that separates the solutions in each half-cell (preventing the production of heat energy); this internal connection allows the flow of ions to complete the circuit. The **anions** from the salt bridge flow towards the anode to balance out the production of metal **cations**, whilst the cations flow towards the cathode to balance out the production of solid metal.

MISCONCEPTION

'Electrons flow through the salt bridge.'

The External circuit (wires) allows Electrons to flow and the Internal circuit (salt bridge) allows Ions to flow.

USEFUL TIP

If the two half cells were not physically separated by the salt bridge and were instead allowed to mix, heat energy would be produced instead of electrical energy.

USEFUL TIP

The fact that salt bridges maintain a neutral overall charge of the cell is not a function of the salt bridge – it is simply a physics phenomenon.

How can redox reactions be used to generate electrical energy?

As we learned in lesson 11A, redox reactions always involve the transfer of electrons. In a galvanic cell, the electrons lost during oxidation flow through the wire connecting the two electrodes, providing electrical power to a component such as a light bulb. Since electrons are negatively charged, and opposite charges attract, we can deduce that electrons are always moving towards the positively charged cathode, as shown in figure 2.

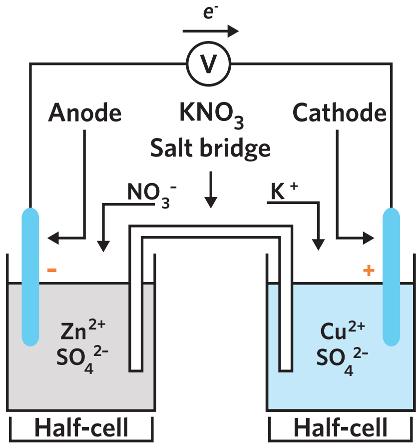

Figure 2 Flow of electrons in the Daniell cell

Given oxidation is the loss of electrons, we can say that oxidation always occurs at the anode, as the electrons will pass through the wire towards the cathode, powering the light bulb as they go. However, as we know, oxidation never occurs without reduction. Therefore, reduction must occur at the cathode to balance out the transfer of electrons. For example, the redox reactions in figure 2 can be summarised as:

Anode

Oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

Cathode

Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Overall

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

How can reactions in galvanic cells be predicted?

Using the reactivity series of metals that we saw in lesson 11B, we know that certain metals are stronger reducing agents, and that certain metal cations are stronger oxidising agents than others. Using this, we can predict what reactions may occur, and at which electrode they will occur. Consider the galvanic cell shown in figure 3.

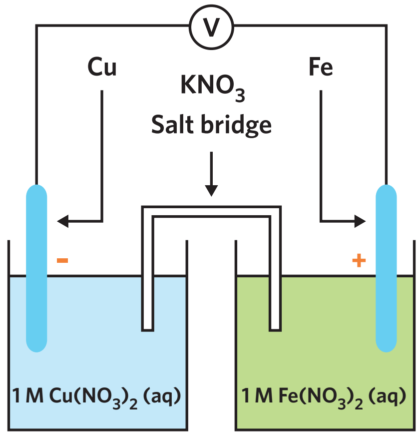

Figure 3 Copper (II) nitrate and iron (II) nitrate galvanic cell

USEFUL TIP

Electrons always flow from the site of oxidation to the site of reduction.

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From the reactivity series of metals shown in figure 4, we know that Fe(s) (bottom right) is a stronger reducing agent than Cu(s). Therefore, we know that Fe(s) will undergo oxidation at the anode, and $Cu^{2+}(aq)$ ions (top left) will undergo reduction at the cathode to balance the transfer of electrons. Therefore, we can write the two half-reactions, and the overall reaction, as:

Anode

Oxidation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Cathode

Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Overall

$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$

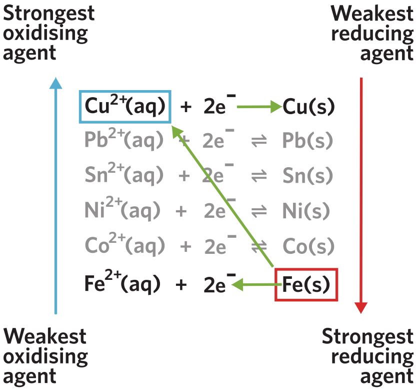
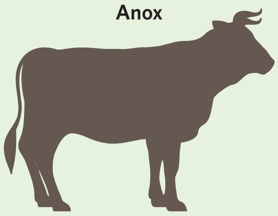


Figure 4 Reactivity series of metals

STRATEGY

A useful mnemonic for remembering the sites of oxidation and reduction in a galvanic cell is 'Anox Redcat':



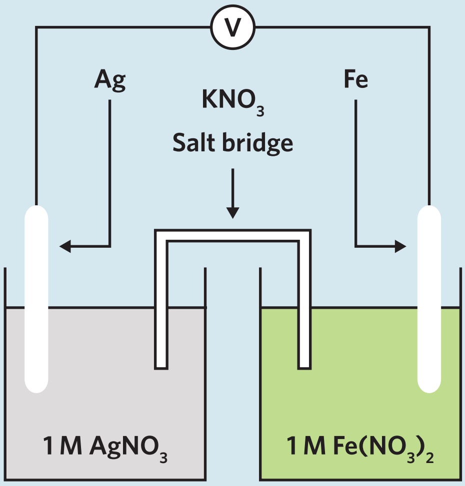
Images (left to right): grmarc, Mary Erskine/Shutterstock.com

Anox – Anode is the site of oxidation

Redcat – Reduction occurs at the cathode

WORKED EXAMPLE 1

Consider the diagram of the galvanic cell shown. Give the site and balanced half equations for the reaction occurring at each electrode, as well as the equation for the overall redox reaction.



What information is presented in the question?

A diagram of a galvanic cell.

What is the question asking us to do?

Identify the half-equations for the reaction occurring.

Give the equation for the overall redox reaction.

What strategies do we need in order to answer the question?

1. Determine which half-equations will produce a spontaneous reaction based on the reactivity series of metals.
2. Write balanced half-equations.
3. Write the balanced overall reaction (including balanced charge).

Continues →

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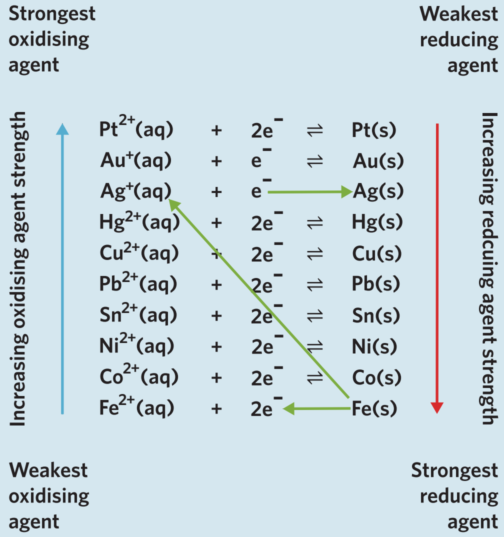
Answer

Using the reactivity series of metals, Fe(s) is a stronger reducing agent than Ag(s).



Therefore Fe(s) will be oxidised, and $Ag^+(aq)$ ions will be reduced.

Anode
Oxidation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Cathode
Reduction: $Ag^+(aq) + e^- \rightarrow Ag(s)$

Overall
$Fe(s) + 2Ag^+(aq) \rightarrow Fe^{2+}(aq) + 2Ag(s)$

Progress questions

Question 1

In a galvanic cell, the cathode is

A. positively charged.

B. negatively charged.

Question 2

Electrons flow from the __________ to the __________ in a galvanic cell.

A. anode, cathode

B. cathode, anode

Question 3

In a galvanic cell, oxidation always occurs at the cathode.

A. True

B. False

Question 4

Which of the following statements is correct about the function of a salt bridge in a galvanic cell?

A. The salt bridge completes the circuit.

B. The salt bridge allows the flow of electrons.

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Corrosion 2.1.12.2

When a metal corrodes, a redox reaction has occurred.

Why do some metals corrode?

Metals are used in many industries worldwide. However, as with most elements and compounds, metals often react to some degree with the environment around them. Some metals will corrode when they come in contact with oxygen, in a process called **dry corrosion**. Sodium, for example, is so reactive with oxygen that it needs to be stored under oil to prevent any contact with oxygen or water. However, in the dry corrosion of aluminium, a strong layer of aluminium oxide, Al_2O_3, forms, as shown in figure 5. This layer protects the pure aluminium beneath from further corrosion, making aluminium a suitable material to use in outdoor conditions.

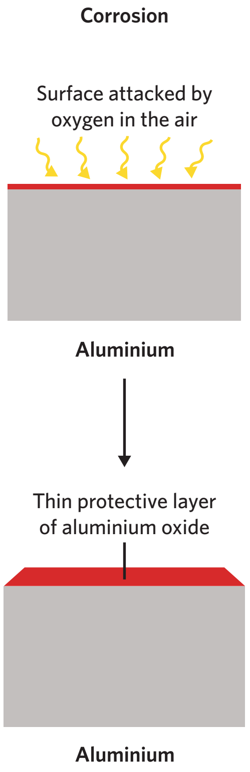

Figure 5 The dry corrosion of aluminium

Unlike sodium and aluminium, iron does not corrode quickly with only oxygen present. However, in the presence of water, iron will corrode more quickly in a process known as **wet corrosion**. The corrosion of iron occurs as follows:

Step 1:

- Solid iron is oxidised: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
- At the same time, the electrons produced from the oxidation of iron are used in the reduction of oxygen: $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$
- Overall equation: $2Fe(s) + O_2(g) + 2H_2O(l) \rightarrow 2Fe^{2+}(aq) + 4OH^-(aq)$

Step 2:

- Solid iron (II) hydroxide is formed: $Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$

Step 3:

- Iron (II) hydroxide is oxidised: $4Fe(OH)_2(s) + O_2(g) + 2H_2O(l) \rightarrow 4Fe(OH)_3(s)$
- Iron (III) hydroxide loses water molecules, forming hydrated iron (III) oxide, $Fe_2O_3{\cdot}xH_2O$, more commonly known as **rust**.

This process is summarised in figure 6.

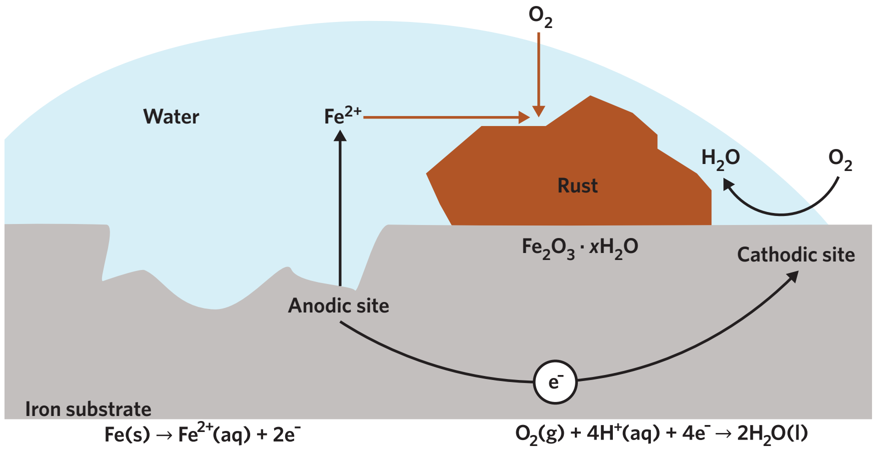

Figure 6 The rusting of iron

The layer of rust that forms will flake away from the chunk of metal, exposing fresh layers of iron. The rusting process can continue until the iron sample has completely corroded.

MISCONCEPTION

'All metals rust.'

Rust is only formed when iron (III) reacts with both oxygen and water to produce hydrated iron (III) oxide. Other metals do not rust – they simply corrode, or oxidise.

USEFUL TIP

The wet corrosion of iron will occur more quickly if the water contains dissolved ions, such as in regions where saltwater is sprayed onto iron structures. This is due to the increased conductivity of water containing ions.

How can we prevent the corrosion of some metals?

Sacrificial protection is a form of protection in which iron acts as a cathode, and a more readily oxidised metal, such as zinc, forms a **sacrificial anode** to protect the iron from corrosion. **Galvanised** iron – iron coated with a zinc layer – is commonly used in the steel hulls of ships[1] to prevent corrosion. It's considered to be a very useful form of surface protection as even if the zinc coating is damaged, the zinc is still preferentially oxidised over the iron, as shown in figure 7.

KEEN TO INVESTIGATE?

[1] **How are ships protected from corrosion?**
Search YouTube: Ship Hull Protection System

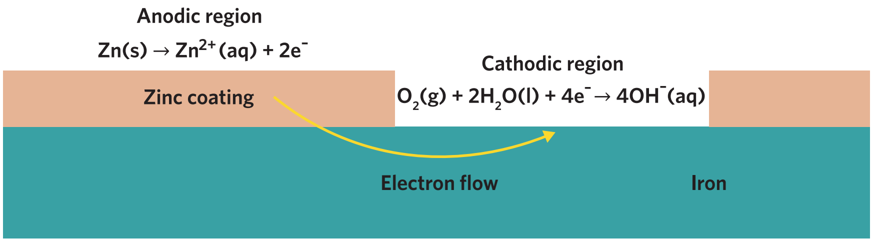

Figure 7 Redox reactions in galvanised iron

Progress questions

Question 5

Dry corrosion occurs in the presence of oxygen and water.

A. True

B. False

Question 6

Which of the following would not be a suitable covering for the storage of sodium?

A. Oil

B. Water

Question 7

In the formation of rust, oxygen is

A. reduced.

B. oxidised.

Question 8

A sacrificial ___________ can be used to protect iron from corrosion.

A. cathode

B. anode

Theory summary

- A galvanic cell consists of two half-cells, each with a metal electrode immersed in a solution of its cations.
  - The two electrodes in a galvanic cell are connected by an external wire.
- The salt bridge separates the two half-cells and this internal circuit allows the flow of ions to complete the overall circuit.
- Electrons flow through the external wire from the anode to the cathode.
- Half-equations and overall equations in galvanic cells can be determined using the reactivity series of metals.
- Top left (strongest oxidising agent) reacts with bottom right (strongest reducing agent) on the reactivity series.
- Dry corrosion is the oxidation of metals in the presence of oxygen only.
- Wet corrosion is the oxidation of metals in the presence of oxygen and water.
  - The presence of moisture accelerates corrosion.
- Sacrificial anodes can be used to prevent the corrosion of metals.

The content in this lesson is considered fundamental prior knowledge to primary galvanic cells and fuel cells as sources of energy (Unit 3 AOS 1).

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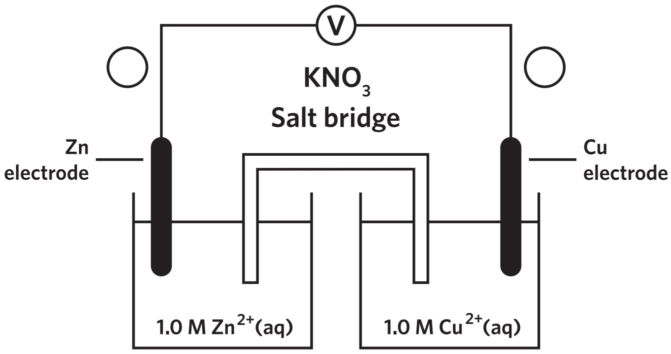
11C Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 9-11.

The Daniell cell, a type of galvanic cell, was first constructed in the mid-1800s and this type of cell is still in use today. A diagram of the Daniell cell is shown.



Question 9 (1 MARK) 🌶

Which of the following options correctly describes the Zn(s) electrode?

A. Positive cathode

B. Positive anode

C. Negative cathode

D. Negative anode

Question 10 (1 MARK) 🌶🌶

Consider the following statements regarding the Daniell cell.

Statement I – The positive electrode is made of copper.

Statement II – K^+(aq) ions would flow towards the zinc solution.

Statement III – Zinc is a stronger reducing agent than copper.

Which of the statements are correct?

A. I only

B. I and II

C. I and III

D. I, II and III

Question 11 (3 MARKS) 🌶🌶

Give the reduction and oxidation equations occurring in the Daniell cell, and hence give the overall equation for the redox reaction occurring (including state symbols).

Adapted from VCAA 2018 exam Short answer Q6b

Exam-style

Question 12 (4 MARKS) 🌶

Complete the following equations for each step in the wet corrosion of iron.

a. $Fe(s) \rightarrow$ (1 MARK)

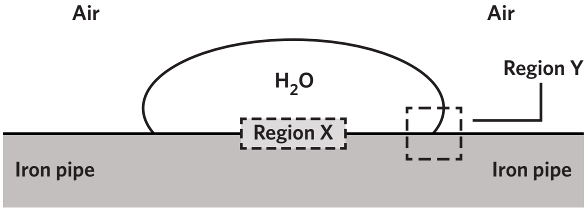
b. $O_2(g) + 2H_2O(l) + 4e^- \rightarrow$ (1 MARK)

c. $Fe^{2+}(aq) + 2OH^-(aq) \rightarrow$ (1 MARK)

d. $4Fe(OH)_2(s) + O_2(g) + 2H_2O(l) \rightarrow$ (1 MARK)

Question 13 (1 MARK)

Iron pipes are used to transport natural gas to cities. Corrosion occurs when water droplets sit on the outer surface of the iron pipe. Miniature galvanic cells are created, with regions such as those shown below, that act as anodes and cathodes.



The type of region and reaction occurring at X in the cell is

	Region	Reaction
A.	anode	$Fe(s) \rightarrow Fe^{2+} + 2e^-$
B.	cathode	$Fe(s) \rightarrow Fe^{2+} + 2e^-$
C.	anode	$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$
D.	cathode	$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$

VCAA 2009 Exam 2 Multiple choice Q13

Question 14 (5 MARKS)

Galvanised iron is commonly used in the construction of the hulls of ships due to its resistance to rusting.

a. Identify the conditions required for rusting to occur. (2 MARKS)

b. Explain how the protective zinc coating prevents galvanised iron from rusting. Include a relevant equation in your answer. (3 MARKS)

Question 15 (1 MARK)

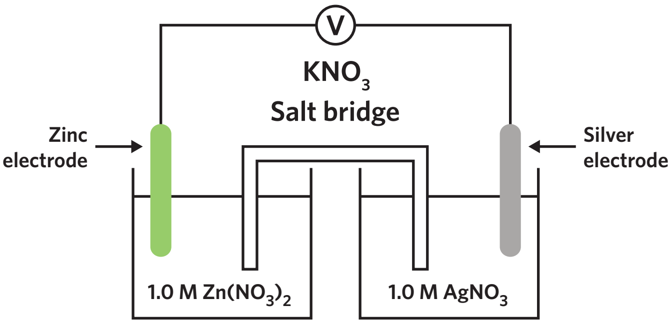
Which of the following statements about the cell shown is true?

A. The silver electrode is the anode.

B. The concentration of Zn^{2+} ions will increase.

C. K^+ ions will flow towards the zinc electrode.

D. Electrons in the external circuit will flow from the silver electrode to the zinc electrode.



Adapted from VCAA 2014 exam Multiple choice Q27

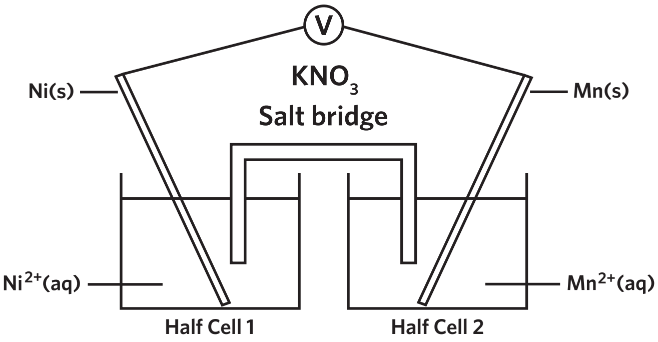
Question 16 (3 MARKS)

A galvanic cell is constructed from the following two half cells under standard conditions.

Half cell 1: a nickel electrode in a solution of 1.0 M nickel (II) nitrate

Half cell 2: a manganese electrode in a solution of 1.0 M manganese (II) nitrate

A sketch of the cell is shown.



Using the reactivity series of metals, in which direction will electrons flow in the external circuit of this galvanic cell? (1 MARK)

a. Give the equation for the half reaction that takes place at the anode of this cell. (1 MARK)

b. What type of energy transformation is occurring in this cell? (1 MARK)

Adapted from VCAA 2008 Exam 2 Short answer Q7

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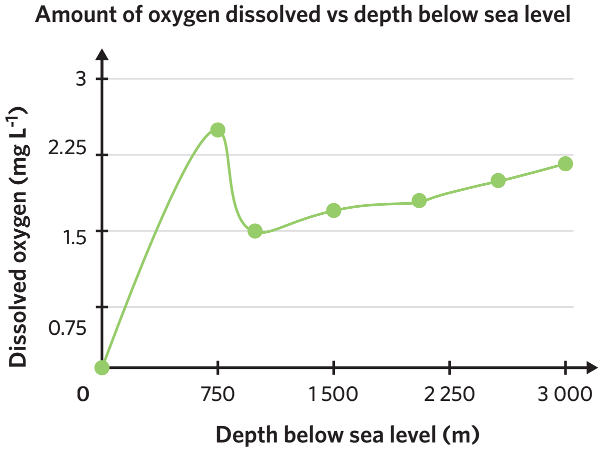
Key science skills

Question 17 (5 MARKS)

A group of scientists were studying deep-sea aquatic life. They undertook an expedition into the Pacific Ocean to measure the oxygen content at several depths below sea level. The graph shown summarises their findings.



a. What type of data was collected by the group of scientists? (1 MARK)

b. Using the information presented in the graph, compare the corrosion of shipwrecks made of iron at 1000 m and 3000 m below sea level. (3 MARKS)

One of the deep sea divers observed two different shipwrecks as she was testing the oxygen levels at different depths. She told the lead scientist that a ship at 2500 m below sea level was significantly less corroded than a ship at 1500 m below sea level.

c. Would this be classified as an anecdote or evidence? (1 MARK)

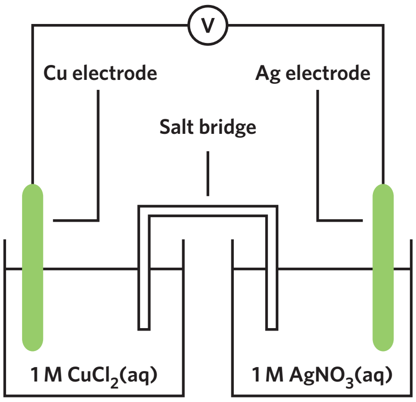
FROM LESSONS 16A, 16B & 16C

Questions from multiple lessons

Question 18 (6 MARKS)

Anna decided to create a galvanic cell as shown.

She decided to use a salt bridge with sodium chloride.



a. After connecting the two half cells together, Anna soon discovered that the galvanic cell wasn't running well, and a solid was beginning to form. Explain this observation. (2 MARKS)

b. In light of this, Anna removed the salt bridge entirely, but no electricity was generated. Why did this happen? (1 MARK)

c. Anna decided to use a salt bridge with potassium nitrate instead. Write out the equations occurring at the anode and cathode, as well as the overall equation occurring in the cell. (3 MARKS)

FROM LESSONS 4C & 11B

Hints

12a.	Recall step 1 in the corrosion of iron summarised in figure 5.
12b.	Recall step 1 in the corrosion of iron summarised in figure 5.
12c.	Recall step 2 in the corrosion of iron summarised in figure 5.
12d.	Recall step 3 in the corrosion of iron summarised in figure 5.
13.	Region X is not exposed to air.
14a.	Rusting is a type of wet corrosion.
14b.	The stronger reducing agent is more readily oxidised.
15.	The oxidation of solid metal produces metal cations.
16a.	Electrons flow from the anode to the cathode.
16b.	The stronger reducing agent is oxidised at the anode.
16c.	Recall the type of energy released from a redox reaction.
17a.	Data is classified based on whether it is numeric or descriptive.
17b.	The rate of corrosion is increased by increased presence of oxygen.
17c.	A story told by someone would be anecdotal evidence.
18a.	Consider the solubility of the potential compounds formed in each half-cell.
18b.	Recall the function of a salt bridge in a galvanic cell.
18c.	The potassium nitrate salt bridge will allow a redox reaction to occur.

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Chapter 11 review

Multiple choice (10 MARKS)

Question 1 (1 MARK)

Which of the following unbalanced redox half-equations does NOT represent a reduction half-reaction?

A. $O_2(g) \rightarrow O^{2-}(s)$

B. $HClO_4(g) \rightarrow HCl(g) + H_2O(l)$

C. $NH_3(aq) \rightarrow NO_2(g)$

D. $MnO_2(s) \rightarrow Mn_2O_3(s)$

Question 2 (1 MARK)

Consider the following redox reaction between manganese (II) ions, $Mn^{2+}(aq)$ and sodium bismuthate, $NaBiO_3(s)$ in acidic solution.

$14H^+(aq) + 2Mn^{2+}(aq) + 5NaBiO_3(s) \rightarrow 7H_2O(l) + 2MnO_4^-(aq) + 5Bi^{3+}(aq) + 5Na^+(aq)$

Which of the following identifies the correct oxidising agent and reducing agent for this reaction?

	Oxidising agent	Reducing agent
A.	$Mn^{2+}(aq)$	$NaBiO_3(s)$
B.	$NaBiO_3(s)$	$Mn^{2+}(aq)$
C.	$Bi^{3+}(aq)$	$MnO_4^-(aq)$
D.	$MnO_4^-(aq)$	$Bi^{3+}(aq)$

Question 3 (1 MARK)

Which of the following best describes the given unbalanced redox half-reaction?

$Mn_2O_3(aq) + H_2O(l) \rightarrow 2MnO_2(s) + 2H^+(aq)$

A. This is a reduction reaction as the oxidation number of oxygen decreases from −2 to −4.

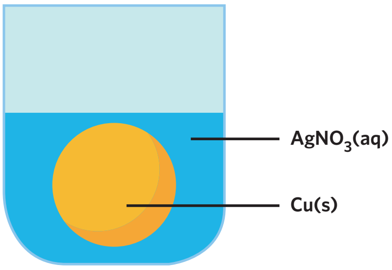
B. This is a reduction reaction as the oxidation number of manganese increases from +3 to +4.

C. This is an oxidation reaction as the oxidation number of oxygen decreases from −2 to −4.

D. This is an oxidation reaction as the oxidation number of manganese increases from +3 to +4.

Question 4 (1 MARK)

A chemistry student was interested in testing the properties of a mesmerising 99.9% solid copper sphere she had recently purchased. In order to do so, the student immersed the sphere in a solution of silver nitrate, $AgNO_3(aq)$.



Would a spontaneous redox reaction be expected to occur in this scenario?

A. Yes, because Cu(s) is a stronger oxidising agent than Ag(s).

B. Yes, because Cu(s) is a stronger reducing agent than Ag(s).

C. No, because Cu(s) is a weaker oxidising agent than Ag(s).

D. No, because Cu(s) is a weaker reducing agent than Ag(s).

Question 5 (1 MARK)

In which one of the following compounds is sulfur in its lowest oxidation state?

A. SO_3

B. HSO_4^-

C. SO_2

D. Al_2S_3

VCAA 2015 exam Multiple choice Q6

Question 6 (1 MARK)

Consider the following reaction between iodine, $I_2(s)$, and lead (IV) oxide, $PbO_2(s)$:

$I_2(s) + 5PbO_2(s) + 8H^+(aq) \rightarrow 2IO_3^-(aq) + 5Pb^{2+}(aq) + 4H_2O(l)$

I. This reaction is unbalanced with respect to charge.

II. Electrons are lost by the $H^+(aq)$ ions and accepted by $H_2O(l)$.

III. $PbO_2(s) + 4H^+(aq) + 2e^- \rightarrow Pb^{2+}(aq) + 2H_2O(l)$ is the balanced reduction half-equation for this reaction.

IV. The oxidation number of iodine increases from 0 to +5 in this reaction.

Which of the above statements about the given reaction are true?

A. III and IV only

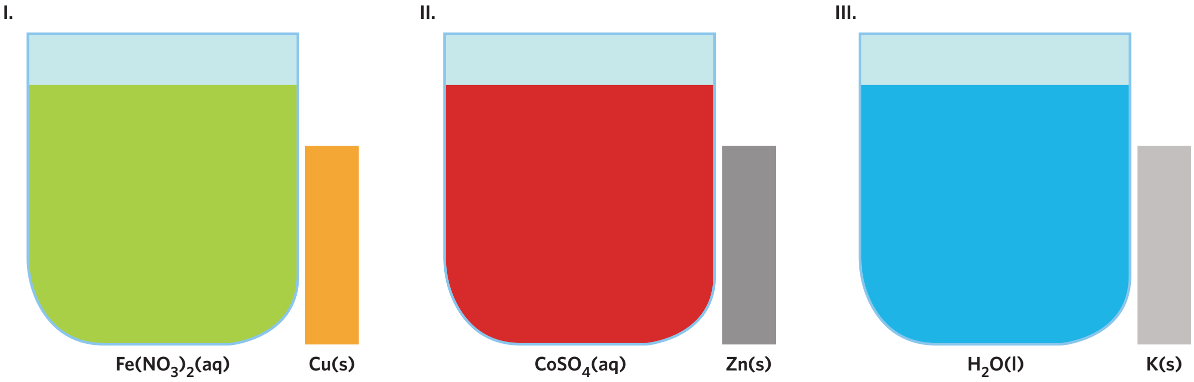
B. I and IV only

C. I, II, and III only

D. II, III, and IV only

Question 7 (1 MARK)

Consider the following combinations of beakers and solid metal strips.



In which beakers, if any, would a reaction be expected to occur if the metal strip was added to the beaker?

A. I and II only

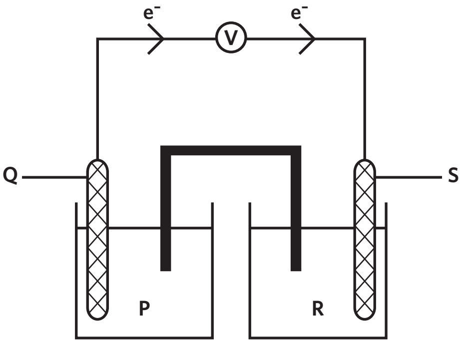
B. II and III only

C. II only

D. No reaction would occur in any of the beakers.

Question 8 (1 MARK)

A diagram of an electrochemical cell is shown.



Which of the following gives the correct combination of the electrode in the oxidation half-cell and the solution in the reduction half-cell?

	Electrode (oxidation half-cell)	Solution (reduction half-cell)
A.	S	P
B.	S	R
C.	Q	R
D.	Q	P

Adapted from VCAA 2020 exam Multiple choice Q3

Question 9 (1 MARK)

Which of the following metals would be a suitable sacrificial anode in preventing the corrosion of iron?

A. Mg(s)

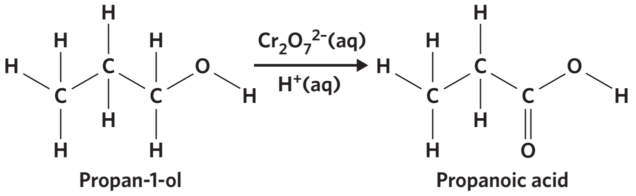
B. Co(s)

C. Ni(s)

D. Ag(s)

Question 10 (1 MARK)

The dichromate ion, $Cr_2O_7^{2-}(aq)$, is a powerful oxidising agent when dissolved in acidic solution. As shown, $Cr_2O_7^{2-}(aq)$ can be used to oxidise the alcohol propan-1-ol to the carboxylic acid propanoic acid. During this redox reaction, $Cr_2O_7^{2-}(aq)$ ions are reduced to $Cr^{3+}(aq)$ ions.



Which of the following gives the balanced overall equation for this reaction?

A. $C_3H_8O(aq) + H_2O(l) \rightarrow C_3H_6O_2(aq) + 4H^+(aq) + 4e^-$

B. $C_3H_8O(aq) + Cr_2O_7^{2-}(aq) + 10H^+(aq) \rightarrow C_3H_6O_2(aq) + 2Cr^{3+}(aq) + 6H_2O(l)$

C. $2C_3H_8O(aq) + 3Cr_2O_7^{2-}(aq) + 34H^+(aq) \rightarrow 2C_3H_6O_2(aq) + 6Cr^{3+}(aq) + 12H_2O(l)$

D. $3C_3H_8O(aq) + 2Cr_2O_7^{2-}(aq) + 16H^+(aq) \rightarrow 3C_3H_6O_2(aq) + 4Cr^{3+}(aq) + 11H_2O(l)$

Short answer (30 MARKS)

Question 11 (6 MARKS)

Oxalic acid, $H_2C_2O_4(aq)$, and permanganate, $MnO_4^-(aq)$, will react together in a redox reaction to produce $CO_2(g)$ and $Mn^{2+}(aq)$ ions in an acidic solution.

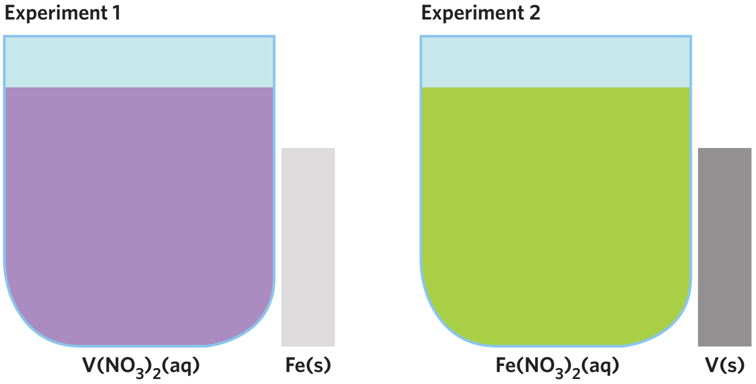
a. Determine the oxidation states for all carbon and manganese atoms in this reaction. (2 MARKS)

b. Which chemical species is undergoing reduction and which chemical species is undergoing oxidation? Explain. (2 MARKS)

c. Write balanced half-equations for the oxidation and reduction half-reactions. (2 MARKS)

Question 12 (8 MARKS)

A student is trying to determine the reactivity of the metal vanadium, V(s), relative to iron, Fe(s). To accomplish this, she sets up the two experiments shown and records her observations in the following table:



Experiment 1	No reaction observed
Experiment 2	Fizzing, solution became a darker colour, vanadium strip appears discoloured/lighter

a. What is the name given to the reaction occurring in Experiment 2? Is this a spontaneous or non-spontaneous redox reaction? (2 MARKS)

b. Give the balanced oxidation and reduction half-equations for this reaction. Hence, give the balanced overall equation for this reaction. (2 MARKS)

c. What do the results of Experiments 1 and 2 suggest about the oxidising and/or reducing strength of vanadium metal, V(s), relative to iron metal, Fe(s)? Justify your answer with reference to the experimental observations. (2 MARKS)

d. If a coil of vanadium metal were instead added to a solution of copper (II) sulfate, $CuSO_4(aq)$, would a reaction be expected to occur? Explain. (2 MARKS)

Question 13 (10 MARKS)

Viktória was investigating the corrosion of the metal parts on their fishing boat. They took two iron bolts, one unused, and one that had been exposed to seawater for a number of years. On the bolt that had been exposed to seawater, they noticed an orange solid had formed.

a. Give the chemical formula and common name for this solid. (2 MARKS)

b. Explain why the corrosion of iron occurs more quickly in seawater than in freshwater. (3 MARKS)

Viktória had a chemistry colleague, Hana, galvanise the unused iron bolt to see if it would be protected from corrosion.

c. Describe what Viktória would observe after a few months of using the newly galvanised bolt on their boat. (2 MARKS)

d. Viktória decided to see if the aluminium parts on their car had corroded to the same extent as the iron bolt they had used for a number of years. What would Viktória expect to see? Justify your answer. (3 MARKS)

Question 14 (6 MARKS)

Consider the given unbalanced redox reaction between nitrate ions, $NO_3^-(aq)$ and hypobromous acid, $HBrO(aq)$. This reaction results in the production of dinitrogen tetroxide, $N_2O_4(g)$, a colourless gas and bromate ions, $BrO_3^-(aq)$.

$NO_3^-(aq) + HBrO(aq) \rightarrow N_2O_4(g) + BrO_3^-(aq)$

a. Give the oxidation numbers of all atoms present in this reaction. (2 MARKS)

b. Determine the balanced overall equation for this redox reaction, given that it occurs in acidic solution. (2 MARKS)

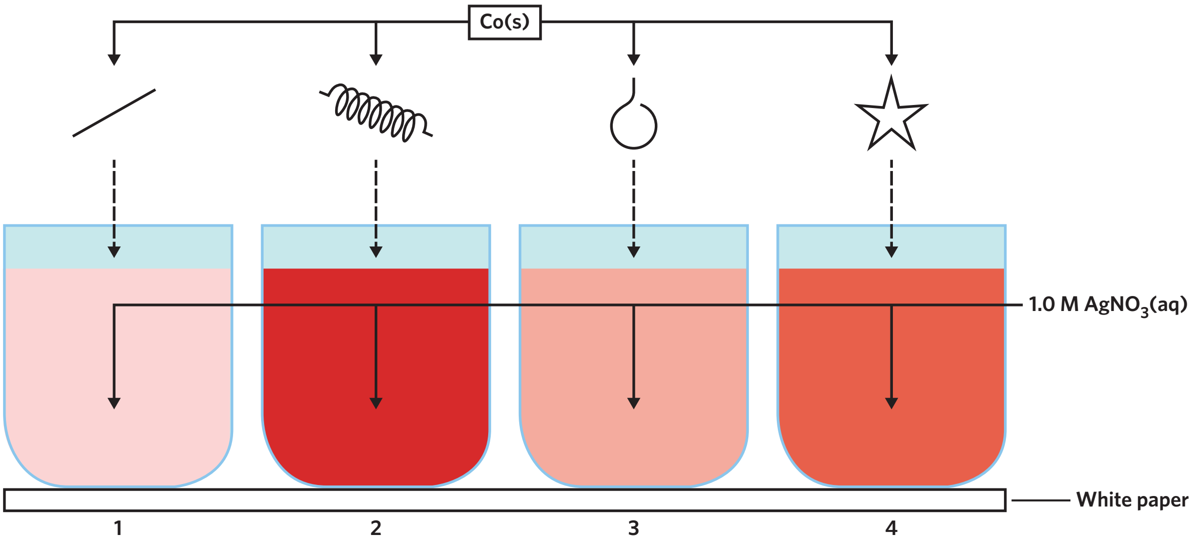
c. If 2.78 mol of $N_2O_4(g)$ is produced in this reaction, calculate the total number of electrons that have been transferred between the nitrate ions and hypobromous acid. (2 MARKS)

Key science skills (10 MARKS)

Question 15 (5 MARKS)

Cristina is a chemistry student interested in exploring the relationship between a metal's surface area and how it behaves in a metal displacement reaction. To do this, she places a number of differently shaped pieces of cobalt wire into separate 1.0 M $AgNO_3(aq)$ solutions.

Cristina knows that $AgNO_3(aq)$ forms a colourless solution, whereas $Co^{2+}(aq)$ ions in low concentrations form a light red solution. She measures the surface area of each piece of cobalt wire, as well as the amount of light that is absorbed by each solution after each piece of wire is left in the solution for 120 seconds. Cristina also asks her friend, Meredith, to describe and record the colour of each solution.



Beaker	Surface area (cm^2)	Absorption (%)	Description
1	7.9	24	Salmon pink
2	37.7	73	Ladybug-red
3	12.6	41	Fire-engine red
4	23.4	62	Like sunset

a. Cristina placed a sheet of white paper behind each beaker whenever its colour was recorded. Suggest a possible reason for this part of the experimental method. (2 MARKS)

b. Give the balanced overall equation for the redox reaction occurring in this experiment. (1 MARK)

c. Meredith wants to estimate the surface area of a new piece of cobalt wire that she has taken out of her mobile phone. After replicating Cristina's experimental conditions and procedure perfectly, she writes down her results:

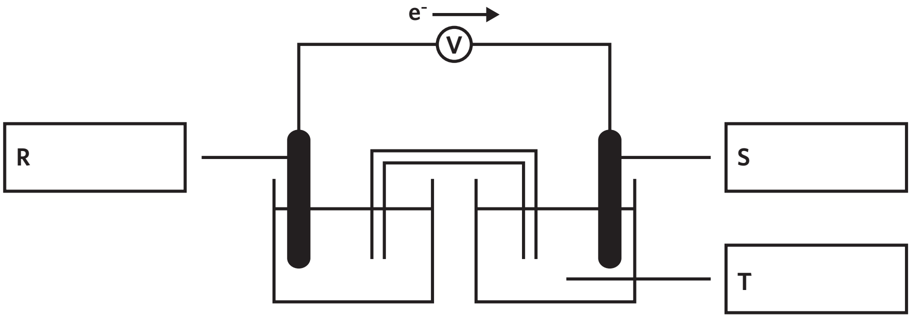
"The solution appears to be the shade of red similar to a ripe tomato."

Can Meredith use this observation to estimate the surface area of the cobalt wire? Explain. (2 MARKS)

FROM LESSON 16B

Question 16 (5 MARKS)

Batteries made up of primary galvanic cells have traditionally been used in small electrical devices. The setup shown consists of a galvanic cell with a zinc electrode in a solution of zinc cations, and a copper electrode in a solution of copper cations.



a. Identify which electrode belongs in each half-cell. Explain. (3 MARKS)

b. Investigations conducted using galvanic cells require a number of safety precautions to be taken. Identify two of them, including a justification for each. (2 MARKS)

Adapted from VCAA 2019 (NHT) exam Short Answer Q5a,b

FROM LESSON 16B

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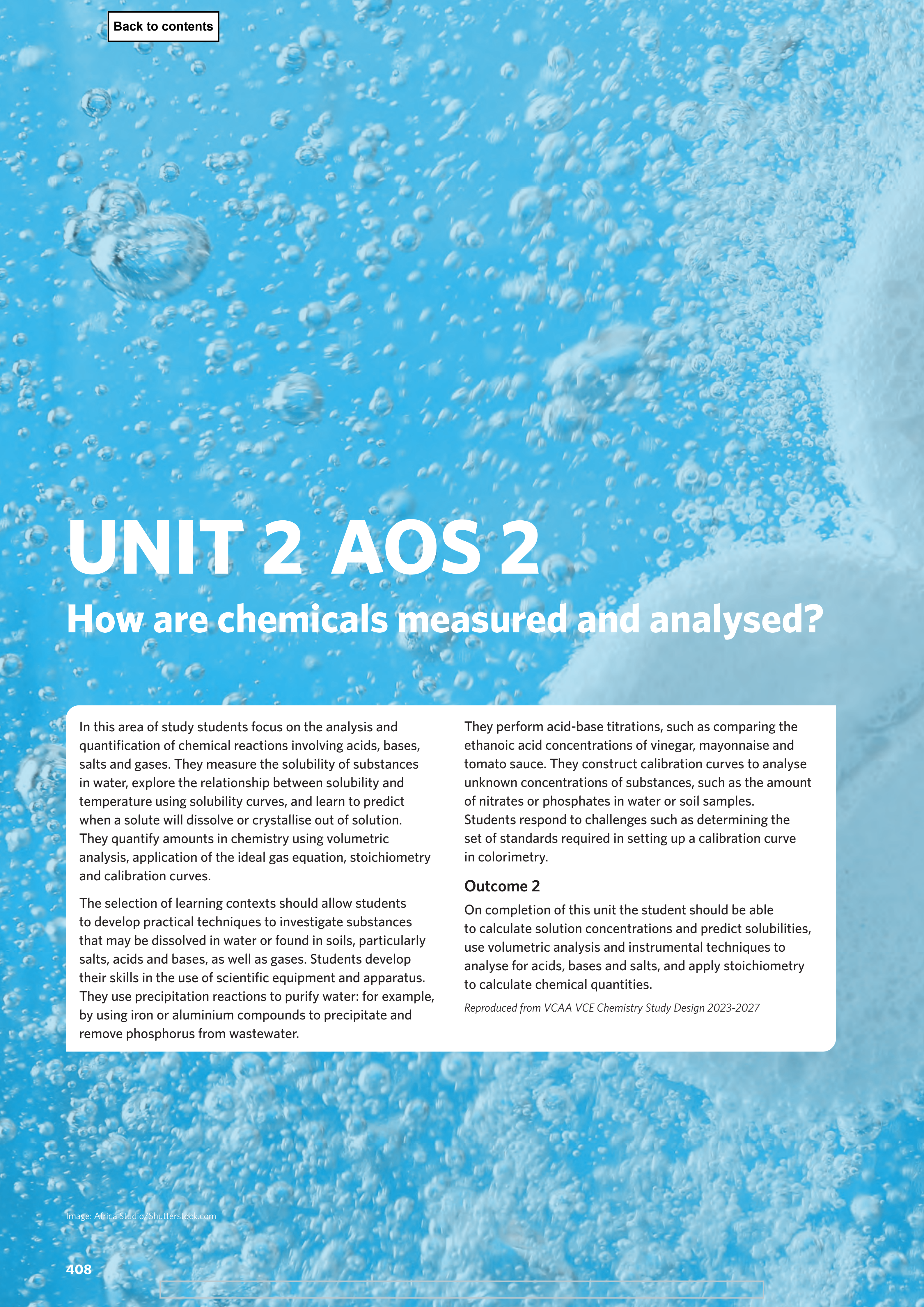

UNIT 2 AOS 2

How are chemicals measured and analysed?

In this area of study students focus on the analysis and quantification of chemical reactions involving acids, bases, salts and gases. They measure the solubility of substances in water, explore the relationship between solubility and temperature using solubility curves, and learn to predict when a solute will dissolve or crystallise out of solution. They quantify amounts in chemistry using volumetric analysis, application of the ideal gas equation, stoichiometry and calibration curves.

The selection of learning contexts should allow students to develop practical techniques to investigate substances that may be dissolved in water or found in soils, particularly salts, acids and bases, as well as gases. Students develop their skills in the use of scientific equipment and apparatus. They use precipitation reactions to purify water: for example, by using iron or aluminium compounds to precipitate and remove phosphorus from wastewater. They perform acid-base titrations, such as comparing the ethanoic acid concentrations of vinegar, mayonnaise and tomato sauce. They construct calibration curves to analyse unknown concentrations of substances, such as the amount of nitrates or phosphates in water or soil samples. Students respond to challenges such as determining the set of standards required in setting up a calibration curve in colorimetry.

Outcome 2

On completion of this unit the student should be able to calculate solution concentrations and predict solubilities, use volumetric analysis and instrumental techniques to analyse for acids, bases and salts, and apply stoichiometry to calculate chemical quantities.

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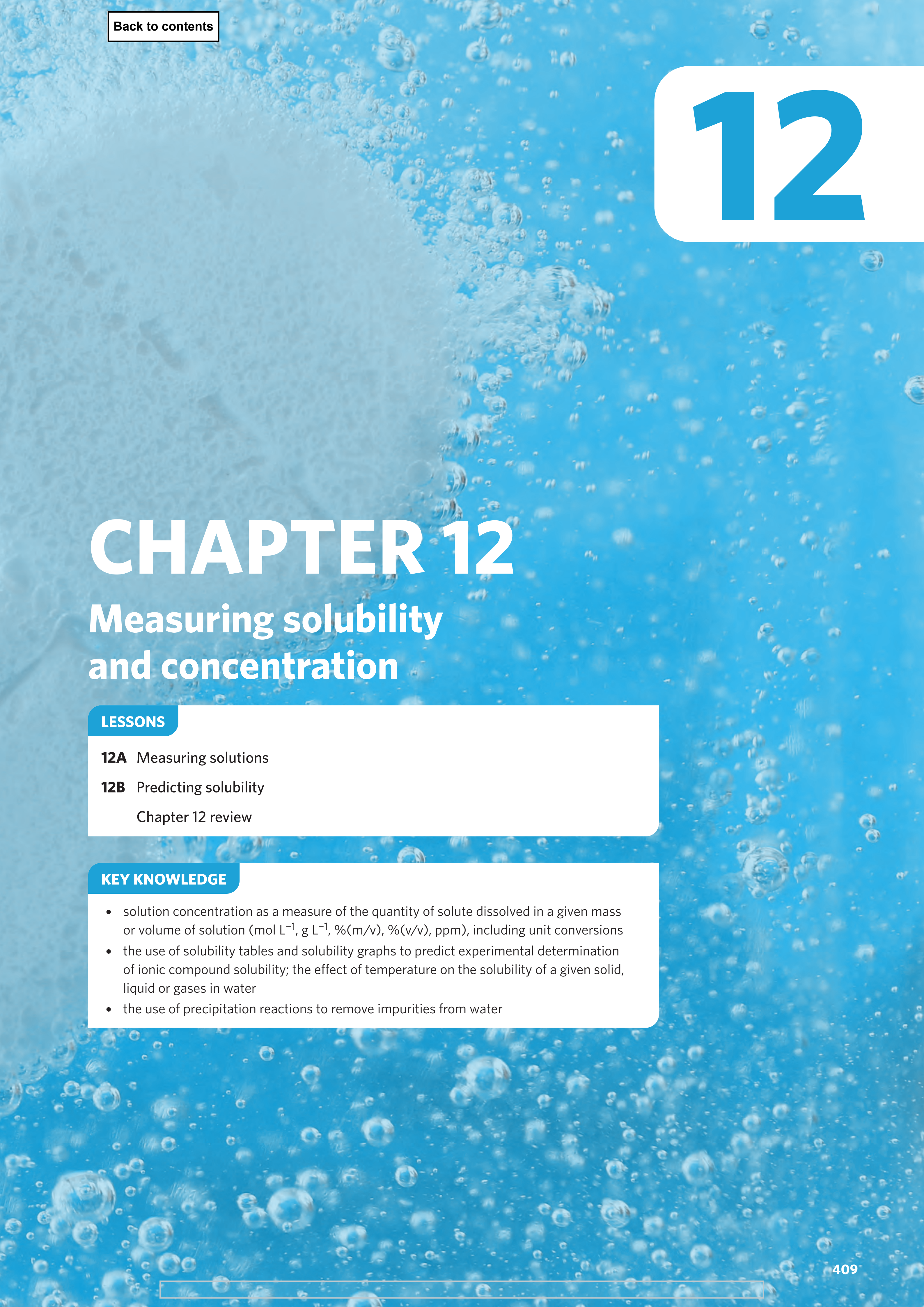

12

CHAPTER 12

Measuring solubility and concentration

LESSONS

KEY KNOWLEDGE

- solution concentration as a measure of the quantity of solute dissolved in a given mass or volume of solution (mol L^{-1}, g L^{-1}, %(m/v), %(v/v), ppm), including unit conversions
- the use of solubility tables and solubility graphs to predict experimental determination of ionic compound solubility; the effect of temperature on the solubility of a given solid, liquid or gases in water
- the use of precipitation reactions to remove impurities from water

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12A Measuring solutions

STUDY DESIGN DOT POINT

- solution concentration as a measure of the quantity of solute dissolved in a given mass or volume of solution (mol L^{-1}, g L^{-1}, %(m/v), %(v/v), ppm), including unit conversions

12A 12B

2.2.1.1 Concentration

2.2.1.2 Units of concentration of solutions

2.2.1.3 Dilution

ESSENTIAL PRIOR KNOWLEDGE

5A Solutions

6B The mole

See questions 79–80.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: Microgen/Shutterstock.com

How can the concentrations of toxic metals like mercury be determined in fish?

There are government standards regarding what **concentration** of mercury can be present in a fish before it is considered unsafe for human consumption. In this lesson, we will learn about the concept of concentration and how the concentration of substances can be expressed in different ways.

KEY TERMS AND DEFINITIONS

Concentration amount of chemical substance in a unit of volume

Concentrated solution solution with a relatively high amount of solute in comparison to the volume of solution

Dilute solution solution with a relatively low amount of solute in comparison to the volume of solution

Dilution process of decreasing concentration of a solution by adding more solvent

Molarity (M) also called molar concentration, is a measure of the concentration of a chemical species in terms of amount of substance per unit volume of solution, measured in moles per litre (mol L^{-1})

Parts per billion (ppb) one in one billion parts of a solution

Parts per million (ppm) one in one million parts of a solution

Solute substance that dissolves in a solute

Solution homogeneous mixture in which a substance is dissolved in another

%(m/m) mass percent concentration of solute in solution

%(w/w) weight percent concentration of solute in solution

%(m/v) mass percent concentration of solute per unit of volume of solution

%(v/v) volume percent concentration of solute in solution

Concentration 2.2.1.1

Concentration is a measure of the amount of **solute** in a given volume of **solution**.

How can the quantity of a substance in solution be described?

As we learned in lesson 5A, at the most basic level, solutions are made up of a solute and solvent. The concentration of a solution is a measure of the amount of substance per unit of volume of a solution.

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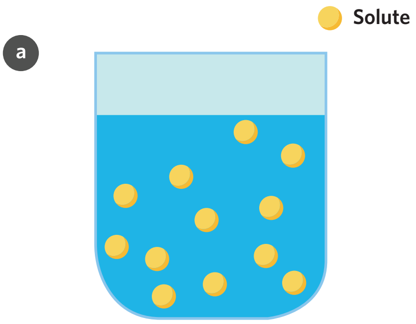
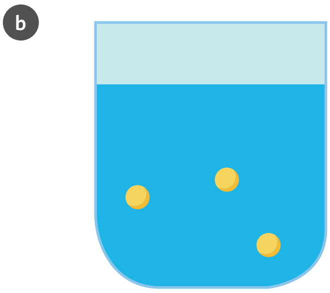
A solution with a relatively high amount of solute per unit of volume is described as being a **concentrated solution**, whilst a solution with relatively little solute per unit of volume is referred to as being a **dilute solution**, as shown in figure 1.

An important distinction to make is that a solution with a high number of total solute particles is not necessarily a concentrated solution. Whether a solution is concentrated or dilute depends on the ratio of solute particles to the total volume of the solution. For example, in figure 1, we can see that the more concentrated solution contains 12 solute particles per litre of solution, whilst the dilute solution only contains 3 solute particles per litre of solution.



Solution A
Concentrated 1 L solution



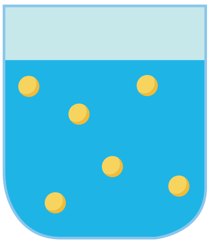
Solution B
Dilute 1 L solution

Figure 1 A concentrated solution **(a)** and a dilute solution **(b)**

Progress questions

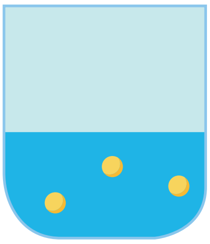
Question 1

Which of the beakers, A or B, would have the same concentration as the solution in the reference beaker?

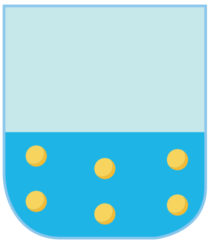


Reference beaker

A.

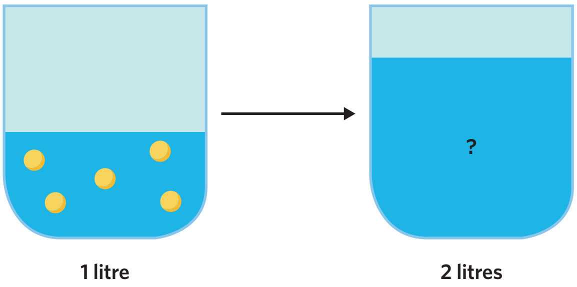


B.



Question 2

How many particles of solute would be found in the second beaker if both solutions have the same concentration?



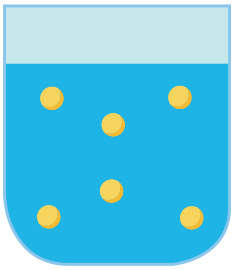
A. 5

B. 10

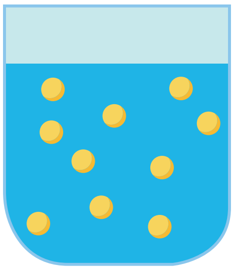
Question 3

Which of the following beakers is more dilute?

A.



B.



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Units of concentration of solutions 2.2.1.2

There are a number of different units used to express the concentration of a solution.

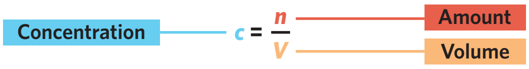
How can the amount of substance in a solution be quantified?

USEFUL TIP

The units mol L^{-1} and M can be used interchangeably. They both mean moles per litre!

The concentration of a solution can be quantified by calculating the amount of substance per unit of volume. In chemistry, one common way to determine the concentration of solutions is by calculating the amount, in mol, of solute in a volume of solution. This is known as the **molarity** of a solution, and is represented by the units mol L^{-1} or M.

To calculate the concentration of a solution in mol L^{-1}, we can use the following formula:



$$c = \frac{n}{V}$$

Suppose we added 6 mol of sodium chloride, NaCl(s), to one litre of water, as shown in figure 2.

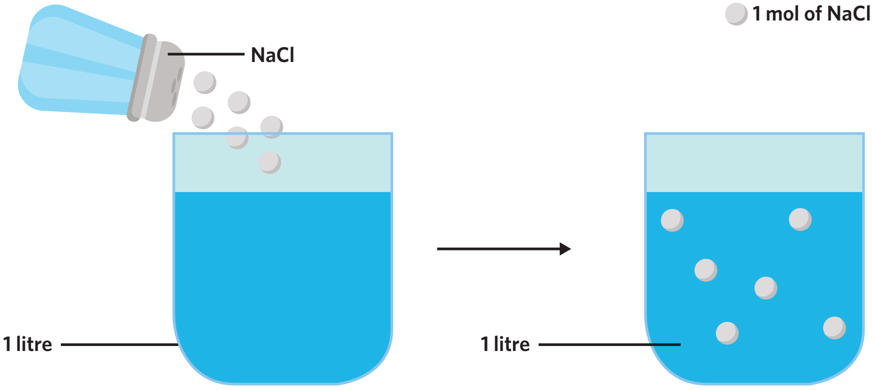


Figure 2 Creating a sodium chloride, NaCl(s), solution

USEFUL TIP

When attempting to calculate the concentration of a solution in mol L^{-1}, always remember to convert the volume of solution into litres.

Therefore, to calculate the concentration of sodium chloride in the solution.

$$c(\text{NaCl}) = \frac{n(\text{NaCl})}{V(\text{water})}$$

$$c(\text{NaCl}) = \frac{6 \text{ mol}}{1 \text{ L}}$$

$$c(\text{NaCl}) = 6 \text{ mol L}^{-1}$$

WORKED EXAMPLE 1

Calculate the concentration, in mol L^{-1}, of a 100 mL solution containing 0.25 mol of NaCl.

What information is presented in the question?

The volume of the solution is 100 mL.

The amount of NaCl is 0.25 mol.

What is the question asking us to do?

Calculate the concentration of the solution in mol L^{-1}.

What strategies do we need in order to answer the question?

1. Convert the volume into appropriate units.
2. Use the formula for the concentration of a solution.

Answer

$$V = \frac{100 \text{ mL}}{1000}$$

$$V = 0.100 \text{ L}$$

$$c = \frac{n}{V}$$

$$c = \frac{0.25 \text{ mol}}{0.100 \text{ L}}$$

$$c = 2.5 \text{ mol L}^{-1}$$

Now that we know how to calculate the concentration of a solution in mol L^{-1}, we can use this formula to learn other information about a solution. We can calculate the amount of solute in a solution of known volume and known concentration, or we can calculate the volume of a solution with a known concentration and known amount of solute.

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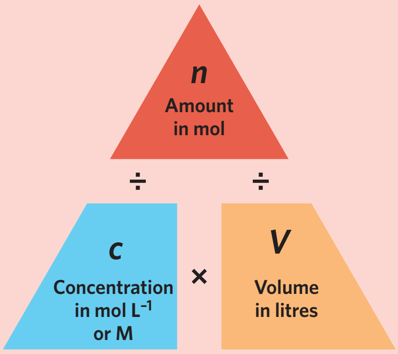
STRATEGY

When doing calculations involving concentration, amount, and volume, we can use the following concept triangle to work out what formula to use.



To use this triangle, cover up the value you want to find with your finger, and calculate using the remaining variables and sign. For example, to find the amount, cover the 'amount' section of the triangle and calculate $c \times V$.

As we learned in 6B, the mole concept is connected to many different quantitative values, and can be calculated when given the mass of a substance. We can therefore also calculate concentration in mol L^{-1} when we are given the mass of solute in a solution.

WORKED EXAMPLE 2

5.0 g of HCl was dissolved in 200 mL of water to create a solution. Find the concentration in mol L^{-1} of the HCl solution.

What information is presented in the question?

The mass of HCl dissolved in solution is 5.0 g.

The volume of water is 200 mL.

What is the question asking us to do?

Calculate the concentration of the solution in mol L^{-1}.

What strategies do we need in order to answer the question?

1. Calculate the amount of HCl present.
2. Convert the volume into appropriate units.
3. Use the formula for the concentration of a solution.

Answer

$$n(\text{HCl}) = \frac{m}{M}$$

$$n(\text{HCl}) = \frac{5.0\text{ g}}{(1.0 + 35.5)\text{ g mol}^{-1}}$$

$$n(\text{HCl}) = 0.137\text{ mol}$$

$$V = \frac{200\text{ mL}}{1000}$$

$$V = 0.200\text{ L}$$

$$c = \frac{0.137\text{ mol}}{0.200\text{ L}}$$

$$c = 0.68\text{ mol L}^{-1}$$

WORKED EXAMPLE 3

A 200 mL sample of a 1.7 M solution of HCl was used in a reaction. What amount, in mol, of HCl is present?

What information is presented in the question?

The volume of a 1.7 mol L^{-1} HCl solution is 200 mL.

What is the question asking us to do?

Calculate the amount, in mol, of HCl present.

What strategies do we need in order to answer the question?

1. Convert the volume into appropriate units.
2. Use the formula for the concentration of a solution.

Answer

$$V(\text{HCl}) = \frac{200\text{ mL}}{1000}$$

$$V(\text{HCl}) = 0.200\text{ L}$$

$$c = \frac{n}{V}$$

Therefore, $n = c \times V$

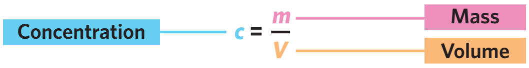
$$n(\text{HCl}) = 1.7\text{ mol L}^{-1} \times 0.200\text{ L}$$

$$n(\text{HCl}) = 0.34\text{ mol}$$

By law, the products that we purchase from supermarkets must include information about the concentration of certain ingredients. For example, liquid medications must have the concentration of their active ingredients clearly labelled, as shown in figure 3. The concentration is typically expressed as a unit of mass per unit of volume, such as grams per litre (g L^{-1}) or milligrams per litre (mg L^{-1}).

Figure 3 A liquid medication with a zidovudine concentration of 10 mg mL^{-1}

To calculate the concentration of a solution in a unit of mass per a unit of volume, we can use the following formula:



$$c = \frac{m}{V}$$

where the volume is expressed in litres, L, or millilitres, mL, and the mass is expressed in grams, g, or milligrams, mg.

WORKED EXAMPLE 4

Calculate the concentration in mg L^{-1} of a liquid medication that contains 40.0 mg of paracetamol in 100 mL.

What information is presented in the question?

The mass of paracetamol is 40.0 mg.

The volume of medication is 100 mL.

What is the question asking us to do?

Calculate the concentration of the solution in mg L^{-1}.

What strategies do we need in order to answer the question?

1. Convert the volume into the appropriate units.
2. Use the appropriate formula for the concentration of a solution.

Answer

$$V = \frac{100 \text{ mL}}{1000}$$

$$V = 0.100 \text{ L}$$

$$c = \frac{m}{V}$$

$$c = \frac{40.0 \text{ mg}}{0.100 \text{ L}}$$

$$c = 400 \text{ mg L}^{-1}$$

For substances that have very small quantities, we measure concentration in **parts per million (ppm)** or **parts per billion (ppb)**. This refers to a single part of a substance per one million or billion parts, respectively, of another substance, as shown in figure 4.

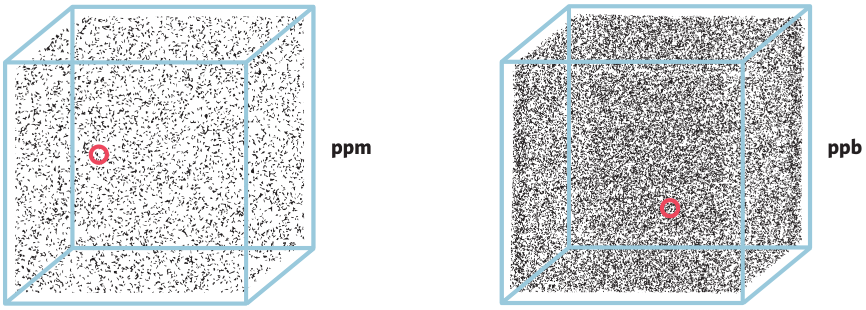


Figure 4 Illustration of parts per million and parts per billion

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These units are used to detect the concentrations of active ingredients in drugs, dangerous metals in food products and drinking water, and even organic pollutants in pesticides. For example, in Victoria, the concentration of fluoride in drinking water cannot exceed 1 ppm, a level recommended by the World Health Organisation (WHO). Fluoride is added into water[1] and toothpastes to help strengthen the structure of teeth.

Ppm and ppb can be expressed using different, equivalent units, as shown in figures 5 and 6. For example 1 kg consists of 1 million mg. As a result, 1 mg in a 1 kg sample represents 1 ppm.

$$1\text{ kg} = 1\,000\,000\text{ mg} \qquad \frac{\text{Mass of solute (mg)}}{\text{Mass of solute (kg)}} = \text{ppm} = \frac{\text{Mass of solute (mg)}}{\text{Volume of solution (L)}} \qquad 1\text{ L (of solution)} = 1\,000\,000\text{ mg}$$

Figure 5 Different ways to express ppm

$$1\text{ kg} = 1\,000\,000\,000\ \mu\text{g} \qquad \frac{\text{Mass of solute } (\mu\text{g})}{\text{Mass of solute (kg)}} = \text{ppb} = \frac{\text{Mass of solute } (\mu\text{g})}{\text{Volume of solution (L)}} \qquad 1\text{ L (of solution)} = 1\,000\,000\,000\ \mu\text{g}$$

Figure 6 Different ways to express ppb

KEEN TO INVESTIGATE?

[1] **How much fluoride is in the water we drink?**
Search: Water fluoridation in Victoria

USEFUL TIP

ppb is not part of the study design for units 1 and 2 VCE chemistry, however, it is an important unit to understand when exploring the concept of concentration. It has a number of applications in chemistry, such as when describing the concentration of gases in the atmosphere or when measuring the concentration of contaminants in soils.

WORKED EXAMPLE 5

A solution has a concentration of 1.010 g L^{-1}. Express this concentration in ppm.

What information is presented in the question?

The concentration of the solution is 1.010 g L^{-1}.

What is the question asking us to do?

Convert the concentration of the solution to ppm.

What strategies do we need in order to answer the question?

1. Determine the reference unit.
2. Determine the unit that is required for ppm.
3. Convert the mass of solute accordingly.

Answer

Based on the concentration given, we need to express our answer 'per L'. To express this concentration in ppm requires the units to be changed from g L^{-1} to mg L^{-1}.

$m = 1.010\text{ g}$

$m = 1.010 \times 1000\text{ mg}$

$m = 1010\text{ mg}$

Therefore, the concentration of the solution is 1010 ppm.

WORKED EXAMPLE 6

Express the concentration of a 2.0×10^{-9} mol L^{-1} $MgCl_2$ solution in ppm.

What information is presented in the question?

The concentration of the solution is 2.0×10^{-9} M.

What is the question asking us to do?

Convert the concentration of the solution to ppm.

What strategies do we need in order to answer the question?

1. Determine the units required for ppm.
2. Convert any values to the appropriate units.
3. Convert the concentration to ppm.

Answer

The solution is expressed in mol L^{-1}. Ppm can be expressed as mg L^{-1}. Considering both are expressed 'per litre', we need to convert the mol value to mg.

A concentration of 2.0×10^{-9} mol L^{-1} means that there are 2.0×10^{-9} moles of solute per litre of solution. We need to convert the mole value to mass of $MgCl_2$ in mg:

$m = n \times M$

$m(MgCl_2) = 2.0 \times 10^{-9}\text{ mol} \times (24.3 + (2 \times 35.5))\text{ g mol}^{-1}$

$m(MgCl_2) = 1.906 \times 10^{-7}\text{ g}$

$m(MgCl_2) = 1.906 \times 10^{-7}\text{ g} \times 1000\text{ mg}$

$m(MgCl_2) = 1.91 \times 10^{-4}\text{ mg}$

Therefore, the concentration of the solution is

1.9×10^{-4} ppm.

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Concentration can also be expressed as a percentage of solute per solvent. This can be done in a number of ways, depending on the state of the solute and solvent. For example, for a solution where both solute and solution quantities are expressed as masses, we can calculate the concentration as % mass/mass (**%(m/m)**) or % weight/weight (**%(w/w)**). Table 1 outlines the common percentage concentrations.

Table 1 Percentage concentrations

Concentration	Description	Calculation
% mass/mass* %(m/m)	Both solute and solution are expressed as a unit of mass eg. mg, g, kg	$c = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$ *masses must be in same units
% weight/weight* %(w/w)	Both solute and solution are expressed as a unit of mass eg. mg, g, kg	$c = \frac{\text{weight of solute}}{\text{weight of solution}} \times 100\%$ * weights must be in same units
% mass/volume **%(m/v)**	The solute is expressed in grams whereas the solution is expressed in millilitres (g mL^{-1})	$c = \frac{\text{mass of solute g}}{\text{volume of solution mL}} \times 100\%$
% volume/volume* **%(v/v)**	Both solute and solution are expressed as a unit of volume (eg. mL, L)	$c = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\%$ *volumes must be in same units

As these values are expressed as a percentage, the percentage concentration value can be interpreted as an amount per 100. For example, an alcohol with a concentration of 75.0 %(v/v) has 75 mL of alcohol in every 100 mL of the beverage, or 75 mL/100 mL. Since the density of water is approximately 1.0 g mL^{-1}, we can also convert volumes to masses.

USEFUL TIP

%(m/m) and %(w/w) are not part of the study design for units 1 and 2 VCE chemistry, however, they are useful units to understand when investigating percentage concentrations. They can be used when quantities of solute and solution are expressed in units of mass.

WORKED EXAMPLE 7

A 1.5 L solution contains 21.2 g of NaCl. What is the %(m/v) concentration of the NaCl solution?

What information is presented in the question?

The volume of the solution is 1.5 L.

The mass of NaCl in the solution is 21.2 g.

What is the question asking us to do?

Convert the concentration of the solution to %(m/v).

What strategies do we need in order to answer the question?

1. Determine the units used to express %(m/v).
2. Convert the given values to the appropriate units.
3. Calculate the concentration by %(m/v).

Answer

%(m/v) is expressed as g per mL, however our values are in g and L. Therefore, we need to convert the 1.5 L solution to mL:

$V = 1.5 \text{ L}$

$V = 1.5 \times 1000 \text{ mL}$

$V = 1500 \text{ mL}$

$\%(\text{m/v}) = \frac{21.1 \text{ g}}{1500 \text{ mL}} \times 100\%$

$\%(\text{m/v}) = 1.4\%$

USEFUL TIP

When converting units, the VCE Data Book can be used to determine the metric prefixes as well as the relevant multiplying factors.

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Progress questions

Question 4

The molarity of a solution is expressed as __________ per __________.

A. grams, litre

B. moles, litre

Question 5

Which of the following units of concentration is equivalent to 1 ppm?

A. 1 mg g^{-1}

B. 1 mg L^{-1}

Question 6

To convert molarity to ppm

A. the units for the amount of solute should be converted from mol to mg.

B. the units for the volume of solution should be converted from L to mL.

Question 7

To calculate concentration as %(v/v),

A. any unit can be used so long as it fits the unit of measurement.

B. the volumes must be in the same units.

Question 8

2.35 g of $CaCO_3$ was dissolved to make a 30.0 mL solution. Which of the following units would be most suitable to express the concentration of this solution?

A. %(v/v)

B. %(m/v)

Dilution 2.2.1.3

Diluting a solution makes it less concentrated.

How can we change the concentration of a solution?

Earlier in the lesson, we came across the concept of dilute solutions. Since the concentration of a solution measures the amount of solute in a solution, adding more solvent into the solution without changing the amount of solute will make the solution more dilute, as shown in figure 7. The process of decreasing the concentration of a solution is known as **dilution**.

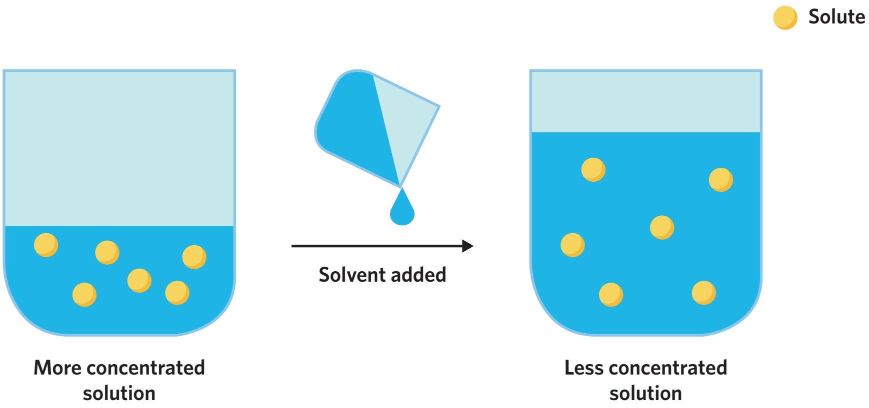

Figure 7 Process of diluting a solution

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MISCONCEPTION

'When a solution is diluted, the amount of solute changes.'

When a solution is diluted, the concentration of solute in the solution changes – not the total amount of solute (moles) present in the solution.

As we can see in figure 7, there is no change in the amount of solute in the solution when more solvent is added. Therefore, given the amount of solute in the solution before dilution is the same after the dilution, we can use the formula for the amount of solute in a solution we learned earlier to create a new formula, as shown:

Original solution before dilution

$$c_1 \times V_1 = n = c_2 \times V_2$$

Solution after more solvent has been added

Here c_1 and V_1 represent the respective concentration and volume of the solution before dilution, and c_2 and V_2 represent the respective concentration and volume of the solution after dilution respectively. We can simplify our equation by removing *n*:

Before dilution

$$c_1 \times V_1 = c_2 \times V_2$$

After dilution

WORKED EXAMPLE 8

50.0 mL of water was added to 20.0 mL of 2.0 mol L^{-1} KCl solution. Calculate the concentration of the diluted solution.

What information is presented in the question?

The initial volume of the solution is 20.0 mL.

The initial concentration of the solution is 2.0 mol L^{-1}.

50 mL of water is added to the initial solution.

What is the question asking us to do?

Calculate the concentration of the diluted solution.

What strategies do we need in order to answer the question?

1. Determine the final volume of water in the diluted solution.
2. Determine the values for c_1, V_1, c_2, and V_2.
3. Calculate the concentration of the diluted solution.

Answer

Since 50 mL of water was added, the final volume of the diluted solution is:

$$V_2 = 50.0\ \text{mL} + 20.0\ \text{mL}$$

$$V_2 = 70.0\ \text{mL}$$

$$V_2 = \frac{70.0}{1000}\ \text{L}$$

$$V_2 = 0.0700\ \text{L}$$

$$c_1 = 2.0\ \text{mol L}^{-1}$$

$$V_1 = 20.0\ \text{mL}$$

$$V_1 = \frac{20.0}{1000}\ \text{L}$$

$$V_1 = 0.0200\ \text{L}$$

$$c_2 = ?$$

$$V_2 = 0.0700\ \text{L}$$

$$c_1 \times V_1 = c_2 \times V_2$$

$$2.0\ \text{mol L}^{-1} \times 0.0200\ \text{L} = c_2 \times 0.0700\ \text{L}$$

$$c_2 = \frac{2.0\ \text{mol L}^{-1} \times 0.0200\ \text{L}}{0.0700\ \text{L}}$$

$$c_2 = 0.57\ \text{mol L}^{-1}$$

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WORKED EXAMPLE 9

How much water would need to be added to 150 mL of a 2.0 mol L^{-1} solution to achieve a final concentration of 1.5 mol L^{-1}?

What information is presented in the question?

The initial volume of the solution is 150 mL.

The initial concentration of the solution is 2.0 mol L^{-1}.

The final concentration of the solution must be 1.5 mol L^{-1}.

What is the question asking us to do?

Calculate the volume of water required to dilute the solution to 1.5 mol L^{-1}.

What strategies do we need in order to answer the question?

1. Determine the values for c_1, V_1, c_2, and V_2.
2. Calculate the final volume of the diluted solution.
3. Calculate the volume of water required.

Answer

$c_1 = 2.0 \text{ mol L}^{-1}$

$V_1 = 150 \text{ mL}$

$V_1 = \frac{150}{1000} \text{ L}$

$V_1 = 0.150 \text{ L}$

$c_2 = 1.5 \text{ mol L}^{-1}$

$V_2 = ?$

$c_1 \times V_1 = c_2 \times V_2$

$2.0 \text{ mol L}^{-1} \times 0.150 \text{ L} = 1.5 \text{ mol L}^{-1} \times V_2$

$V_2 = \frac{2.0 \text{ mol L}^{-1} \times 0.150 \text{ L}}{1.5 \text{ mol L}^{-1}}$

$V_2 = 0.20 \text{ L}$

Volume required $= V_2 - V_1$

Volume required $= 0.20 \text{ L} - 0.150 \text{ L}$

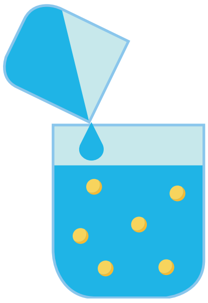
Volume required $= 0.050 \text{ L}$

Progress questions

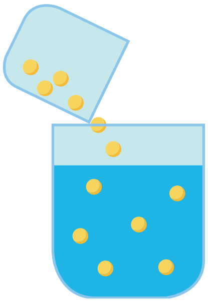
Question 9

Which of the following scenarios illustrates the process of dilution?

A.



B.



● Solute

Question 10

When comparing a diluted solution with its original solution,

A. both solutions have the same concentration.

B. both solutions contain the same amount of solute.

Theory summary

- Concentration measures the amount of solute per unit of solution.
- Molarity is a common unit of concentration, expressed in mol L^{-1} (or M).
- ppm and ppb measure the amount of solute per million or billion parts of solvent respectively.
- Concentration can also be expressed as %(m/m), %(w/w), %(m/v) and %(v/v).
- When solutions are diluted, the amount of solute (moles) remains the same.

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12A Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 11–13.

A student was asked to make a 500 mL sample of a 2.5 mol L^{-1} NaOH solution.

Question 11 (1 MARK) 🌶

The solution would contain

A. 2.5 mol of NaOH for every 1.0 L of solution.

B. 2.5 g of NaOH for every 500 mL of solution.

C. 500 mol of NaOH for every 2.5 g of solution.

D. 500 g of NaOH for every 500 g of solution.

Question 12 (1 MARK) 🌶

Which of the following formulas would be used to calculate the molarity of the solution?

A. $\%(\text{m/v}) = \dfrac{\text{mass (in g)}}{\text{volume (in mL)}}$

B. $c = \dfrac{\text{amount (in mol)}}{\text{volume (in L)}}$

C. $\text{ppm} = \dfrac{\text{mass (in µg)}}{\text{mass (in kg)}}$

D. $\%(\text{v/v}) = \dfrac{\text{volume (in mL)}}{\text{volume (in L)}}$

Question 13 (2 MARKS) 🌶🌶

Calculate the mass of NaOH found in the solution.

Exam-style

Question 14 (17 MARKS) 🌶

There are different ways to express concentration.

a. Give the concentration of each of the following solutions in the units indicated.

  i. Molarity of a 200 mL solution containing 1.47 g of $CaCl_2$. (2 MARKS)

  ii. 200 mL of a NaOH solution containing 2.0 g of NaOH, in ppm. (2 MARKS)

  iii. 200 mL of a NaOH solution containing 2.0 mol of NaOH, in ppm. (2 MARKS)

  iv. 100 mL of solution containing 0.50 mol of $FeSO_4$ in %(m/v). (2 MARKS)

b. Convert the units of each solution accordingly.

  i. 1337 ppm → g L^{-1} (1 MARK)

  ii. 981 ppm KCl → mol L^{-1} (2 MARKS)

  iii. 1.5 mol L^{-1} PbI_2 → %(m/v) (2 MARKS)

c. Calculate the following concentrations.

  i. The initial concentration of a 20.0 mL sample that was diluted to a 50.0 mL solution with a concentration of 2.0 mol L^{-1}. (2 MARKS)

  ii. The final concentration when 200 mL of water is added to 150 mL of a 1.05 mol L^{-1} solution. (2 MARKS)

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Question 15 (3 MARKS)

An iced coffee brand has a nutritional label as shown.

Nutrition Information	
Servings per pack: 1.5	**Serving size: 500 mL**
	Average quantity per 100 mL
Protein	3.4 g
Carbohydrate	9.0 g
Sugars	8.7 g

a. What is the concentration of sugar in the iced coffee in %(m/v)? (1 MARK)

b. Given sugar has a molar mass of 180 g mol^{-1}, calculate the concentration of sugar in the iced coffee in mol L^{-1}. (2 MARKS)

Question 16 (9 MARKS)

Fish that are sold in Australia cannot have a mercury concentration of more than 1 ppm.

a. What is the maximum mass of mercury that can be found in a snapper fish weighing 500 g? (2 MARKS)

b. After net fishing in Port Melbourne, Ganesh caught a small fish. All of the mercury was collected and made into a 50 mL solution with a concentration of 1.76 mol L^{-1}. What mass of mercury was present in the fish? (2 MARKS)

c. Given that the fish had an initial weight of 2.67 kg and a mercury concentration of 1.76 mol L^{-1}, would this meet the Australian standards of mercury levels? Explain. (3 MARKS)

d. How much water would be required to dilute the mercury solution so that it has a concentration of 1.0 mol L^{-1}? (2 MARKS)

Question 17 (5 MARKS)

Miguel bought a bottle of cordial that was shown to have a 50 %(m/v) concentration of sugar, $C_6H_{12}O_6$.

a. Calculate the mass of sugar in every litre of the cordial solution. (1 MARK)

b. Calculate the concentration of sugar in the cordial in mol L^{-1}. (2 MARKS)

c. Miguel poured a 50 mL sample of the cordial into a cup and wanted to dilute it so that the final sample had a volume of 200 mL. Calculate the concentration, in mol L^{-1}, of sugar in the final solution. (2 MARKS)

Question 18 (6 MARKS)

Pesticides such as aldicarb can leak into waterways and cause serious health risks when consumed by humans. It has been advised that a safe concentration of aldicarb in the human bloodstream is 7.0 ppm.

a. What mass of aldicarb, in mg, is safe in the bloodstream of a toddler weighing 13.6 kg? (1 MARK)

b. The toddler drank 200 mL of water containing 1.82×10^{-5} mol L^{-1} aldicarb. Has the toddler exceeded the safe concentration of aldicarb in their bloodstream? Include calculations in your answer (M(aldicarb) = 190.3 g mol^{-1}). (5 MARKS)

Key science skills

Question 19 (5 MARKS)

A student prepared a solution for a reaction by following the method below:

1. Add 250 mL of water into a beaker.
2. Weigh 4.00 g of solid NaCl.
3. Add the NaCl sample into the beaker of water and stir.

a. On the way back from the scale, the student tripped over their shoelace and spilled a small amount of NaCl. Since the amount lost was minimal, they decided to proceed to make the solution anyway. What type of error is this? (1 MARK)

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b. How would this affect the concentration of the solution? (2 MARKS)

c. What is the theoretical concentration of the solution? (2 MARKS)

FROM LESSON 16D

Questions from multiple lessons

Question 20 (5 MARKS)

Tran wanted to make a 2.5 mol L^{-1} KCl solution.

a. What type of compound is KCl? (1 MARK)

b. Write the equation to show the dissociation of KCl. (1 MARK)

c. What mass of solid KCl needs to be added to 100 mL of water to produce the desired solution? (2 MARKS)

d. Balance the equation for the reaction between silver nitrate ($AgNO_3$) and potassium chloride. (1 MARK)

FROM LESSONS 4B & 5A

Hints

14ai. Molarity is equivalent to mol L^{-1}.
14aii. ppm is equivalent to mg L^{-1}.
14aiii. ppm is equivalent to mg L^{-1}.
14aiv. Convert the amount of $FeSO_4$ to g.
14bi. ppm is equivalent to mg L^{-1}.
14bii. Convert ppm to mg L^{-1}.
14biii. Calculate the amount, in grams, of PbI_2.
14ci. Solve $c_1 \times V_1 = c_2 \times V_2$ for c_1.
14cii. Solve $c_1 \times V_1 = c_2 \times V_2$ for c_2.
15a. The mass and volume need to be expressed in particular units of measurement.
15b. Molarity depends on the amount of solute per unit of volume.
16a. ppm can be determined by identifying the mass : volume ratio.
16b. The amount of solute present depends on the volume of the sample.
16c. The mass of substance depends on the amount of substance.
16d. Dilution changes the concentration of the solution, but not the total amount of solute.
17a. %(m/v) depends on the mass of substance per unit of volume.
17b. Values presented as a % can be viewed as an amount per 100.
17c. Dilution changes the concentration of the solution, but not the total amount of solute.
18a. The total mass of albicarb depends on the total weight.
18b. To compare data, the values need to be expressed in the same unit of measurement.
19a. Errors are categorised based on their effect on the results.
19b. Concentration depends on the amount of solute present.
19c. Concentration depends on the amount of solute present per unit of volume.
20a. Compounds are generally named based on the type of intramolecular bonding.
20b. Dissociation equations do not include water.
20c. Concentration depends on the amount of solute per unit of volume.
20d. A reaction between ionic compounds results in the 'swapping' of ion pairs.

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12B Predicting solubility


Image: Opachevsky Irina/Shutterstock.com

Why do salt crystals form at the surface of the dead sea?

As water in the dead sea evaporates, it becomes highly saturated with salt, causing crystals to form readily on the water's surface. In this lesson, we will explore the effect of temperature on the **solubility** of substances, the use of **solubility curves**, and how **precipitation reactions** can be used to purify water.

STUDY DESIGN DOT POINTS

- the use of solubility tables and solubility graphs to predict experimental determination of ionic compound solubility; the effect of temperature on the solubility of a given solid, liquid or gases in water
- the use of precipitation reactions to remove impurities from water

12A 12B

2.2.2.1 Effect of temperature on solubility

2.2.2.2 Solubility curves

2.2.3.1 Purifying water

ESSENTIAL PRIOR KNOWLEDGE

4C	Precipitation reactions
4C	Solubility tables
5A	Solubility

See questions 81–83.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

KEY TERMS AND DEFINITIONS

Alum	inorganic chemical compound composed of a hydrated aluminium salt
Crystallisation	process by which solute molecules leave solution to form a highly organised structure known as a crystal
Eutrophication	excessive algal growth due to increased nutrient concentration in water
Precipitation reaction	chemical reaction resulting in the formation of a solid from an aqueous solution
Saturated solution	solution containing the maximum amount of dissolved solute
Solubility	ability to be dissolved in a given solvent
Solubility curve	graph plotting the solubility of a substance (vertical axis) against temperature (horizontal axis)
Solubility table	table listing the solubilities of compounds formed from different ions when dissolved in water
Supersaturated solution	solution containing more than the maximum amount of solute that can be dissolved
Unsaturated solution	solution containing less than the maximum amount of dissolved solute

Effect of temperature on solubility 2.2.2.1

Changes in temperature can impact the extent to which a solute is able to dissolve in a given solvent.

How does solubility change with temperature?

As we learned in lesson 4C, **solubility tables** can be used to predict the occurrence of precipitation reactions. Whilst solubility tables provide a general overview of whether an ionic compound will dissolve in water, there are also other factors to be considered. One such factor that affects the solubility of a substance is temperature. For example, figure 1 shows how the solubility of sugar increases with increasing temperature.

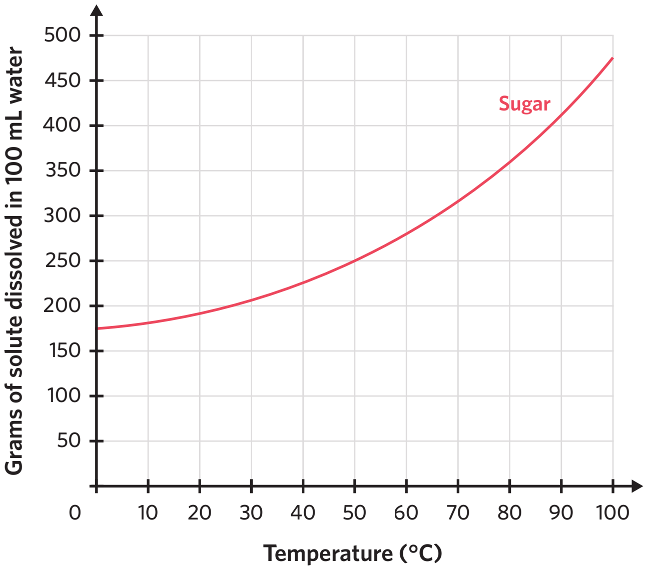


Figure 1 Solubility of sugar in water

USEFUL TIP

Whilst solubility tables provide information on the ability of a substance to dissolve at a given temperature, they do not show the effect of temperature on solubility.

As in the case of sugar (figure 1), an increase in temperature results in an increase in the solubility of many solid and liquid substances. However, increasing the temperature does not always increase solubility; in some cases, increasing the temperature can even decrease the solubility of a solute. For example, increasing the temperature of a cerium (III) sulfate solution reduces the amount of cerium (III) sulfate that can dissolve. Nonetheless, it is important to note that – for the majority of substances – there is a correlation between the temperature and the solubility of a solute.

This proportional relationship between solubility and temperature typically only applies for solid and liquid solutes. When considering the properties of gaseous solutes in a solution, as the temperature of the solution increases, the solubility of these substances will decrease, as shown in figure 2.

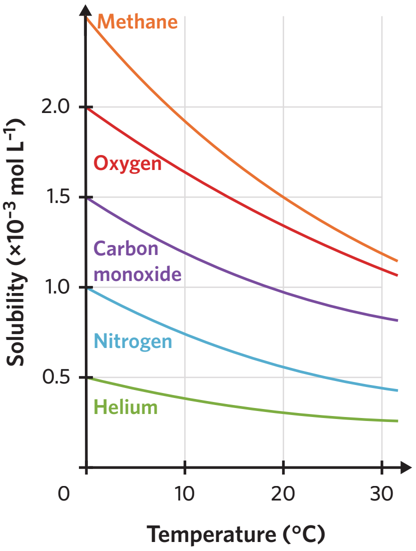


Figure 2 Solubility of common gases

MISCONCEPTION

'The solubility of a substance always increases with temperature.'

Whilst many substances do become more soluble as the temperature of the solution increases, substances such as cerium (III) sulfate and a number of gases instead become less soluble with increasing temperature.

Progress questions

Question 1

The term solubility refers to

A. the number of particles dissolved in a solution.

B. the extent to which a solute is able to dissolve in a solvent.

Question 2

Solubility tables

A. show the effect of temperature on solubility.

B. can be used to predict the solubility of ionic compounds.

Continues →

Question 3

Which of the following statements is most accurate?

A. The effect of temperature on solubility depends on the substance.

B. Increasing temperature increases solubility.

Question 4

For gaseous solutes, increasing temperatures generally

A. increases solubility.

B. decreases solubility.

Solubility curves 2.2.2.2

Solubility curves provide a method of visualising and comparing the solubility of substances in water at different temperatures.

How do we interpret solubility graphs?

So far, we have learned that the solubility of a given substance is not fixed – it varies with temperature. We can visualise the relationship between temperature and solubility using a solubility curve, as demonstrated in the graph shown in figure 3.

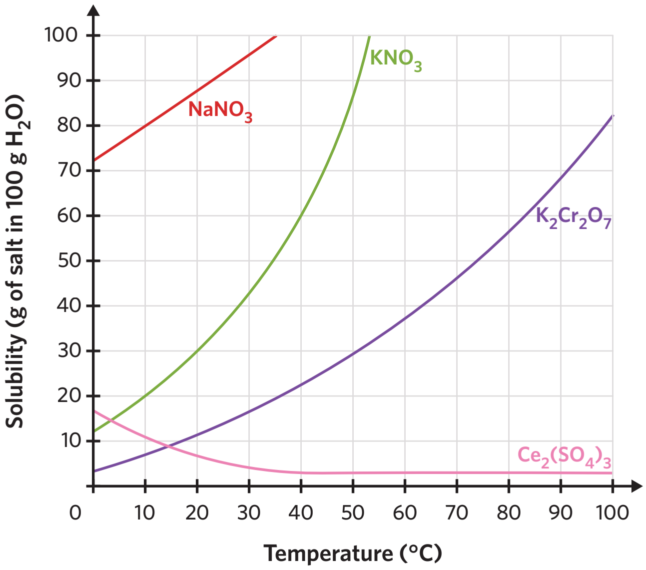


Figure 3 Solubility curves of some common ionic compounds

In figure 3, we can see the solubility curves of four different compounds. The advantage of a graph like this is that it enables us to compare the solubility of substances at a given temperature. For example, in figure 3, 30 g of KNO_3 will dissolve in 100 g of water at 20 °C. However, only around 12 g of $K_2Cr_2O_7$ will dissolve in 100 g of water at 20 °C. Therefore, we can say that at 20 °C, KNO_3 is more soluble in water than $K_2Cr_2O_7$.

The other key advantage of a solubility curve is that it enables us to visualise the general trend in the solubility of a compound as temperature changes. In figure 3, we can see that the solubility of $NaNO_3$ in water generally increases with increasing temperature. On the other hand, the solubility of $Ce_2(SO_4)_3$ in water generally decreases with increasing temperature.

At any given temperature, the corresponding solubility (in g per 100 g of H_2O) indicates the amount of solute that would be dissolved in a **saturated solution**. Therefore, at a given temperature, solutions with solubilities that lie above the curve indicate a **supersaturated solution**, whilst solutions with solubilities that lie below the curve indicate an **unsaturated solution**. Consider the solubility curve for $K_2Cr_2O_7$ shown in figure 4.

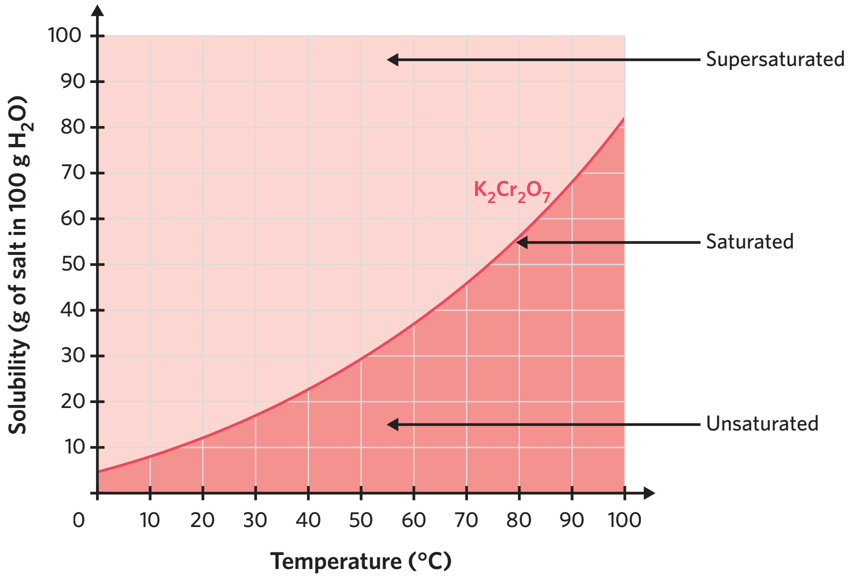


Figure 4 Solubility curve for $K_2Cr_2O_7$

From figure 4, we can see that a saturated $K_2Cr_2O_7$ solution at 50 °C would contain 30 g of solute per 100 mL of water. Therefore, a solution at 50 °C that contains more than 30 g of dissolved solute would be considered supersaturated, whilst a solution at 50 °C that contains less than 30 g of dissolved solute would be considered unsaturated.

As shown in figure 5, unsaturated solutions still have the potential to dissolve more solute, whilst saturated solutions cannot dissolve any more solute. When a supersaturated solution is formed and subsequently disturbed, crystals will form[1] (figure 5).

KEEN TO INVESTIGATE?

[1] **What happens when sodium acetate is added to a saturated solution?**
Search YouTube: Supersaturated Solutions – Working with Sodium Acetate

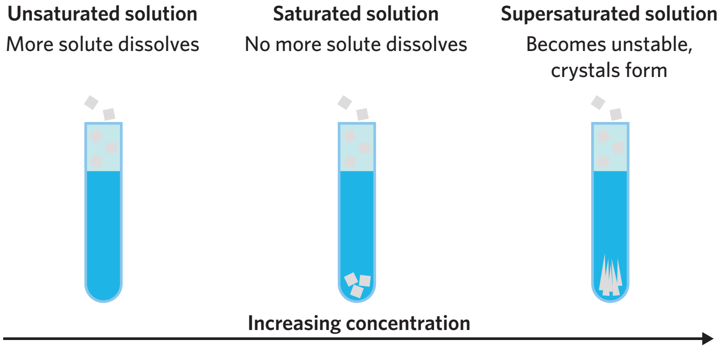


Figure 5 Unsaturated, saturated, and supersaturated solutions

The key pieces of information that we can obtain from a solubility curve are:

- the relative solubility of a substance,
- the general trend in solubility of a substance as temperature changes,
- how much solute can be dissolved to create a saturated solution,
- how much dissolved solute creates a supersaturated solution,
- and, how much solute would be dissolved in an unsaturated solution.

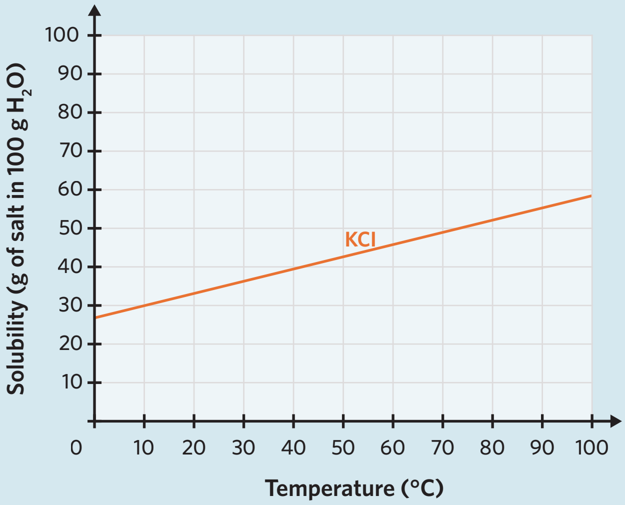
How can we use solubility curves to predict the solubility of substances?

Solubility curves enable us to visualise and predict the relationship between temperature changes and the solubility of different substances in a solvent. This means solubility curves can be used to:

- predict the amount of solute that can be dissolved at a given temperature,
- predict the amount of solute that won't dissolve when added to a solvent,
- and, predict the mass of crystals formed in a supersaturated solution.

WORKED EXAMPLE 1

Using the solubility curve provided, predict the mass, in grams, of KCl that would dissolve in 300 g of H_2O at a temperature of 90 °C.



What information is presented in the question?

The identity of the solute is KCl.

The mass of solvent, H_2O, is 300 g.

The temperature of the solution is 90 °C.

What is the question asking us to do?

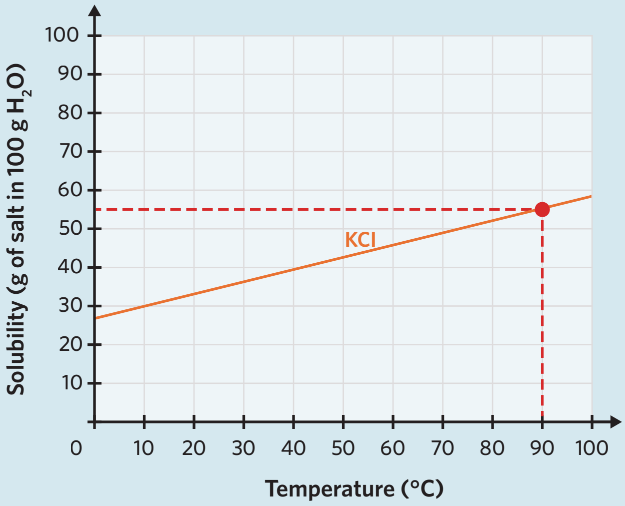
Predict the amount of KCl, in grams, that will dissolve in 300 g of H_2O at 90 °C.

What strategies do we need in order to answer the question?

1. Using the curve, determine the number of grams of KCl dissolved in 100 g of water at 90 °C.
2. Use this value to find the amount of KCl dissolved in 300 g of water.

Answer

According to the solubility curve, 55 grams of KCl can dissolve in 100 g of H_2O at 90 °C.



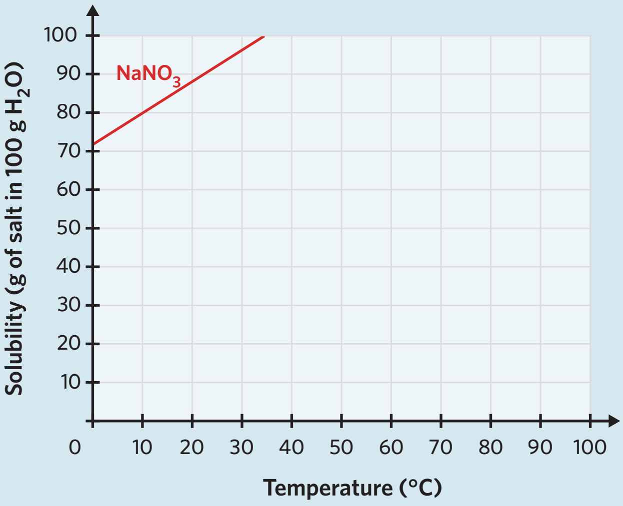
Therefore, to find the amount of KCl that would dissolve in 300 g of H_2O at 90 °C, this value must be multiplied by 3.

$55 \text{ g} \times 3 = 165 \text{ g}$

Therefore, 165 g of KCl would dissolve in 300 g of H_2O at 90 °C.

WORKED EXAMPLE 2

Using the solubility curve provided, predict the mass of crystals produced when 120 g of $NaNO_3$ is added to 100 g of H_2O at 10 °C.



What information is presented in the question?

120 g of $NaNO_3$ is added to 100 g of H_2O at 10 °C.

What is the question asking us to do?

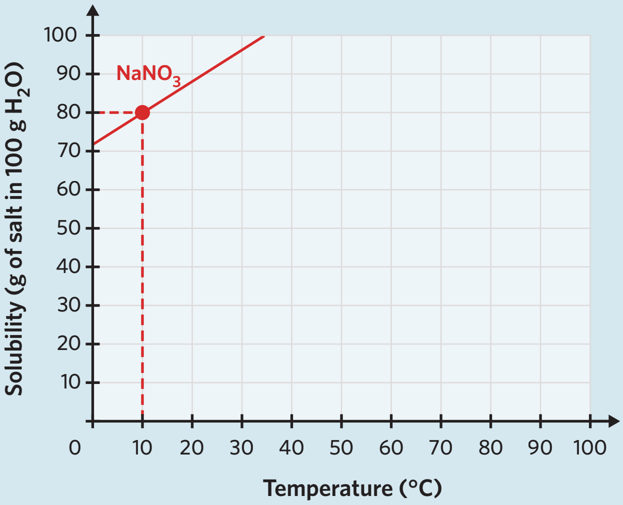
Predict the mass of $NaNO_3$ crystals that would form.

What strategies do we need in order to answer the question?

1. Using the curve, find the mass of $NaNO_3$ that would dissolve in 100 g of H_2O at 10 °C.
2. Find the difference between the mass of $NaNO_3$ that can be dissolved and the amount added.

Answer

According to the solubility curve, 80 g of $NaNO_3$ can dissolve in 100 g of H_2O at 10 °C.



Crystals form when a supersaturated solution is disturbed. Therefore, the mass of crystals formed is the difference between the amount of $NaNO_3$ added to the water and the amount that can dissolve in 100 g of H_2O at 10 °C.

$120\text{ g} - 80\text{ g} = 40\text{ g}$

Therefore, 40 g of $NaNO_3$ crystals will be formed.

Solubility curves have a number of different domestic and industrial applications. An example of this is in the pharmaceutical industry, where solubility curves are often used to determine the temperature at which a reaction should take place. Solubility curves are also used in the pharmaceutical industry to determine the point at which a solution becomes supersaturated, as there are some pharmaceuticals that require **crystallisation**.

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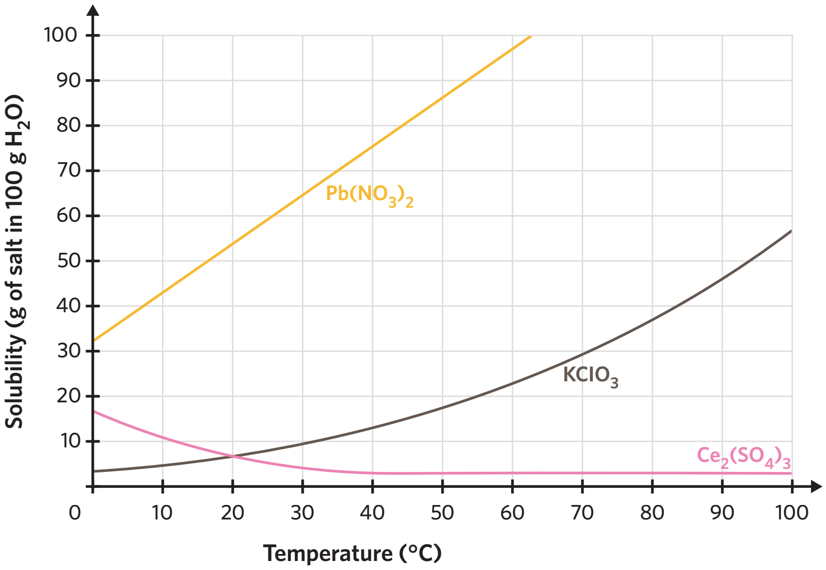
Progress questions

Question 5



Using the solubility curve shown, determine which of the following statements is true?

A. The solubility of $KClO_3$ increases exponentially with increasing temperature.

B. The solubility of $Pb(NO_3)_2$ increases exponentially with increasing temperature.

Question 6

A saturated solution is

A. a solution that can dissolve more solute.

B. a solution containing the maximum amount of dissolved solute.

Question 7

The formation of crystals takes place when

A. a supersaturated solution is disturbed.

B. a saturated solution is heated.

Question 8

An unsaturated solution

A. is able to dissolve more solute.

B. will form crystals.

Question 9

Concentrations in the area above a solubility curve indicate

A. saturated solutions.

B. supersaturated solutions.

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Purifying water 2.2.3.1

Water can be purified using the principles of solubility.

How can precipitation reactions be used to purify water?

As we learned in lesson 4C, precipitation reactions involve the formation of an insoluble salt from two aqueous reactants. Precipitation reactions are an important tool in the wastewater treatment industry, as they can be used to purify large bodies of water.

In nature, one of the biggest threats to water purity is **eutrophication**, the process wherein a body of water becomes oversaturated with nutrients, leading to an increase in the growth of algae. Excess algal growth significantly deteriorates the quality of water and depletes the amount of dissolved oxygen in bodies of water. Phosphate is known to stimulate algal growth, and therefore, increased phosphate concentration is associated with decreased water purity. Phosphate can enter waterways from a number of sources, such as human and animal waste, cleaning wastewater, industrial sewage, and fertiliser runoff. Once it enters a body of water, the phosphates settle at the bottom.

USEFUL TIP

Other soluble anions, such as nitrates (NO_3^-), also contribute to eutrophication. Nitrates are removed from water through processes such as distillation and reverse osmosis.

Alum – a hydrated salt of aluminium – is commonly used in water treatment plants to reduce the concentration of phosphate in water.[2] A crystal of potassium alum is shown in figure 6.

KEEN TO INVESTIGATE?

[2] **Is alum safe to use?**
Search: This chemical turns polluted green lakes clear. Is it safe?

Image: SOORACHET KHEAWHOM/Shutterstock.com

Figure 6 A crystal of potassium alum

When added to water, alum reacts with H_2O to form aluminium hydroxide. Phosphate released from lake beds reacts with aluminium hydroxide to form an insoluble aluminium phosphate precipitate. This precipitate then settles at the bottom of lakes and no longer contributes to eutrophication, thereby improving the quality of the water. The process of water purification using alum is shown in figure 7.

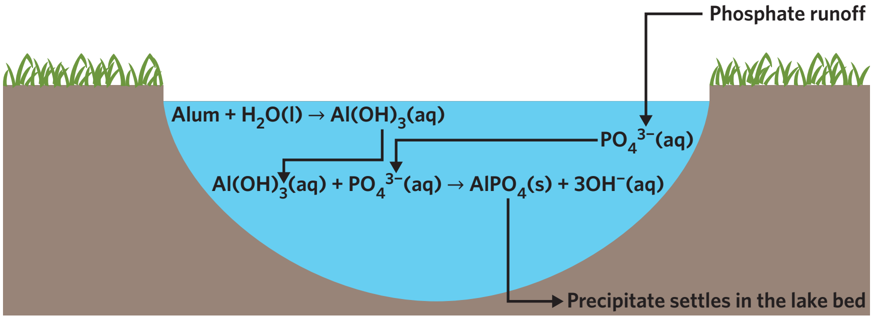

Figure 7 Using alum to precipitate excess phosphate in water purification

Progress questions

Question 10

Eutrophication is the overgrowth of algae due to excess nutrients in water.

A. True

B. False

Question 11

Which of the following is a source of phosphate in large bodies of water?

A. Fertiliser runoff

B. Alum

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Theory summary

- Increasing temperature generally increases the solubility of solids and decreases the solubility of gases.
- Solubility curves can be used to visualise the relationship between temperature and solubility.
- Solutions can be unsaturated, saturated, or supersaturated.
  - Supersaturated solutions form crystals when disturbed.
- Water can be purified using the precipitation reaction that occurs between aluminium hydroxide and phosphate ions:
  - $Al(OH)_3(aq) + PO_4^{3-}(aq) \rightarrow AlPO_4(s) + 3OH^-(aq)$

12B Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 12–14.

Johnson is a chemical engineer looking to find a method to industrialise the large-scale crystallisation of a solute in order to use them for the formation of pharmaceutical solids. To achieve this, Johnson has to know how to use solubility curves to determine the point of crystallisation.

Question 12 (1 MARK) 🌶

What is a supersaturated solution?

A. An unstable solution with less solute dissolved than a saturated solution.

B. A stable solution with more solute dissolved than a saturated solution.

C. An unstable solution with more solute dissolved than a saturated solution.

D. A stable solution with more solute dissolved than an unsaturated solution.

Question 13 (1 MARK) 🌶

Which of the following options correctly describes a supersaturated solution?

A. A solution is supersaturated when the mass of dissolved solute lies above the solubility curve.

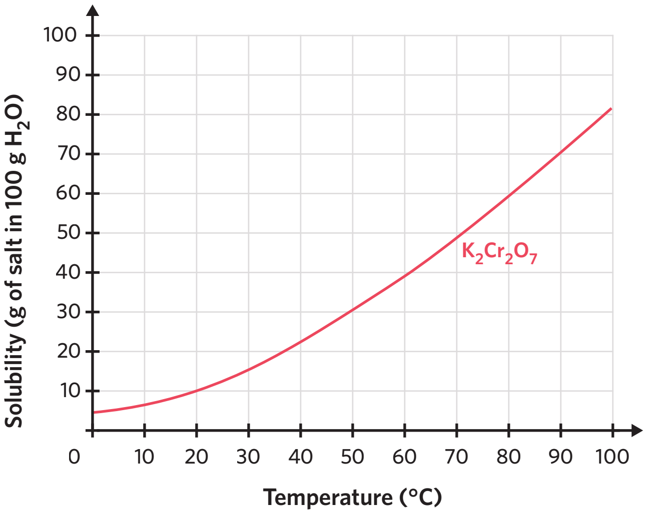
B. A solution is supersaturated when a mass of dissolved solute is achieved at a temperature higher than that shown on the solubility curve.

C. A solution is supersaturated when the mass of dissolved solute is higher than 100 g per 100 g of H_2O.

D. Any mass of dissolved solute above the solubility of a substance at 0 °C is considered supersaturated.

Question 14 (3 MARKS) 🌶🌶

Johnson is attempting to make a pharmaceutical solid from potassium dichromate crystals, $K_2Cr_2O_7$. Using the given solubility curve, explain if and why crystals will be formed if Johnson dissolves 180 g of $K_2Cr_2O_7$ in 200 g of H_2O at 90 °C. If crystals are formed, calculate the mass of the crystals formed.



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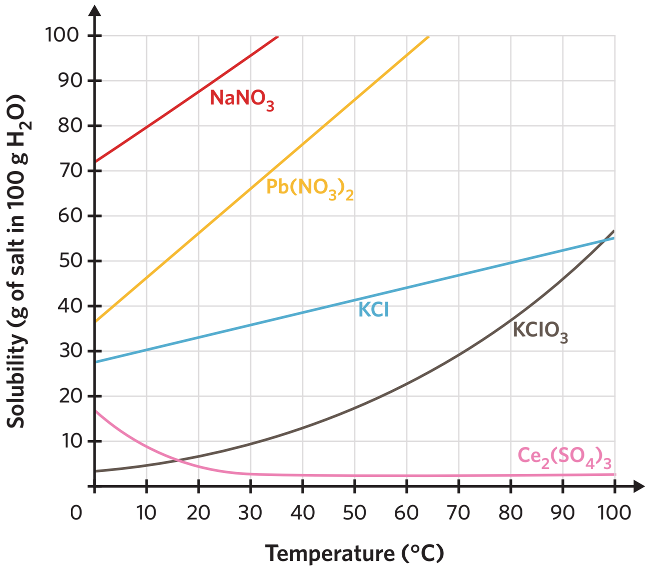
Exam-style

Question 15 (7 MARKS)

A class is using a solubility curve to determine the relationship between temperature and the solubility of ionic compounds.



a. Based on the solubility curve provided, what is the most common trend between temperature and solubility of ionic compounds? (1 MARK)

b. Using the solubility curve provided, identify any compounds that do not follow the trend from your answer to part a. (1 MARK)

c. Using the solubility curve provided, determine whether the following solutions are unsaturated, saturated or supersaturated.

- **i.** 40 g of KCl in 100 g of H_2O at 70 °C. (1 MARK)
- **ii.** 10 g of $KClO_3$ in 100 g of H_2O at 30 °C. (1 MARK)
- **iii.** 70 g of $Pb(NO_3)_2$ in 100 g of H_2O at 34 °C. (1 MARK)
- **iv.** 80 g of $NaNO_3$ in 95 g of H_2O at 10 °C. (1 MARK)
- **v.** 25 g of $Ce_2(SO_4)_3$ in 450 g of H_2O at 20 °C. (1 MARK)

Question 16 (5 MARKS)

Solubility curves are useful for visualising the relationship between temperature and solubility in the dissolution process. Scientists have discovered a new element that forms the ionic compound XCl_3 with chloride ions, and wish to construct a solubility curve for the compound.

a. Use the data provided to plot the saturation points of a solution of XCl_3 at different temperatures on a graph.

- 10 g of XCl_3 per 100 g of H_2O at 20 °C
- 25 g of XCl_3 per 100 g of H_2O at 40 °C
- 40 g of XCl_3 per 100 g of H_2O at 50 °C
- 100 g of XCl_3 per 100 g of H_2O at 60 °C

(2 MARKS)

b. Draw the solubility curve for XCl_3. (1 MARK)

c. Comment on the shape of the solubility curve. What implications does this have regarding the relationship between temperature and the solubility of XCl_3? (2 MARKS)

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Question 17 (4 MARKS)

Large bodies of water often become oversaturated with nutrients, leading to excessive growth of algae. This can result in a significant decrease in the purity of the water, making it unsafe for human consumption.

a. What is the name of the process outlined above? (1 MARK)

b. Phosphorus (as phosphate) is one such nutrient that can cause excessive algal growth. Describe the steps involved in using alum to prevent phosphate from contributing to algal growth. (3 MARKS)

Key science skills

Question 18 (5 MARKS)

Luke is a year 11 chemistry student who is attempting to construct a solubility curve. In constructing the curve, Luke uses a series of different soluble ionic compounds, and compares their solubilities at different temperatures to produce a curve. He does this by measuring the mass of a solute, dissolving it and heating it in water, and subsequently cooling it and noting the temperature when crystallisation takes place. However, when plotting the curve, Luke notices that some of the values differ greatly from the continuous exponential curve produced by the other values.

a. What type of error is present? (1 MARK)

b. Identify two potential sources of this error. (2 MARKS)

c. After repeating the experiment, Luke noticed that all of his results were centred around the expected value, however the individual results were all different to each other. Comment on the accuracy and precision of Luke's experimental results. (2 MARKS)

FROM LESSONS 16C & 16D

Questions from multiple lessons

Question 19 (12 MARKS)

Dissolution is the process in which a substance gets dissolved in another substance. When referring to the dissolution process, a number of important terms are used.

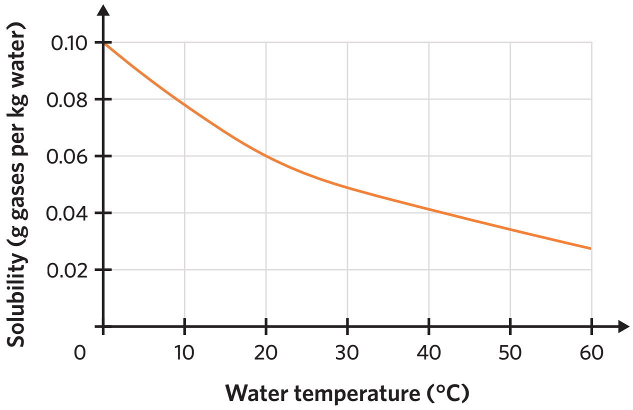
a. Define the following terms.

i. Solute (1 MARK)

ii. Solvent (1 MARK)

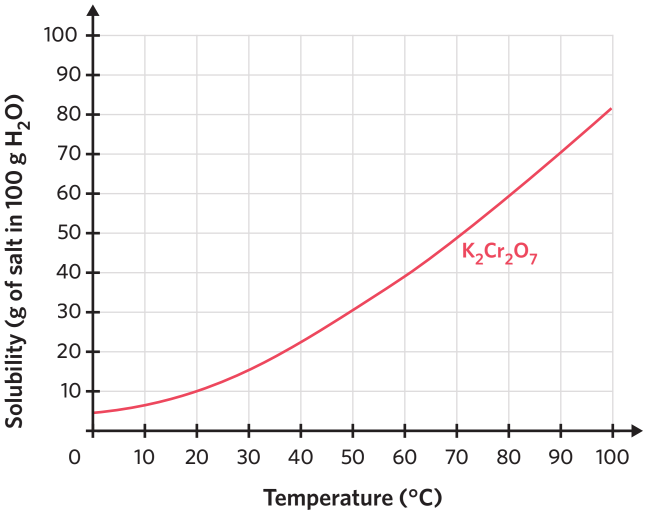
b. Water is sometimes referred to as the universal solvent, as it is able to dissolve such a large variety of substances. Due to the structure of water molecules, water is far better at dissolving polar molecules than non-polar substances. Will ethanol or ethane have a higher solubility in water? Justify your answer with reference to the structures of each molecule. (3 MARKS)

c. Ethane is a hydrocarbon with a boiling point of −89 °C. The solubility curve of ethane in a solution of water is shown.



Does the solubility curve for ethane follow the expected relationship between temperature and solubility? Explain your answer. (3 MARKS)

d. $K_2Cr_2O_7$ is a substance whose solubility increases as the temperature increases, as shown in the graph.



i. Using the graph, determine the mass of $K_2Cr_2O_7$ that will be dissolved in 500 g of H_2O at 50 °C. (1 MARK)

ii. Calculate the amount, in mol, of $K_2Cr_2O_7$ that will be dissolved in 500 g of H_2O at 50 °C. (1 MARK)

iii. If the density of water is 1.00 g mL^{-1}, determine the concentration, in mol L^{-1}, of a saturated solution of $K_2Cr_2O_7$ in 500 g of H_2O at 50 °C. (2 MARKS)

FROM LESSONS 5A, 6B, 7A & 12A

Hints

15a. Graphs show the relationship between the independent and dependent variables.

15b. Not all ionic compounds follow the same trend regarding temperature and solubility

15ci. Examine the curve for KCl at 70 °C.

15cii. Examine the curve for $KClO_3$ at 30 °C.

15ciii. Examine the curve for $Pb(NO_3)_2$ at 34 °C.

15civ. Examine the curve for $NaNO_3$ at 10 °C.

15v. Examine the curve for $Ce_2(SO_4)_3$ at 20 °C.

16a. The mass of solute dissolved is given on the y-axis. Temperature is given on the x-axis.

16b. The solubility curve is a continuous curve connecting the plotted points.

16c. The effect of temperature on the solubility of the solute changes.

17a. There is a specific term given to the growth of algae due to oversaturation of nutrients.

17b. Alum ultimately forms a precipitate.

18a. If some values differ greatly from the expected value, then only they are being impacted by the error.

18b. Random changes to any controlled variables in the experiment can cause this type of error.

18c. Accuracy and precision describe two different properties of experimental results.

19ai. A solution is a mixture of one substance that has dissolved in another.

19aii. A solution is a mixture of one substance that has dissolved in another.

19b. Ethanol molecules possess a hydroxyl functional group.

19ci. Solubility depends on the strength of intermolecular forces.

19cii. Substances in different states tend to have different trends in solubility as temperature changes.

19di. Examine the curve for $K_2Cr_2O_7$ at 50 °C.

19dii. Convert your answer from part i to moles.

19diii. Convert the amount of water to litres.

Chapter 12 review

Multiple choice (10 MARKS)

Question 1 (1 MARK)

Which of the following statements regarding solubility is false?

A. Stirring a solution increases the solubility of the solute.

B. Cooling a mixture decreases the solubility of the solute.

C. Changing the solvent may affect the solubility of the solute.

D. A supersaturated solution is unstable.

Question 2 (1 MARK)

400 g of ammonium chloride, NH_4Cl, is dissolved in 2.00 L of water. Which of the following accurately gives the concentration of the NH_4Cl solution in ppm?

A. 0.200 ppm **B.** 2 ppm **C.** 200 ppm **D.** 2.00×10^5 ppm

---

Question 3 (1 MARK)

Which of the following is not a source of phosphate in large bodies of water?

A. Animal waste

B. Wastewater

C. Fertiliser runoff

D. Acid rain

Question 4 (1 MARK)

A student diluted a sample of DNA by 1/30 (i.e. the volume of diluted DNA solution is 30 times the volume of the original solution). They then placed it in a device called a spectrophotometer, which calculated a concentration of 0.00130 %(m/v). What is the concentration of DNA in the original solution in $\mu g\ mL^{-1}$?

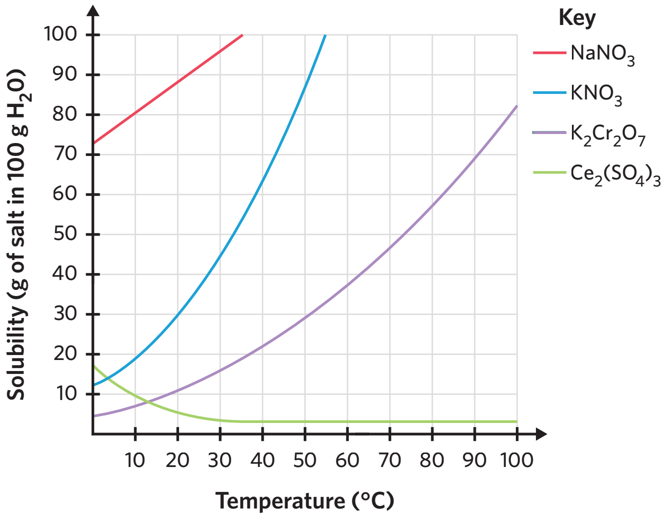
A. $0.0390\ \mu g\ mL^{-1}$

B. $0.390\ \mu g\ mL^{-1}$

C. $3.90\ \mu g\ mL^{-1}$

D. $390\ \mu g\ mL^{-1}$

Use the following solubility curve to answer questions 5–6.



Question 5 (1 MARK)

A compound has a solubility of 30 g per 100 g of H_2O at 20 °C and a solubility of 65 g per 100 g of H_2O at 40 °C. Which of the following correctly identifies this compound?

A. $NaNO_3$
B. KNO_3
C. $K_2Cr_2O_7$
D. $Ce_2(SO_4)_3$

Question 6 (1 MARK)

Which solution is supersaturated?

A. 45 g of KNO_3 in 150 g of H_2O at 30 °C
B. 120 g of $K_2Cr_2O_7$ in 200 g of H_2O at 90 °C
C. 9 g of $Ce_2(SO_4)_3$ in 100 g of H_2O at 10 °C
D. 50 g of $NaNO_3$ in 50 g of H_2O at 20 °C

Question 7 (1 MARK)

Which of the following concentrations is equivalent to 0.25 g L^{-1}?

A. 2.5×10^{-4} mg L^{-1}
B. 0.25 kg L^{-1}
C. 2.5 g kL^{-1}
D. 250 mg L^{-1}

Question 8 (1 MARK)

A student has 0.25 g of $Ca(OH)_2(s)$ and would like to generate a calcium hydroxide solution of 0.25 M. What volume of distilled water is closest to the required amount to prepare this solution?

A. 25 mL **B.** 100 mL **C.** 13 mL **D.** 22 mL

Question 9 (1 MARK)

A research scientist was investigating the solubility of a substance at different temperatures. She knows that the following steps are involved:

I. Record the temperature at which crystals start to form.
II. Weigh out several samples, ensuring that each sample contains the same mass of solute and has a constant volume of water.
III. Cool the test tubes slowly, stirring constantly.
IV. Heat the test tubes until all solute is dissolved.

Which of the following correctly lists the order of these steps?

A. III, I, II, IV
B. II, III, I, IV
C. II, IV, III, I
D. II, III, IV, I

Question 10 (1 MARK)

Consider the neutralisation reaction that occurs between sodium hydroxide and sulfuric acid.

$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

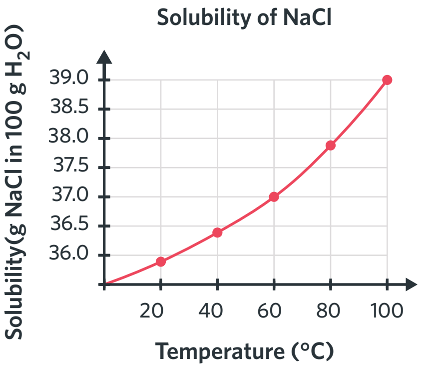
What volume of 0.50 mol L^{-1} sodium hydroxide would be required to completely neutralise 50 mL of 0.25 mol L^{-1} sulfuric acid?

A. 25 mL **B.** 50 mL **C.** 75 mL **D.** 100 mL

Short answer (30 MARKS)

Question 11 (7 MARKS) 🌶🌶

Sodium chloride is a very common ionic compound that is found in table salt. The solubility curve of sodium chloride is shown.



a. What is a supersaturated solution? (1 MARK)

b. Identify the mass of NaCl required in 100 mL of H_2O to create a supersaturated solution at 80 °C. (1 MARK)

c. Calculate the concentration, in mol L^{-1}, of NaCl required to create a saturated solution at 60 °C. (3 MARKS)

d. Based on your answer to part c, how many grams of NaCl are required to create a saturated solution with a volume of 193 mL at 60 °C? (2 MARKS)

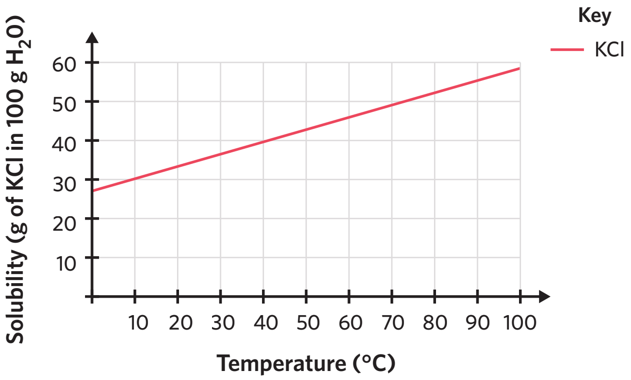
Question 12 (12 MARKS) 🌶🌶

A particular 1.0 L bottle of water contains 2.0 ppm of chloride ions, which come from dissolved potassium chloride.

a. Calculate the mass, in grams, of potassium chloride in this bottle of water. (4 MARKS)

b. The water in the bottle is sourced from a lake with a total volume of 900 ML. Assuming the concentration of potassium chloride in the bottle is equivalent to the concentration of potassium chloride in the lake, what is the total mass, in kg, of potassium chloride in the lake? (3 MARKS)

c. Consider the solubility curve of potassium chloride shown.



Based on your understanding of the trends in solubility of compounds in different states, would KCl be expected to be a solid or a gas for the temperature range shown on the curve above? (1 MARK)

d. At what temperature would 45 g of potassium chloride dissolve in 100 g of H_2O to make a saturated solution? (1 MARK)

e. Calculate the final concentration, in mol L^{-1}, if 750 mL of H_2O was added to 1.5 L of 2.0 mol L^{-1} potassium chloride solution. (3 MARKS)

Question 13 (11 MARKS) 🌶🌶🌶

The table provided shows the solubility of sodium carbonate, Na_2CO_3, in water at different temperatures.

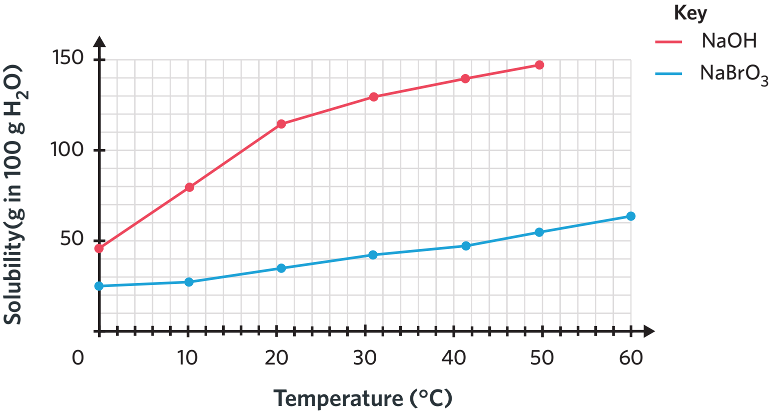
Temperature (°C)	0	10	20	30	40	50	60
Solubility (g per 100 g of water)	12.0	21.5	29.0	39.0	49.0	47.0	44.0

a. Describe the trend in the solubility of sodium carbonate in water. (1 MARK)

b. Construct a solubility curve for sodium carbonate using the data points provided. (3 MARKS)

c. Use the solubility curve from your answer to part b to calculate the mass of crystals that would form if a saturated solution of sodium carbonate in 300 g of water were cooled from 40 °C to 20 °C. (3 MARKS)

d. The graph provided shows the solubility curves of sodium bromate and sodium hydroxide.



List the two sodium compounds from most soluble to least soluble at 50 °C. (1 MARK)

e. A student places 200 g of sodium bromate in 200 g of water at 60 °C and stirs until no more dissolves. Calculate the mass of undissolved sodium bromate remaining. (3 MARKS)

Key science skills (10 MARKS)

Question 14 (10 MARKS)

Po is investigating the effect of temperature changes on the solubility of helium gas in aqueous solution. He would like to create a solubility curve in order to interpret the results of his experiment. Therefore, he measures the solubility of helium gas at different temperatures.

a. What are the independent and dependent variables in this experiment? (2 MARKS)

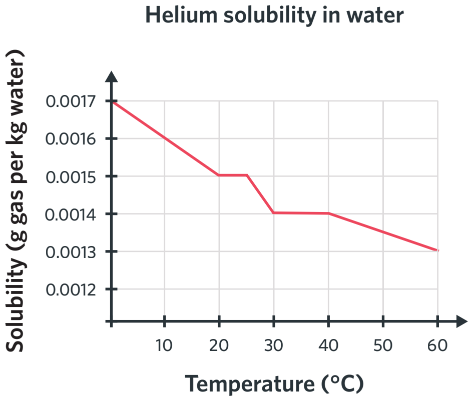
b. What variables should be controlled throughout the experiment? (2 MARKS)

An excerpt from the SDS (Safety Data Sheet) of helium gas is shown.

Signal Words:
Warnings
Hazard statement(s):
H280: Contains gas under pressure;
may explode if heated.

c. Identify two precautions that should be taken in the laboratory to minimise the danger of working with helium gas? (2 MARKS)

d. Po uses his results to construct the solubility curve for helium shown.



According to the curve, at what temperatures would a solution of 0.0016 g of helium dissolved in 1 kg of H_2O be supersaturated? Justify your answer by explaining what a supersaturated solution is. (2 MARKS)

e. After his experiment, his results show that the solubility of helium gas in water decreases as the temperature increases. Po notes that this is the opposite to solid solutes. Suggest a plausible explanation for why the solubility of gases decreases with increasing temperature. (2 MARKS)

FROM LESSONS 16A & 16B

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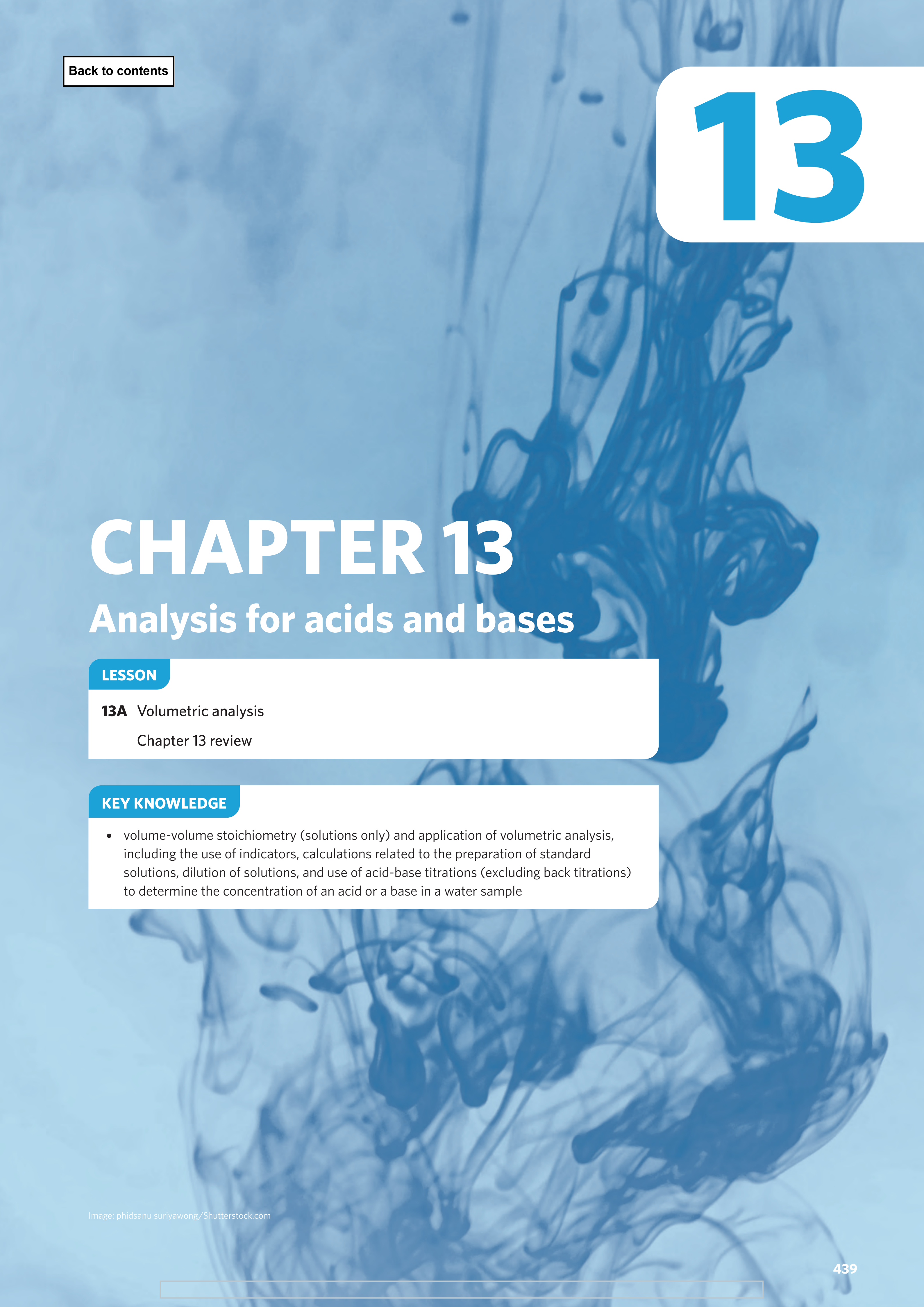

13

CHAPTER 13

Analysis for acids and bases

LESSON

13A Volumetric analysis

Chapter 13 review

KEY KNOWLEDGE

- volume-volume stoichiometry (solutions only) and application of volumetric analysis, including the use of indicators, calculations related to the preparation of standard solutions, dilution of solutions, and use of acid-base titrations (excluding back titrations) to determine the concentration of an acid or a base in a water sample

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13A Volumetric analysis

STUDY DESIGN DOT POINT

- volume-volume stoichiometry (solutions only) and application of volumetric analysis including the use of indicators, calculations related to preparation of standard solutions, dilution of solutions and use of acid-base titrations to determine the concentration of an acid or a base in a water sample

13A

2.2.4.1 Volume-volume stoichiometry

2.2.4.2 Principles of titrations

2.2.4.3 Acid-base titrations

ESSENTIAL PRIOR KNOWLEDGE

6B	The mole
10D	Indicators
12A	Dilutions

See questions 84–86.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: Rabbitmindphoto/Shutterstock.com

How are titrations used to check the quality of food?

The regulation of food quality is a billion-dollar industry. **Titration**, a form of **volumetric analysis**, is often used in this area to check the amount of various ingredients present in a food sample. In this lesson, we will learn about the principles of titration, and how they can be applied.

KEY TERMS AND DEFINITIONS

Aliquot portion of a sample solution used for analysis during titration

Concentration amount of chemical substance per unit of volume

Concordant titres three or more titres that fall within approximately 0.1 mL of each other

End point point at which a colour change occurs during titration

Equivalence point point at which the reaction has occurred as per the stoichiometric ratio of the balanced equation

Indicator a substance that changes colour to indicate a pH range

Meniscus curved upper surface of liquid

Primary standard substance used to make a primary standard solution

Primary standard solution solution of precisely known concentration

Titration quantitative technique used to find the concentration and/or amount of substance in a solution

Titration curve graphic representation of an acid-base titration, with titre volume on the *x*-axis and pH on the *y*-axis.

Titre volume of the solution delivered from the burette to reach the end point of a titration

Volumetric analysis quantitative analytical technique used to determine unknown concentrations of solutions

Volume-volume stoichiometry 2.2.4.1

Volume-volume stoichiometry relates the **concentrations** and volumes of solutions to the amount of chemical present.

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e use the equation $n = cV$ to relate the number of moles (n), concentration (c) and volume (V) of solutions?

As we learned in lesson 6B, the mole is a fundamental concept in chemistry. In conjunction with Avogadro's constant, it defines a certain number of particles. Whenever quantitative chemistry is performed, it is important to work with the mole concept as chemicals react in fixed, whole number mole ratios determined by chemical equations. For example, consider the following acid-base reaction:

$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$

1 mole of hydrochloric acid, $HCl(aq)$, reacts with 1 mole of sodium hydroxide, $NaOH(aq)$, to produce 1 mole of water, $H_2O(l)$, and 1 mole of sodium chloride, $NaCl(aq)$.

However, $HCl(aq)$ and $NaOH(aq)$ are not present as pure substances but rather as aqueous solutions. This is because they both dissolve readily in water, as shown in figure 1.

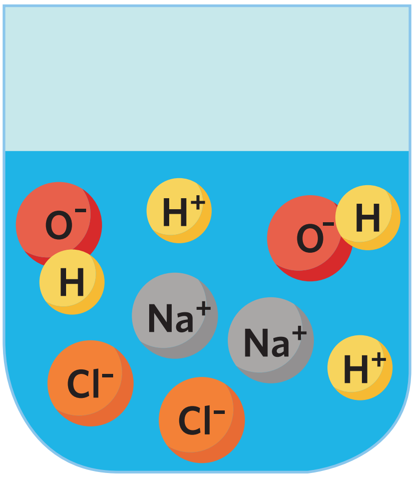

Figure 1 Solute particles dissolved in water

MISCONCEPTION

'When 1 mole of NaCl is fully dissolved in water, there will be 0.5 moles of sodium ions and 0.5 moles of chloride ions present in solution.'

Recall that a mole is a measure of quantity. There is 1 sodium ion and 1 chloride ion in each NaCl, so 1 mole of each will be present in solution upon the dissolving of NaCl. Similarly, if 1 mole of $Mg(OH)_2$ is dissolved in water, 1 mole of magnesium ions and 2 moles of hydroxide ions will be present in the solution.

The number of moles of solute present in solution depends on two factors: concentration and volume. As we learned previously, concentration (c) is the amount of moles of a chemical substance in a defined volume (V). Both of these variables are related to the number of moles of particles present by the equation:

$n = c \times V$

where n is the number of moles of a substance present, c is the concentration of the solution in mol L^{-1} (M) and V is the volume of solution present in litres (L). This equation is found in the VCE Data Book.

This equation can also be usefully represented as 'the magic triangle', shown and explained in figure 2.

MISCONCEPTION

'M and *M* mean the same thing.'

Capital M is the molarity of a solution which can also be shown as mol L^{-1}. Italicised capital *M* represents the molar mass and has the units g mol^{-1}.

STRATEGY

We can use the following concept triangle to help us figure out the type of calculation to perform to find the concentration, amount of solute or even volume of solutions.

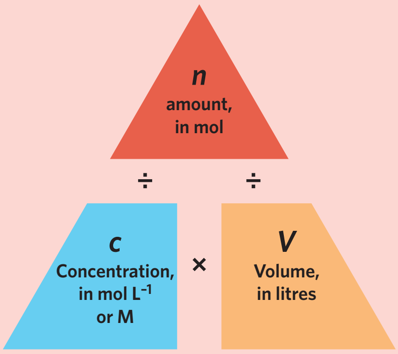

Figure 2 'The magic triangle'

To use this, use your finger to cover up the value you are trying to find. Calculate using the remaining variables as indicated by the signs. For example, to find the amount, cover the section of the triangle that says 'amount', which leaves us to perform the calculation $c \times V$.

USEFUL TIP

Concentrations and volumes can be expressed using different units. Therefore, it is important to ensure that the values are in the same appropriate units (mol L^{-1}, or its shorthand 'M', for concentration, L for volume) before the equation can be used to determine the number of moles of a substance present

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PLE 1

What is the number of moles of $Cl^-(aq)$ present in 200 mL of 0.25 mol L^{-1} $CaCl_2(aq)$ solution?

What information is presented in the question?

Volume of solution, V: 200 mL.

Concentration of solution, c: 0.25 mol L^{-1}.

What is the question asking us to do?

Find the number of moles of $Cl^-(aq)$ dissolved in the solution.

What strategies do we need in order to answer the question?

1. Substitute the volume and concentration of $CaCl_2(aq)$, using the appropriate units, into the equation $n = c \times V$.
2. Use the dissolution equation $CaCl_2(s) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$ to determine the number of moles of $Cl^-(aq)$ dissolved in the solution.

Answer

$n = c \times V$

$c = 0.25 \text{ mol L}^{-1}$

$V = \left(\frac{200}{1000}\right) \text{ L}$

$V = 0.200 \text{ L}$

$n(CaCl_2) = 0.25 \text{ mol L}^{-1} \times 0.200 \text{ L}$

$n(CaCl_2) = 0.050 \text{ mol}$

$n(CaCl_2) : n(Cl^-)$

$1 \quad : \quad 2$

$n(Cl^-) = \frac{2}{1} \times n(CaCl_2)$

$n(Cl^-) = 2 \times 0.050 \text{ mol}$

$n(Cl^-) = 0.10 \text{ mol}$

WORKED EXAMPLE 2

What is the volume in mL of a 0.750 mol L^{-1} solution containing 0.500 moles of NaCl?

What information is presented in the question?

Concentration of solution, c: 0.750 mol L^{-1}.

Number of moles in solution, n: 0.500 moles

What is the question asking us to do?

Find the volume of the solution.

What strategies do we need in order to answer the question?

1. Rearrange the equation $n = c \times V$, to make V the subject.
2. Substitute the concentration and number of moles of NaCl(aq), using the appropriate units, into the rearranged equation.
3. Convert volume into mL.

Answer

$n = c \times V$

Rearranged: $V = \frac{n}{c}$

$c = 0.750 \text{ mol L}^{-1}$

$n = 0.500 \text{ mol}$

$V = \frac{0.500 \text{ mol}}{0.750 \text{ mol L}^{-1}}$

$V = 0.670 \text{ L}$

Conversion: $0.670 \text{ L} \times 1000 = 670 \text{ mL}$

$V = 670 \text{ mL}$

USEFUL TIP

If volume is given in mL, the equation $n = c \times \frac{V}{1000}$ can be used, where V is in mL.

Progress questions

Question 1

What is the unit of c in the equation $n = c \times V$?

A. g mL^{-1}

B. mol L^{-1}

Question 2

Which of the following can be used to calculate V?

A. $V = \frac{n}{c}$

B. $V = \frac{c}{n}$

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Principles of titrations 2.2.4.2

Titration reactions are used to determine the unknown concentration of a solution and the amount of a substance in a solution.

Why do we use titrations?

Sometimes chemists deal with solutions of unknown concentrations. This may be because they are testing a random sample of water from a stream, or they want to see if they have prepared a solution correctly. The basic principle of titration is that by reacting a solution of known concentration with a solution of unknown concentration, chemists can use their understanding of volume-volume stoichiometry to calculate the concentration of the unknown solution.

As part of this process, a solution of accurately known concentration is used to measure the amount or concentration of a substance in a sample solution of unknown concentration. A basic titration setup is shown in figure 3.

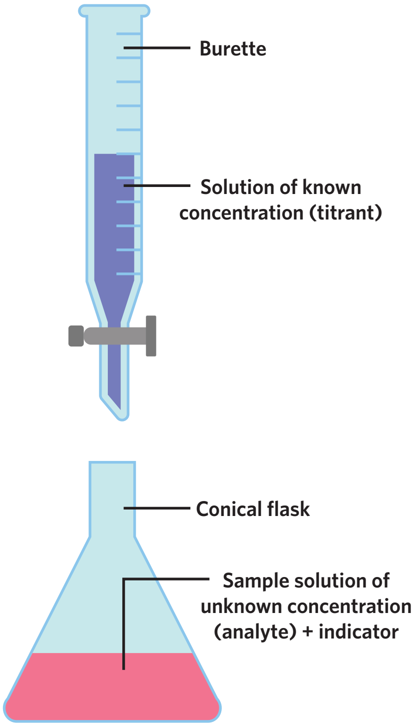

Figure 3 The basic setup for a titration reaction

How do we prepare and calculate the concentration of a standard solution?

The solution of accurately known concentration – usually found in the burette – in titration is known as a standard solution. This solution is often standardised (its exact concentration has been determined) by a solution known as the **primary standard solution**.

To prepare the primary standard solution, a **primary standard** is accurately weighed and dissolved in deionised water in a volumetric flask according to the steps outlined in figure 4. The amount of primary standard used depends on the concentration of the primary standard solution we want to prepare.

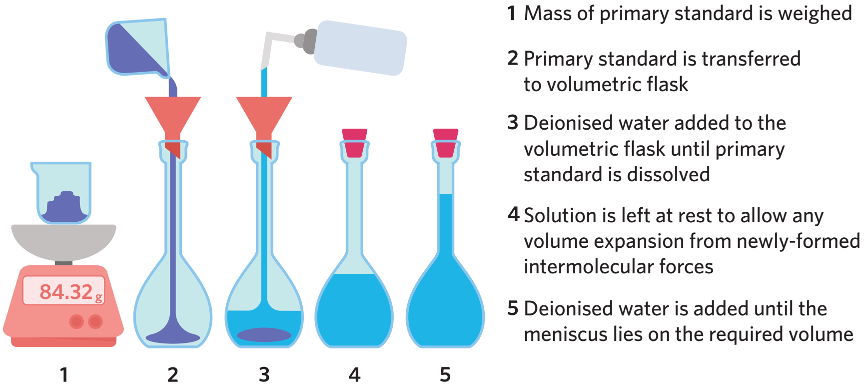

Figure 4 Preparation of a primary standard solution

Given the primary standard solution is used to determine the concentration of other secondary standard solutions through titration, there are some criteria that need to be met before substances can be considered a suitable primary standard. These include:

- The substance must be pure and found in a stable form.
- The substance must not react with the atmosphere e.g. moisture, carbon dioxide, etc.
- The substance must be soluble in solution.
- The substance must be readily available and inexpensive.
- The substance must be able to be weighed accurately (i.e. have a high, known molar mass to reduce error).

An example of a primary standard would be the ionic compound anhydrous sodium carbonate Na_2CO_3, which fits all of the above criteria. Common reagents such as NaOH and HCl cannot be used as primary standards as they both will react with the atmosphere. Sodium hydroxide reacts with $CO_2(g)$ from the atmosphere to produce sodium carbonate according to the following equation:

$2NaOH(s) + CO_2(g) \rightarrow Na_2CO_3(s) + H_2O(l)$

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This leads to a significant mass increase that will affect any calculations made using the initial concentration of sodium hydroxide. Any concentration of a HCl solution will also change gradually over time, as it has a tendency to vaporise (spontaneously convert to gaseous state) in small amounts when exposed to standard laboratory conditions, again affecting any calculations made using its initial concentration.

WORKED EXAMPLE 3

Tom is conducting a titration in order to determine the concentration of ethanoic acid, CH_3COOH, in a sample of water. Outline the steps required to make a 0.0500 mol L^{-1} Na_2CO_3 primary standard solution in a 100 mL volumetric flask.

What information is presented in the question?

Volume: 100 mL.

Concentration of $Na_2CO_3 = 0.0500$ mol L^{-1}.

What is the question asking us to do?

Write out the steps needed to prepare a primary standard solution and the quantities involved.

What strategies do we need in order to answer the question?

1. Calculate the amount, in mol, of $Na_2CO_3(s)$ required to be present in the 100 mL volumetric flask with the equation $n = c \times V$.
2. Determine the mass of $Na_2CO_3(s)$ required to be weighed on the analytical balance with the equation $m = n \times M$.
3. Write out all the steps required to create this standard solution.

Answer

$$n = c \times \frac{V}{1000}$$

$$n(Na_2CO_3) = 0.0500 \text{ mol L}^{-1} \times \left(\frac{100}{1000}\right) \text{ L}$$

$$n(Na_2CO_3) = 0.00500 \text{ mol}$$

$$m = n \times M$$

$$M(Na_2CO_3) = (2 \times 23.0 \text{ g mol}^{-1}) + 12.0 \text{ g mol}^{-1}$$

$$+ (3 \times 16.0 \text{ g mol}^{-1})$$

$$M(Na_2CO_3) = 106.0 \text{ g mol}^{-1}$$

$$m(Na_2CO_3) = 0.00500 \text{ mol} \times 106.0 \text{ g mol}^{-1}$$

$$m(Na_2CO_3) = 0.530 \text{ g}$$

Therefore, the steps required to produce this primary standard solution are:

1. Weigh 0.530 g of anhydrous sodium carbonate using an analytical balance.
2. Transfer the sodium carbonate to a 100 mL volumetric flask.
3. Add deionised water until sodium carbonate is fully dissolved.
4. Swirl volumetric flask with stopper on and let the solution rest.
5. Add more deionised water until the bottom of the **meniscus** is at the required mark.

Any factor that may compromise the concentration of the solution will affect the results obtained during the titration reaction. Therefore, we must be careful when choosing a primary standard. Furthermore, we need to be just as careful during the preparation of the standard solution to minimise the chances of errors occurring. For example, it is important to avoid:

- Incorrect weighing of substances
- Contamination of instruments or substances used
- Rinsing of apparatus with the inappropriate substance
- Incomplete dissolving of the primary standard

All of these errors can impact the accuracy of the standard solution produced, which will then affect the results obtained.

MISCONCEPTION

'The solution of unknown concentration is always in the conical flask, and the solution of known concentration is always in the burette.'

Regardless of whether the unknown solution is in the conical flask or the burette, the principles of titration still apply.

How do we carry out a titration?

To prepare a titration experiment, a pipette is used to transfer an accurate volume, called an **aliquot**, of an unknown sample into the conical flask, along with the relevant **indicator**. The conical flask is placed under the burette, and the burette is usually then filled with a solution of known concentration. This solution is dispensed incrementally from the burette into the conical flask, resulting in a chemical reaction in the conical flask. The titration process is summarised in figure 5.

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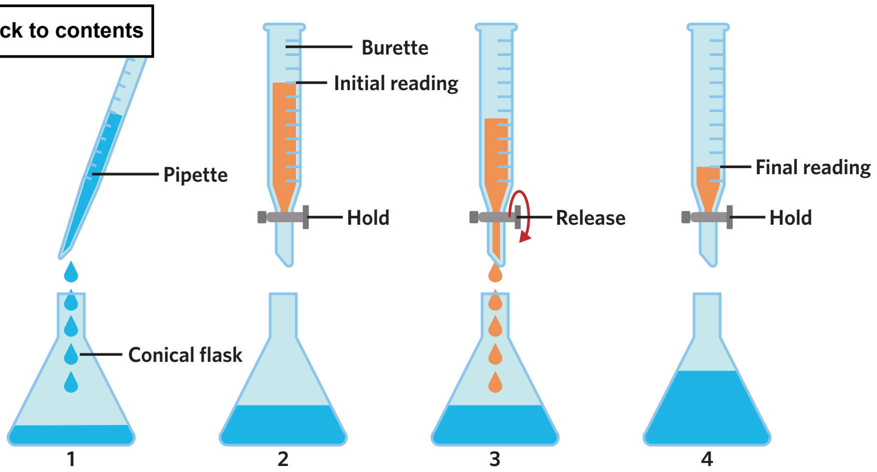
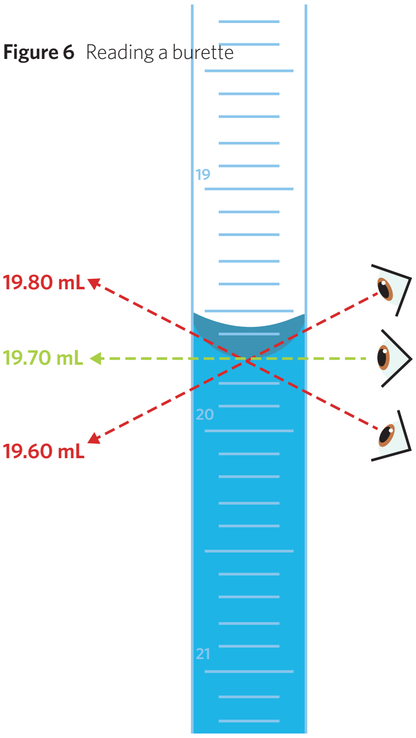


Figure 5 A summary of the titration process

Depending on the experiment, the reaction will stop at a certain point, at which the appropriate experimental data is recorded and is used to perform calculations to determine the concentration of an unknown substance.

It is important to note how to read these values from a burette correctly. Due to the way liquids interact with the walls of the burette, we will usually see a dip in the solution in the burette. Therefore, when we read the scale on a burette, we need to position ourselves so that our line of sight is horizontal to the bottom dip of the meniscus. If we don't align ourselves appropriately, the volume read will be higher or lower than the true value as illustrated in figure 6 (depending on whether we are lower or higher than the meniscus), making our results inaccurate.

Figure 6 Reading a burette



The experiment is repeated multiple times until **concordant titres** (usually all **titres** within ± 0.1 mL) are obtained to increase the reliability and precision of the results by minimising the effects of random errors. Table 1 shows an example of the type of data that is recorded during a titration experiment.

Table 1 Sample values from a titration

Titration number	1	2	3	4
Initial burette reading	3.50	7.30	13.35	0.00
Final burette reading	22.90	26.30	32.40	19.10
Titre volume (mL)	19.40	19.00	19.05	19.10

In this case, titrations 2, 3 and 4 are the concordant titres as they are all titres that differ with a ± 0.1 mL range from the highest value (19.10 mL) to the lowest value (19.00 mL). The average titre, which is used in calculations, is the average of the concordant titres. Therefore, in this example:

$$\text{Average titre: } \frac{19.00\text{ mL} + 19.05\text{ mL} + 19.10\text{ mL}}{3} = 19.05\text{ mL}$$

Note that sometimes the initial reading from the burette is lower than the final reading. This is due to the fact that some burettes measure the total volume that has been dispensed and therefore '0.00' is located at the very top of the burette. As such, the more volume that is used during the reaction, the less solution that will be left in the burette, and therefore the larger the reading of the final value, resulting in a larger titre volume.

MISCONCEPTION

'A titre of 19.02 mL can be read from a standard burette.'

The resolution of a standard burette is ± 0.05 mL, so the reading would be 19.00 mL. Note that multiple readings from a burette could have an average of 19.02 mL, but not a single reading.

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What are some common experimental errors during titrations?

Due to the practical nature of titration, experimental errors can be introduced in a variety of ways. Some common errors and their effects are described in table 2.

Table 2 Common errors during titrations and their effects

Situation	Effect on accuracy of calculated concentration, if unknown solution is in:		Explanation
	Conical flask	**Burette**	
Burette rinsed with water	Overestimated	Underestimated	Solution in burette becomes diluted
Pipette rinsed with water	Underestimated	Overestimated	Solution in conical flask becomes diluted
Conical flask rinsed with solution to fill it	Overestimated	Underestimated	Solution in conical flask has higher concentration of solute
Missed **indicator end point**	Overestimated	Underestimated	Discrepancy between end point and **equivalence point**

Progress questions

Question 3

A primary standard must be

A. unreactive with the atmosphere.

B. insoluble.

Question 4

In order to make a standard solution, a primary standard is weighed on an analytical balance, which is a very accurate measurement device. Why is it important to use an analytical balance?

A. The mass of primary standard used is important in determining the concentration of the standard solution.

B. The mass of primary standard used has to always be the same in standard solutions.

Question 5

Common strong acids and bases like HCl and NaOH make excellent standard solutions.

A. True

B. False

Question 6

When reading the burette shown below, the value recorded is the ___________ of the meniscus.

A. top

B. bottom

Continues →

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Question 7

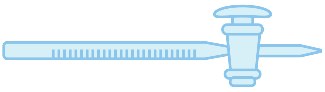
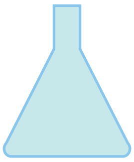
Which of the following options contains three concordant titres?

A. 17.65 mL, 17.50 mL, 18.05 mL, 18.00 mL, 17.65 mL

B. 17.65 mL, 17.80 mL, 18.05 mL, 17.65 mL, 17.60 mL

Question 8

Which option correctly matches the type of glassware used in titration with its function?

	Glassware	Function
A.		Dispenser of the indicator
B.		Container for the acid-base reaction

Acid-base titrations 2.2.4.3

Acid-base titrations are used to determine the concentration of acids and bases through a neutralisation reaction between the two substances.[1]

How do we calculate an unknown in acid-base titrations?

As we learned in lesson 10B, acid-base reactions between acids and metal hydroxides are neutralisation reactions in which hydroxide ions, OH^-, from a base react with hydrogen ions, H^+, from an acid to produce an ionic salt and water. In an acid-base reaction, the hydrogen ion is transferred from the acid to the base. Acid-base titrations are used to determine the concentration of an unknown base or acid.

Building on our knowledge of titration reactions, there are a few key features of acid-base titrations that we need to understand. As discussed in lesson 10D, there are a number of indicators to help signal a change in pH values during an acid-base reaction. We can use the colour change experienced by the indicator in a titration to help signal a particular point in the acid-base reaction.

When preparing an acid-base titration, we need to add an appropriate indicator to the solution in the conical flask prior to dispensing the burette solution into the flask. After the addition of a particular volume of the burette solution, the solution in the conical flask will change colour. The acid-base titration reaction is summarised in figure 7.

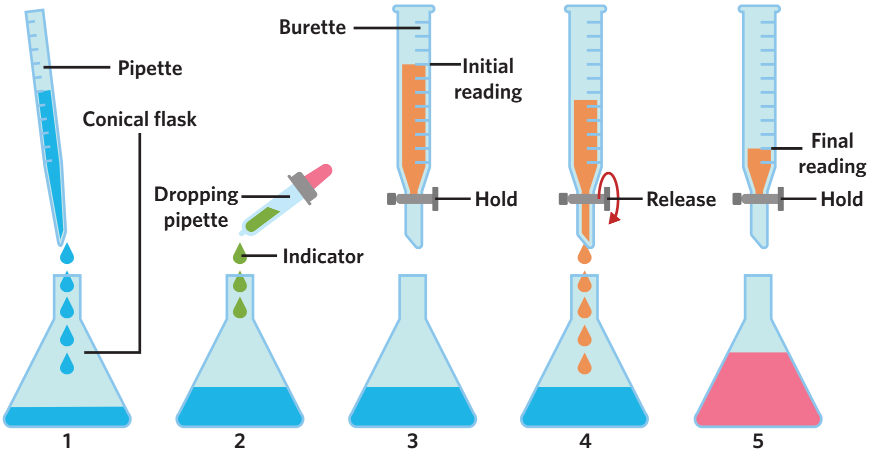


Figure 7 Steps of an acid-base titration

KEEN TO INVESTIGATE?

[1] **What other types of titrations are there?**
Search YouTube: Redox Titration between MnO_4^- and Fe^{2+}

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'The equivalence point and the end point are the same.'

The equivalence point represents the end of the chemical reaction. The end point, however, describes the colour change that follows the equivalence point.

The colour change that occurs represents the end point of the titration. This indicates that the reactants and products have reached a balanced stoichiometric ratio according to the balanced chemical equation for the equation; this is referred to as the equivalence point. Due to this, it is very important that we choose an indicator in which the colour change occurs within the range of the equivalence point for the reaction. It is important to note that the end point generally occurs just after the equivalence point. However, for calculation purposes, we can assume that the equivalence point is approximately the same as the end point.

The volume of solution emitted from the burette used to reach the end point – the titre – is used to calculate the amount or concentration of the unknown solution. This volume is calculated by taking the difference between the initial and final volumes in the burette. After multiple repetitions of the titration, the average value of concordant titres, (i.e. average titre), obtained during the reaction is used in calculations to determine the unknown value.

The values obtained from a titration reaction can be used to determine unknown concentrations of any of the substances involved in the reaction.

Consider the reaction between nitric acid, HNO_3, and potassium hydroxide, KOH:

$HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$

Suppose a student wanted to find the concentration of an unlabeled bottle of potassium hydroxide. Using titration, an average titre of 24.10 mL of a 0.50 mol L^{-1} solution of nitric acid was used to neutralise a 20.00 mL aliquot of potassium hydroxide.

Assuming the titration was carried out with an appropriate choice of indicator, the average titre of nitric acid represents the volume required to reach the end and equivalence points of the reaction. Therefore, we need to first determine the amount, in moles, of the substance of known concentration used in the reaction:

$n(HNO_3) = c \times V$

$n(HNO_3) = 0.50 \text{ mol L}^{-1} \times (\frac{24.10}{1000}) \text{ L}$

$n(HNO_3) = 0.01205 \text{ mol}$

According to the balanced equation for the reaction,

$n(HNO_3) : n(KOH)$

$\quad 1 \quad : \quad 1$

Therefore, for every 0.01205 mol of nitric acid used in the reaction, 0.01205 mol of potassium hydroxide would be required to balance the reaction equation. Since a 20.00 mL aliquot of KOH(aq) was used, we can now use that information to determine the unknown concentration of the potassium hydroxide solution:

$c(KOH) = \frac{n}{V}$

$c(KOH) = \frac{0.01205 \text{ mol}}{(\frac{20.00}{1000}) \text{ L}}$

$c(KOH) = 0.60 \text{ mol L}^{-1}$

An overall summary of the calculations used following a titration is shown in figure 8.

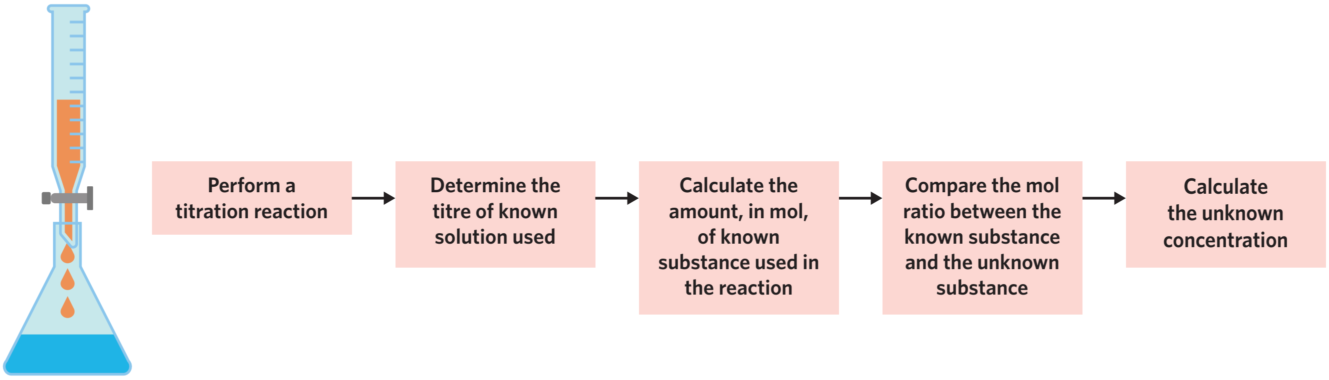

Figure 8 Calculating the concentration of an unknown solution using titration

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AMPLE 4

A chemistry student was asked to determine the concentration of a sample of the basic calcium hydroxide solution, $Ca(OH)_2$. After conducting multiple titrations with a 25.00 mL aliquot of calcium hydroxide, she found that it took an average of 25.60 mL of a 0.50 mol L^{-1} hydrochloric acid, HCl, solution for a colour change to occur. Calculate the concentration of the calcium hydroxide solution.

What information is presented in the question?

25.00 mL aliquot of calcium hydroxide.

25.60 mL of 0.50 mol L^{-1} hydrochloric acid used.

What is the question asking us to do?

Find the concentration of the calcium hydroxide solution.

What strategies do we need in order to answer the question?

1. Determine the amount of hydrochloric acid used in the reaction with the equation $n = c \times V$.
2. Use stoichiometry to determine the mole ratio of HCl to $Mg(OH)_2$.
3. Determine the amount of calcium hydroxide present.
4. Determine the concentration of the calcium hydroxide solution with the equation $c = \frac{n}{V}$.

Answer

The balanced equation for this reaction is

$$2HCl(aq) + Ca(OH)_2(aq) \rightarrow CaCl_2(aq) + 2H_2O(l)$$

$$n(HCl) = 0.50 \text{ mol L}^{-1} \times \left(\frac{25.60}{1000}\right) \text{ L}$$

$$n(HCl) = 0.0128 \text{ mol}$$

$$n(HCl) : n(Ca(OH)_2)$$

$$2 \quad : \quad 1$$

$$n(Ca(OH)_2) = 0.0128 \text{ mol} \times \frac{1}{2}$$

$$n(Ca(OH)_2) = 0.0064 \text{ mol}$$

$$c(Ca(OH)_2) = \frac{0.0064 \text{ mol}}{\left(\frac{25.00}{1000}\right) \text{ L}}$$

$$c(Ca(OH)_2) = 0.26 \text{ mol L}^{-1}$$

How can we choose the best indicator for an acid-base titration?

Acid-base reactions can be represented through a **titration curve** which shows the change in pH of the unknown solution as a titre is added. As we know, the change in pH throughout the reaction depends on the concentration of hydroxide and hydrogen (or hydronium H_3O^+) ions found in the unknown solution.

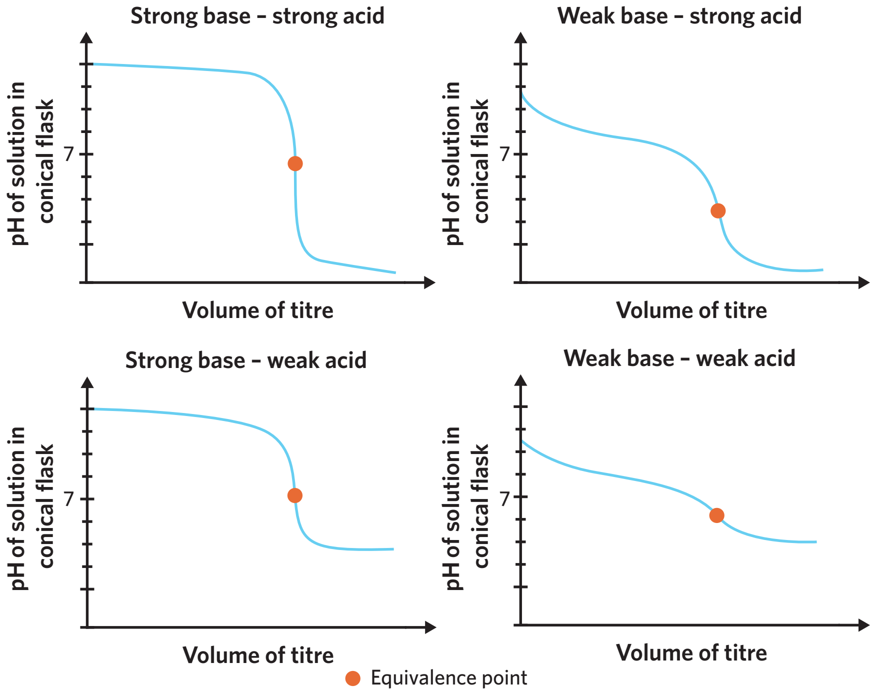

Figure 9 Example of a titration curve involving the titration of different strengths of acids and bases

As shown in figure 9, the equivalence point is the point at which the curve starts to inflect. Hence, this can sometimes be referred to as the point of inflection. The starting point and shape of the titration curves can help us to determine the nature of the unknown solutions. In figure 9, we can see that the unknown solutions were initially basic, as the curves started at a high pH and progressively became more acidic. From this, we can deduce that the unknown solutions were basic and the titres were acidic. If the unknown solutions were acidic and the titres basic, the curves would shift from a low pH to a high pH.

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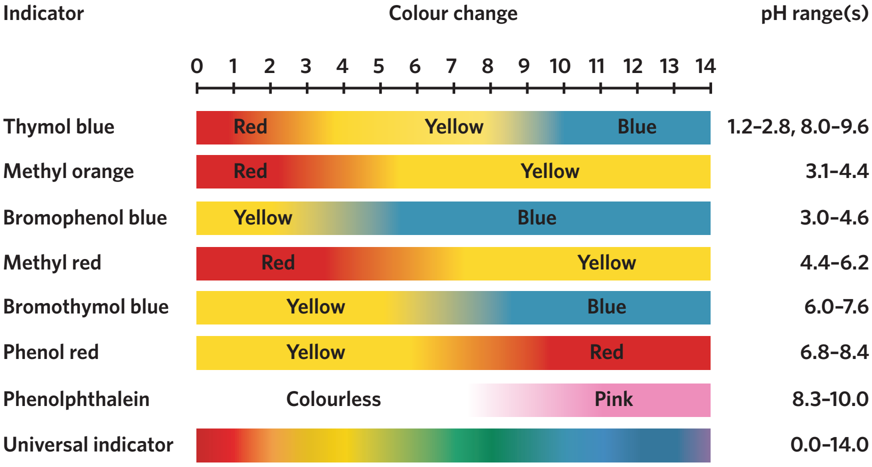
When choosing an appropriate acid-base indicator, it is important to know the approximate pH of the equivalence point. This is because we need to make sure that the equivalence point falls within the pH range for the colour change of the indicator – the observed end point of the titration. For example, when a strong acid reacts with a weak base, we would expect that the pH of the solution at the equivalence point would be lower than 7. Therefore, we would need to pick an acid-base indicator that will change colour within this range. Figure 10 displays the list of indicators in the VCE Data Book, their colour changes and what pH range they change in.



Indicator	Colour change	pH range(s)
Thymol blue	Red Yellow Blue	1.2–2.8, 8.0–9.6
Methyl orange	Red Yellow	3.1–4.4
Bromophenol blue	Yellow Blue	3.0–4.6
Methyl red	Red Yellow	4.4–6.2
Bromothymol blue	Yellow Blue	6.0–7.6
Phenol red	Yellow Red	6.8–8.4
Phenolphthalein	Colourless Pink	8.3–10.0
Universal indicator		0.0–14.0

Figure 10 List of indicators used in titrations

Consider the reaction between a strong acid and a weak base. Since the equivalence point will occur at a pH lower than 7, bromophenol blue could be an appropriate choice of indicator. This is because it changes from blue to yellow in the 3.0–4.6 pH range. As shown in figure 11, the titration can be used to accurately determine the equivalence point of the titration from the end point signalled by the colour change.

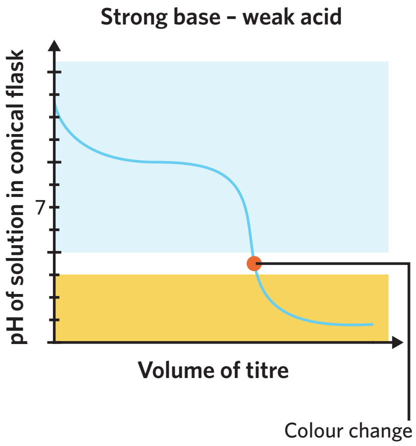


Figure 11 Titration curve of a reaction involving a strong acid and weak base with bromophenol blue

STRATEGY

It is important to note that the equivalence point should lie within the pH range of the indicator, rather than at the ends of its pH range. This is to ensure that a complete colour change occurs as close as possible to the equivalence point, rather than too early or too late.

Progress questions

Question 9

The equivalence point is the same as the end point in every titration.

A. True

B. False

Question 10

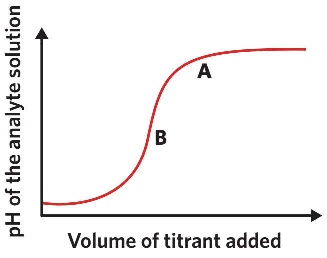
When the indicator changes colour, this signals that the ___________ has been reached.

A. equivalence point

B. end point

Question 11

Which point on the graph represents the equivalence point?



A. Point A

B. Point B

Continues →

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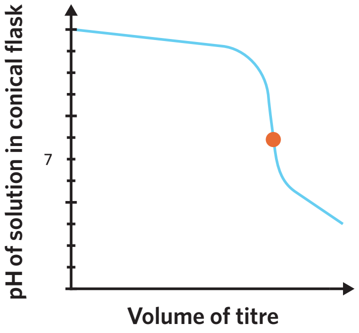
Question 12

Based on the titration curve shown, which of the following indicators would be most appropriate for this acid-base titration?



A. Methyl red

B. Phenol red

Use the following information to answer questions 13–14.

A titration was conducted using a NaOH standard solution and an HCl solution of unknown concentration. 0.006125 mol of NaOH was found to react with a 13.49 mL titre of HCl solution.

Question 13

Based on these results, the concentration of HCl is

A. 0.1322 mol L^{-1}

B. 0.4540 mol L^{-1}

C. 1.322 mol L^{-1}

D. 2.202 mol L^{-1}

VCAA 2016 exam Multiple choice Q7

Question 14

The experimental value of the concentration of HCl obtained from these titrations was less than the actual value. Which one of these actions by the students most likely accounts for the lower than expected result?

A. Rinsing the burette with water

B. Rinsing the pipette with water

C. Rinsing the conical flask with water

D. Rinsing the pipette with NaOH

Adapted from VCAA 2016 exam Multiple choice Q8

Theory summary

- Titration is a quantitative technique to determine the unknown concentration of a solution and/or the amount of a compound in a solution.
- Titrations use standard solutions of known concentration to determine the concentration of unknown solutions.
- The equivalence point is where the exact stoichiometric ratio of acid to base is present, whereas the end point is when the indicator changes colour.
- Primary standards need to be readily available, pure and not react with components in the air as it may compromise the concentration of the standard solution.
- Titration curves of acid-base reactions show pH changes and can be used to determine the appropriate choice of indicator.

The content in this lesson is considered fundamental prior knowledge to the laboratory analysis of organic compounds (Unit 4 AOS 2).

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13A Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 15–17.

Fabian, an enthusiastic chemistry student, decided to set up a titration experiment for his practical report. He is investigating the concentration of sodium hydroxide in different brands of household cleaners. He decides to use 1.0 mol L^{-1} hydrochloric acid as the titre solution in order to neutralise the sodium hydroxide present in a 20.00 mL aliquot of household cleaner. His table of results for one cleaner, SutherClean^{TM}, is recorded below:

Titration number	**1**	**2**	**3**	**4**	**5**
Final burette reading	22.50	23.75	32.60	34.90	27.80
Initial burette reading	2.50	3.70	13.00	14.95	8.30
Titre volume (mL)	20.00	20.05	19.60	19.95	19.50

Question 15 (1 MARK) 🌶

What is the neutralisation equation occurring in the conical flask during the titration?

A. $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

B. $HCl(aq) + NaCl(aq) \rightarrow NaOH(aq) + H_2O(l)$

C. $NaCl(aq) + H_2O(l) \rightarrow NaOH(aq) + HCl(aq)$

D. $Cl^-(aq) + Na^+(aq) \rightarrow NaCl(aq)$

Question 16 (1 MARK) 🌶

Which of the following titration numbers produced concordant results?

A. 1, 2, 3

B. 2, 3, 4

C. 2, 4, 5

D. 1, 2, 4

Question 17 (3 MARKS) 🌶🌶

Calculate the concentration of sodium hydroxide present in SutherClean^{TM}.

Exam-style

Question 18 (9 MARKS) 🌶🌶🌶

The Japanese giant hornet, *Vespa mandarinia japonica*, is the largest wasp in the world. Chris was recently stung by one of these wasps and decided to investigate the properties of its sting. He knows that magnesium hydroxide, $Mg(OH)_2$, is present in this wasp's sting. He carries out an acid-base titration to determine the concentration of magnesium hydroxide in the sting in order to design a therapeutic treatment that neutralises the sting. He knows that each wasp sting delivers 0.30 mL of fluid. His choice of titre solution, 0.0010 mol L^{-1} HCl, reacts with $Mg(OH)_2$ in a 2:1 ratio. For the titration, he uses a 20.00 mL aliquot of wasp sting fluid.

His table of results is shown:

Titration number	**1**	**2**	**3**
Final burette reading	24.60	29.50	23.60
Initial burette reading	2.50	7.35	1.50
Titre volume (mL)	22.10	22.15	22.10

a. How many wasp stings does Chris need to extract the venom from to carry out this titration? (1 MARK)

b. He uses a special indicator for this titration. What is the purpose of the indicator in a titration? (1 MARK)

Back to contents concordant titres? Explain why or why not. (2 MARKS)

d. Calculate the concentration of magnesium hydroxide present in the aliquot. (3 MARKS)

e. Calculate the amount of magnesium hydroxide present in each wasp sting. Based on this, what volume of 0.0010 mol L^{-1} HCl should be applied to the wasp sting for treatment? (2 MARKS)

Question 19 (7 MARKS) 🌶🌶🌶

Biodiesel is a viscous fuel that will be studied in more detail in VCE chemistry units 3 & 4. Biodiesel can be produced from waste vegetable oils. It is important for biodiesel producers to determine the pH of waste vegetable oils which are slightly acidic, however, normal techniques of measuring pH like litmus paper or a pH probe are not able to be used.

a. Suggest a reason why titration is used to determine the pH of the vegetable oils rather than the traditional methods. (1 MARK)

b. An excerpt of the experimental method is shown below to determine the pH of biodiesel:

1. Add 1 mL of waste vegetable oil to the conical flask, accompanied by 10 mL of propanol.
2. Add a few drops of the bromothymol blue indicator.
3. While swirling the conical flask, add 1.0 mol L^{-1} NaOH from the burette drop by drop until the indicator permanently transitions from yellow to blue colour.

With respect to this experimental method, explain the difference between the equivalence point and the end point. (4 MARKS)

c. The analyst performing this experiment did not use a primary standard solution to standardise the NaOH titre solution used in the experiment. Why is the use of a primary standard solution important? (2 MARKS)

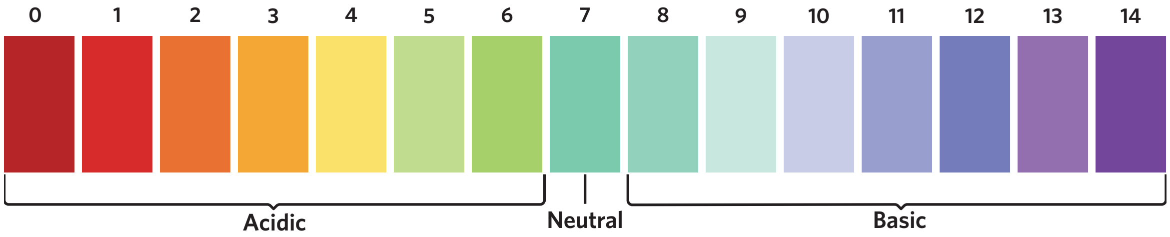
Question 20 (8 MARKS) 🌶🌶🌶

Acid-base indicators are used to determine the point a solution is acidic, basic or neutral through a colour change. This colour change is caused by the fact that most indicators are actually weak acids and bases themselves and actually react with the acid and base in the titration. Therefore, when they react, the conversion to the conjugate acid/base of the indicator is what causes the colour change.

a. Given that acid-base indicators are weak acids and bases themselves, would adding more indicator increase or decrease the accuracy of the experiment? Explain why or why not. (3 MARKS)

b. For a given titration, the equivalence point is at a pH of approximately 6.8. What would be the best choice of indicator to use? Justify your answer with reference to the definitions of the equivalence and end points. (3 MARKS)

c. Universal indicator is an acid-base indicator that has a colour spectrum that ranges from very acidic conditions to very basic conditions.



It contains thymol blue, methyl orange, methyl red, bromothymol blue and phenolphthalein. Describe a possible reason for how the rainbow of colours above could be created from just five indicators. (2 MARKS)

Question 21 (11 MARKS) 🌶🌶🌶

Sodium carbonate, Na_2CO_3, is a chemical compound that is commonly used as a primary standard. It is commercially available in a form that has some water present (i.e. it is not completely dry).

a. In order for sodium carbonate to be effective as a primary standard, what needs to happen to it between being purchased from a commercial source and its use. Explain your answer. (2 MARKS)

b. List three properties that define a suitable primary standard. (1 MARK)

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ation of a sodium carbonate primary solution, a volumetric flask is used. Why is a volumetric flask used as opposed to other types of glassware? (1 MARK)

d. Potassium hydrogen phthalate ($M = 204.2$ g mol^{-1}) is commonly used as a primary standard to standardise sodium hydroxide solutions. If 250.00 mL of a 1.000 mol L^{-1} potassium hydrogen phthalate solution is required, outline the steps involved in the preparation of this standard solution, including all relevant calculations. (7 MARKS)

Key science skills

Question 22 (8 MARKS)

It is important in titrations to wash the glassware before using it. This is because the glassware may not have been washed correctly or at all before being packed away. In one titration, a student is titrating 1.0 mol L^{-1} HCl with a 20.0 mL aliquot of NaOH solution of unknown concentration.

a. If the burette is washed with water, what effect would this have on the accuracy of the calculated concentration of NaOH solution? Justify your answer. (3 MARKS)

b. Why is the conical flask washed with distilled water rather than the NaOH solution? (2 MARKS)

c. Glassware is fragile and can be easily broken. Unfortunately, a student accidentally knocks over a retort stand and the burette, along with the acid inside, is knocked on the floor and smashes. Suggest three safety precautions that have to be taken to minimise the effect of this incident. (3 MARKS)

FROM LESSONS 16B & 16D

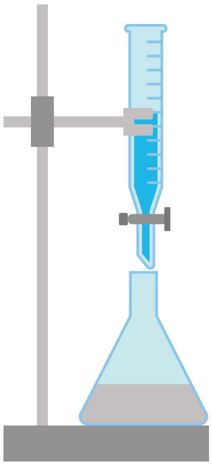
Questions from multiple lessons

Question 23 (8 MARKS)

Food is carefully monitored and regulated in Australia to ensure that products are safe for human consumption. Therefore, vinegar and its acidity is carefully controlled. The acidity of vinegar is caused by ethanoic acid, CH_3COOH. Vinegar acidity is closely monitored because of its importance in food preservation where a certain %(v/v) of acetic acid is required to kill harmful microbes and bacteria.

The Food Safety Board decided to investigate the pH of a new brand of vinegar by titration with the following procedure:

1. 25.00 mL aliquot of vinegar was pipetted into a conical flask.
2. 3 drops of phenolphthalein indicator were added to the vinegar.
3. The burette was filled with standardised 1.0 mol L^{-1} NaOH.
4. NaOH was dripped into the conical flask until the indicator turned from colourless to pink.
5. The burette values were recorded and the experiment repeated until concordant titres were obtained.



The results are shown below:

Titration number	1	2	3	4	5
Titre volume (mL)	21.55	21.35	21.60	21.85	21.65

a. Write the full balanced equation for the reaction between NaOH and ethanoic acid. (1 MARK)

b. What is the name of the functional group that gives ethanoic acid its acidic properties? (1 MARK)

c. Calculate the average titre volume. (1 MARK)

d. Calculate the concentration of ethanoic acid present in the vinegar. (2 MARKS)

e. Ethanoic acid has a density of 1.05 g mL^{-1} at 25 °C. It is regulated that the %(v/v) of ethanoic acid has to be greater than 4%. Calculate the %(v/v) of ethanoic acid in the vinegar. Does this new brand of vinegar meet the regulatory standards? (3 MARKS)

FROM LESSONS 7B & 12A

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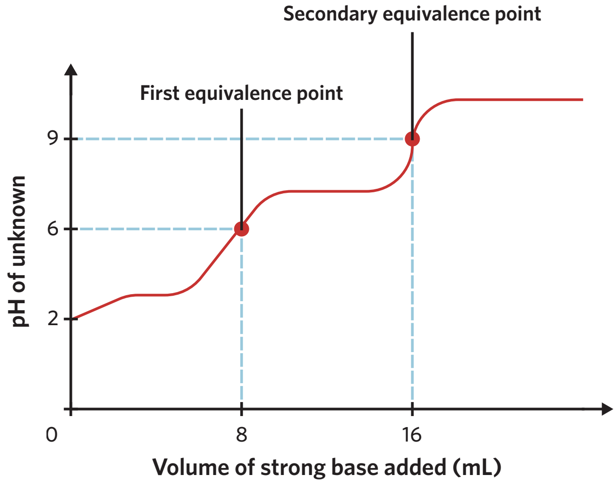
(8 MARKS) 🌶🌶🌶

There are many different types of acids, each with different strengths and properties.

a. What is a polyprotic acid? (1 MARK)

b. The titration curve for a diprotic acid is shown.



Why are there two equivalence points? (2 MARKS)

c. What choice of indicator(s) will be suitable to identify both equivalence points? (2 MARKS)

d. When performing this titration, the concentration of acid in the conical flask solution was too large for the base, and therefore, was diluted. The solution was diluted by a factor of 10.

If this acid-base titration had a stoichiometric ratio of 1:1 of acid to base, and 16.00 mL of a 1.0 mol L^{-1} basic solution was used to neutralise a 20.00 mL aliquot of the solution of acid, after the above dilution, what is the concentration of acid in the undiluted solution? (3 MARKS)

FROM LESSON 10A

Hints

18a. The number of wasps required depends on the volume of venom required and how much venom each wasp sting delivers.

18b. Since most acids and bases are colourless, the experimenter needs to know when to stop adding titre soluion to the unknown solution.

18c. Concordant titres differ from each other by a certain value.

18d. Titrations determine the number of moles of the unknown compound and then the concentration of the unknown solution.

18e. The aliquot of $Mg(OH)_2$ contains many wasp stings.

19a. pH measurements usually take place in aqueous solutions.

19b. The end point and the equivalence point are different.

19c. NaOH is not a primary standard.

20a. Only a few drops of indicator are added in a titration.

20b. Indicators indicate the end point of a reaction.

20c. There are many different types of indicators.

21a. Water adds to the weight of substances.

21b. Primary standards have specific required properties to perform their function.

21c. Volumetric flasks are highly accurate pieces of glassware.

21d. Primary standards solutions are prepared to ensure solutions of an accurately known concentration.

22a. The concentration of the titre has to be accurate in order to calculate the concentration of the unknown solution correctly.

22b. The amount of solute in the conical flask is what the calculation is based upon, not the concentration in the conical flask, as this is unknown.

22c. The type of precautions taken depends on the nature of the safety hazard.

23a. Acid-base reactions with metal hydroxides produce water and an ionic salt.

23b. Functional groups are special arrangements or groups of atoms that give compounds certain chemical and physical properties.

23c. To calculate the average titre, the concordant titres need to be determined.

23d. The stoichiometric ratio of acid to base according to the chemical equation is important in titration.

23e. The %(v/v) relates to the volume of acetic acid per volume of vinegar solution.

24a. "Poly" is a prefix that means "many".

24b. Equivalence points relate to when an acid is fully deprotonated.

24c. Different indicators are used depending on the pH of the equivalence point.

24d. Dilution factors relate the concentration of the diluted solution to the undiluted solution.

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Chapter 13 review

Multiple choice (10 MARKS)

Question 1 (1 MARK)

During a titration, a ___________ is used to deliver a solution into a ___________.

A. burette, measuring cylinder

B. burette, conical flask

C. conical flask, burette

D. conical flask, measuring cylinder

Question 2 (1 MARK)

Which of the following is not a necessary criterion for a standard solution in volumetric analysis?

A. The substance must be pure and in a stable form.

B. The substance must be non-reactive with the atmosphere.

C. The substance must be inexpensive.

D. The substance must have a low molecular mass.

Question 3 (1 MARK)

The concentration of a propanoic acid, C_2H_5COOH, solution was determined by titration with standardised 0.100 M sodium hydroxide, NaOH, solution at 25 °C. The reaction for this titration is shown.

$C_2H_5COOH(aq) + NaOH(aq) \rightarrow C_2H_5COONa(aq) + H_2O(l)$

A 0.10 M solution of sodium propanoate, C_2H_5COONa, in water has a pH of 8.9 at 25 °C.

The most appropriate indicator to use for this titration would be

A. thymol blue.

B. methyl red.

C. phenol red.

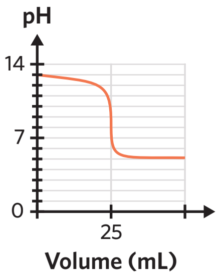
D. bromothymol blue.

VCAA (NHT) 2018 exam Multiple choice Q26

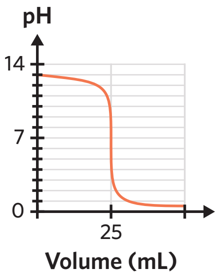
Question 4 (1 MARK)

Which one of the following graphs represents the pH change when a weak acid is added to a strong base?

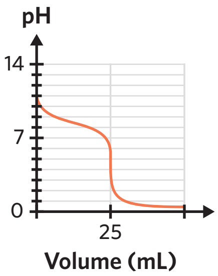
A.



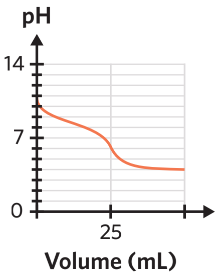
B.



C.



D.



VCAA 2015 exam Multiple choice Q1

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Question 5 (1 MARK)

In a titration, the solution of unknown concentration

A. should always be in the conical flask.

B. may be found in the conical flask or the burette.

C. should always be dispensed from the burette.

D. should never contain an indicator.

Question 6 (1 MARK)

Which of the following is accurate regarding the reading of a titre volume from a standard burette?

A. The volume should be read at eye-level to avoid the effects of parallax error.

B. The volume should be read at the bottom of the solution meniscus.

C. The volume can only be read to the nearest 0.05 mL.

D. All of the above

Question 7 (1 MARK)

Two titrations were performed as shown. Which one of the following statements is true?

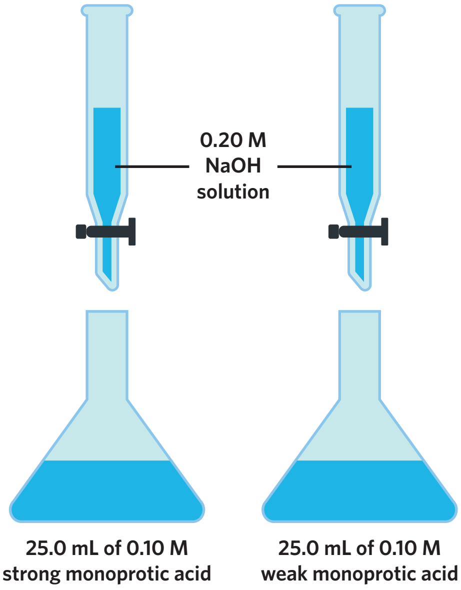
A. The weak acid will require a greater volume of NaOH solution than the strong acid to reach the equivalence point.

B. The weak acid will require a smaller volume of NaOH solution than the strong acid to reach the equivalence point.

C. The weak acid will require the same amount of NaOH solution as the strong acid to reach the equivalence point.

D. The equivalence point in a titration of a weak monoprotic acid with NaOH solution cannot be determined.

VCAA 2011 exam Multiple choice Q11



Question 8 (1 MARK)

In a titration, a 25.00 mL titre of 1.00 M hydrochloric acid neutralised a 20.00 mL aliquot of sodium hydroxide solution. If, in repeating the titration, a student failed to rinse one of the pieces of glassware with the appropriate solution, the titre would be

A. equal to 25.00 mL if water was left in the conical flask after final rinsing.

B. less than 25.00 mL if the final rinsing of the burette is with water rather than the acid.

C. greater than 25.00 mL if the final rinsing of the 20.00 mL pipette is with water rather than the base.

D. greater than 25.00 mL if the conical flask had been rinsed with the acid prior to the addition of the aliquot.

Adapted from VCAA 2013 exam Multiple choice Q3

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Question 9 (1 MARK)

What volume of 0.25 M hydrochloric acid (HCl) is required to react completely with 40 mL of 0.50 M calcium hydroxide ($Ca(OH)_2$)?

A. 40 mL

B. 80 mL

C. 120 mL

D. 160 mL

VCAA 2014 exam Multiple choice Q14

Question 10 (1 MARK)

A student titrated 25 mL aliquots of three different concentrations of an organic acid against a standardised potassium hydroxide, KOH, solution. The student's results are shown in the table.

	KOH titre for Sample 1 (mL)	KOH titre for Sample 2 (mL)	KOH titre for Sample 3 (mL)
Titration 1	20.35	19.85	21.55
Titration 2	20.45	19.65	21.45
Titration 3	20.30	20.45	21.65
Average titre	20.37	19.98	21.55

Which one of the following statements is consistent with the results shown in the table?

A. Sample 2 is the most concentrated acid.

B. Sample 3 is the most concentrated acid.

C. There is not enough information to draw a valid conclusion.

D. The averages in the table are correct as the results are concordant.

VCAA 2021 exam Multiple choice Q23

Short answer (30 MARKS)

Question 11 (5 MARKS)

Mahesh is a chemistry student learning how to conduct volumetric analyses. He wants to determine the concentration of oxalic acid in a sample of beetroot juice. The concentration of oxalic acid ($C_2H_2O_4$) in the sample is important as high concentrations may induce vomiting.

Before attempting the analytical technique, he tries to familiarise himself with some key terms.

a. Complete the following table to assist Mahesh. (3 MARKS)

Term	Definition
Equivalence point	
End point	
	Titres within a range of 0.10 mL

b. Mahesh adds beetroot juice with an unknown concentration of oxalic acid to a burette and adds 25 mL of 1.0 M sodium hydroxide (NaOH) to a pipette. After carrying out the reaction a number of times, the average titre of beetroot juice was found to be 27.40 mL. Calculate the concentration of oxalic acid in the beetroot juice. (Note: there is a 2:1 ratio between NaOH and oxalic acid.) (2 MARKS)

Question 12 (10 MARKS)

The following table lists a student's summary notes about food chemistry. It contains some correct and incorrect statements.

Number	Statement
1	The equivalence point and the end point will always occur at exactly the same time.
2	Methyl orange is an appropriate indicator for use in a titration between a weak acid and a strong base.
3	Number of moles, concentration, and volume are related by the equation $n = cV$.
4	Before a titration, the conical flask should be rinsed with water.
5	When 2 moles of KOH are fully dissolved in water, there will be 1 mole of potassium ions and 1 mole of hydroxide ions present in solution.
6	NaOH is a good primary standard.
7	Titrations are usually used to find an unknown concentration of a solution.
8	Titre volumes within ± 0.2 mL are considered concordant.
9	The volume on a burette should be read at eye level.
10	Primary standard solutions must have precisely known concentrations, so it is important to be careful when preparing them.

a. Identify four correct statements. (4 MARKS)

b. Identify three incorrect statements by writing the statement numbers. Explain why each statement is incorrect. (6 MARKS)

Question 13 (5 MARKS)

A student is attempting to accurately determine the concentration of a solution of sodium hydrogen carbonate in a titration against a standard solution of hydrochloric acid, HCl. The first step in this experiment is to accurately dilute 100.0 mL of a 1.00 M HCl stock solution to a 0.100 M solution using a 1.00 L volumetric flask. However, instead of using distilled water in the dilution, the student mistakenly adds 900.0 mL of 0.0222 M sodium hydroxide, NaOH, solution.

a. Write an equation for the reaction that occurs in the 1.00 L volumetric flask. (1 MARK)

b. Calculate the concentration of the hydrochloric acid in the 1.00 L volumetric flask after the student added the sodium hydroxide solution. Give your answer to correct significant figures. (2 MARKS)

c. The student then uses this contaminated hydrochloric acid solution to determine the accurate concentration of the unknown sodium hydrogencarbonate solution.

Will the calculated concentration of sodium hydrogencarbonate solution be greater or smaller than the true value? Justify your answer. (2 MARKS)

Adapted from VCAA 2009 Exam 1 Short answer Q3

Question 14 (10 MARKS)

Some students were set the task of determining the concentration of acetic acid in a particular brand of vinegar.

An outline of the method they used is given.

1. A burette is filled with a standard solution of sodium hydroxide.
2. The vinegar is diluted by a factor of 10 in a volumetric flask. A pipette is used to transfer 20.00 mL of diluted vinegar to a conical flask and a few drops of phenolphthalein indicator is added.
3. The diluted vinegar is titrated with the base. Titrations are repeated until three concordant results are obtained.

The equation for the reaction is:

$NaOH(aq) + CH_3COOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$

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a. The volumetric flask, the burette, the pipette and the conical flask are all rinsed before they are used. Indicate which solution should be used to finally rinse each of these pieces of glassware by ticking your responses in the appropriate boxes provided. (4 MARKS)

Glassware used	**Rinse with water**	**Rinse with diluted vinegar solution**	**Rinse with NaOH solution**
Volumetric flask			
Burette			
20.00 mL pipette			
Conical flask			

b. Why is the vinegar diluted before titrating? (1 MARK)

c. Explain why titrations are repeated until three concordant results are obtained. (1 MARK)

d. One student's results are given. The data shown in the student's laboratory book was:

Concentration of NaOH(aq) = 0.11 mol L^{-1}

Volume of undiluted vinegar = 10.00 mL

Total volume of diluted vinegar = 100.00 mL

Volume of diluted vinegar used in each titration = 20.00 mL

Average titre of NaOH = 15.35 mL

Based on these results, calculate the concentration, in mol L^{-1}, of acetic acid in the undiluted vinegar solution. Be careful to use the correct number of significant figures in your answer. (4 MARKS)

VCAA 2002 Exam 1 Short answer Question 6

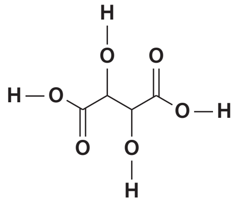
Key science skills (10 MARKS)

Question 15 (10 MARKS)

A group of students designed and carried out an experiment to investigate if tartaric acid, $C_4H_6O_6$, that was bought commercially is 99% pure, as claimed by the manufacturer. The experiment involved titrating $C_4H_6O_6$ with sodium hydroxide, NaOH, solution, calculating the percentage purity of $C_4H_6O_6$ and comparing the experimental value to the manufacturer's stated value. Part of the report submitted by one of the students is shown.

Research:
Tartaric acid is a diprotic acid that occurs naturally in grapes and other fruit.

Skeletal diagram for $C_4H_6O_6$



Equation for reaction
$C_4H_6O_6(aq) + 2NaOH(aq) \rightarrow Na_2C_4H_4O_6(aq) + 2H_2O(l)$

Aim: To determine the percentage purity of the commercial sample of tartaric acid by titration to verify the stated value of 99.0%

Calculations of predicted titre (in mL) (assuming 100% purity of sample)

$[C_4H_6O_6]$ solution $= \frac{n}{V}$

$[C_4H_6O_6]$ solution $= \frac{30 \text{ g}}{150 \text{ g mol}^{-1}} \times \frac{1}{0.50}$ L

$[C_4H_6O_6]$ solution = 0.40 mol L^{-1}

$n(C_4H_6O_6)$ in 10.00 mL = $c \times V$

$n(C_4H_6O_6)$ in 10.00 mL = 0.40 mol $L^{-1} \times \frac{10}{1000}$ L

$n(C_4H_6O_6)$ in 10.00 mL = 0.0040 mol

$n(NaOH) : n(C_4H_6O_6)$

2 : 1

$n(NaOH) = 2 \times 0.0040$ mol

$n(NaOH) = 0.0080$ mol

$V(\text{NaOH}) \text{ titre} = \frac{n}{c}$

$V(\text{NaOH}) \text{ titre} = \frac{0.0080 \text{ mol}}{0.5 \text{ M}}$

$V(\text{NaOH}) \text{ titre} = 0.016 \text{ L}$

$V(\text{NaOH}) \text{ titre} = 16.00 \text{ mL}$

Method

Part A – Preparation of tartaric acid solution:

1. Purchase tartaric acid, $C_4H_6O_6$, powder.
2. Prepare a solution of $C_4H_6O_6$ by accurately measuring 30.0 g of the powder, placing it in a 500.00 mL volumetric flask and then making it up to 500.00 mL with de-ionised water.

Part B – Titration:

3. Collect a stock solution of 0.5 M sodium hydroxide, NaOH, and use this to fill a burette.
4. Deliver a 10.00 mL aliquot of $C_4H_6O_6$ solution into a conical flask. Add four drops of phenolphthalein indicator.
5. Carefully titrate 0.5 M NaOH into the $C_4H_6O_6$ solution until a permanent pink colour remains.
6. Record the volume of the titre.
7. Repeat the titration until concordant titres are obtained.

Results

Trial number	Volume of aliquot of $C_4H_6O_6$ (mL)	Volume of titre of NaOH (mL)
1	10.00	14.95
2	10.00	14.80
3	10.00	14.60
4	10.00	14.75
5	10.00	14.80

Calculations

$\text{Average titre} = \frac{(14.80 + 14.75 + 14.80) \text{ mL}}{3}$

Average titre = 14.78 mL

$n = c \times V$

$n(\text{NaOH}) = 0.5 \text{ M} \times \frac{14.78}{1000} \text{ L}$

$n(\text{NaOH}) = 0.00739 \text{ mol}$

$n(C_4H_6O_6) : n(\text{NaOH})$

$1 \quad : \quad 2$

$n(C_4H_6O_6) = \frac{1}{2} \times n(\text{NaOH})$

$n(C_4H_6O_6) = \frac{0.00739 \text{ mol}}{2}$

$n(C_4H_6O_6) = 0.00370$ mol in 10.00 mL of $C_4H_6O_6$ solution

Percentage purity

$\% \text{ purity of } C_4H_6O_6 = \frac{\text{actual } n}{\text{predicted } n} \times 100\%$

$\% \text{ purity of } C_4H_6O_6 = \frac{0.0037 \text{ mol}}{0.0040 \text{ mol}} \times 100\%$

$\% \text{ purity of } C_4H_6O_6 = 92.4\%$

Conclusion

Through direct titration of tartaric acid with sodium hydroxide solution, the percentage purity of the commercial supply of tartaric acid was found to be 92.4%. This is less than the stated value of 99% purity. Consequently, the manufacturer's claim is wrong.

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a. Is the value the student used for the average titre in the 'Calculations' section appropriate? Explain your reasoning. (2 MARKS)

b. Consider the method undertaken by the student in this experiment to determine the percentage purity of $C_4H_6O_6$ powder. Identify how specific steps in the method affect the accuracy and reliability of the data. (4 MARKS)

c. Identify a limitation of the student's conclusion. How could this limitation be addressed? (2 MARKS)

d. The material safety data sheet (MSDS) for $C_4H_6O_6$ powder includes the following statement: 'Warning! This product causes eye, skin and respiratory tract irritation.' Apart from a laboratory coat, what personal protective equipment (PPE) should be used by the students in each of the following situations?

  i. Preparing the $C_4H_6O_6$ solution (1 MARK)

  ii. Conducting the titration (1 MARK)

Adapted from VCAA 2017 exam Short answer Q9c,d,e,f

FROM LESSONS 16B, 16D & 16E

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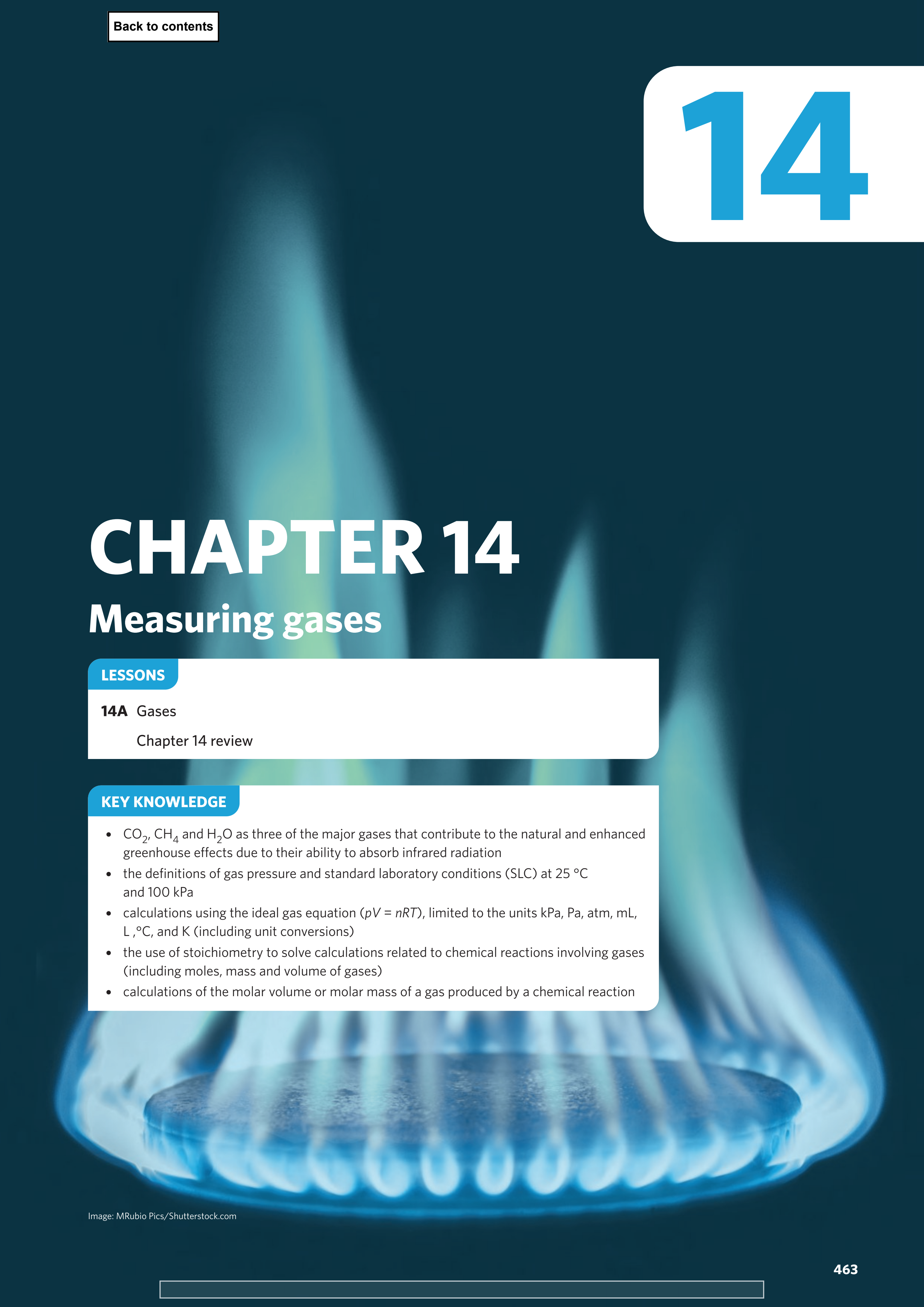

CHAPTER 14

Measuring gases

LESSONS

KEY KNOWLEDGE

- CO_2, CH_4 and H_2O as three of the major gases that contribute to the natural and enhanced greenhouse effects due to their ability to absorb infrared radiation
- the definitions of gas pressure and standard laboratory conditions (SLC) at 25 °C and 100 kPa
- calculations using the ideal gas equation ($pV = nRT$), limited to the units kPa, Pa, atm, mL, L ,°C, and K (including unit conversions)
- the use of stoichiometry to solve calculations related to chemical reactions involving gases (including moles, mass and volume of gases)
- calculations of the molar volume or molar mass of a gas produced by a chemical reaction

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14A Gases

STUDY DESIGN DOT POINTS

- CO_2, CH_4 and H_2O as three of the major gases that contribute to the natural and enhanced greenhouse effects due to their ability to absorb infrared radiation
- the definitions of gas pressure and standard laboratory conditions (SLC) at 25 °C and 100 kPa
- calculations using the ideal gas equation ($pV = nRT$), limited to the units kPa, Pa, atm, mL, L, 25 °C, and K (including unit conversions)
- the use of stoichiometry to solve calculations related to chemical reactions involving gases (including moles, mass and volume of gases)
- calculations of the molar volume or molar mass of a gas produced by a chemical reaction

14A

2.2.5.1 Natural and enhanced greenhouse effects
2.2.6.1 & 2.2.6.2 Behaviour of gases & defining pressure
2.2.7.1 & 2.2.8.1 The universal gas law & calculations involving gases
2.2.9.1 Molar volume

ESSENTIAL PRIOR KNOWLEDGE

2A Molecule
13A Stoichiometry
See questions 87–88.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: Vaclav Volrab/Shutterstock.com

How cold would Earth be without greenhouse gases?

Gases play an important role in regulating Earth's temperature and human activities continue to change the concentrations of these gases in the atmosphere. In this lesson, we will cover the general **greenhouse gases** and their properties, as well as calculations involving **pressure**, temperature, volume and the number of moles in chemical reactions.

KEY TERMS AND DEFINITIONS

Gas state of matter of a substance with no fixed shape or volume that fills the available space

Global warming a long-term increase in the average global temperature of Earth's climate system since the industrial revolution due to human activities

Greenhouse effect the natural warming of the earth that results when gases in the atmosphere trap heat from the sun that would otherwise escape into space

Greenhouse gases molecules that absorb infrared radiation from the sun

Ideal gas substance assumed to be composed of molecules of negligible volume with no intermolecular interactions

Infrared radiation heat energy emitted from the sun

Pressure force a gas exerts on the walls of its container

Standard Laboratory Conditions (SLC) set of conditions used as a standard for experiments (298 K and 100 kPa)

Natural and enhanced greenhouse effects 2.2.5.1

Without the **greenhouse effect**, Earth would have a similar temperature to that of a household freezer.

How do natural greenhouse gases make Earth warmer?

The natural greenhouse effect is a phenomenon caused by gases naturally present in the atmosphere like carbon dioxide (CO_2), methane (CH_4) and water (H_2O). These gases have the property of being able to absorb and re-emit **infrared radiation** by the following process:

- Infrared radiation hits a molecule like carbon dioxide[1] and causes the bonds to vibrate and bend at a greater rate.
- The molecule then re-emits this infrared heat energy in all directions, with some energy escaping into space, but less than would have escaped if the atmosphere and its greenhouse gases were not present (figure 1).
- This results in some of the sun's heat energy becoming trapped in the lower atmosphere, making the Earth much warmer than it would be otherwise (figure 2).

KEEN TO INVESTIGATE?

[1] **How exactly does carbon dioxide cause global warming?**
Search YouTube: How Exactly Does Carbon Dioxide Cause Global Warming?

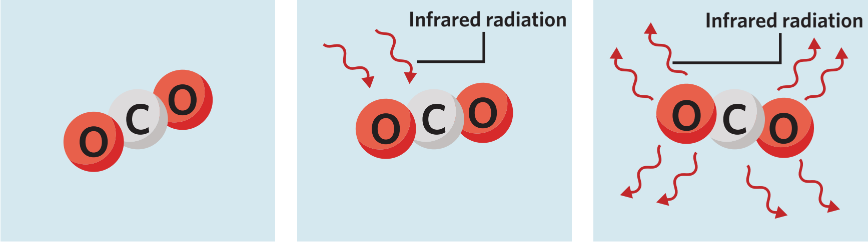

Figure 1 A CO_2 molecule absorbing and re-emitting infrared heat energy

MISCONCEPTION

'All gases in the atmosphere can act as a greenhouse gas.'

Simple molecules like oxygen and nitrogen do not absorb and re-emit infrared radiation.

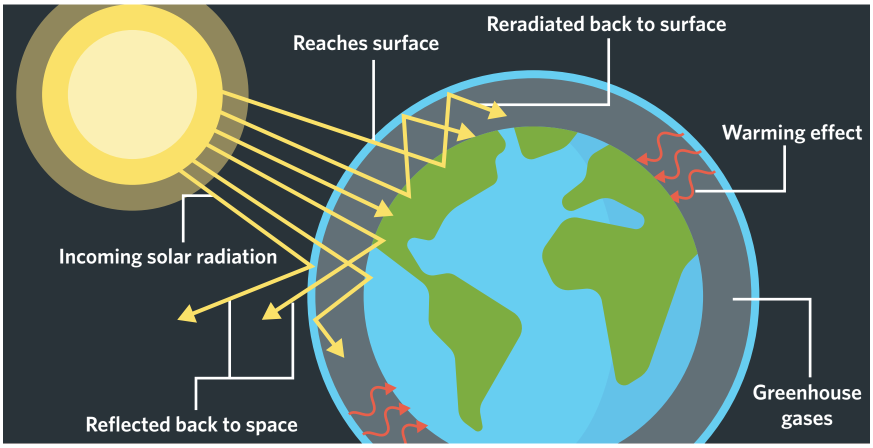

Figure 2 Natural greenhouse effect

How are humans enhancing the greenhouse effect?

Human activities have led to a buildup of extra greenhouse gases in the atmosphere. These include nitrous oxide (N_2O), methane (CH_4) and many other industrial gases. As a result, average surface temperatures are rising at a faster rate than ever before in a phenomenon called **global warming**.[2] Figure 3 illustrates how more heat is trapped due to the increased amount of greenhouse gases present in the atmosphere as a direct result of human activities (figure 4).

KEEN TO INVESTIGATE?

[2] **What is climate change: evidence and causes?**
Search: rsc video evidence climate change

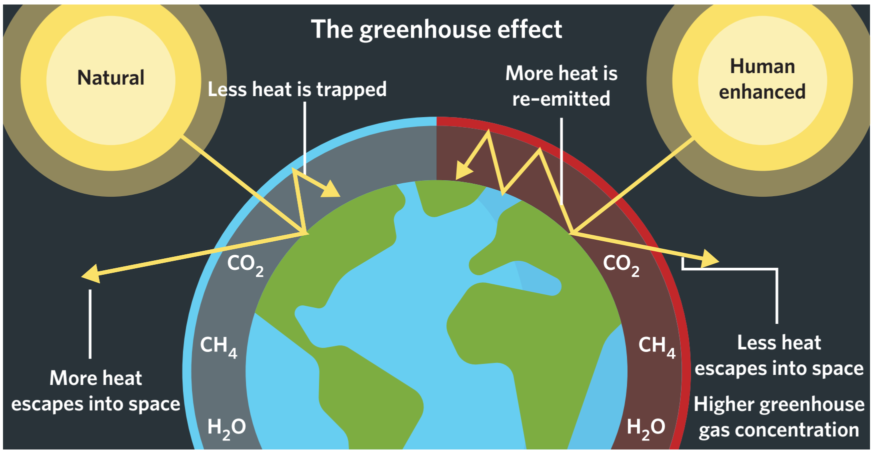

Figure 3 Natural versus enhanced greenhouse effect

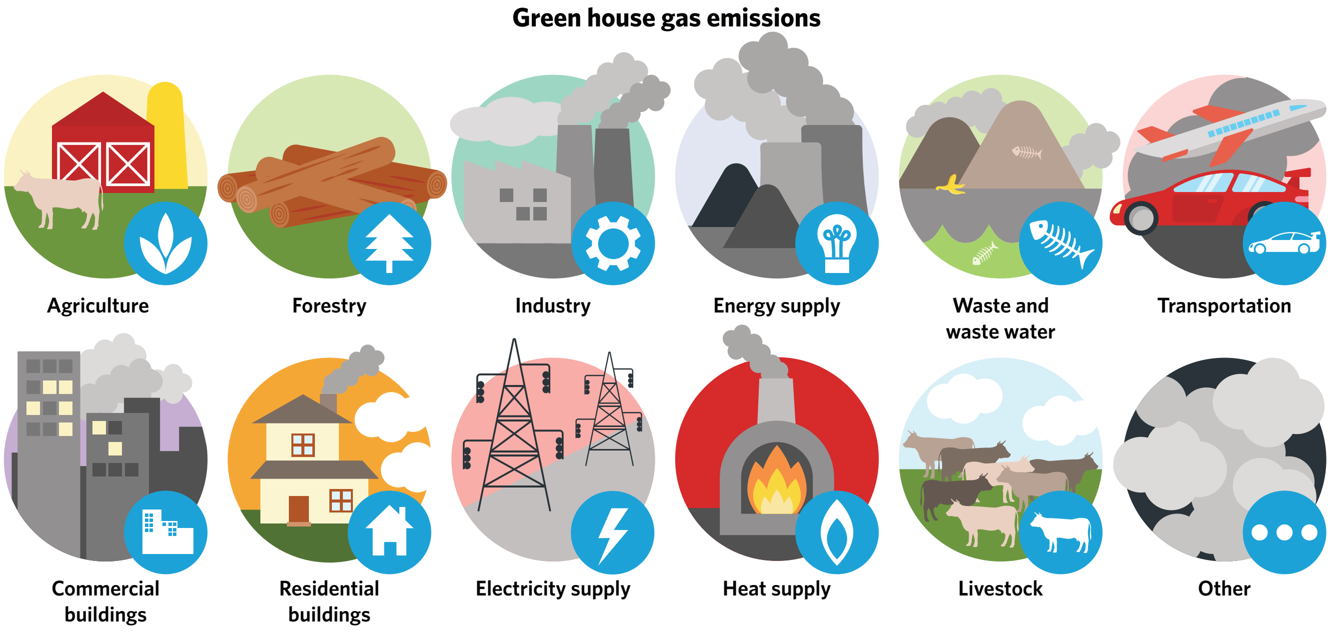

Image: In Art/shutterstock.com

Figure 4 Examples of human activities that produce greenhouse gases

Progress questions

Question 1

Water vapour (cloud) is a potent greenhouse gas.

A. True

B. False

Question 2

Infrared radiation is a form of

A. sound energy.

B. heat energy.

Question 3

All the infrared radiation emitted from the sun gets trapped by Earth's atmosphere.

A. True

B. False

Behaviour of gases & defining pressure 2.2.6.1 & 2.2.6.2

The particles within a gas can be modelled by the **ideal gas** approximation. Pressure (p) is a property of a gas and can be expressed in a number of different units.

What is an ideal gas?

A gas can be described as a substance with no fixed shape or volume. For example, the air in the atmosphere around us consists of gases and does not have a constant shape. It can expand or condense to fill the available volume. For the purposes of VCE chemistry, we will assume all gases are ideal gases, as real gases are quite complicated to work with.

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This means we assume that molecules within a gas:

- experience no intermolecular forces,
- have zero volume individually,
- and exchange no energy when a collision occurs.

This assumption makes our calculations far simpler, and allows us to approximate the properties of gases precisely.

How can pressure be defined and modelled?

Pressure is defined as the force exerted by a gas on the walls of the container enclosing it. The molecules within a gas are constantly moving around and colliding both with each other and with the walls of the container or vessel they are in, as shown in figure 5. The collisions of molecules with the container wall collectively result in a force we define as pressure.

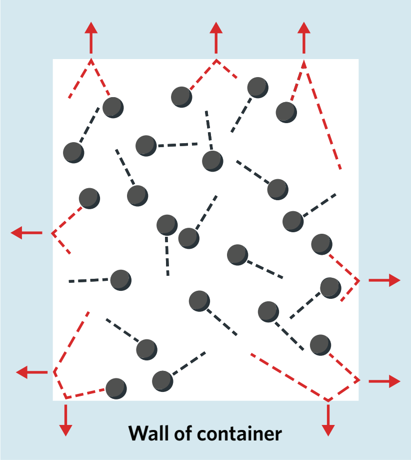

Figure 5 Representation of collisions within a closed volume

We quantify pressure by identifying the force exerted by the gas per unit area of the container. Therefore:

$$\text{Pressure (Pa)} = \frac{\text{force (N)}}{\text{area (m}^2\text{)}}$$

where area is the interior surface area of the container that the gas collides with.

In this equation, we conventionally use force measured in newtons (N), and area measured in metres squared (m^2) meaning pressure has the units of newtons per metre squared (Nm^{-2}). This unit, Nm^{-2}, is more commonly referred to in chemistry as the Pascal (Pa).

Progress questions

Question 4

Ideal gases actually exist.

A. True

B. False

Question 5

Pressure is best defined as

A. the amount of gas per unit area of its container.

B. the force per unit area of a gas on its container.

Question 6

Nm^{-2} is the same unit as a Pascal.

A. True

B. False

The universal gas law & calculations involving gases 2.2.7.1 & 2.2.8.1

We can quantify the properties of a gas using an equation known as the universal gas law.

How can we convert between units used for pressure?

Pressure can also be expressed in units other than Pascals (Pa), particularly kilopascals (kPa) and atmospheres (atm). Refer to figure 6 for conversions of common units.

Pa	:	kPa	:	atm
100 000 Pa	:	100 kPa	:	0.987 atm

Figure 6 The conversion of units of pressure

USEFUL TIP

All the information required to convert pressure units can be found in the VCE Data Book.

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WORKED EXAMPLE 1

An ideal gas is stored in a cylinder at a pressure of 15 atm. Express this pressure in pascals.

What information is presented in the question?

A gas is stored at a pressure of 15 atmospheres.

What is the question asking us to do?

Convert atmospheres to pascals.

What strategies do we need in order to answer the question?

1. Use the VCE Data Book to find the relationship between atmospheres and pascals.
2. Apply this to our pressure value.

Answer

The VCE Data Book states 0.987 atm is equivalent to 100 kPa.

$0.987 \text{ atm} = 100 \times 10^3 \text{ Pa}$

$\text{Pa} = \dfrac{15 \text{ atm}}{0.987 \text{ atm}} \times (100 \times 10^3)$

$\text{Pa} = 1.5 \times 10^6$

What is the universal gas law, *pV* = *nRT* ?

The properties of a gas such as temperature, pressure, volume and number of moles are related and dependent on each other. We can represent this relationship with the equation given in figure 7. This equation holds true for all ideal gases. Note that 'R' is a constant number known as the universal gas constant. It can be found in the VCE Data Book and has a value of 8.31 J mol^{-1} K^{-1}.

$$pV = nRT$$

Symbol	Quantity	Unit
p	Pressure	kPa
V	Volume	L
n	Number of moles	mol
R	Universal gas constant (8.31)	J mol^{-1} K^{-1}
T	Temperature	K

Figure 7 The universal gas equation and the units

The units of the variables used in this equation are also given in figure 7 and are a guide to the calculations we require to obtain the above units. The conversions for pressure were covered previously, and the calculation required for temperature conversion is given in figure 8.

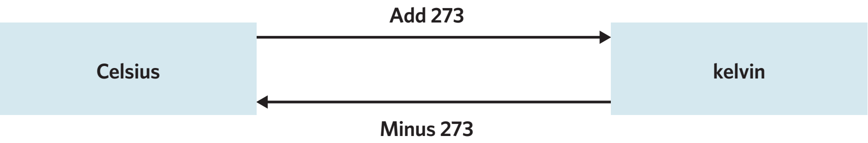

Figure 8 The conversion for units of temperature

MISCONCEPTION

'Any units can be used in *pV* = *nRT*.'

The units used in the universal gas equation must cancel out the unit of the universal gas constant, which is J mol^{-1} K^{-1}. Therefore, in VCE, the units kPa, L and K should be input into the equation.

USEFUL TIP

0 °C = 273 K and −273 °C = 0 K (absolute zero).

USEFUL TIP

IUPAC recommends using a lowercase *p* as the symbol for pressure, but it is also commonly represented by a capital *P*.

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WORKED EXAMPLE 2

15 L of gas is stored in a cylinder at pressure 2.0 kPa. If the temperature of the cylinder is 25 °C, how many moles of gas are contained?

What information is presented in the question?

Volume = 15 L

Pressure = 2.0 kPa

Temperature = 25.0 °C

What is the question asking us to do?

Find the number of moles.

What strategies do we need in order to answer the question?

1. Convert all quantities to the required units.
2. Rearrange the universal gas law equation for the required quantity.
3. Apply the gas law.

Answer

Volume = 15 L

Pressure = 2.0 kPa

Temperature 25.0 °C = (25 + 273) K = 298 K

The universal gas equation is $pV = nRT$

We can rearrange this for n, giving $n = \frac{pV}{RT}$

$$n = \frac{2.0 \text{ kPa} \times 15 \text{ L}}{8.31 \times 298 \text{ K}}$$

$n = 0.012$ mol or 1.2×10^{-2} mol

In order to compare ideal gases, and experiments with gases, it is useful to have a standard set of conditions under which experiments are conducted. The scientific community has agreed upon a set of conditions known as the **Standard Laboratory Conditions (SLC)** that provide this consistency. The SLC conditions are:

- Temperature of 298 K (25 °C)
- Pressure of 100 kPa (100 000 Pa or 0.987 atm)

These numbers can be found in the VCE Data Book.

Progress questions

Question 7

Standard laboratory conditions (SLC) change depending on the laboratory that the investigation is carried out in.

A. True

B. False

Question 8

A sealed flask of oxygen is at pressure 0.987 atm. This can be given in kPa as

A. 100 kPa

B. 1.00 kPa

Molar volume 2.2.9.1

Molar volume is very useful in gas calculations when the conditions are at SLC.

Why is molar volume useful in calculations?

The molar volume is the volume of one mole of substance specified at a given temperature and pressure. Alternatively, at the same temperature and pressure, an equal number of molecules of gas are contained in the same volume.

USEFUL TIP

The volume of one mole of ideal gas is always 24.8 L mol^{-1} at SLC, regardless of the formula or mass of the gaseous substance.

WORKED EXAMPLE 3

1.00 mol of a gas is stored at SLC. Calculate the volume the gas occupies.

What information is presented in the question?

Pressure = 100 kPa

Temperature = 25 °C

Number of moles = 1.00 mol

What is the question asking us to do?

Find the volume.

What strategies do we need in order to answer the question?

1. Convert all quantities to required units.
2. Rearrange the universal gas law equation for the required quantity.
3. Apply the gas law.

Answer

Pressure = 100 kPa

Temperature 25 °C, (25 + 273) = 298 K

Number of moles = 1.00 mol

The universal gas equation is $pV = nRT$

We can rearrange this for V, giving $V = \frac{nRT}{p}$

$$V = \frac{1.00 \text{ mol} \times 8.31 \times 298 \text{ K}}{100 \text{ kPa}}$$

$V = 24.8$ L

MISCONCEPTION

'V_m in VCE chemistry can change.'

The value of V_m depends on the temperature and pressure. However, in VCE chemistry, it will always be taken as 24.8 L mol^{-1} as the conditions will always be at SLC, which are fixed at 298 K and 100 kPa.

In worked example 3, we determined that one mole of any ideal gas at SLC occupies 24.8 L. From this, we can infer that the relationship between the number of moles and volume for an ideal gas is

$$n = \frac{V}{V_m}$$

V_m is the value 24.8 L mol^{-1}, which was calculated in worked example 3. This value, known as the molar volume of an ideal gas, can be found in the VCE Data Book.

WORKED EXAMPLE 4

A sample of the anaesthetic gas, halothane, occupies a volume of 10.2 L at 25.0 °C and 100 000 Pa. How many moles of halothane are present in the sample?

What information is presented in the question?

Volume = 10.2 L

Temperature = 25.0 °C

Pressure = 100 000 Pa

What is the question asking us to do?

Find the number of moles.

What strategies do we need in order to answer the question?

1. Convert the units Pa to kPa and °C to K.
2. Use the molar volume formula.

Answer

Temperature 25.0 °C, (25.0 + 273) = 298 K (SLC)

Pressure 100 000 Pa, $\frac{100\,000}{100} = 100$ kPa (SLC)

Therefore the molar volume equation can be used where $n = \frac{V}{V_m}$.

$$n = \frac{10.2 \text{ L}}{24.8 \text{ L mol}^{-1}}$$

n(halothane) = 0.411 mol

WORKED EXAMPLE 5

What is the volume of 0.025 g of hydrogen gas at SLC?

What information is presented in the question?

The mass of hydrogen gas is 0.025 g.

What is the question asking us to do?

Find the volume.

What strategies do we need in order to answer the question?

1. Calculate the molar mass of 1 mole of hydrogen gas.
2. Calculate the moles of hydrogen gas in 0.025 g.
3. Use the molar volume formula.

Answer

Hydrogen gas has a molar mass of 2.0 g mol^{-1}.

$$n(H_2) = \frac{0.025\text{ g}}{2.0\text{ g mol}^{-1}}$$

$$n(H_2) = 0.0125\text{ mol}$$

Rearrange $n = \frac{V}{V_m}$ to $V = n \times V_m$

$$V(H_2) = 0.0125\text{ mol} \times 24.8\text{ L mol}^{-1}$$

$$V(H_2) = 0.31\text{ L}$$

How can molar volume and stoichiometric ratios from chemical equations be used together to calculate the volume of gases?

In chemical equations the stoichiometric ratios are often different ratios between reactants and products. In order to determine volumes these ratios must be taken into account.

WORKED EXAMPLE 6

Calculate the volume of carbon dioxide given off at SLC when an excess of dilute hydrochloric acid is added to 2.00 g of calcium carbonate ($M = 100.1$ g mol^{-1}) using the given equation.

$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

What information is presented in the question?

The mass of calcium carbonate (2.00 g) and the balanced equation of the reaction.

What is the question asking us to do?

Find the volume.

What strategies do we need in order to answer the question?

1. Calculate the moles of calcium carbonate in 2.00 g.
2. Use the stoichiometric ratio from the equation.
3. Use the molar volume formula.

Answer

$$n(CaCO_3) = \frac{n}{M}$$

$$n(CaCO_3) = \frac{2.00\text{ g}}{100.1\text{ g mol}^{-1}}$$

$$n(CaCO_3) = 0.01998\text{ mol}$$

The stoichiometric ratio from the balanced equation is

$n(CaCO_3) : n(CO_2)$

1 : 1

$$n(CO_2) = 0.01998\text{ mol}$$

Rearrange $n = \frac{V}{V_m}$ to $V = n \times V_m$

$$V(CO_2) = 0.01998\text{ mol} \times 24.8\text{ L mol}^{-1}$$

$$V(CO_2) = 0.496\text{ L}$$

WORKED EXAMPLE 7

Calculate the volume of carbon dioxide evolved at SLC when an excess of propane gas is added to 0.100 g of oxygen gas ($M = 32.0\ \text{g mol}^{-1}$) using the given equation.

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

What information is presented in the question?

The mass of oxygen (0.100 g) and the balanced equation of the reaction.

What is the question asking us to do?

Find the volume.

What strategies do we need in order to answer the question?

1. Calculate the moles of oxygen in 0.100 g.
2. Use the stoichiometric ratio from the equation.
3. Use the molar volume formula.

Answer

$$n(O_2) = \frac{m}{M}$$

$$n(O_2) = \frac{0.100\ \text{g}}{32.0\ \text{g mol}^{-1}}$$

$n(O_2) = 0.003125$ mol (or 3.125×10^{-3} mol)

The stoichiometric ratio from the balanced equation is

$$n(O_2) : n(CO_2)$$

$$5 \quad : \quad 3$$

$$n(CO_2) = 0.003125\ \text{mol} \times \frac{3}{5}$$

$$n(CO_2) = 0.001875\ \text{mol}$$

Rearrange $n = \frac{V}{V_m}$ to $V = n \times V_m$

$$V(CO_2) = 0.001875\ \text{mol} \times 24.8\ \text{L mol}^{-1}$$

$V(CO_2) = 0.0465$ L (or 4.65×10^{-2} L)

USEFUL TIP

To calculate volume, the combined equation is $V = \frac{m}{M} \times \frac{RT}{p}$

To calculate pressure, $p = \frac{m}{M} \times \frac{RT}{V}$

To calculate molar mass, $M = m \times \frac{RT}{pV}$

How can $pV = nRT$ and $n = \frac{m}{M}$ be used together to calculate the volume, mass and molar mass of gases in chemical reactions?

Given that $n = \frac{m}{M}$ and the number of moles can be calculated for a gas using $pV = nRT$, then either mass (m) or molar mass (M) can also be calculated. If the molar mass of the gas is known, then the mass can be calculated using $m = M \times n$. If the mass and moles are known then the molar mass can be calculated using $M = \frac{m}{n}$, which can also help identify the gas. It is often convenient to combine these two equations to make

$$pV = \frac{m}{M} \times RT$$

WORKED EXAMPLE 8

What pressure is exerted by 1.65 g of oxygen gas ($M = 32.0\ \text{g mol}^{-1}$) in a volume of 4.00 L at a temperature of 75 °C? ($R = 8.31\ \text{J mol}^{-1}\ \text{K}^{-1}$).

What information is presented in the question?

Mass of oxygen = 1.65 g

Volume = 4.00 L

Temperature = 75 °C

What is the question asking us to do?

Calculate the pressure.

What strategies do we need in order to answer the question?

1. Convert Celsius to kelvin by adding 273.
2. Calculate the moles of oxygen in 1.65 g.
3. Use the rearranged ideal gas equation equalling p.

Answer

Temperature in kelvin, (75 + 273) = 348 K

$$p = \frac{m}{M} \times \frac{RT}{V}$$

$$p(O_2) = \frac{1.65\ \text{g}}{32.0\ \text{g mol}^{-1}} \times \frac{8.31 \times 348\ \text{K}}{4.00\ \text{L}}$$

$$p(O_2) = 37.3\ \text{kPa}$$

Progress questions

Question 9

1.3 mol of neon gas at 17 °C and 380 kPa would have __________ a sample of 1.3 mol of nitrogen gas under the same conditions.

A. the same volume as

B. a smaller volume than

Question 10

A sample of argon gas occupies a 37 L sealed container and is at 35 °C at 1 atm. Under these same conditions, it would be expected that a different gas would have

A. the same number of moles.

B. a greater number of moles.

Question 11

Which of the following is the correct combined ideal gas, mass and molar mass equation?

A. $M = m \times \frac{RT}{P}$

B. $V = \frac{m}{M} \times \frac{RT}{P}$

Theory summary

- CO_2, CH_4 and H_2O are three of the major gases that contribute to the natural and enhanced greenhouse effects.
- CO_2, CH_4 and H_2O are considered greenhouse gases due to their ability to absorb and re-emit infrared radiation.
- Human activities are increasing the concentrations of greenhouse gases in the atmosphere, which is in turn accelerating the rate of global warming.
- A gas is a state of matter with no fixed shape or volume where molecules fill the available space.
- Pressure is the force a gas exerts on its surrounding container.
- Pressure is given by force per unit area.
- The universal gas law $pV = nRT$ can be used to determine the volume, pressure, temperature or number of moles of a gas, assuming an ideal gas.
- SLC is a set of conditions used to standardise scientific experiments.
- Stoichiometric ratios from equations are used to determine amounts of unknowns.
- We can apply $n = \frac{V}{V_m}$ when a gas is at SLC.

The content in this lesson is considered fundamental prior knowledge to measuring changes in chemical reactions (Unit 3 AOS 1).

14A Questions

Mild Medium Spicy

Deconstructed

Use the following information to answer questions 12–14.

Nitrous oxide N_2O is a colourless, non-flammable gas commonly used as an anaesthetic. Nitrous oxide is administered to a patient from a tank of fixed volume of 50 L in a hospital room at a consistent 25 °C. The patient requires 0.0455 moles of nitrous oxide, and the entire volume of the tank is delivered. The hospital needs to determine the minimum pressure that the tank should be designed to hold.

Question 12 (1 MARK)

Which of the following equations is best suited to perform this calculation?

A. $p = \frac{m}{M} \times \frac{RT}{V}$

B. $M = m \times \frac{RT}{pV}$

C. $V = \frac{nRT}{pV}$

D. $p = \frac{nRT}{V}$

Question 13 (1 MARK)

What is the temperature conversion required?

A. Subtract 273

B. Divide by 273

C. Add 273

D. Multiply by 273

Question 14 (2 MARKS)

What minimum pressure should the tank be designed to hold?

Exam-style

Question 15 (2 MARKS)

Identify two human activities that increase the levels of methane and carbon dioxide in the atmosphere.

Question 16 (3 MARKS)

Scientists have shown that the concentration of carbon dioxide in the atmosphere is increasing. Describe how increased levels of atmospheric carbon dioxide are causing enhanced global warming.

Question 17 (1 MARK)

Under SLC, assuming gas ideality, which of the following statements is true?

A. One mole of chlorine gas occupies the same volume as one mole of neon gas.

B. 20 grams of chlorine gas occupies exactly double the volume of 10 grams of neon gas.

C. One mole of chlorine gas occupies a greater volume than one mole of neon gas.

D. 20 grams of chlorine gas occupies the same volume as 10 grams of neon gas.

Question 18 (3 MARKS)

Two cylinders with the same volume, labelled D and G, are at SLC and contain 10.0 g of Cl_2 ($M = 71.0$ g mol^{-1}) gas and 2.85 grams of an unknown monatomic gas respectively. What is the identity of the unknown gas?

Question 19 (5 MARKS) 🌶🌶🌶

A 250 mL aerosol can contains 4.60 g of carbon monoxide (CO). The label contains a warning that, at temperatures exceeding 65.0 °C, the can may explode.

a. At the point of explosion, what pressure is the gas inside the can? (4 MARKS)

b. Express the pressure at the point of explosion in atmospheres. (1 MARK)

Question 20 (6 MARKS) 🌶🌶🌶

Zoe is a chemical engineer investigating the properties of a range of different gases. 5.00 gram samples of each of the following gases are analysed in separate containers at SLC: carbon dioxide ($M(CO_2) = 44.0 \text{ g mol}^{-1}$), sulfur dioxide ($M(SO_2) = 64.0 \text{ g mol}^{-1}$), and nitrogen gas ($M(N_2) = 28.0 \text{ g mol}^{-1}$).

a. Will the three gases occupy the same volume? Justify your answer with calculations. (4 MARKS)

b. Zoe decides to further analyse the carbon dioxide sample. A cylinder of the sample is removed from SLC and the temperature increased to 40 °C. Would you expect the pressure to increase, decrease, or remain constant given that the volume is kept constant? Justify your answer using the universal gas law. (2 MARKS)

Question 21 (3 MARKS) 🌶🌶🌶

55 mL of a pure, unknown gas was measured at a temperature of 363 K and a pressure of 100 kPa and was found to weigh 0.135 g. Find the molar mass of this gas.

Question 22 (3 MARKS) 🌶🌶🌶

At what temperature (in kelvin) will 2.00 g of pure Cl_2 gas, $M(Cl_2) = 71.0 \text{ g mol}^{-1}$, in a volume of 1.00 L, exert a pressure of 100 kPa?

Question 23 (4 MARKS) 🌶🌶🌶

Calculate the volume of oxygen gas required at SLC to completely combust 0.150 g of pure methane gas, $M(CH_4) = 16.0 \text{ g mol}^{-1}$, according to the equation $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$.

Key science skills

Question 24 (8 MARKS) 🌶🌶🌶

Zahra was given the task of determining the molar mass of an unknown gas using the hypodermic syringe shown.

Image: Mediarts/shutterstock.com

a. What information does Zahra need to know (using the ideal gas equation) in order to work out the pressure of the laboratory they are working in? (1 MARK)

b. What equation will they need to use in order to calculate the molar mass of the unknown gas? (1 MARK)

c. Zahra was provided with the following instructions

- Extract 0.20 mL of propanone into a hypodermic syringe and then measure the mass of this syringe
- Remove a gas syringe from the oven and note the volume of air already in the barrel – about 5 mL.
- Inject the propanone through the self-seal cap into the barrel. The plunger will move straight away.
- Put the gas syringe back into the oven.
- Measure the mass of the empty hypodermic syringe immediately.
- After a few minutes measure the volume of the gas in the gas syringe, record the temperature of the oven shelf and the pressure of the room.

Identify two different hazards and corresponding precautions that Zahra must follow. (4 MARKS)

d. Consider the scale shown on the hypodermic syringe. In terms of resolution and precision, identify the limitation of the scale and explain the effect this could have on the calculated value of the molar mass. (2 MARKS)

FROM LESSONS 16B & 16D

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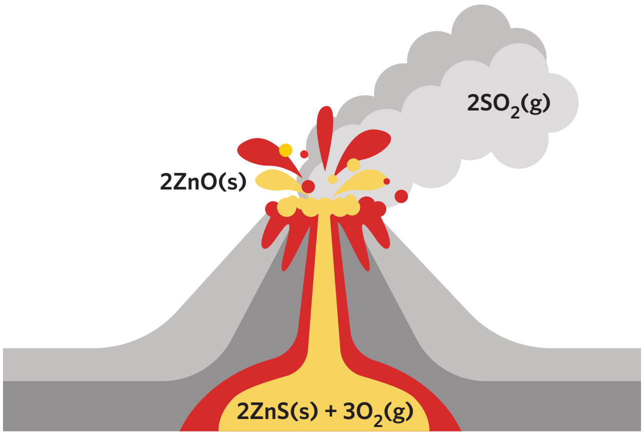
Questions from multiple lessons

Question 25 (5 MARKS)

Sulfur dioxide, $SO_2(g)$, is a toxic gas released naturally during volcanic eruptions. One of the main reactions responsible for this release of thousands of tonnes of $SO_2(g)$ is the roasting of sphalerite ores, also known as zinc sulfide, $ZnS(s)$.
The balanced chemical reaction for this process is shown.

$2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$



a. In the conversion of $ZnS(s)$ to $SO_2(g)$, does sulfur undergo oxidation or reduction? Explain. (2 MARKS)

b. Write the balanced half-equation for the conversion of $S^{2-}(s)$ to $SO_2(g)$ as it would occur in acidic aqueous solution. (1 MARK)

c. Write the balanced half-equation for the conversion of $O_2(g)$ to $O^{2-}(s)$. (1 MARK)

d. For every mole of $ZnS(s)$ consumed, how many moles of $ZnO(s)$ will be produced? (1 MARK)

FROM LESSON 11A

Hints

15. Recall the human activities that add to the natural greenhouse gases.

16. Carbon dioxide behaves like the glass in a greenhouse.

17. Molar volume is fixed at SLC.

18. The number of moles will be the same for each of the gases.

19a. Convert units and find *p* using the ideal gas equation.

19b. The databook always contains any conversions that are required in the study design.

20a. The volume of a gas at SLC depends on the number of moles.

20b. The universal gas equation demonstrates the relationship between temperature and pressure.

21. Choose the correct formula to find molar mass.

22. Rearrange the universal gas formula and then substitute in the values.

23. Calculate the moles of methane, use the stoichiometric ratio from the equation and then use the molar volume equation.

24a. Pressure is affected by a change in energy of the particles.

24b. Find the formula that makes the calculation more direct.

24c. Read the instructions from the method very closely.

24d. The definitions for these terms need to be well understood.

25a. Changes in oxidation numbers help determine whether electrons have been lost or gained.

25b. Half-equations must be balanced according to a specific sequence of steps using KOHES.

25c. Half-equations must be balanced according to a specific sequence of steps using KOHES.

25d. A reaction's stoichiometric ratio describes the ratios of reactants and products.

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Chapter 14 review

Multiple choice (10 MARKS)

Question 1 (1 MARK)

Which of the following gases is not considered a potent greenhouse gas?

A. O_2

B. CO_2

C. CH_4

D. H_2O

Question 2 (1 MARK)

Which of the following gases found in the Earth's atmosphere is considered a potent greenhouse gas?

A. N_2

B. He

C. O_3

D. Ar

Question 3 (1 MARK)

Which of the following human activities is known to be a major cause in the increase in the concentration of carbon dioxide in the atmosphere?

A. Growing crops

B. Farming livestock

C. Burning fossil fuels

D. Manufacturing solar panels

Question 4 (1 MARK)

What type of energy is absorbed and re-emitted by greenhouse gases?

A. Kinetic

B. Infrared radiation

C. Motion

D. Gravitational

Question 5 (1 MARK)

Which of the following is not a property of an ideal gas?

A. Ideal gases experience no intermolecular forces.

B. Ideal gas molecules have zero volume individually.

C. Ideal gases exchange no energy when a collision occurs.

D. Ideal gases can only occupy a certain volume.

Question 6 (1 MARK)

Which of the following are the correct units that can be used in the ideal gas equation, $pV = nRT$?

A. Pa, m^3, mol, °C

B. kPa, L, mol, K

C. Pa, L, g, °C

D. kPa, mL, g, K

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Question 7 (1 MARK)

Which of the following formulas show the correct rearrangement of $pV = nRT$ to make the subject of the equation n?

A. $n = \frac{RT}{pV}$

B. $n = \frac{pV}{RT}$

C. $n = \frac{pV}{R}$

D. $n = \frac{pV}{T}$

Question 8 (1 MARK)

A 0.33 mol sample of an ideal gas at 17 °C occupies a volume of 4.2 L. What is the pressure of the gas?

A. 1.3×10^1 kPa **B.** 1.1×10^2 kPa **C.** 1.7×10^2 kPa **D.** 1.9×10^2 kPa

---

Question 9 (1 MARK)

One possible reaction that occurs when trinitrotoluene (TNT), $C_7H_5N_3O_6$, explodes is shown.

$2C_7H_5N_3O_6(s) \rightarrow 2C(s) + 12CO(g) + 5H_2(g) + 3N_2(g)$

When one mol of TNT explodes, the total volume of the gases produced from this reaction, measured at 27.0 °C and 1.00×10^2 kPa, is closest to

A. 0.249 L

B. 22.7 L

C. 249 L

D. 274 L

VCAA 2010 Exam 1 Multiple choice Q5

Question 10 (1 MARK)

A helium balloon is inflated to a volume of 5.65 L and a pressure of 10.33 atm at a temperature of 25.0 °C. The amount of helium, in moles, in the balloon is

A. 0.023

B. 0.276

C. 2.39

D. 27.95

Adapted from VCAA 2012 Exam 2 Multiple choice Q16

Short answer (30 MARKS)

Question 11 (3 MARKS)

Calculate the volume that 5.00 mol of gas will occupy at 26.0 °C and 520 kPa.

Question 12 (3 MARKS)

When hydrochloric acid is added to aluminium sulfide, the highly toxic gas hydrogen sulfide is evolved. The equation for this reaction is

$Al_2S_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2S(g)$

Excess hydrochloric acid was added to 0.200 mol of aluminium sulfide.

a. Calculate the moles of hydrogen sulfide produced. (2 MARKS)

b. What is the volume of hydrogen sulfide produced if this reaction was carried out under standard laboratory conditions? (1 MARK)

Adapted from VCAA 2014 exam Multiple choice Q8

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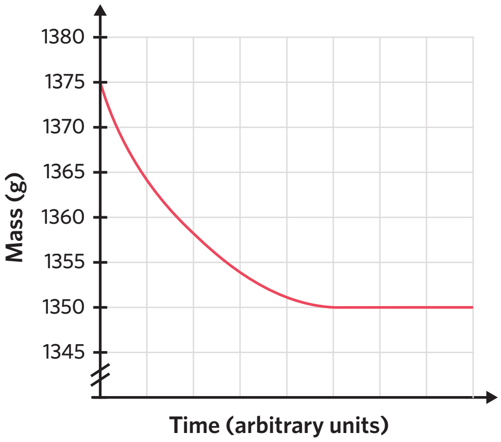
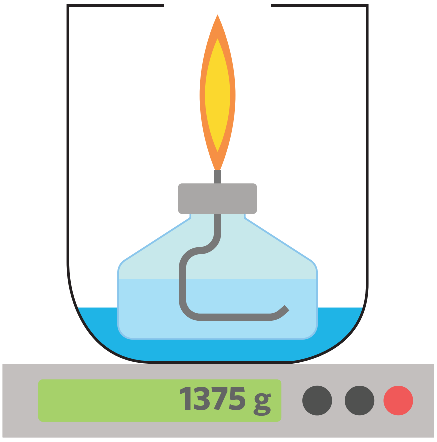
Question 13 (9 MARKS)

Vlad conducted an experiment to determine the changes that occur as a result of the combustion of ethanol. In this experiment, a spirit burner containing ethanol was placed into a cooled and insulated open container. During the course of the experiment, Vlad measured the change in mass of the container. The result can be seen in the graph. The following word equation outlines this complete combustion reaction.

ethanol + oxygen gas → liquid water + carbon dioxide gas





a. What assumption is made regarding a product of combustion, as shown in the diagram? (1 MARK)

b. Identify the cause of the change in mass of the container. (1 MARK)

c. Assuming the experiment was conducted at SLC, calculate the volume of product lost during the combustion reaction. (4 MARKS)

d. The temperature of the room increased to 30 °C. If the pressure was measured at 400 kPa, what amount of carbon dioxide, in moles, would need to be consumed in the reaction to produce the same volume as calculated in part c? (3 MARKS)

Question 14 (6 MARKS)

The emergency oxygen system in a passenger aircraft uses the decomposition of sodium chlorate to produce oxygen. At 76.0 kPa and 292 K, each adult passenger needs about 1.60 L of oxygen per minute. The equation for the reaction is

$2NaClO_3(s) \rightarrow 2NaCl(s) + 3O_2(g)$

$M(NaClO_3) = 106.5 \text{ g mol}^{-1}$

a. Calculate the number of moles of oxygen required per adult per minute. (2 MARKS)

b. Calculate the number of moles of sodium chlorate required to produce the number of moles of oxygen you calculated in part a. (2 MARKS)

c. Calculate the mass of sodium chlorate required to provide the required volume of oxygen for each adult passenger per 10 minutes. (2 MARK)

Adapted from VCAA 2015 exam Multiple choice Q4

Question 15 (9 MARKS)

Airbags are an important safety feature of today's cars. The airbag contains a mixture of solid sodium azide, NaN_3, and potassium nitrate, KNO_3. In the event of an accident, trip sensors send an electric signal to an igniter. The heat generated causes the reactants to decompose completely according to the following equation.

$10NaN_3(s) + 2KNO_3(s) \rightarrow 5Na_2O(s) + K_2O(s) + 16N_2(g)$

A particular car's airbag was found to inflate with a volume of 62.0 L at a pressure of 100 kPa when the temperature reached 36.6 °C. The molar mass of NaN_3 is 65.0 g mol^{-1}.

a. What assumptions are made regarding the conditions that this volume of 62.0 L nitrogen gas is created in? (2 MARKS)

b. Why is it not possible to use a V_m of 24.8 L mol^{-1} in this calculation? (1 MARK)

c. Calculate the number of moles of nitrogen produced. (2 MARKS)

d. Calculate the number of moles of sodium azide consumed. (2 MARKS)

e. Hence, calculate the mass (in grams) of sodium azide contained in the car's airbag. (2 MARKS)

Adapted from VCAA 2011 Exam 1 Multiple choice Q15

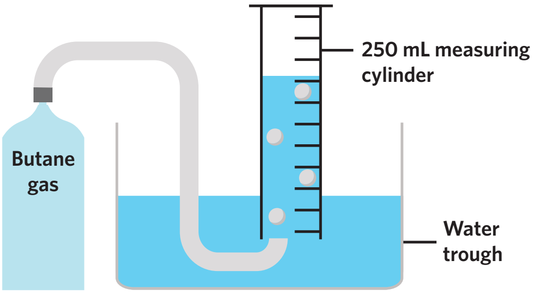
Key science skills (10 MARKS)

Question 16 (10 MARKS)

Medhi was given the task of confirming the molar mass of butane gas through experimentation. The purpose of this experiment is to measure the volume and mass of a sample of butane gas, C_4H_{10}, $M = 58.0$ g mol^{-1}, and to use these values to calculate an experimental value for the molar mass. By filling an inverted measuring cylinder with a known mass and volume of pure butane gas, it is possible to calculate the molar mass of butane. This can be done either by using the ideal gas equation or the assumption that a mole of gas will occupy 24.8 L mol^{-1} at SLC.

Practical requirements:

- 250 mL measuring cylinder
- Glass trough filled with water
- Butane gas with delivery tube
- Balance with an accuracy of at least ± 0.001 g
- Retort stand and clamp



Procedure:

1. Fill the 250 mL measuring cylinder with water and then invert into the trough of water.
2. Clamp the measuring cylinder in place to prevent it from falling over.
3. Weigh a canister of butane camping gas and record in your results table.
4. Add approximately 100 mL of gas to the inverted measuring cylinder.
5. Record the exact volume in your results table.
6. Remove the delivery tube from the canister and re-weigh, record this value in your results table.

Mass of butane gas container before releasing gas	Mass of butane gas container after releasing gas	Temperature of water	Volume of water in measuring cylinder before butane added	Volume of water in measuring cylinder after butane added
16.500 g	16.300 g	17.1 °C	250 mL	162 mL

Pressure of the butane gas (found using Dalton's Law of Partial Pressure): 95.0 kPa.

Volume of gas: 250 mL − 162 mL = 88.0 mL or 0.0880 L

Temperature of the water: (17.1 + 273) K = 290.1 K

Mass of butane in cylinder: 0.200 g

a. Using the data obtained from this experiment, calculate the molar mass from

   i. the molar volume equation, $n = \frac{V}{V_m}$. (3 MARKS)

   ii. the ideal gas equation, $pV = nRT$. (2 MARKS)

b. What is the percentage difference between the molar mass calculated using the ideal gas equation and the molar gas volume equation? (2 MARKS)

c. Explain why this percentage difference exists. (1 MARK)

d. In step 1, why is it important to not allow any air into the measuring cylinder? (1 MARK)

e. What safety precaution must be taken with the butane gas when the experiment is finished? (1 MARK)

FROM LESSONS 16A, 16C, 16D & 16E

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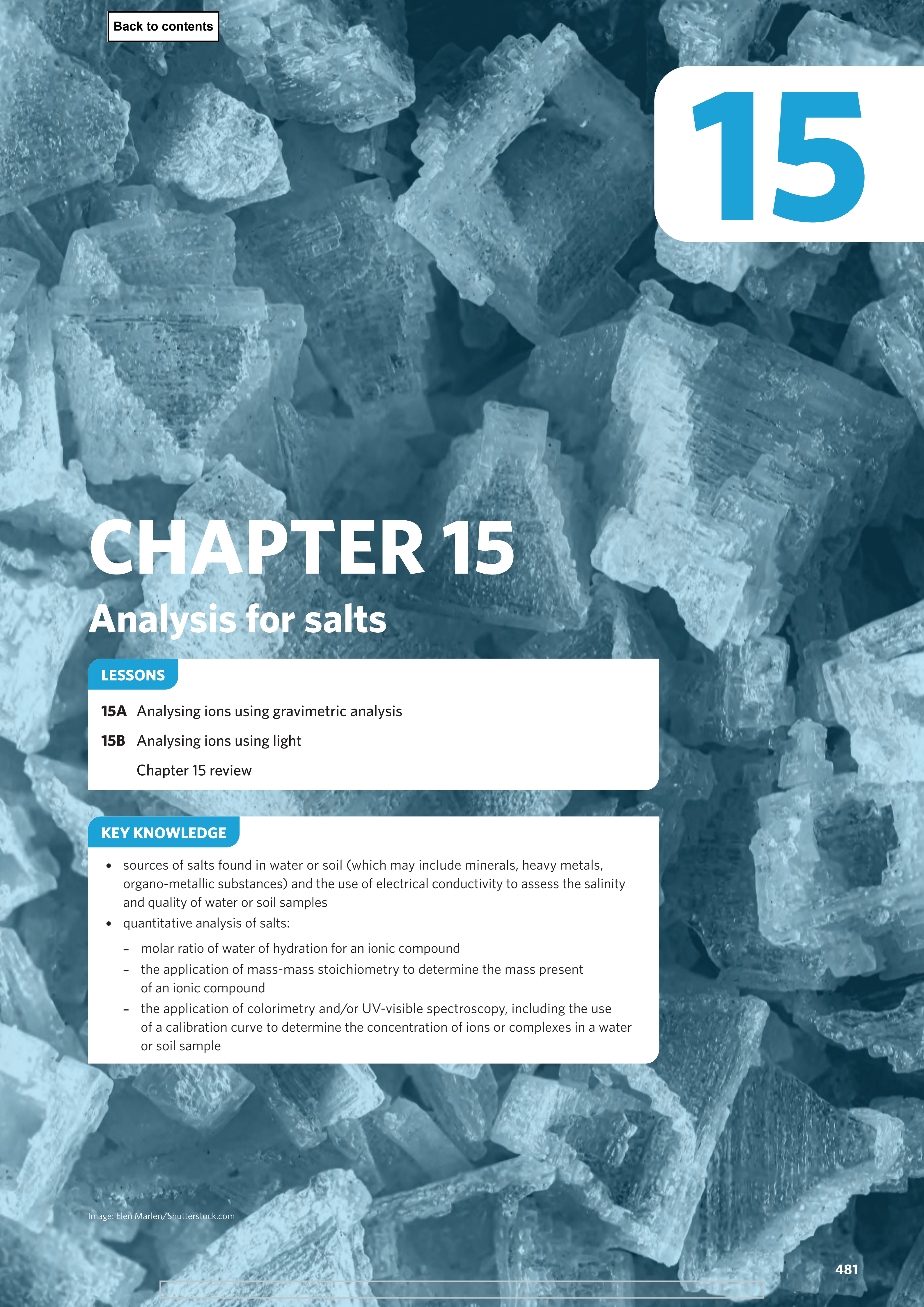

Image: Elen Marlen/Shutterstock.com

15

CHAPTER 15

Analysis for salts

LESSONS

KEY KNOWLEDGE

- sources of salts found in water or soil (which may include minerals, heavy metals, organo-metallic substances) and the use of electrical conductivity to assess the salinity and quality of water or soil samples
- quantitative analysis of salts:
  - molar ratio of water of hydration for an ionic compound
  - the application of mass-mass stoichiometry to determine the mass present of an ionic compound
  - the application of colorimetry and/or UV-visible spectroscopy, including the use of a calibration curve to determine the concentration of ions or complexes in a water or soil sample

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15A Analysing ions using gravimetric analysis

STUDY DESIGN DOT POINTS

- sources of salts found in water or soil (which may include minerals, heavy metals, organo-metallic substances) and the use of electrical conductivity to assess the salinity and quality of water or soil samples
- quantitative analysis of salts:
  - molar ratio of water of hydration for an ionic compound
  - the application of mass-mass stoichiometry to determine the mass present of an ionic compound
  - the application of colorimetry and/or UV-visible spectroscopy, including the use of a calibration curve to determine the concentration of ions or complexes in a water or soil sample

15A 15B

2.2.10.1 Determination of salinity

2.2.11.1 Hydrated compounds

2.2.11.2 Principles of gravimetric analysis

2.2.11.3 Calculations involving gravimetric analysis: mass to mass stoichiometry

ESSENTIAL PRIOR KNOWLEDGE

4C Precipitation
6A The mole
6B Empirical and molecular formula
13A Concentration
See questions 89–92.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

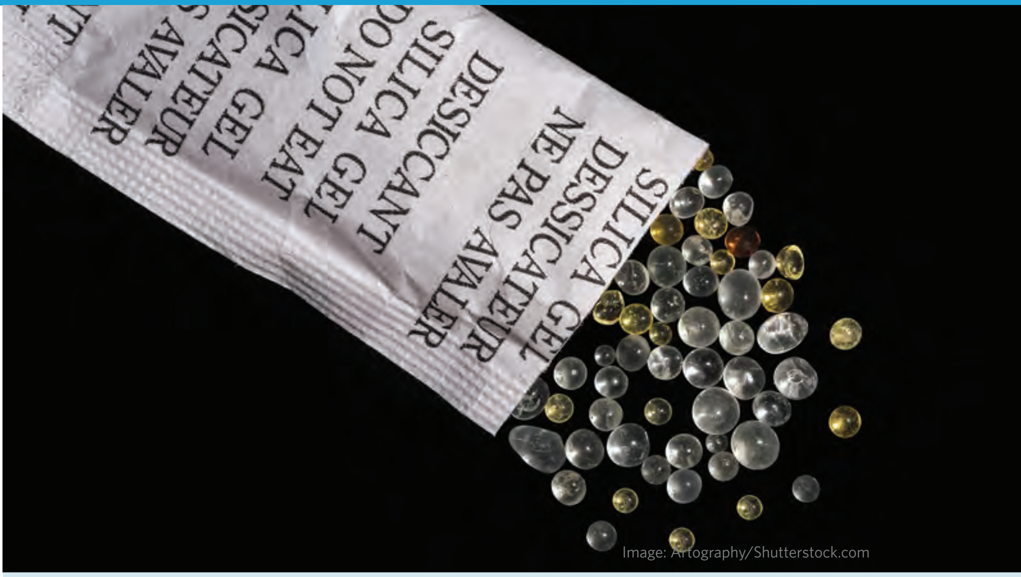

Image: Artography/Shutterstock.com

Why is silica gel found in food, electronics, and clothing packaging?

It is possible for crystals of sodium carbonate to have 51% of its total mass be made up of water molecules and not be aqueous. In this lesson, we will be learning about the **water of crystallisation** and the use of **gravimetric analysis** and salinity testing in the chemical analysis of ions in both water and soil samples.

KEY TERMS AND DEFINITIONS

Analyte	substance which is quantified or identified
Anhydrous	without water
Gravimetric analysis	chemical analysis involving the weighing of a precipitate formed from an analyte
Hydrated salt	contains water of crystallisation molecules
Hygroscopic	able to absorb or adsorb water from its surroundings
Law of conservation of mass	law that states that in a closed system, the total mass of reactants equals the total mass of products
Mass-mass stoichiometry	calculation of reactants and products in a chemical reaction using the law of conservation of mass
Stoichiometry	the relationship between relative quantities of reactants and products in a chemical reaction
Water of crystallisation	water that is bonded into a crystal structure

Determination of salinity 2.2.10.1

Water and soil analysis testing for salinity must be carried out to ensure appropriate usage for the water or the land.

Why must the salinity of water and soil be monitored?

As discussed in lesson 9A, only water that contains very small concentrations of salt ions is safe for human consumption. Water and soils containing high concentrations of salt ions can be very toxic to plants, animals and humans.

Salinity testing involves determining the concentration of dissolved ions in a water or soil sample, which is done by measuring the electrical conductivity of a sample of water. When measuring the salinity of soil, the soil must first be dissolved in water and then filtered before testing the soil-water solution (figure 1). This is an effective means of analysis as pure water is a poor conductor of electricity, while saline water with large quantities of dissolved ions will have a higher electrical conductivity. The concentration of ions present in a water sample is indicated by how well it conducts electricity, as samples which are more electrically conductive must have a higher concentration of salt ions.

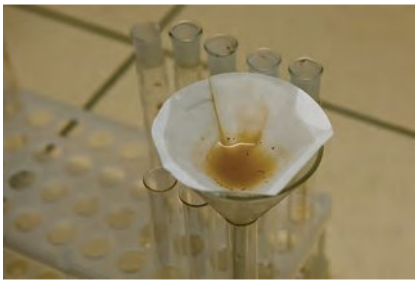

Image: Leo_nik/Shutterstock.com

Figure 1 Simple filtration of a soil sample dissolved in water

The test to measure electrical conductivity requires many parts, as shown in figure 2, and they include:

- two electrodes
- a battery
- a lightbulb
- an ammeter

As shown in figure 2, electricity needs to be able to flow through the water from one electrode to the other in order to complete the circuit and thus power the lightbulb. The more electricity that is able to pass through the lightbulb, the brighter the light will appear. Therefore, the intensity of the light produced can be used as a measure of the salinity of the water. In order to obtain a more accurate measurement, an ammeter is typically used in the circuit to measure the amount of current passing through.

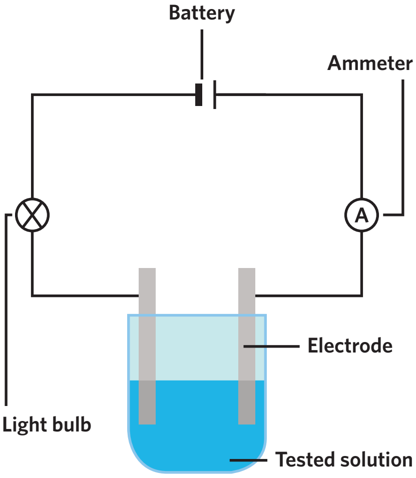

Figure 2 Solution electrical conductivity test

As seen in figure 3, the conductivity of the water sample increases linearly as the concentration of ions increases. Conductivity can be measured using light intensity or ammeter readings and its unit of measurement is microsiemens per centimetre (μS cm^{-1}).

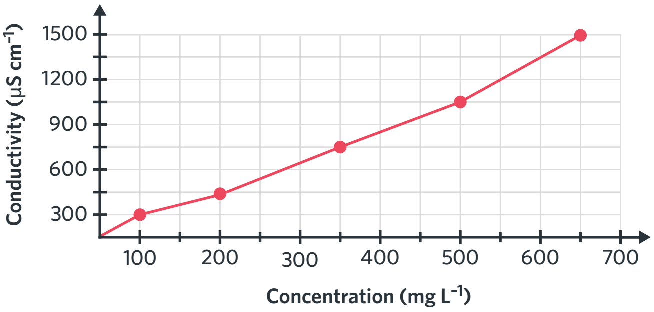

Figure 3 Relationship between the concentration and conductivity of ions in a water sample

It should also be noted that the conductivity of a solution increases as the temperature increases, so it is standard practice to use electrical conductivity readings at 25 °C to ensure that the results produced are consistent. In addition, conductivity readings simply provide an indication of the concentration of ions found in the solution, and do not provide any qualitative information about the types of ions present.

Therefore, although it does not provide a comprehensive analysis of the types of ions found in a water sample, electrical conductivity testing can provide researchers with an indication of the salt concentration in a water or soil sample. This enables researchers to determine whether a water source or soil sample is safe for a particular usage, like as drinking water or for growing a particular crop (figure 4).

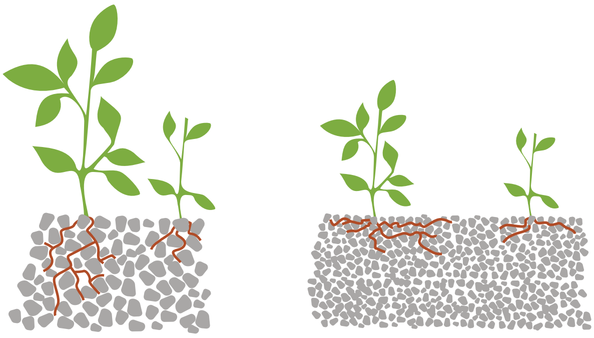

Figure 4 Increased salt concentration inhibits plant growth.

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Image: Sitthipong Pengjan/Shutterstock.com

Figure 5 Portable digital salinity meter

Salinity metres are handheld devices for assessing water salinity outdoors, as shown in figure 5. These are often used by the agricultural industry and water authorities to obtain an estimate of a water source's salinity. However, as conditions are not as controlled outside of a laboratory, salinity meters must be calibrated before use.

Progress questions

Question 1

Electrical conductivity testing provides information about

A. the concentration of ions in water.

B. the amount of precipitate in water.

Question 2

Water samples with a high salt concentration have

A. higher electrical conductivity.

B. lower electrical conductivity.

Question 3

Lightbulbs receiving more electrical current will

A. shine for a longer duration of time.

B. have a higher light intensity.

Hydrated compounds 2.2.11.1

Many salts are **hygroscopic** and have water molecules within their chemical formula.

What is water of crystallisation?

Water of crystallisation refers to molecules of water bonded into crystals[1] of a salt, which have been allowed to crystallise (figure 6) out of water. As a result, the **hydrated salt** of the compound is formed.

KEEN TO INVESTIGATE?

[1] **How does silica gel work?**
Search YouTube: how does silica gel work

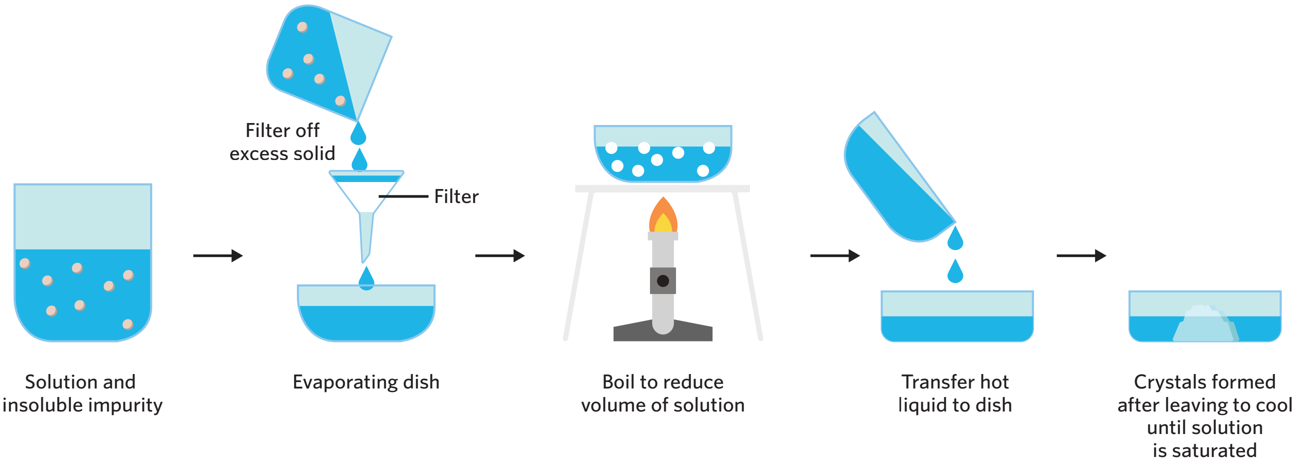

Figure 6 The crystallisation process

- If copper (II) sulfate crystals are formed in the absence of water, the salt is considered **anhydrous**. It has a white colour and its formula is $CuSO_4(s)$, $M = 159.6 \text{ g mol}^{-1}$.
- If copper (II) sulfate crystals are formed in the presence of water, the salt is considered hydrated and it has a blue colour with the formula $CuSO_4 \bullet 5H_2O(s)$, $M = 249.7 \text{ g mol}^{-1}$. Each copper sulfate can be bonded to five water molecules and is therefore called copper (II) sulfate pentahydrate.

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The dot (•) in hydrated salts is used to indicate that water of crystallisation molecules are bonded to the salt. This water of crystallisation can be removed (as this is a reversible process) by heating the hydrated salt, making it anhydrous (figure 7).

$$CuSO_4\overset{\downarrow}{\bullet}5H_2O(s)$$

Images (left to right): RHJPhtotos/krolya25/Shutterstock.com

Figure 7 Anhydrous **(a)** and hydrated **(b)** copper (II) sulfate have different coloured crystals.

USEFUL TIP

Most hydrated ionic compounds can have more than one stoichiometric amount of moles of water bonded to it. For example, copper (II) sulfate can have either five ($CuSO_4$•$5H_2O$) or seven ($CuSO_4$•$7H_2O$) moles of water per mole of copper (II) sulfate.

MISCONCEPTION

'Water of crystallisation does not count towards the molar mass of a hydrated compound.'

All the mass of the water molecules is included in the overall molar mass of hydrated compounds.

How do we perform mole ratio calculations in hydrated salts?

In lesson 6B, calculations involving empirical and molecular formulas were introduced. The strategy of using mole ratios to determine empirical formulas is the same for working out the mole ratio of water in hydrated salts.

WORKED EXAMPLE 1

12.50 g of hydrated copper (II) sulfate, otherwise known as $CuSO_4$•xH_2O (x is an unknown mole ratio), was heated to a mass of 8.00 g of anhydrous copper (II) sulfate. What is the value of x and the molecular formula of the hydrated salt? ($CuSO_4(s)$, $M = 159.6\ \text{g mol}^{-1}$.)

What information is presented in the question?

The molecular formula and initial/final mass.

The mass of hydrated salt = 12.50 g

The mass of anhydrous salt after heating = 8.00 g

What is the question asking us to do?

Calculate the mole ratio of water molecules in 1 mole of the compound and give the molecular formula.

What strategies do we need in order to answer the question?

1. Calculate the change in mass.
2. Find the ratio of mass of compound to mass of water in the compound.
3. Divide these masses by the relevant relative molar mass.
4. Calculate the molar ratio.

Answer

Mass of water in the salt 12.50 g – 8.00 g = 4.50 g

The mass ratio of $CuSO_4 : H_2O$

is $8.00 : 4.50$

The mole ratio of $CuSO_4 : H_2O$ using $n = \frac{m}{M}$

The mole ratio of $CuSO_4 : H_2O$

is $\frac{8.00\ \text{g}}{159.6\ \text{g mol}^{-1}} : \frac{4.50\ \text{g}}{18.0\ \text{g mol}^{-1}}$

The mole ratio of $CuSO_4 : H_2O$

is $5.01 \times 10^{-2}\ \text{mol} : 2.5 \times 10^{-1}\ \text{mol}$

Mole ratio of $H_2O : CuSO = \frac{2.5 \times 10^{-1}\ \text{mol}}{5.01 \times 10^{-2}\ \text{mol}}$

Mole ratio of $H_2O : CuSO = 5$

$x = 5$

$CuSO_4$•$5H_2O(s)$

Progress questions

Question 4

Which of the following refers to a salt crystal that contains no water of crystallisation?

A. Anhydrous

B. Hydrated

Continues →

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Question 5

Water of crystallisation molecules can be removed by heating a hydrated salt.

A. True

B. False

Question 6

The ___________ of water molecules to the salt compound are used to calculate the formula of the hydrated salt crystals.

A. mole ratios

B. mass ratios

Principles of gravimetric analysis 2.2.11.2

Gravimetric analysis is an analytical technique used to determine the quantity of a substance in a water sample.

What are the principles of gravimetric analysis?

Another method of determining the concentration of salt in a sample of water is through gravimetric analysis. In gravimetric analysis, salt ions are precipitated out of a solution to be weighed, allowing the quantity of the substance in a sample to be determined. The primary advantage of gravimetric analysis, as opposed to electrical conductivity testing, is that it can be performed to specifically measure the amount of a certain ion in a solution, rather than measuring the total ion concentration in the water sample.

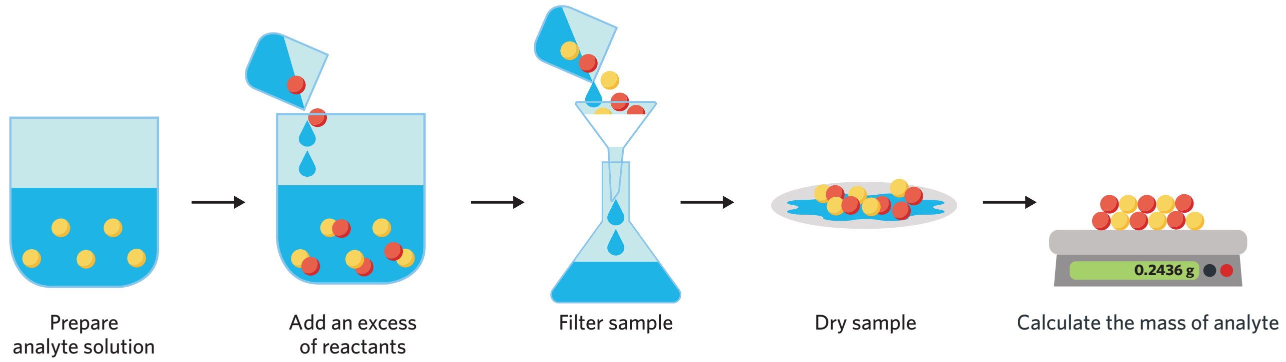
In gravimetric analysis, the substance whose quantity is being measured is referred to as the **analyte**. Depending on the identity of the analyte, a reactant is selected that will specifically react with the analyte. The amount of an analyte in a solution can be determined by weighing the precipitate formed by the analyte and the reactant and then determining the percentage of the mass of the precipitate compared to the original sample.

USEFUL TIP

If other ions, that are not being measured, are also present in the water sample, the reactant must be chosen so that it doesn't form a precipitate with them. This is done as inaccuracies in measuring the amount of analyte in the water sample may otherwise occur.

When adding a reactant to the solution to form a precipitate, it is important to ensure that the reactant is added in excess. Since we are trying to determine the mass of the analyte in a water sample, we want to make sure that all of the analyte will be found in the precipitate. By adding the reactant in excess, this ensures that there will be a sufficient amount of the reactant present for all of the analyte to react with it. The leftover reactant will remain dissolved in the solution, without contributing to the mass of the precipitate.

Figure 8 is a summary of the method involved in gravimetric analysis.



1. Prepare a sample with the analyte in solution.

2. Add an excess amount of a reactant to the water sample to react with the analyte.

3. Separate and collect the precipitate from the solution.

4. Wash the precipitate with distilled water to remove any impurities.

5. Filter the sample to remove water.

6. Dry the sample to ensure that any excess water is removed. This is generally done through multiple rounds of heating, cooling and drying of the sample until a constant mass is achieved.

7. Calculate the percentage of the analyte using the mass of the precipitate and mass-mass stoichiometry.

Figure 8 Gravimetric analysis technique

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The precipitate should remain stable while being heated, as decomposition of the precipitate during heating will mean that the amount of analyte in the water sample cannot be accurately measured.

The formation of a precipitate from the analyte and reactant is important for isolating and measuring the amount of analyte in a substance which may be either a solid or in a solution (aqueous):

- If the analyte is found as a solid by itself, then the process of gravimetric analysis is unnecessary since the sample of pure analyte can be weighed by itself to determine how much of it is present.
- If the sample that the analyte is from contains multiple substances, then it will be impossible to determine the mass of a single analyte from the substance. However, by dissolving all the contents of the sample in a solvent and selecting a reactant that only reacts with the analyte of interest, a precipitate containing only the analyte and the reactant will be formed. This will then allow for a quantitative analysis of the analyte to be performed.
- If you have a sample of analyte already dissolved in a solvent, it is important to precipitate the analyte out so that it can be weighed without the solvent.

Gravimetric analysis serves a number of important functions for chemists. Given it is relatively inexpensive to find an ion to form a precipitate and measure the mass of the precipitate, gravimetric analysis can be carried out rapidly and easily. This has led to it being used in the analysis of water impurities, foods, and other industrial applications.

Progress questions

Question 7

What is gravimetric analysis?

A. A chemical analysis technique involving interactions of ions with light

B. A chemical analysis technique involving precipitation of ions out of solution to determine their quantity

Question 8

In gravimetric analysis, a poorly dried precipitate will result in a higher mass.

A. True

B. False

Question 9

A precipitate is formed in gravimetric analysis through

A. the introduction of a reactant to an analyte in solution.

B. supersaturating a solution to form crystals.

Question 10

What type of data does gravimetric analysis provide?

A. Qualitative

B. Quantitative

Calculations involving gravimetric analysis: mass to mass stoichiometry 2.2.11.3

Mass-mass stoichiometry utilises the knowledge of ratios of ions to determine the mass of an analyte in gravimetric analysis.

How do we use mass-mass stoichiometry in gravimetric calculations?

Having produced a precipitate from the sample solution, the next step is to use mass-mass **stoichiometry** to determine the amount of analyte present in the precipitate. This is necessary since the precipitate consists of both the analyte and the reactant that was added, which cannot be individually weighed. In order to do so, we first need to revise our understanding of the concepts of moles.

As discussed in chapter 6, the mole is a unit of measurement for the amount of a substance. By knowing the mass and molar mass of a substance, the number of moles can be determined. Conversely, by knowing the number of moles of a substance, one can determine the mass of the substance present. These conversions between units can be carried out using the following formula:

$$\text{Number of moles} = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1}\text{)}}$$

Another important concept in mass-mass stoichiometry is mole ratios. Take the equation below as an example.

$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

The coefficients (the numbers in front of each molecule) indicate how many of that molecule will react each time the reaction occurs. In the example given, each time the reaction takes place, one $H_2SO_4(aq)$ molecule will react with two NaOH(aq) molecules. More importantly, the coefficients also indicate the mole ratios of reactants and products in a reaction. This concept of using mole ratios to determine quantitative data about any of the species involved in a reaction is referred to as stoichiometry.

USEFUL TIP

When comparing substances, use the mole ratio as the point of comparison.

For the reaction between $H_2SO_4(aq)$ and NaOH(aq), the mole ratio between the two reactants, $n(H_2SO_4) : n(NaOH)$ is 1 : 2. This illustrates that for every mole of $H_2SO_4(aq)$ used in the reaction, two moles of NaOH(aq) will be consumed. Mole ratios can be used to predict the amount of substances in reactions as long as the amount of one molecule is known, referred to as the 'known chemical'.

Another important concept to understand in stoichiometric calculations is that the total mass in the equation does not change, meaning that the total mass of the reactants will be equal to the total mass of the products. This is referred to as the **law of conservation of mass**.

Mass-mass stoichiometry is a method of calculating the mass of reactants and products in a reaction using stoichiometry as described above. Let's apply this to gravimetric analysis by introducing a reaction that we may see in gravimetry as shown:

$Pb(NO_3)_2(aq) + 2KCl(aq) \rightarrow PbCl_2(s) + 2KNO_3(aq)$

As we can see, the reaction between the two solutions, lead nitrate ($Pb(NO_3)_2(aq)$) and potassium chloride (KCl(aq)), results in the formation of a lead chloride precipitate ($PbCl_2(s)$). If we compare the mole ratios between the reactants and precipitate, we will identify different relationships, as seen in table 1.

Table 1 Breakdown of the mole ratios between the reactants and products

Reference substance	Substance being compared to	Mole ratio
$Pb(NO_3)_2(aq)$	$PbCl_2(s)$	$n(Pb(NO_3)_2 : n(PbCl_2)$ 1 : 1
KCl(aq)	$PbCl_2(s)$	$n(KCl) : n(PbCl_2)$ 2 : 1

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We can use this information to perform appropriate calculations from our reaction involving lead nitrate $Pb(NO_3)_2(aq)$ and $KCl(aq)$. Let's assume that 127.0 g of precipitate was formed. If we were required to calculate the mass of $KCl(aq)$ needed to produce this mass of $PbCl_2(s)$ precipitate, we would use the steps shown in figure 9.

1	Calculate the amount of the known substance.	**1**	$n(PbCl_2) = \frac{m}{M}$ $n(PbCl_2) = \frac{127.0 \text{ g}}{(207.2 + (35.5 \times 2)) \text{ g mol}^{-1}}$ $n(PbCl_2) = 0.4565 \text{ mol}$
2	Use the balanced equation to find the mole ratio between the known and unknown substance.	**2**	$n(PbCl_2) : n(KCl)$ 1 : 2 0.4565 : ?
3	Calculate the amount of the unknown.	**3**	$n(KCl) = \frac{2}{1} \times 0.4565 \text{ mol}$ $n(KCl) = 0.9130 \text{ mol}$
4	Calculate the mass of the unknown.	**4**	$m(KCl) = n(KCl) \times M(KCl)$ $m(KCl) = 0.9130 \text{ mol} \times (39.1 + 35.5) \text{ g mol}^{-1}$ $m(KCl) = 68.1 \text{ g}$

Figure 9 Application of mass-mass stoichiometry to calculate an unknown mass

WORKED EXAMPLE 2

A reaction occurred between $MgCl_2$ and Na_2CO_3 forming a precipitate as shown:

$$MgCl_2(aq) + Na_2CO_3(aq) \rightarrow MgCO_3(s) + 2NaCl(aq)$$

If 21.1 g of $MgCl_2$ reacts with an excess of Na_2CO_3, determine the mass of $MgCO_3$ that would be formed.

What information is presented in the question?

The mass of $MgCl_2$ and the identity of the reactants and products.

What is the question asking us to do?

Determine the mass of $MgCO_3(s)$ precipitate formed.

What strategies do we need in order to answer the question?

1. Determine the amount of $MgCl_2$ reacted.
2. Compare the mol ratio between $MgCl_2$ and $MgCO_3$.
3. Determine the number of moles of $MgCO_3$ that would be formed.
4. Determine the mass of $MgCO_3$ using the molar mass and number of moles.

Answer

$m(MgCl_2) = 21.1 \text{ g}$

$M(MgCl_2) = (24.3 + (2 \times 35.5)) \text{ g mol}^{-1}$

$M(MgCl_2) = 95.3 \text{ g mol}^{-1}$

$n(MgCl_2) = \frac{m}{M}$

$n(MgCl_2) = \frac{21.1 \text{ g}}{95.3 \text{ g mol}^{-1}}$

$n(MgCl_2) = 0.2214 \text{ mol}$

$n(MgCl_2) : n(MgCO_3)$

1 : 1

$n(MgCO_3) = 0.2214 \text{ mol}$

$M(MgCO_3) = (24.3+ 12.0 + (3 \times 16.0)) \text{ g mol}^{-1}$

$M(MgCO_3) = 84.3 \text{ g mol}^{-1}$

$m(MgCO_3) = n \times M$

$m(MgCO_3) = 0.221 \text{ mol} \times 84.3 \text{ g mol}^{-1}$

$m(MgCO_3) = 18.7 \text{ g}$

By using the amount of a known substance, we can use stoichiometry to determine the amount of a different substance in the reaction, and use that information to find the mass of the substance.

USEFUL TIP

It is important that all precipitates and hydrated salts are dried to a constant mass otherwise overestimations will be made due to the mass of water remaining.

How can mass-mass stoichiometry be used to calculate the percentage composition of substances?

We can also use our understanding of the general concepts of mass-mass stoichiometry to find the composition or concentrations of substances.

Continuing with our example shown in figure 9, suppose we were told that the $PbCl_2(s)$ precipitate was collected from an unknown substance that had a total mass of 219.6 g. We can calculate the percentage mass of chlorine found in the original sample using the steps shown in figure 10.

5	Determine the molar ratio between the known substance and the unknown substance.	5	$n(\mathrm{KCl}) : n(\mathrm{Cl})$ $1 : 1$
6	Determine the amount, in mol, of the unknown substance.	6	$n(\mathrm{KCl}) = n(\mathrm{Cl})$ $n(\mathrm{Cl}) = 0.9130$ mol (as calculated in figure 9)
7	Determine the mass of the substance we're looking for.	7	$m(\mathrm{Cl}) = n(\mathrm{Cl}) \times M(\mathrm{Cl})$ $m(\mathrm{Cl}) = 0.9130\ \mathrm{mol} \times 35.5\ \mathrm{g\ mol^{-1}}$ $m(\mathrm{Cl}) = 32.4\ \mathrm{g}$
8	Calculate the percentage.	8	$\%(\mathrm{Cl}) = \frac{32.4\ \mathrm{g}}{219.6\ \mathrm{g}} \times 100$ $\%(\mathrm{Cl}) = 14.8\%$

Figure 10 Determining the percentage of a substance in a sample using mass – mass stoichiometry

The assumption in this scenario is that all the chlorine atoms present existed in the form of KCl(aq). This will generally be the case unless we are told otherwise.

By using this method, we can quantitatively determine the presence of a particular substance in a sample.

WORKED EXAMPLE 3

A 3.00 g sample of limestone ($CaCO_3$) was dissolved, where the calcium ions were precipitated using excess ammonium oxalate solution $(NH_4)_2C_2O_4(aq)$. The precipitate formed was calcium oxalate, $CaC_2O_4(s)$ according to the equation:

$$Ca^{2+}(aq) + C_2O_4^{2-}(aq) \rightarrow CaC_2O_4(s)$$

Determine the percentage, by mass, of calcium present in limestone if 3.65 g of precipitate was formed.

What information is presented in the question?

The precipitation reaction occuring, the mass of limestone (3.00 g) and the mass of precipitate formed (3.65 g).

What is the question asking us to do?

Determine the percentage, by mass, of calcium present in limestone.

What strategies do we need in order to answer the question?

1. Calculate the amount of precipitate formed.
2. Compare the molar ratio between the precipitate and calcium.
3. Determine the amount of calcium present.
4. Calculate the mass of calcium present.
5. Calculate the percentage by mass of calcium in the limestone sample.

Answer

$$n(CaC_2O_4) = \frac{3.65\ \mathrm{g}}{(40.1 + (2 \times 12.0) + (4 \times 16.0))\ \mathrm{g\ mol^{-1}}}$$

$$n(CaC_2O_4) = 0.02849\ \mathrm{mol}$$

$$n(CaC_2O_4) : n(Ca^{2+})$$

$$1 : 1$$

$$n(Ca^{2+}) = 0.02849\ \mathrm{mol}$$

$$m(Ca^{2+}) = n \times M$$

$$m(Ca^{2+}) = 0.02849\ \mathrm{mol} \times 40.1\ \mathrm{g\ mol^{-1}}$$

$$m(Ca^{2+}) = 1.14\ \mathrm{g}$$

$$\%(Ca^{2+}) = \frac{1.14\ \mathrm{g}}{3.00\ \mathrm{g}} \times 100$$

$$\%(Ca^{2+}) = 38.1\%$$

Therefore, the limestone sample contains 38.1% calcium.

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How can mass-mass stoichiometry be used to calculate mass in reactions involving excess and limiting reactants?

With a deeper level of understanding of chemical reactions, we now know that chemical reactions occur as per a given ratio. Therefore, not all of the substance that we add into a reaction will react. Sometimes, we have more of one reactant than the other, resulting in an excess reactant and a limiting reactant. The limiting reactant is completely consumed in the reaction and is the limiting factor to how much product we can make, whereas the excess reactant will be left over after the reaction is considered to be complete.

USEFUL TIP

To find the limiting and excess reactant, we need to compare the amount, in mol, of both reactants. The reactant that is always 'left over' is the excess reactant.

WORKED EXAMPLE 4

A student is carrying out a reaction between sulfuric acid and sodium hydroxide, which has the following equation: $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

If there is 90.0 g of H_2SO_4 and 70.0 g of NaOH present in the sample, determine which is the limiting reactant and which is the excess reactant.

What information is presented in the question?

The chemical equation and the masses of reactants.

What is the question asking us to do?

Determine the limiting and excess reactant in the reaction.

What strategies do we need in order to answer the question?

1. Determine the mole ratio of the reactants.
2. Determine the number of moles of each reactant.
3. Compare the number of moles using the mole ratio.
4. Determine which reactant is in excess and which is limiting.

Answer

$n(H_2SO_4) : n(NaOH)$

$1 \quad : \quad 2$

$n(H_2SO_4) = \frac{m}{M}$

$n(H_2SO_4) = \frac{90.0\ \text{g}}{((2 \times 1.0) + 32.1 + (4 \times 16.0))\ \text{g mol}^{-1}}$

$n(H_2SO_4) = 0.917\ \text{mol}$

$n(NaOH) = \frac{m}{M}$

$n(NaOH) = \frac{70.0\ \text{g}}{(23.0 + 16.0 + 1.0)\ \text{g mol}^{-1}}$

$n(NaOH) = 1.75\ \text{mol}$

If we use H_2SO_4 as a reference point:

$n(H_2SO_4) : n(NaOH)$

$1 \quad : \quad 2$

$0.917\ \text{mol} : x$

$n(NaOH) = \frac{2}{1} \times 0.917\ \text{mol}$

$n(NaOH) = 1.83\ \text{mol}$

To use all 0.917 mol of H_2SO_4, a total of 1.83 mol of NaOH is required. Considering there is only 1.75 mol of NaOH available, NaOH is the limiting reactant and H_2SO_4 is the excess reagent.

When there is a limiting reactant, we must use the limiting reactant as our point of reference to calculate the amount of all other substances in the reaction.

WORKED EXAMPLE 5

A 5.23 g sample of $AgNO_3$ and 1.62 g sample of NaCl was used in the following reaction:

$AgNO_3(aq) + NaCl(aq) \rightarrow NaNO_3(aq) + AgCl(s)$

Determine the mass of AgCl precipitate that would be formed.

What information is presented in the question?

The chemical equation and the masses of reactants.

What is the question asking us to do?

Determine the mass of precipitate that would be formed.

What strategies do we need in order to answer the question?

1. Determine the amount of each reactant present.
2. Determine the limiting reactant.
3. Apply stoichiometry by using the limiting reactant to find the mass of precipitate formed.

Continues →

Answer

$n(AgNO_3) = \frac{m}{M}$

$n(AgNO_3) = \frac{5.23\text{ g}}{(107.9 + 14.0 + (3 \times 16.0))\text{ g mol}^{-1}}$

$n(AgNO_3) = 0.03078\text{ mol}$

$n(NaCl) = \frac{1.62\text{ g}}{(23.0 + 35.5)\text{ g mol}^{-1}}$

$n(NaCl) = 0.02769\text{ mol}$

If we use $AgNO_3$ as a reference point to determine the limiting reactant:

$n(AgNO_3)$	:	$n(NaCl)$
1	:	1
0.03078 mol	:	?

$n(NaCl) = \frac{1}{1} \times 0.03078\text{ mol}$

Therefore to use 0.03078 mol of $AgNO_3$, we would require 0.03078 mol of NaCl. However, we only have a total of 0.02769 mol of NaCl, which is not enough. NaCl is our limiting reactant and will be our point of reference to calculate the mass of precipitate formed.

$n(NaCl)$	:	$n(AgCl)$
1	:	1
0.02769 mol	:	?

$n(AgCl) = 0.02769\text{ mol}$

$m(AgCl) = n \times M$

$m(AgCl) = 0.02769\text{ mol} \times (107.9 + 35.5)\text{ g mol}^{-1}$

$m(AgCl) = 3.97\text{ g}$

In this reaction, 3.97 g of AgCl precipitate will be formed.

Progress questions

Question 11

What is the molar ratio of NH_3:O_2 in the following reaction:

$4NH_3(g) + 3O_2(g) \rightarrow 2N_2(g) + 6H_2O(l)$

A. 2:6

B. 4:3

Question 12

If five mol of H_2 reacts with three mol of O_2 in the following reaction,

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

which is the limiting reactant?

A. H_2

B. O_2

Question 13

If the molar ratio of ion A:ion B is 2:1, then the mass of ion A is

A. double the mass of ion B.

B. unable to be determined with the information provided.

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Theory summary

- Salinity testing uses electrical conductivity principles of pure water and ions to determine the salinity of a water sample.
- Many salts contain water of crystallisation and the amount can be calculated using mole ratios.
- Gravimetric analysis involves the formation and weighing of a precipitate to determine the amount of an analyte.
- Mass-mass stoichiometry uses the stoichiometric ratios in an equation to determine the mass or percentage composition of other substances.
- Limiting reactants will completely react and determine the extent of a reaction, whilst excess reactants are left over at the end of the reaction.
- When calculating amounts of substance, always use the limiting reactant as the point of reference.

The content in this lesson is considered fundamental prior knowledge to measuring changes in chemical reactions (Unit 3 AOS 1).

15A Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 14–16.

Joseph is a chemistry professor assessing the chemical composition of a sample of seawater. He decides to use either electrical conductivity testing or gravimetric analysis as a means of measuring the concentration of sodium in the sample of seawater.

Question 14 (1 MARK) 🌶

Which option correctly describes the method used for gravimetric analysis?

A. Measuring the amount of electricity flowing through a circuit connected to a water sample

B. Inserting a sample into a flame and observing the colour produced to determine the chemical composition

C. Determining the amount of a substance dissolved in water through weighing a precipitate formed with it

D. Weighing a sample of water and subtracting the mass of water to determine the concentration of salts

Question 15 (1 MARK) 🌶

Which option correctly describes the method used for electrical conductivity testing?

A. Measuring the amount of electricity flowing through a circuit connected to a water sample

B. Inserting a sample into a flame and observing the colour produced to determine the chemical composition

C. Determining the amount of a substance dissolved in water through weighing a precipitate formed with it

D. Weighing a sample of water and subtracting the mass of water to determine the concentration of salts

Question 16 (4 MARKS) 🌶🌶

Suggest some advantages and disadvantages of using gravimetric analysis and electrical conductivity testing. Justify your reasoning.

Exam-style

Question 17 (4 MARKS) 🌶🌶

10.00 g of hydrated sodium carbonate $Na_2CO_3 \cdot xH_2O$ was heated to a constant mass and after re-weighing the sample it had a new mass of 3.70 g. Calculate the value of x and give the formula of the hydrated salt.

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Question 18 (16 MARKS)

Sometimes, not all reactants are used up in a chemical reaction.

a. For each of the following reactions, determine the limiting reactant.

- **i.** $MgSO_4(aq) + Ca(NO_3)_2(aq) \rightarrow CaSO_4(s) + Mg(NO_3)_2(aq)$
  $n(MgSO_4(aq)) = 0.332$ mol and $n(Ca(NO_3)_2(aq)) = 1.98$ mol (2 MARKS)
- **ii.** $2AgNO_3(aq) + CaCl_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2AgCl(s)$
  $n(AgNO_3(aq)) = 3.71$ mol and $n(CaCl_2(aq)) = 1.94$ mol (2 MARKS)
- **iii.** $Pb(NO_3)_2(aq) + 2NaCl(aq) \rightarrow PbCl_2(s) + 2NaNO_3(aq)$
  $m(Pb(NO_3)_2(aq)) = 0.866$ g and $m(NaCl(aq)) = 2.34$ g (2 MARKS)
- **iv.** $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$
  $m(Al(s)) = 1.45$ g and $m(O_2(g)) = 3.11$ g (2 MARKS)

b. For each of the reactions above, calculate the mass of precipitate that would be formed. (8 MARKS)

Key science skills

Question 19 (5 MARKS)

In chemical analysis, it is vitally important to ensure that all collected values are accurate and precise in order to gather correct information about the sample being analysed.

a. Define the terms accuracy and precision. (2 MARKS)

b. A student conducting gravimetric analysis of a solution of copper ions forms a precipitate and weighs it. The student repeats this same method across a number of trials.

- **i.** Looking at his results the student notices that all of the measurements are close to each other but not close to the expected value. Comment on the precision and accuracy of these results. (1 MARK)
- **ii.** What type of error has the student most likely made? Justify your answer. (2 MARKS)

FROM LESSON 16D

Question 20 (17 MARKS)

Vinh is a year 11 chemistry student who wants to put together his knowledge of chemical analytical techniques with other principles of chemistry. As a result, Vinh decides to set up the following experiment:

1. Add a sample of pure octane to water.
2. Add a reactant to form a precipitate with octane.
3. Dry the precipitate with a paper towel and weigh it.
4. Calculate the amount of octane through mass-mass stoichiometry.

a. What term is used to refer to the substance that is being quantitatively analysed in gravimetric analysis? (1 MARK)

b. Will it be possible to use octane in the experiment described? Justify your answer. (3 MARKS)

c. Identify a potential change to the solvent that could be made to overcome the issue identified in part b. Justify your answer. (2 MARKS)

d. Is it necessary to carry out the experiment at all? Explain your answer. (3 MARKS)

e. Vinh decides to maintain the experimental technique outlined above, but replaces the analyte with a solution containing $MgCl_2$. Through reacting the solution with $AgNO_3$, a precipitate of AgCl was formed.

- **i.** With reference to the experimental technique outlined above, identify at least two potential sources of error. (2 MARKS)
- **ii.** Write the equation of the reaction between $MgCl_2$ and $AgNO_3$. (1 MARK)
- **iii.** The solution contains 33.2 g of $MgCl_2$, and 108.0 g of $AgNO_3$ is added to form a precipitate. Using calculations, identify the limiting and excess reactant in this reaction. (3 MARKS)
- **iv.** Calculate the mass of precipitate that will be formed. (2 MARKS)

FROM LESSON 16D

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Questions from multiple lessons

Question 21 (6 MARKS)

A scientist wants to determine the mass of aluminium nitrate in a solution by adding sodium phosphate. The equation for this reaction is:

$Al(NO_3)_3(aq) + Na_3PO_4(aq) \rightarrow AlPO_4(s) + 3NaNO_3(aq)$

In his experiment, the scientist finds that 7.78 g of precipitate is obtained.

a. What is the scientific name of the precipitate formed? (1 MARK)

b. Calculate the number of moles of precipitate formed. (2 MARKS)

c. Determine the molar ratio of the reactants in the reaction. (1 MARK)

d. Calculate the mass of $Al(NO_3)_3$ in the solution. (2 MARKS)

FROM LESSONS 4B, 4C & 6A

Question 22 (10 MARKS)

In the chemistry lab at his school, Alan is attempting to determine the concentration of chloride ions in a sample of seawater. In the lab with him Alan has a standard solution of silver nitrate. When the silver nitrate is mixed with the seawater in excess, a white precipitate of silver chloride (AgCl) is formed.

a. Why must the silver nitrate be added in excess? Explain your reasoning. (2 MARKS)

b. What must be done before the precipitate can be weighed? Justify your answer. (4 MARKS)

c. Alan removes the precipitate from the solution following correct experimental technique and records the mass to be 167.4 g.

  i. What is the mass of chloride in the sample of seawater? Show your calculations. (Assume all chloride ions were precipitated). (3 MARKS)

  ii. The sample of seawater contains 2.0 litres of water. What is the concentration of chloride ions in the seawater? (1 MARK)

FROM LESSONS 4C & 13A

Question 23 (14 MARKS)

Hard water is water that has a high concentration of cations present, which prevents soaps from forming a lather when introduced to water. Because of this, the efficacy of soaps in hard water is greatly reduced.

a. Explain why soap is less effective in hard water, and why a precipitate is formed. (3 MARKS)

b. A scientist suggests that in order to determine the concentration of ions that are present in a sample of hard water, the precipitate could be collected and weighed.

  i. Assuming that the scientist did not know the identity of the ions in the precipitate, would this be an effective method? Explain your answer. (2 MARKS)

  ii. Assuming that the scientist did know the identity of the ions in the precipitate, would this be an effective method for measuring the number of soap ions in the sample? Explain your answer. (3 MARKS)

  iii. The scientist carries out the experiment correctly. Determine whether the analyte or the added reactant would be the limiting reactant. (1 MARK)

c. A 500 mL solution containing stearate anions ($C_{18}H_{35}O_2^-$) was produced from adding 300.0 g of soap. Magnesium cations (Mg^{2+}) were added using correct technique, leading to the formation of a solid with a mass of 61.0 g. The equation of the reaction taking place is:

$Mg^{2+}(aq) + 2C_{18}H_{35}O_2^-(aq) \rightarrow Mg(C_{18}H_{35}O_2)_2(s)$

  i. Determine the concentration of stearate anions in the solution. (3 MARKS)

  ii. Determine the percentage by mass of stearate anions in the soap. (2 MARKS)

FROM LESSON 13A

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Question 24 (8 MARKS)

A chemistry teacher is setting up an experiment for her class to practice stoichiometric calculations. She begins by adding 150.0 g of gypsum ($CaSO_4{\cdot}2H_2O$) to a beaker containing water, with all of the gypsum dissolving. To determine the amount of calcium in the solution, phosphoric acid (H_3PO_4) is added to the solution to carry out the following reaction:

$3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$

Note: $CaSO_4{\cdot}2H_2O$ is a hydrate, which means it is a substance that contains water.

a. If this experiment is carried out correctly, which reactant should be the limiting reactant? Explain your answer. (2 MARKS)

b. After allowing the reaction to take place fully, the precipitate was removed from the solution.

  i. Identify one factor that must be accounted for before weighing the sample. (1 MARK)

  ii. The class measured the mass of the precipitate to be 87.4 g. Identify the number of moles of calcium that have reacted. (3 MARKS)

  iii. Calculate the %(m/m) of calcium in gypsum. (2 MARKS)

FROM LESSON 6B

Hints

17. Compare the mole ratios of the salt and water molecules.
18a. Compare the relative amounts of reactants to mole ratios.
18b. The limiting reactant determines the quantity of products.
19a. These variables measure different aspects of data.
19bi. Precision is a measure of the closeness of data to one another, whilst accuracy is a measure of the closeness of data to the true value.
19bii. Errors differ based on their consistency.
20a. The substance under analysis has a specific name.
20b. Polarity affects the ability of substances to interact.
20c. Octane is a non-polar molecule.
20d. Gravimetric analysis precipitates a substance out of a solution to be weighed.
20e. All of the substance must be weighed for gravimetric analysis to be effective.
21a. Ionic compounds are named based on the ions present.
21b. Number of moles equals mass divided by molar mass.
21c. Molar ratio is illustrated in the chemical equation.
21d. The mass formed depends on the number of moles and the molar mass.
22a. The reactant is not the substance being measured.
22b. Sample contamination should be avoided where possible.
22c. Molar masses are important for determining how much of a precipitate is made up by an element.
23a. This molecule is often found in a compound with a metal ion.
23b. Soap ions work best as ions.
23c. Soap ions are the analyte, and salt ions are the reactant.
24a. Consider which reactant is being measured.
24b. A precipitate should not be contaminated when being weighed.

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15B Analysing ions using light


Image: Dmitry Kalinovsky/Shutterstock.com

How can the concentration of coloured and colourless ions in these solutions be quantified?

In previous lessons, we explored how electrical conductivity and gravimetric analysis can be employed to determine the concentration of ions in solution. Even though some ions are not coloured, they can still absorb light. In this lesson, we will be learning about **colorimetry** and **UV-visible spectroscopy** and how these techniques can be used to find out the concentrations of known substances.

STUDY DESIGN DOT POINT

- quantitative analysis of salts:
  - molar ratio of water of hydration for an ionic compound
  - the application of mass-mass stoichiometry to determine the mass present of an ionic compound
  - the application of colorimetry and/or UV-visible spectroscopy, including the use of a calibration curve to determine the concentration of ions or complexes in a water or soil sample

15A 15B

2.2.11.4 & 2.2.11.6 Analysis of ions using colorimetry & determining the concentration of metal ions using calibration curves

2.2.11.5 & 2.2.11.6 Analysis of ions using UV-visible spectroscopy & determining the concentration of metal ions using calibration curves

ESSENTIAL PRIOR KNOWLEDGE

1A Atomic structures

See question 93.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

KEY TERMS AND DEFINITIONS

Calibration curve graph depicting the relationship between concentration and light absorbance for a particular substance

Colorimetry method of determining the concentration of a known substance which exploits the light absorbance properties of substances

Colorimeter instrument used to undertake colorimetric analysis

Complementary colours colours opposite to each other on the colour wheel

Electromagnetic radiation waves of the electromagnetic field including visible and ultraviolet light

Monochromator device used to select a particular wavelength of light

Spectroscopy analysis of the interaction between matter and electromagnetic radiation

Standard solution solution with a precisely known concentration

UV-visible spectroscopy spectroscopy method which uses the ultraviolet light region of the electromagnetic spectrum

UV-visible spectrophotometer instrument used in UV-visible spectroscopy

Analysis of ions using colorimetry & determining the concentration of metal ions using calibration curves 2.2.11.4 & 2.2.11.6

Colorimetry is a method of determining the concentration of a known substance in a sample using the light absorbance properties of substances.

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How can colour be used to analyse and calculate the concentration of ions?

In this section, we will be introducing another method for chemical analysis, known as colorimetry. Colorimetry involves measuring the intensity of colour in a solution, which is directly proportional to the concentration of ions in the solution. This method is often simple to conduct and is done in a non-destructive manner, meaning that the sample of water remains undamaged throughout the experiment.

Although the colour of a solution can be observable by eye, the concentration of ions in a solution can be measured more accurately using a device known as a **colorimeter**. However, in order to understand the function of a colorimeter, we first need to investigate some fundamental principles of light.

What part does light play?

Visible light is a form of **electromagnetic radiation**, which refers to waves of the electromagnetic spectrum. Other forms of electromagnetic radiation include *x*-rays, ultraviolet light and radio waves. The different types of light radiation are shown in figure 1.

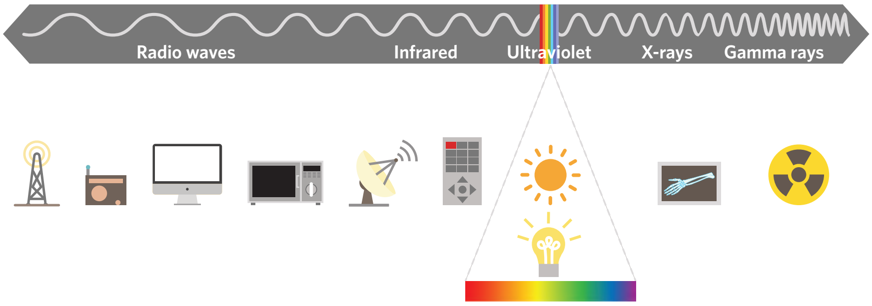

Figure 1 Electromagnetic spectrum

Electromagnetic radiation, such as visible light, interacts with atoms. By distinguishing between different types of atomic interactions with light, an experimenter is able to ascertain the atomic make-up of the solution being tested. This form of chemical analysis is referred to as **spectroscopy**. The principles of spectroscopy can be applied to colorimetry as the technique allows for the analysis of coloured substances which correspond to wavelengths between 400 nm and 700 nm on the electromagnetic spectrum.

Depending on their chemical composition, different materials are able to absorb and reflect different wavelengths of light. The colour reflected by the substance is the colour that the observer can see. For example, a copper solution will absorb all colours except blue, which will be reflected into the observer's eyes. Hence, the solution will appear blue. This principle is demonstrated in figure 2, using a ladybug as an example.

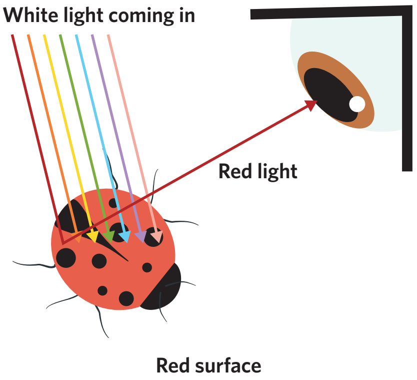

Figure 2 Red is the only main colour of white light that is not absorbed by the beetle's shell.

The observed colour of a material and the colour most strongly absorbed are referred to as **complementary colours** – these are found on opposite sides of the colour wheel shown in figure 3. For example, orange is a complementary colour to blue. This means that a solution that appears blue will absorb orange light the most strongly.

Figure 3 Colour wheel

One method of learning the complementary colours is by remembering the order in which colours appear on the electromagnetic spectrum – from the largest wavelength to the smallest wavelength. ROYGBV is an acronym commonly used to remember complementary colours and it stands for red, orange, yellow, green, blue, and violet. The pairs of complementary colours produced from the two acronyms can be lined up as shown in figure 4.

If the substance being tested is opaque (cloudy), the colour that is reflected – and therefore seen – is complementary to the colour most strongly absorbed. However, if the substance is transparent, it will freely transmit the colour complementary to the colour most strongly absorbed. The colour being transmitted is what we are able to see. These different interactions of light with a material are shown in figure 5.

R O Y G B V
G B V R O Y

Figure 4 Alignment of complementary colours

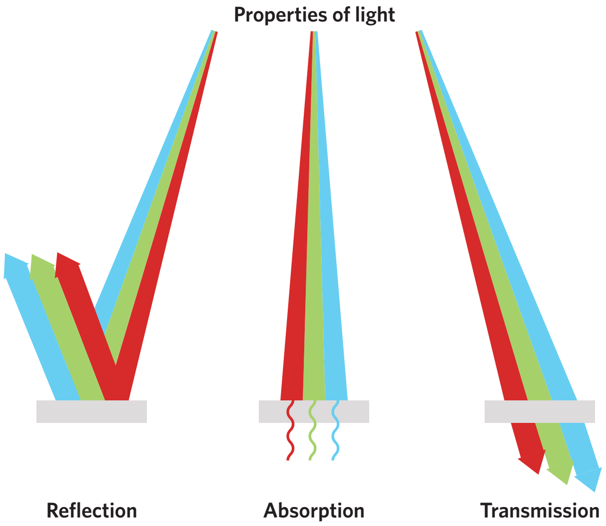

Figure 5 Reflection, absorption and transmission of light

How does colorimetry work?

A substance's ability to reflect, absorb, or transmit light can be examined to determine the concentration of ions in solution; this is the premise of colorimetry.

Colorimeters function by measuring the absorbance of visible light as the light passes through a solution. This enables the concentration of solute in solution to be established. Each solute has its own unique atomic structure, which means that every solution best absorbs light at a different wavelength. Furthermore, more concentrated solutions absorb more electromagnetic radiation. This is known as the Beer-Lambert Law, which states that absorbance of light is proportional to the concentration of a substance. Hence, the more light that is absorbed, the greater the concentration of the sample.

Colorimetry is both a qualitative and quantitative technique. It is used qualitatively to determine the identity of a solute in a solution and quantitatively to determine the concentration of a solute in solution. A colorimeter consists of three main parts that are used to measure the intensity of colour in a sample solution, as shown in figure 6.

1. Light from a light source is transmitted through a filter to select the wavelength of light most readily absorbed by the solute in the solution.

2. Solute ions in the test tube solution absorb a particular amount of light.

3. An electronic detector is used to measure the percentage of light absorbed by the solution.

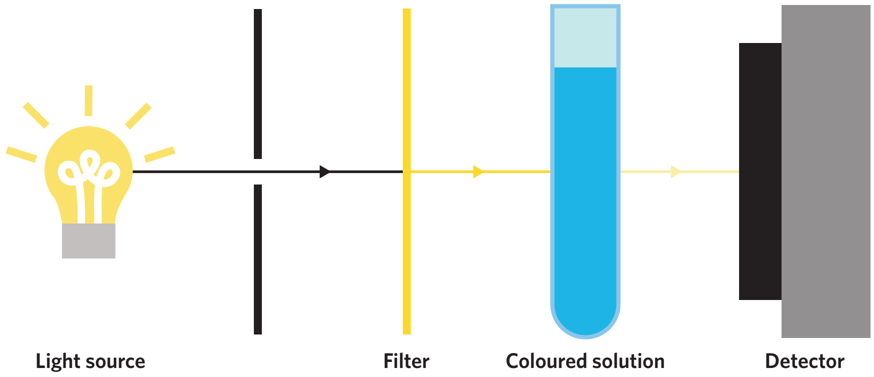

Figure 6 Analysis of solutions using colorimetry

First, the light produced is filtered in order to select for the wavelength that will be most strongly absorbed by the sample. For the blue copper solution in figure 6, the filter allows only the complementary colour of light, orange, to pass through. Then, a recorder detects how much light passes through the solution to generate a value which represents how much light was absorbed by the solution. We know that the greater the concentration of copper ions in the solution, the greater the absorbance of the orange light will be. This will, therefore, lead to a higher percentage of light absorbed by the solution.

How can standard solutions and calibration curves be used to quantify a known substance?

To accurately determine the concentration of a solution using colorimetry, the absorbance results from the electronic detector must be compared to a series of **standard solutions**. Standard solutions are solutions with a precisely known concentration. Developing a **calibration curve** from standard solutions allows us to identify the concentration of the sample solution. That is, the absorbance data from the standard solutions can be compared with the sample under analysis to determine the concentration of solute in the sample.

A calibration curve is used to visualise the relationship between concentration and absorbance. These calibration curves are a plot of concentrations of the solutions on the x-axis, and their corresponding absorbance values on the y-axis. This is shown in figure 7, which depicts a linear relationship between concentration and absorbance in accordance with the Beer-Lambert Law.

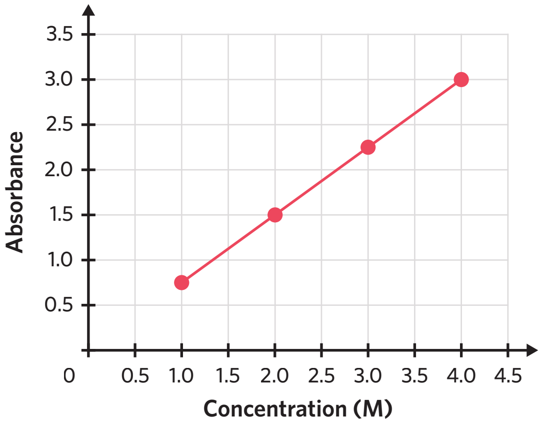

Figure 7 Calibration curve

USEFUL TIP

Unlike concentration, absorbance has no units as it is dependent on the type of spectrophotometer used and its individual setup.

By constructing a calibration curve, we are able to ascertain the concentration of a sample solution using the absorbance value produced by the detector. To determine the concentration, the x-value (concentration) that corresponds to the y-value (absorbance) from the detector needs to be identified. For example, if we take a solution with a recorded absorbance of 2.5, we can then find its location on the curve as shown in figure 8. As can be seen, the concentration corresponding to an absorbance of 2.5 is approximately 3.35 M.

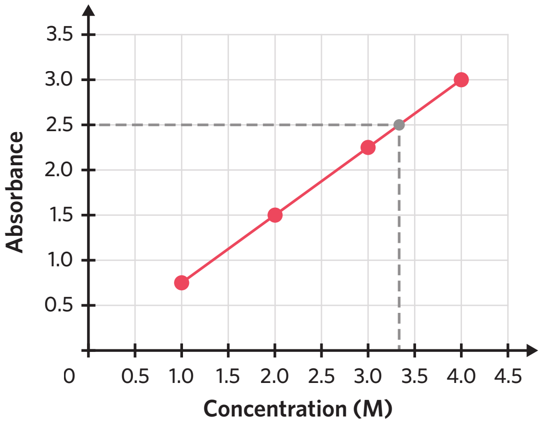

Figure 8 Calibration curve

The following steps are used to accurately perform colorimetric analysis of an unknown substance:

- Prepare a set of standard solutions of the substance at known concentrations.
- Measure the absorbance of the standard solutions at a particular wavelength of visible light.
- Plot a calibration curve using the collected data, with concentration on the x-axis and absorbance on the y-axis.
- Measure the absorbance of the solution of unknown concentration using a colorimeter.
- Determine the concentration of the solution using the calibration curve.

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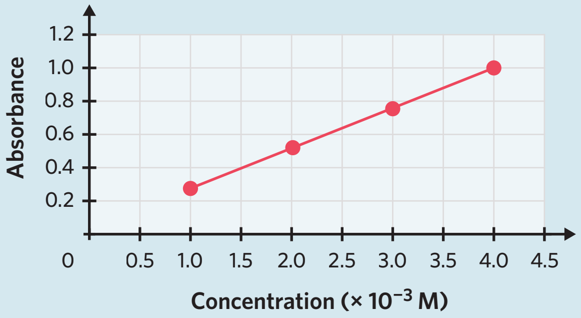
When using a calibration curve to determine the concentration of a solution, the absorbance value obtained must not be greater than that of the standard solution with the highest concentration. If the absorbance value obtained is greater than the maximum value or lower than the minimum value on the calibration curve, the curve cannot be used to accurately determine the concentration of the solution. Instead, more standards need to be measured.

USEFUL TIP

Most samples are diluted when tested in a colorimeter, and it is important that these dilutions are taken into account when calculating the original concentration.

WORKED EXAMPLE 1

A researcher is using colorimetry to determine the amount of copper ions in a water sample. Having already constructed a calibration curve, the absorbance of orange light passing through the sample is measured, producing a result of 0.62. Using the calibration curve shown, determine the concentration of copper ions in the water sample.



What information is presented in the question?

The absorbance of the sample and the calibration curve.

What is the question asking us to do?

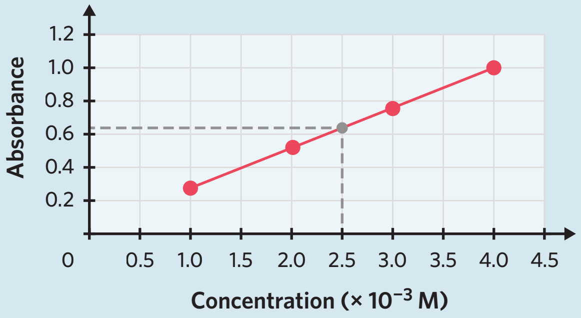
Determine the concentration of copper ions in the water sample.

What strategies do we need in order to answer the question?

1. Use the absorbance value to plot a point on the calibration curve.
2. Determine the concentration of the sample by observing the x-axis value.

Answer

The absorbance of the sample is 0.62.



Based on the calibration curve, it can be seen that the concentration of copper ions is 2.5 × 10^{-3} M.

Progress questions

Question 1

The electromagnetic spectrum is mostly made of visible light.

A. True

B. False

Question 2

The complementary colour of the wavelength of light most strongly absorbed is

A. strongly absorbed.

B. reflected.

Question 3

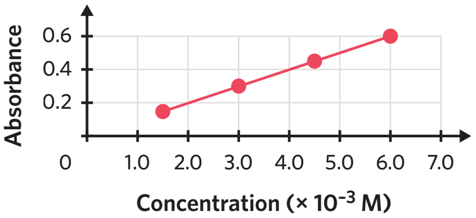
A calibration curve is used in colorimetry to

A. determine the concentration of an ion with known absorbance of light.

B. visualise the inverse relationship between ion concentration and light absorbance.

Continues →

Question 4



What is the concentration of a substance with an absorbance of 0.25?

A. 1.25×10^{-3} M

B. 2.50×10^{-3} M

Question 5

The colour of an object corresponds to

A. the colour most strongly absorbed by the material.

B. the colour reflected or transmitted through the material.

MISCONCEPTION

'Metal ion complexes that are colourless do not absorb any light.'

Visible light is not the only light that can be absorbed by metal ion complexes – UV can also be absorbed. Humans are unable to see UV, but spectrophotometers can detect it.

Analysis of ions using UV-visible spectroscopy & determining the concentration of metal ions using calibration curves 2.2.11.5 & 2.2.11.6

UV-visible spectroscopy is an analytical technique using ultraviolet and visible light to determine the concentration of a known substance in a sample of water.

How can we analyse and calculate the concentration of ions that are not coloured?

Another form of chemical analysis using many of the same principles as colorimetry and spectroscopy is UV-visible spectroscopy. This technique is more accurate than the others we have explored due to the use of a more sophisticated instrument – a **UV-visible spectrophotometer**.[1] This instrument uses a **monochromator** to select the specific wavelength of light to be used in spectroscopic analysis.

In UV-visible spectroscopy, the solution being analysed is first scanned across multiple wavelengths by the spectrophotometer to assess which wavelengths are absorbed to the greatest extent by the solution. The wavelength that is most absorbed by the solution is the one that is selected for spectroscopy.

In the scan shown in figure 9, strong absorbance is shown at a wavelength of 250 nm, which is in the ultraviolet range of the electromagnetic spectrum. There is also a smaller peak around 525 nm, which corresponds to the colour green in the visible light region. Generally, the analysis would be conducted at the wavelength corresponding to the highest absorbance. However, we also need to make sure that there isn't significant absorbance from other substances dissolved in the solution. Therefore, the wavelength chosen may not be the one with the strongest absorbance by the solution. For the scan shown in figure 9, the wavelength of 525 nm may be used if other substances in the solution also readily absorb light with a wavelength of 250 nm.

KEEN TO INVESTIGATE?

[1] **How do you use a UV-Vis spectrophotometer?**
Search YouTube: Spectroscopy in a Suitcase

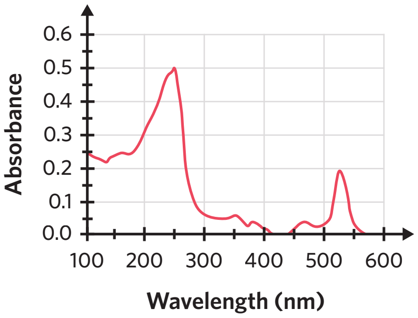


Figure 9 Spectrophotometer scan

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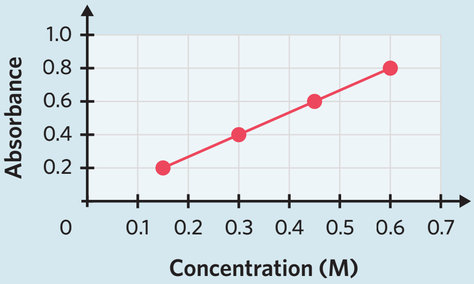
WORKED EXAMPLE 2

UV-visible spectroscopy makes use of calibration curves to visualise the relationship between concentration and absorbance of light in solutions. A student produced the following calibration curve and determined the absorbance of a solution of unknown concentration to be 0.7. Determine the concentration of the solution.



What information is presented in the question?

The absorbance of the sample and the calibration curve.

What is the question asking us to do?

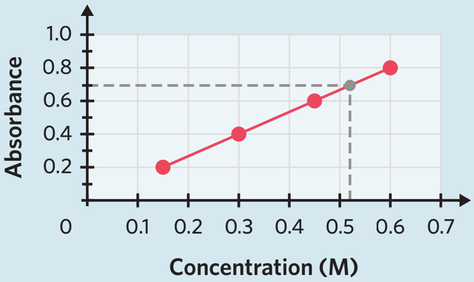
Determine the concentration of the solution.

What strategies do we need in order to answer the question?

1. Use the absorbance value to plot a point on the calibration curve.
2. Determine the concentration of the sample by observing the *x*-axis value.

Answer

The concentration of the sample is roughly 0.52 M.



Like colorimetry, UV-visible spectroscopy involves the construction of a calibration curve from the absorbance of a series of standard solutions. The absorbance from the solution with an unknown concentration can be compared to the curve to determine the concentration of the solution.

USEFUL TIP

If the concentration of the substance under investigation does not fit within the calibration curve then it must be diluted or concentrated. It is not valid to extrapolate or interpolate the calibration curve when trying to find the concentration.

What is a metal complex?

Some metal ions do not absorb UV or visible light effectively. Hence, they require the formation of a metal complex in order to be observed in spectroscopy. For example, solutions containing Fe^{3+}(aq) ions don't absorb visible or ultraviolet light effectively. However, in a complex such as $FeSCN^{2+}$(aq), iron (II) thiocyanate, the solution is brightly coloured and thus, suitable for UV-visible spectroscopic analysis. This process is shown in figure 10.

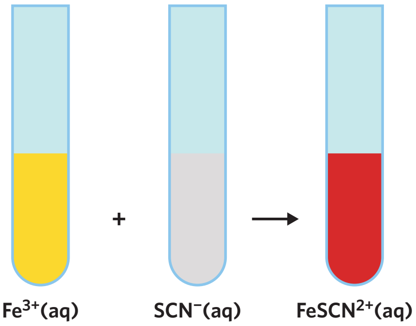


Figure 10 $FeSCN^{2+}$ metal ion complex

How are colorimetry and UV-visible spectroscopy used?

Both colorimetry and UV-visible spectroscopy are extremely useful in determining the concentration of solutes in a given solution. This has a number of real-world applications, including but not limited to:

- Identifying and quantifying the contents of blood including bilirubin, cholesterol, iron, haemoglobin and sugars.
- Assessing the contents for toxicity markers such as lead.
- Carrying out quality assurance of milk and other dairy products.
- Testing air for the presence of chromium and lead.
- Assessing the phosphate concentration in waterways to prevent eutrophication.

Progress questions

Question 6

The benefit of a spectrophotometer over regular colorimeters is that it is

A. able to produce an emission spectrum.

B. able to scan different light wavelengths to assess which has the greatest absorbance.

Continues →

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Question 7

A monochromator is used in UV-visible spectroscopy to

A. reduce the intensity of light being used.

B. select a particular wavelength of light.

Question 8

Calibration curves are constructed by

A. measuring the absorbance of solutions of known concentration.

B. measuring the concentrations of solutions with known absorbance values.

Question 9

One potential use of colorimetry and UV-visible spectroscopy is

A. measuring electrical conductivity of a water sample.

B. assessment of blood contents.

Question 10

Metal complexes can be used for

A. solutions that do not absorb light to enable absorption.

B. solutions that readily absorb light to reduce the intensity of absorption.

Theory summary

- Colorimetry makes use of the interactions between light and ions to determine water sample quality.
- The complementary colour of the colour most strongly absorbed will be the one most strongly reflected.
- UV-visible spectroscopy uses many of the same principles as colorimetry, but with more sophisticated equipment.
- Calibration curves can be used to quantify known substances.

The content in this lesson is considered fundamental prior knowledge to instrumental analysis of organic compounds (Unit 4 AOS 2).

15B Questions

Mild Medium Spicy

Deconstructed

Use the following information to answer questions 11–13.

A chemistry teacher is teaching a class about spectroscopy, most notably colorimetry and UV-visible spectroscopy, and has been asked a series of questions.

Question 11 (1 MARK)

Which option correctly describes colorimetry?

A. Measuring the absorbance of light passing through a solution to establish the concentration of ions.

B. Inserting a sample into a flame and observing the colour produced to determine the chemical composition of the sample.

C. Determining the amount of a substance dissolved in water by weighing the precipitate formed.

D. Producing an emission spectrum to determine the presence of particular ions in a solution.

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Question 12 (1 MARK)

Which option correctly describes the difference between colorimetry and UV-visible spectroscopy?

A. Ultraviolet electromagnetic radiation is used in UV-visible spectroscopy instead of visible light.

B. More advanced equipment is used in UV-visible spectroscopy for more accurate results.

C. More advanced equipment is used in colorimetry for more accurate results.

D. UV-visible spectroscopy produces an emission spectrum whereas colorimetry does not.

Question 13 (3 MARKS)

Explain how colorimetry and UV-visible spectroscopy make use of the interactions between ions in solution and light to provide information about a sample.

Exam-style

Question 14 (9 MARKS)

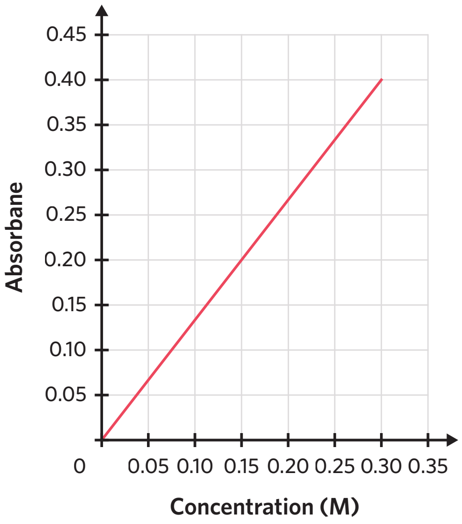
Colorimetry is a form of chemical analysis used to determine the elements present in a solution through their interactions with electromagnetic radiation. In order to determine the concentration of these elements, a calibration curve must be consulted.

a. What is a calibration curve? (1 MARK)

b. Describe the process used to produce a calibration curve. (2 MARKS)

c. Describe how a calibration curve is used to determine the concentration of a sample solution. (2 MARKS)

d. The calibration curve shown was constructed using standard solutions of copper ions in water. Using the calibration curve, determine the approximate concentration of a solution of copper ions with an absorbance of 0.25. (1 MARK)



e. A solution of copper ions has a blue colour. What colour light will be used for colorimetry? Explain your reasoning. (3 MARKS)

Question 15 (10 MARKS)

Ryan has recently learned about colorimetry in his chemistry class, and wants to put his knowledge to use. At the school laboratory, Ryan produces a series of standard solutions of cobalt chloride and measures the absorbance of each.

Concentration (M)	Absorbance
0.5	0.19
1.0	0.40
1.5	0.60
2.0	0.80

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a. What is the unit used for absorbance in colorimetry? (1 MARK)

b. Produce a calibration curve using Ryan's experimental results. Include correctly labelled axes and units. (3 MARKS)

c. Ryan claims that violet light should be used to carry out colorimetry as the cobalt chloride solution is violet. Explain whether Ryan is correct. (3 MARKS)

d. Ryan wishes to use his calibration curve to determine the concentration of another solution containing cobalt chloride.

  i. The absorbance of the unknown solution is 0.7. Using the calibration curve, determine the concentration of the solution. (1 MARK)

  ii. Ryan uses colorimetry to determine the absorbance of yet another unknown solution, which has an absorbance of 0.90. Explain why this calibration curve cannot be used for this solution. (2 MARKS)

Question 16 (8 MARKS)

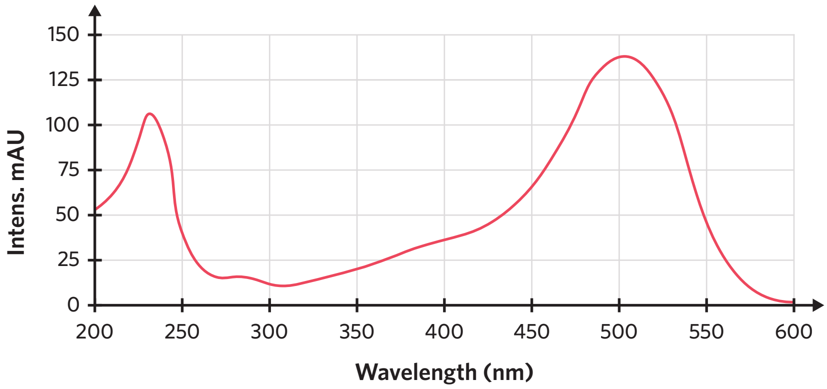
UV-visible spectroscopy is a form of chemical analysis that uses many of the same principles as colorimetry, to observe the interactions between light and molecules in solution and determine concentration.

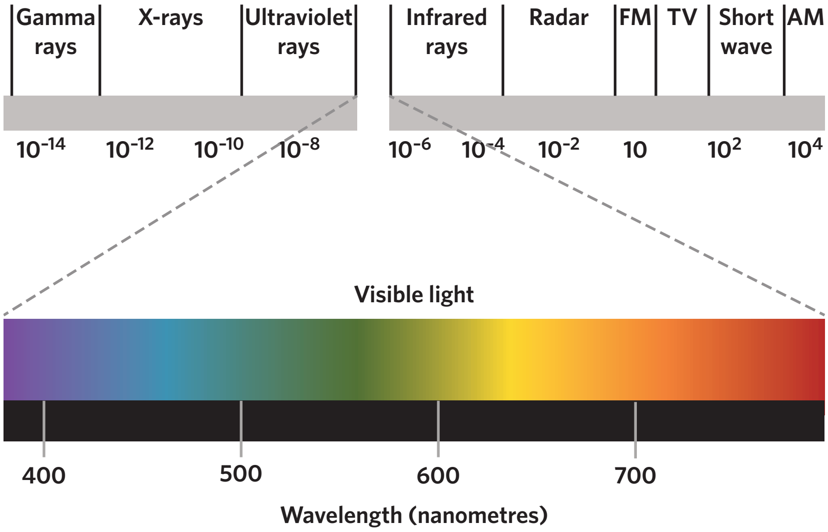
a. A UV-visible spectrophotometer uses a monochromator for spectroscopy. What is the function of a monochromator, and what is used instead in colorimetry? (2 MARKS)

b. A UV-visible spectrophotometer has the ability to scan a number of different wavelengths of light in order to assess for absorbance.

  i. Explain why a UV-visible spectrophotometer is advantageous for UV-visible spectroscopy. (2 MARKS)

  ii. Using the following UV-visible spectrum and electromagnetic spectrum, determine what colour would be best suited for this solution. (1 MARK)





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c. Spectroscopy relies on the interactions between ions in a solution and light, which requires the solution to be coloured. Describe the process utilised to perform UV-visible spectroscopy of a sample of metal ions without colour. (2 MARKS)

d. Identify a real-world application of UV-visible spectroscopy. (1 MARK)

Use the following information to answer questions 17–19.

Mark has a solution of iron ions with an unknown concentration and is debating between colorimetry and UV-visible spectroscopy.

Question 17 (4 MARKS)

Explain one advantage and one disadvantage of colorimetry and UV-visible spectroscopy.

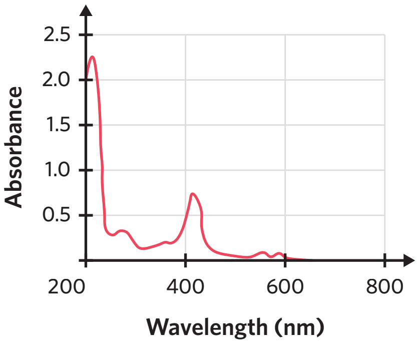
Question 18 (3 MARKS)

Mark decides to use UV-visible spectroscopy for chemical analysis, but wants to clarify his understanding of the method first. Using the UV-visible spectrum, identify which of the following statements are true, justifying your answer(s).

I. The spectrum is produced by measuring the frequencies of light from electrons returning to the ground state from an excited state.

II. The concentration of a solution of this compound can only be determined by UV-visible spectroscopy at 210 nm.

III. The amount of light absorbed by a solution depends on its concentration.



Question 19 (2 MARKS)

The iron solution is transparent. Explain how light used in the colorimetric analysis will interact with the solution.

Key science skills

Question 20 (5 MARKS)

A student is analysing a sample of water that contains copper ions using colorimetry. The student sets up a series of standard solutions in a number of test tubes, some of which are transparent and some of which are translucent, and then carries out colorimetric analysis. When she constructs the calibration curve after the experiment, she discovers that it does not show the expected linear shape.

a. Explain why the calibration curve does not fit the expected linear shape. (2 MARKS)

b. What type of error is this? Explain. (2 MARKS)

c. What should the student do to remove the source of error? (1 MARK)

FROM LESSON 16D

Questions from multiple lessons

Question 21 (7 MARKS)

Annie mixed two test tubes, each containing silver and iodide ions at different concentrations. She then uses the final solution in a colorimetry test to determine the concentration of silver ions in the solution.

a. Comment on the appropriateness of using colorimetry to measure the concentration of silver ions. (3 MARKS)

b. As part of the experiment, Annie mixed 20.0 mL of a 1.0 M silver ion solution with 15.0 mL of a 1.2 M iodide ion solution. Determine the limiting reagent. (3 MARKS)

c. What amount, in mol, of silver ions would be available in the solution to absorb light as part of the experiment? (1 MARK)

FROM LESSONS 12A & 15A

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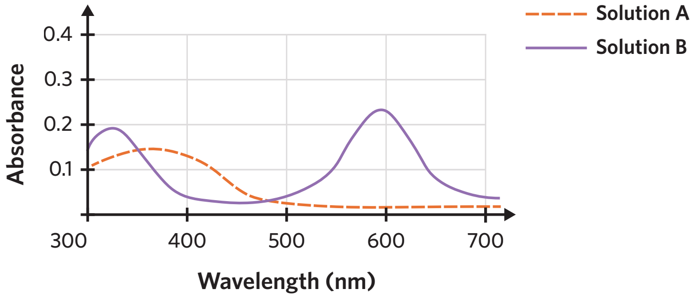
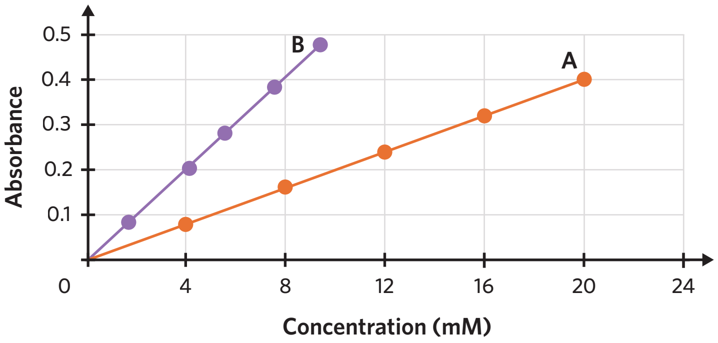
Question 22 (8 MARKS)

UV-visible spectroscopy was used to measure the spectra of solutions A and B. Solution A was orange, and solution B was violet. The following UV-visible spectra were recorded.



a. If 10.00 mL of solution A was mixed with 10.00 mL of solution B, which wavelength should be used to measure the absorbance of solution B? Justify your answer. (3 MARKS)

b. The analyst used two sets of standard solutions and blanks to determine the calibration curves for the two solutions. The absorbances were plotted on the same axes, as shown on the graph.



The analyst found that when measured at an appropriate wavelength, solution B had an absorbance of 0.20. If solution B was cobalt (II) nitrate, $Co(NO_3)_2$, determine its concentration (M). (2 MARKS)

c. A small sample of cobalt (II) nitrate was used and reacted with copper (II) sulfate.

i. Write the balanced equation for this reaction. (1 MARK)

ii. What concentration of a 20.0 mL sample of copper (II) sulfate needs to be used with 10.0 mL of cobalt (II) nitrate so that the same amount of both substances is added to the reaction? (2 MARKS)

FROM LESSON 12A

Hints

14a. Calibration curves are produced from standard solutions.

14b. Calibration curves are produced from standard solutions.

14c. Calibration curves show concentration on the *x*-axis, and absorbance on the *y*-axis.

14d. Every point on the curve has an *x*-value and a *y*-value.

14e. Take ROYGBV, and shift it 3 letters to the right.

15a. Some properties do not have units.

15b. Calibration curves have concentration on the *x*-axis, and absorbance on the *y*-axis.

15c. Colorimetry makes use of complementary colours.

15di. Every point on the curve has an *x*-value and a *y*-value.

15dii. Absorbances must be within specific values on the curve.

16a. A monochromator can increase the accuracy of spectroscopy.

16bi. Spectrophotometers make absorbance take place more effectively.

16bii. Strong absorbance is important for carrying out spectroscopy.

16c. Ionic solutions can become coloured.

16d. UV-visible spectroscopy can be used to analyse the contents of substances.

17. Consider factors such as cost and accuracy.

18. A UV-visible spectrum shows which wavelengths of light have the highest absorbance.

19. Different wavelengths of light have different interactions with ions.

20a. Calibration curves should depict a linear relationship if completed properly.

20b. Errors vary depending on if they affect results consistently or not.

20c. Variables must be controlled in experiments.

21a. Gravimetric analysis involves the addition of a reactant.

21b. The limiting reagent is completely used up in the reaction.

21c. This value is equal to the number of mol of silver ions left over at the end of the reaction.

22a. The wavelength used to measure the absorbance of a solution is the wavelength at which the solution absorbs the most light without interference from other solutions or contaminants.

22b. Every point on the curve has an *x*-value and a *y*-value.

22ci. Consider how cobalt (II) nitrate reacts with copper sulfate.

22cii. Use the formula linking moles (*n*), concentration (*c*) and volume (*V*)

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Chapter 15 review

Multiple choice (10 MARKS)

Question 1 (1 MARK)

In order for a soil sample to be tested for the concentration of salt it must first be

A. broken down by heating.

B. neutralised with acid.

C. dissolved in water and filtered.

D. neutralised with a base.

Question 2 (1 MARK)

Which of the following is not a result of high concentrations of sodium chloride (salt)?

A. Increased plant growth

B. Decreased plant growth

C. Contamination of water supplies used for irrigation

D. Contamination of water supplies used for for human consumption

---

Question 3 (1 MARK)

Water salinity can be tested by

A. placing a lightbulb in a water sample.

B. connecting a water sample to an electrical circuit and measuring the conductivity.

C. adding a solid to a solution and determining the extent of dissolution.

D. observing the colour of a water sample.

Question 4 (1 MARK)

In order to make an anhydrous salt from a hydrated salt, which of the following procedures is carried out?

A. Filtration

B. Neutralisation

C. Evaporation

D. Distillation

Question 5 (1 MARK)

What name would be given to the following formula of a hydrated salt, $Ba(OH)_2 \bullet 8H_2O$?

A. Boron hydroxide octahydrate

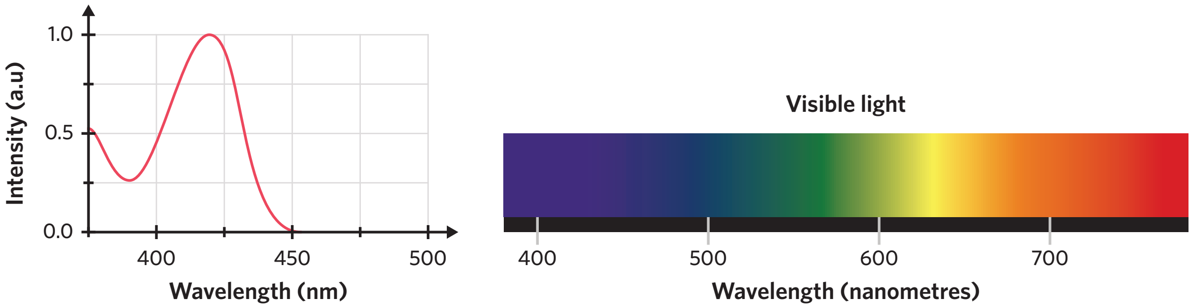
B. Barium hydroxide octahydrate

C. Boron dihydroxide tetrahydrate

D. Barium dihydroxide tetrahydrate

Question 6 (1 MARK)

The following UV-Vis spectrum was formed from the analysis of a solution.



If this sample was to undergo colorimetry analysis, the filter used should select for

A. a purple colour.

B. a yellow/orange colour.

C. a red colour.

D. a green colour.

Question 7 (1 MARK)

An intensely orange coloured solution was analysed using colorimetry. Which of the following statements is not true?

A. The concentration of ions in the solution is inversely proportional to the amount of light detected by the detector.

B. If the colour of the solution is a result of the ions present, a darker solution means that more light will be absorbed.

C. A blue/violet light is used.

D. The solution becomes a lighter orange during analysis.

Question 8 (1 MARK)

Calculate the molar mass of sodium sulfate decahydrate, $Na_2SO_4{\bullet}10H_2O$.

A. 142.3 g mol^{-1}

B. 160.3 g mol^{-1}

C. 180.0 g mol^{-1}

D. 322.1 g mol^{-1}

Question 9 (1 MARK)

If six moles of H_2 were mixed with nine moles of O_2 according to the following reaction,

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

which is the excess reactant and by how much?

A. H_2 and 1 mole excess

B. O_2 and 4 moles excess

C. H_2 and 2 moles excess

D. O_2 and 6 mole excess

Question 10 (1 MARK)

Consider the following reaction, $CaCl_2(aq) + 2KOH(aq) \rightarrow Ca(OH)_2(s) + 2KCl(aq)$.

If 0.42 g of $CaCl_2$ was allowed to react in excess KOH, calculate the mass of precipitate formed.

A. 0.0025 g

B. 0.19 g

C. 0.28 g

D. 0.38 g

Short answer (30 MARKS)

Question 11 (10 MARKS)

Miguel was set the task of trying to determine the formula of a hydrated salt of magnesium sulfate. He collected the following results.

Item	Mass (g)
Mass of empty test tube	29.254
Mass of test tube + hydrated salt	32.637
Mass of test tube + contents during experiment	31.324, 30.931, 30.911, 30.912, 30.911
Mass of hydrated salt	3.383
Mass of anhydrous salt	1.656
Mass of water lost	

a. Why must the sample of hydrated magnesium sulfate be pure? (2 MARKS)

b. Complete the table for the mass of water lost. (1 MARK)

c. Calculate the number of moles of anhydrous $MgSO_4$ left at the end of this experiment. (2 MARKS)

d. Calculate the molar mass of the hydrated magnesium sulfate. (1 MARK)

e. Calculate the mass of water in 1 mole of the hydrated magnesium sulfate. (1 MARK)

f. Calculate the moles of water of crystallisation to the nearest whole number in the sample of hydrated magnesium sulfate. (1 MARK)

g. Give the formula and name for the hydrated salt of magnesium sulfate. (2 MARKS)

Question 12 (9 MARKS)

Antony is a statesman in Rome who has been designated the task of assessing the quality of the city's water supplies. Antony decides to use gravimetric analysis and electrical conductivity testing as potential methods of determining water salinity.

a. When carrying out gravimetric analysis, Antony must ensure that he knows which reactant is limiting and which is in excess.

  i. Should the ion under analysis or the added reactant be the limiting reactant? Justify your answer. (2 MARKS)

  ii. Identify the limiting reactant in the following reaction:

  $Pb(NO_3)_2(aq) + CaCl_2(aq) \rightarrow PbCl_2(s) + Ca(NO_3)_2(aq)$

  Mass of $Pb(NO_3)_2(aq) = 6.34$ g and mass of $CaCl_2(aq) = 2.51$ g

  Show your calculations. (3 MARKS)

  iii. For the reaction shown in part aii, calculate the mass in grams of $PbCl_2$ that will be formed. (1 MARK)

b. Antony also tries to learn about the method of electrical conductivity testing. When comparing a sample of pure water with a river water sample, he notices that the electrical conductivity in the river sample is far greater. Explain this observation. (3 MARKS)

Question 13 (11 MARKS)

Ziggy is a martian investigating the properties of materials on Earth. As the principles of science are consistent throughout all of the observable universe, Ziggy's martian home also makes use of chemical analysis techniques found on Earth, such as colorimetry and UV-Vis spectroscopy. Ziggy scoops up a sample of water from the Yenisei River in Siberia, and compares it to a sample of water taken from the Yangtze River in China.

a. Identify two potential sources of contamination in each of these rivers. (2 MARKS)

b. Describe the processes of colorimetry and UV-Vis spectroscopy, highlighting the similarities and differences. (3 MARKS)

c. Ziggy decides to assess the rivers for the concentration of fluoride in each. He possesses a series of standard solutions of fluoride, and records the following:

Fluoride concentration (ppm)	0.5	1.0	1.5	2.0	2.5
Absorbance	0.13	0.26	0.40	0.53	0.67

i. Plot the data on graph paper, and draw a line of best fit. (4 MARKS)

ii. In order for water to be safe to drink, the concentration of fluoride must be less than 1.5 ppm. In UV-Vis spectroscopy, the sample from the Yenisei River obtained an absorbance value of 0.37, while the sample from the Yangtze River obtained an absorbance value of 0.64. Explain whether these samples are safe to drink. (2 MARKS)

Key science skills (10 MARKS)

Question 14 (10 MARKS)

The amount of calcium carbonate ($CaCO_3$; $M = 100.1$ g mol^{-1}) in the ore dolomite can be determined by gravimetric analysis. The dolomite sample is dissolved in oxalic acid $(COOH)_2$ and the calcium ions (Ca^{2+}) present are precipitated as calcium oxalate (CaC_2O_4; $M = 128.1$ g mol^{-1}). The calcium oxalate is filtered, dried and strongly heated to form calcium oxide (CaO; $M = 56.1$ g mol^{-1}).

a. Write a balanced equation for the complete reaction occurring between calcium carbonate and oxalic acid (including state symbols) to form calcium oxalate. (2 MARKS)

b. What is the name of the type of reaction that occurs between calcium carbonate and oxalic acid? (1 MARK)

c. Write a balanced equation for the complete reaction occurring when calcium oxalate is heated strongly to form calcium oxide (including state symbols). (2 MARKS)

d. In one analysis the mass of dolomite used was 3.72 g. The mass of calcium oxide formed was found to be 1.24 g. Calculate the percentage of calcium carbonate in the dolomite sample. The stoichiometric ratio is $n(CaCO_3)$ in dolomite = $n(CaO)$ produced. (3 MARKS)

e. Suggest two possible sources of error that could lead to an overestimation of the % $CaCO_3$ in the dolomite sample. (2 MARKS)

Adapted from VCAA 2005 Exam 1 Multiple choice Q8,9

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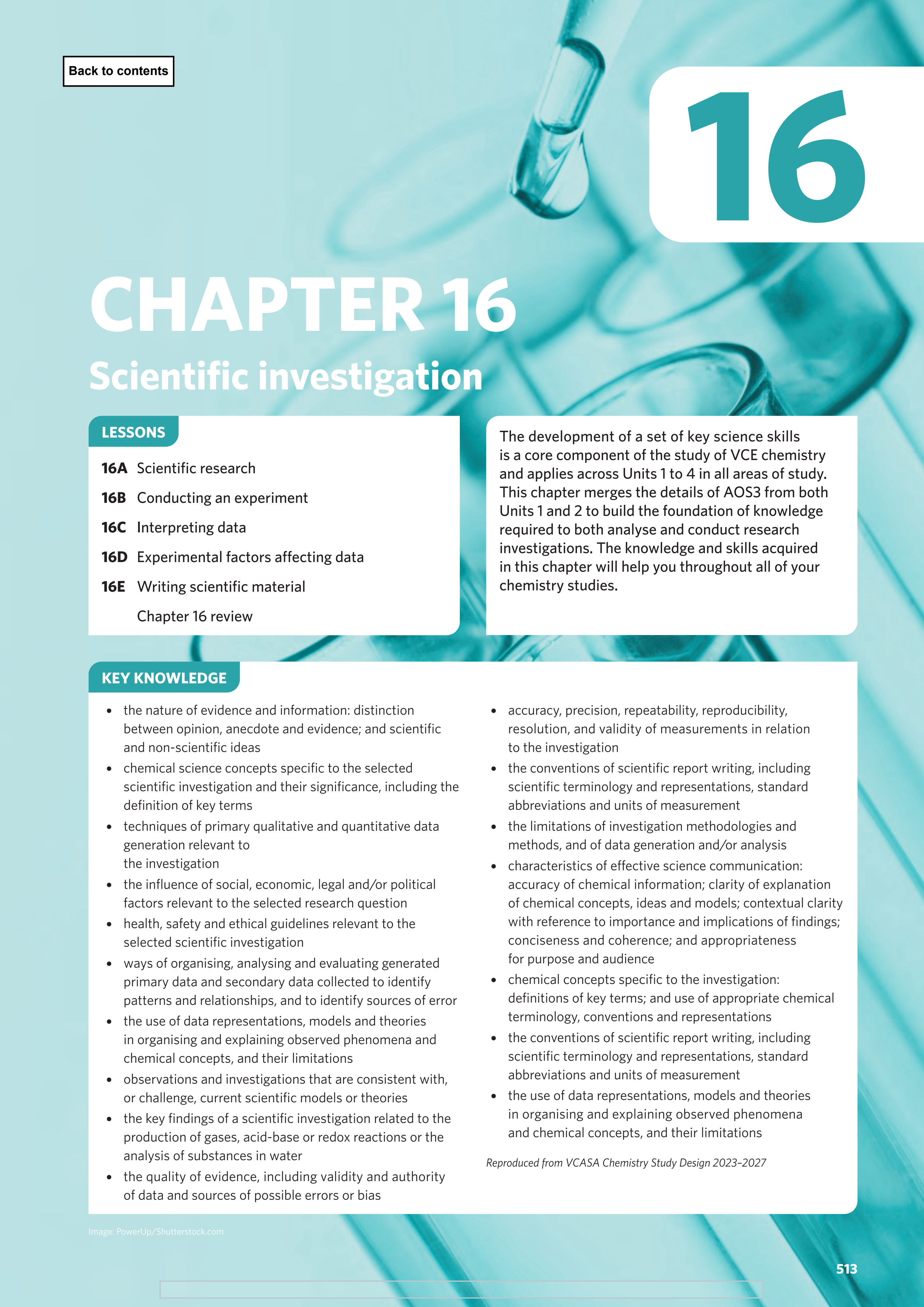

16

CHAPTER 16

Scientific investigation

LESSONS

16A Scientific research

16B Conducting an experiment

16C Interpreting data

16D Experimental factors affecting data

16E Writing scientific material

Chapter 16 review

The development of a set of key science skills is a core component of the study of VCE chemistry and applies across Units 1 to 4 in all areas of study. This chapter merges the details of AOS3 from both Units 1 and 2 to build the foundation of knowledge required to both analyse and conduct research investigations. The knowledge and skills acquired in this chapter will help you throughout all of your chemistry studies.

KEY KNOWLEDGE

- the nature of evidence and information: distinction between opinion, anecdote and evidence; and scientific and non-scientific ideas
- chemical science concepts specific to the selected scientific investigation and their significance, including the definition of key terms
- techniques of primary qualitative and quantitative data generation relevant to the investigation
- the influence of social, economic, legal and/or political factors relevant to the selected research question
- health, safety and ethical guidelines relevant to the selected scientific investigation
- ways of organising, analysing and evaluating generated primary data and secondary data collected to identify patterns and relationships, and to identify sources of error
- the use of data representations, models and theories in organising and explaining observed phenomena and chemical concepts, and their limitations
- observations and investigations that are consistent with, or challenge, current scientific models or theories
- the key findings of a scientific investigation related to the production of gases, acid-base or redox reactions or the analysis of substances in water
- the quality of evidence, including validity and authority of data and sources of possible errors or bias
- accuracy, precision, repeatability, reproducibility, resolution, and validity of measurements in relation to the investigation
- the conventions of scientific report writing, including scientific terminology and representations, standard abbreviations and units of measurement
- the limitations of investigation methodologies and methods, and of data generation and/or analysis
- characteristics of effective science communication: accuracy of chemical information; clarity of explanation of chemical concepts, ideas and models; contextual clarity with reference to importance and implications of findings; conciseness and coherence; and appropriateness for purpose and audience
- chemical concepts specific to the investigation: definitions of key terms; and use of appropriate chemical terminology, conventions and representations
- the conventions of scientific report writing, including scientific terminology and representations, standard abbreviations and units of measurement
- the use of data representations, models and theories in organising and explaining observed phenomena and chemical concepts, and their limitations

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16A Scientific research

STUDY DESIGN DOT POINTS

- the nature of evidence and information: distinction between opinion, anecdote and evidence; and scientific and non-scientific ideas
- chemical science concepts specific to the selected scientific investigation and their significance, including the definition of key terms

16A 16B 16C 16D 16E

1.3.2.1 Scientific vs. non-scientific ideas

2.3.1.1 Linking chemical concepts to investigations

ESSENTIAL PRIOR KNOWLEDGE

- Not all knowledge is scientific

See question 94.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: Blue Planet Studio/Shutterstock.com

Why can something be called scientific?

In this lesson, we will learn about the development of **scientific ideas.** Although we have gained a lot of knowledge over time, only some of that knowledge is considered scientific. Throughout history, humans have been able to gain a lot of knowledge about the world and about ourselves. In early history, discoveries and knowledge were passed on to generations through different methods such as storytelling, dancing, and artefacts. Some of the early knowledge was gained through experiences, intuition, and the senses.

KEY TERMS AND DEFINITIONS

Aim	purpose of an experiment
Controlled variable(s)	variable(s) held constant throughout the experiment
Dependent variable	variable that is measured by the experimenter
Hypothesis	testable statement which predicts the outcome of an experiment
Independent variable	variable that is deliberately manipulated by the experimenter
Non-scientific ideas	ideas that are not developed by following the scientific method
Primary source	source that provides the original raw data
Scientific ideas	ideas that are developed by following the scientific method
Scientific method	procedure used to investigate scientific ideas
Secondary source	source that has interpreted primary sources

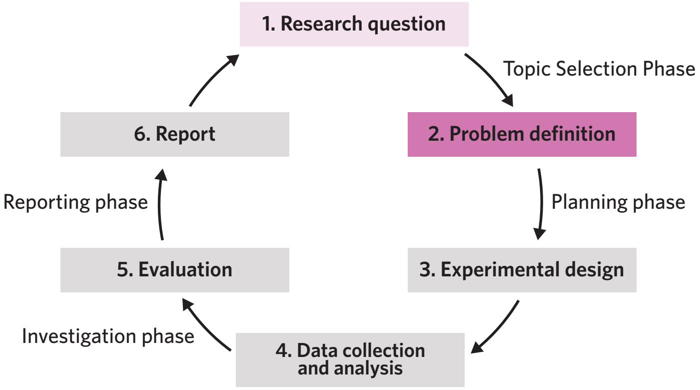

Figure 1 A scientific method

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fic vs. non-scientific ideas 1.3.2.1

Over time, we have developed a more methodical way of developing and explaining new knowledge; a process referred to as the **scientific method**. This method is outlined in the flow chart found on the previous page.

Why is science a process?

The scientific method involves:

- conducting experiments in a way that allows us to have supporting evidence about the conclusions that we make, and therefore the pieces of knowledge that we gain.
- develop predictions and conduct experiments in a controlled environment to collect information that can later be analysed.

One of the key features of the scientific method is that it needs to be reproducible; that is, others should be able to follow the same steps and achieve the same outcomes. The ideas that are developed using the scientific method are known as scientific ideas. The theories that we will study throughout this course have been developed through this process, and can therefore be considered as scientific ideas.

Scientific ideas are presented in a particular format as shown in figure 2. This general structure is followed by all scientists who conduct research and is how scientific journal articles are presented. Developments in scientific knowledge are published in journals and some journals, also known as peer-reviewed journals, are checked by other scientists to ensure that the right processes have been followed and that the conclusions are supported. Ultimately, all of this knowledge contributes to our understanding of the world as shown in figure 3:

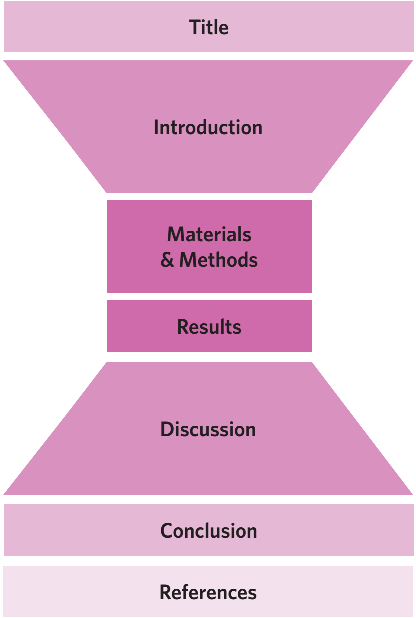

Figure 2 General format used when presenting scientific ideas

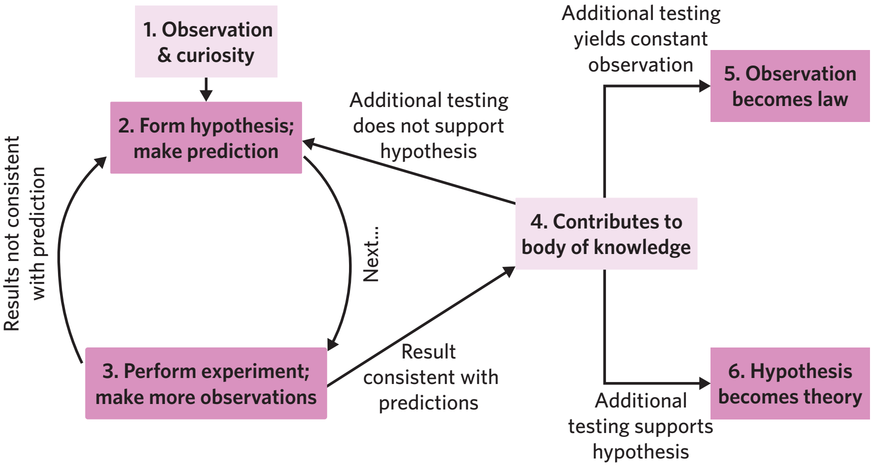

Figure 3 The development of scientific knowledge

There are some ideas that are hard to measure and therefore have not been developed using the scientific method. We call these ideas **non-scientific ideas**. For example, ideas developed through personal experience or tradition that have not been examined can be considered non-scientific.

Progress questions

Question 1

Scientific ideas are purely developed through past experiences.

A. True

B. False

Question 2

Scientific ideas only need to be proven once.

A. True

B. False

Continues →

Question 3

"Art is in the eye of the beholder, and everyone will have their own interpretation", E.A. Bucchianeri.

Art interpretations can be considered to be

A. scientific ideas.

B. non-scientific ideas.

Question 4

I Atoms are the smallest unit of matter...	**II** It can be concluded that atoms have a mass in the centre of their structure
III Based on the results, it can be seen that...	**IV** 1. Set up the cathode ray 2. Place two charged plates around cathode ray

The order in which the above tabs should be placed in order to satisfy the scientific method is

A. I, III, IV, II.

B. I, IV, III, II.

Question 5

Scientific research is built upon previous ideas.

A. True

B. False

Linking chemical concepts to investigations 2.3.1.1

Scientific knowledge is built upon previous ideas. What we know today about chemistry is a result of all of the knowledge that has been developed previously. The discoveries that we make now build upon this knowledge. Therefore, when choosing a research investigation, we need to acknowledge this by linking it to relevant chemical concepts. In doing so, we can ensure that the basis of our research stems from a solid theoretical foundation.

Why is brainstorming an investigation important?

When brainstorming ideas about a specific scientific investigation to pursue, it's useful to choose an area of interest and then conduct background research into what has been already established in the area. Most of the time, this involves finding journal articles of previous research that have investigated questions in the same area/field. By doing this, we are able to gain a better understanding of the current situation, and it also helps to give direction to the kind of research that we can conduct.

For example, say we were interested in investigating water quality. Due to the fact that water is a very important resource, we need to ensure that the water that is available is able to be used for the many different functions in our daily lives. As a result, we need to conduct regular tests to ensure that our water is not contaminated with harmful chemicals and that it has low levels of heavy metals and salts. Figure 4 shows how a research idea can be developed.

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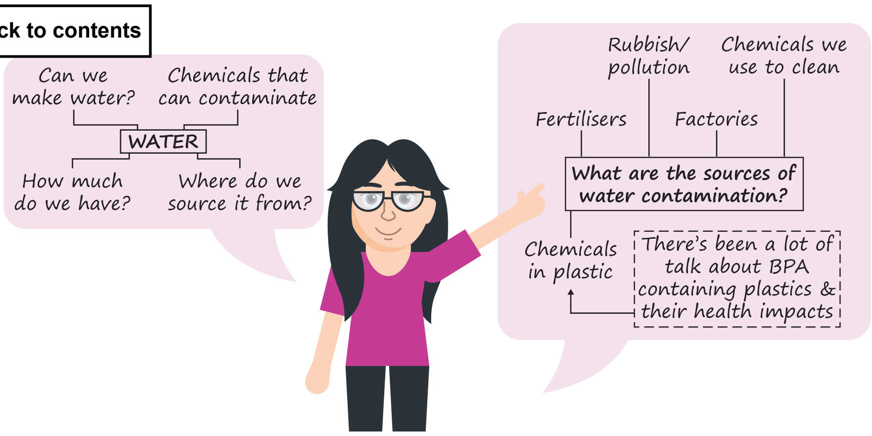


Figure 4 Brainstorming of potential research topics

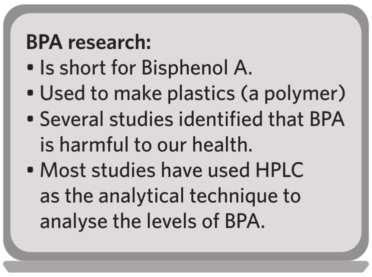
USEFUL TIP

Throughout Chapter 16 you will see the blue bottle icons which are all part of an ongoing investigation that comes together as a poster in lesson 16E.

From this, we can start researching the area to find out more about what is currently understood about the topic as shown in figure 5.

BPA SCIENTIFIC INQUIRY

In this example, we would be conducting research to find out more about BPA.



BPA research:
- Is short for Bisphenol A.
- Used to make plastics (a polymer)
- Several studies identified that BPA is harmful to our health.
- Most studies have used HPLC as the analytical technique to analyse the levels of BPA.

Figure 5 Researching BPA

The best place to gather information about the chemical theory related to different concepts is through reading journal articles and literature reviews related to the topic of interest. These can be original research articles which contain the experimental data and are referred to as **primary sources**, or other resources such as articles or books that have already reviewed and summarised the data and are referred to as **secondary sources**. These types of articles provide a broad overview of the related theory and include related research that can be further explored depending on the specific area of interest. This step is particularly important as it provides us with:

- relevant terminology and definitions that we will need to use throughout the investigation,
- underlying chemical concepts related to the research,
- and, possible methods that can be adapted to our own research.

During this process, it is important to make sure that we use resources that have come from reputable sources to ensure that the information is accurate. This mainly involves asking questions such as:

- Who wrote the article?
  - Are they an expert in the field?
- Where was the article published?
  - Was the article published in a peer-reviewed journal?
- Does the author have affiliations with the relevant industry?
  - Have they been funded by a private company in the industry?

How do we review scientific research?

All scientific research is presented in the format as outlined in figure 2. When reviewing the research, there are some key features that we need to consider.

Although we can get a general overview of the theory and purpose of the study from the introduction, the more important sections of scientific research are the results and methods. By interpreting the data presented in the results section, we are able to further understand the theory being tested. The results section may support the theory by highlighting an expected relationship, or it may question the theory by showing results that do not match the current theoretical understanding. It is important that we are able to evaluate the methodology of the experiment to see whether or not errors may have affected the results, especially in situations where the results deviate from the theory.

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When reviewing the methodology, we need to ask questions like:

- Is this method appropriate for what we are trying to find?
- Are there any potential errors that may occur during the process?

By doing so, we can be more confident with our analysis of the research.

Throughout this chapter, we will take a closer look at what has been discussed above, and how we can make sure that the conclusions from both our own research and the research of others are valid.

How can we develop a scientific investigation?

Once we have a clearer understanding of the topic, we are able to further refine the research question and begin to develop the scope of the research investigation, as well as writing the introduction section of our investigation.

BPA SCIENTIFIC INQUIRY

Is it dangerous to drink water from a plastic water bottle that's been left in a warm environment?

Introduction

Bisphenol A (BPA), as shown in figure 6, is a monomer used to produce a strong plastic polymer known as polycarbonate (Vogel, 2009). Originally developed for medical uses, BPA is now used to produce strong plastics found in everyday products such as bottles, food containers, and pipes. Due to its widespread use, there is growing concern over the health effects of using BPA.

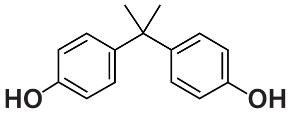

Figure 6 Structure of a bisphenol A (BPA) monomer

In water-containing products, BPA is known to leach into the water over time (Le *et al*, 2008). When consumed, this can result in the disruption of hormone-driven processes in the body (Diamanti-Kandarakis *et al*. 2009). This is potentially dangerous as there are many reactions in the body that are driven by hormones.

In order to be released into the water, BPA needs to be hydrolysed. Due to this, factors such as heat, pH conditions, and microwaves can all affect the extent to which BPA is released into its surrounding environment.

Aim

This experiment aims to test the effect of heat on the release of BPA into water.

Hypothesis

It's expected that if the breakdown of BPA from plastic depends on the heat of the surrounding environment, then an increase in temperature would result in an increase in the concentration of BPA found in the water.

As we can see based on our example experiment, it is important to ensure that our research is based on and is linked to theoretical concepts. Although these links are first introduced here, the chemical ideas need to be woven throughout the whole research investigation. Understanding the underlying scientific theory not only helps us to develop an **aim** and **hypothesis** for the investigation, it also helps us to explain the results of the experiment.

While the aim of an experiment is a statement about the purpose of the research, the hypothesis describes what we expect to happen as a result of the experiment. To develop the hypothesis we first need to understand the main variables in the experiment. More specifically, we need to be able to identify the variable that is being manipulated by the experimenter, the **independent variable**, and the variable that is affected by the change and thus measured, the **dependent variable**. We can then develop a hypothesis that proposes how the independent variable will affect the dependent variable. All variables apart from the independent variable that should be kept consistent throughout the whole experiment are known as **controlled variables**.

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vay to develop a hypothesis can be seen in table 1.

Table 1 Framework for developing a hypothesis. IV stands for independent variable and DV stands for dependent variable.

	Phrase outlining relationship between DV and IV		**trend indicating effect on the DV**		**trend indicating effect on the IV**
If... the DV...	...depends on... ...results from... ...is affected by... ...is directly related to...	**...then...**	...show an increase/ decrease... ...be greater than/ less than... ...be larger/ smaller...	**...when...**	...increased/ decreased... ...greater/ less... ...large/ small...

BPA SCIENTIFIC INQUIRY

For our experiment involving BPA in water bottles, a hypothesis could be:

"If the level of BPA in water is affected by exposure to heat, then an increase in temperature will result in an increase in the level of BPA found in the water."

Progress questions

Question 6

All conclusions from research investigations are completely accurate.

A. True

B. False

Question 7

Textbooks are a

A. primary source.

B. secondary source.

Question 8

The results of every research investigation support the current theory within the topic of interest.

A. True

B. False

Question 9

The hypothesis can help us to determine the

A. controlled variables.

B. independent variable.

Question 10

The observations made during an experiment are mainly due to the

A. controlled variables.

B. independent variable.

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Theory summary

- Scientific ideas are ideas that have been developed through experimentation following the scientific method.
- Non-scientific ideas are ideas that have not been developed as a result of the scientific method.
- Research needs to be conducted prior to an investigation to identify current knowledge within the field of interest.
- The aim and hypothesis identify the purpose and prediction of the research respectively.
- Variables are elements of an experiment that can be changed, controlled or measured.

16A Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 11–13.

The composition of eggshells is very similar to the enamel found on teeth. Col wanted to conduct an experiment to identify factors that contribute to the breakdown of tooth enamel.

Below is a short excerpt of the experimental write-up:

Materials:

- Eggshells
- Cola
- Lemon Juice
- Milk
- Timer
- Four beakers

Method:

1. Label four beakers from A to D.
2. Place a 1 cm piece of eggshell in each beaker.
3. In beaker A, add 20 mL of cola.
4. In beaker B, add 20 mL of lemon juice.
5. In beaker C, add 20 mL of milk.
6. In beaker D, add 20 mL of water.
7. Leave the experiment for 2 hours.
8. Record the results.

Question 11 (1 MARK) 🌶

The main difference between beakers A to D is the

A. volume of substance added.

B. type of liquid used.

C. amount of eggshell used.

D. length of time of the experiment.

Back to contents (1 MARK)

The variable(s) kept constant in this experiment are the

I. volume of substance added.

II. type of liquid used.

III. amount of eggshell used.

IV. duration of the experiment.

A. I only

B. I & II only

C. I & III only

D. I, III & IV only

Question 13 (1 MARK)

Identify the dependent variable in the experiment.

Exam-style

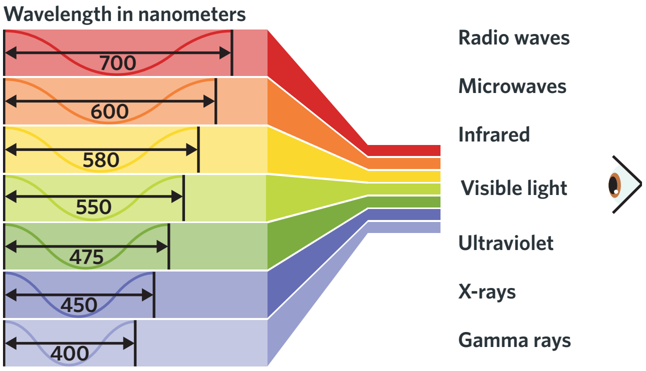
Question 14 (2 MARKS)

It's not quite known where Zodiac signs originated from, however, it is believed that people used to refer to the movement of the sun and stars to predict the future. This eventually led to the development of the zodiac signs, each of which was based on the interaction between constellations and their placement between the sky and the sun during certain periods. Nowadays, many people look to these signs to help them explain current and/or future events. What type of idea would this be considered as? Explain.

Question 15 (3 MARKS)

Bug zappers use visible light to attract bugs.

Edwin wanted to determine which wavelength of the visible light spectrum would be most effective in attracting bugs. The wavelengths of each colour in the visible light spectrum are shown.



a. Identify the independent variable for this experiment. (1 MARK)

b. Write an appropriate hypothesis for this experiment. (2 MARKS)

Hints

14. Ideas do not need to be tested in order to be believed.

15a. The independent variable is deliberately changed in an experiment.

15b. The hypothesis predicts how the independent variable will affect the dependent variable.

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16B Conducting an experiment

STUDY DESIGN DOT POINTS

- techniques of primary qualitative and quantitative data generation relevant to the investigation
- the influence of social, economic, legal and/or political factors relevant to the selected research question
- health, safety and ethical guidelines relevant to the selected scientific investigation

16A 16B 16C 16D 16E

2.3.2.1 Qualitative data collection

2.3.2.2 Quantitative data collection

1.3.11.1 Social, legal, political and economic influences

2.3.5.1 & 2.3.5.2 Ethical considerations & safety guidelines

ESSENTIAL PRIOR KNOWLEDGE

16B Science is a very considered process

16B Data can be both numerical and/or non-numerical

See questions 95–96.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: Sai Thaw Kyar/Shutterstock.com

How important is experimental design?

In this lesson, we will be learning about the factors that we need to take into consideration when selecting and conducting an experiment.

KEY TERMS AND DEFINITIONS

Economic considerations taking into account the associated costs of time and money

Ethical considerations taking into account the effect on other living organisms (e.g. humans and animals)

Legal and/or political factors a combination of factors such as the current political party in power, the efficiency of the current government and their policies, current legal framework, and public attitude towards the economy

Primary data original data collected firsthand by researchers

Qualitative data non-numerical (descriptive) data collected based on observations taken during an experiment

Qualitative analysis technique that determines a non-numerical result

Quantitative analysis technique that identifies the amount of substance present

Quantitative data numerical data collected during experiments

Safety Data Sheet (SDS) document that outlines the health and safety information associated with different materials and chemicals

Secondary data data that has been previously collected that is now accessible to different researchers

Social considerations taking into account the effect on others

Qualitative data collection 2.3.2.1

Different experiments can produce different types of data which means that researchers can select the most appropriate experimental method for their topic of investigation (see figure 1). Some experiments give observational data.

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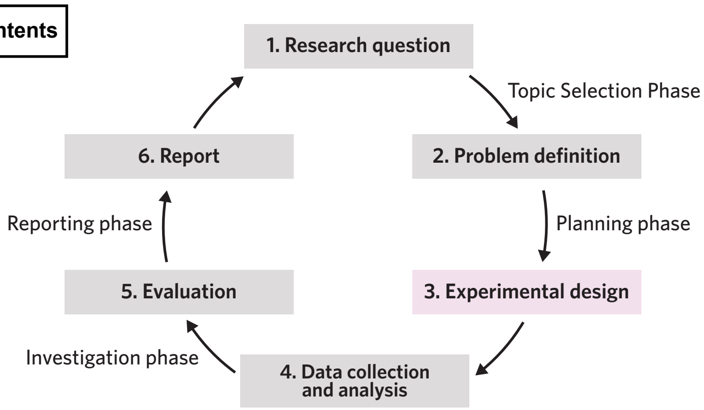

Figure 1 A scientific process

How can observational scientific data be collected?

One of the main ways in which we can continue to develop scientific theory is through conducting experiments to obtain data, which is used to develop conclusions about the area of study. The data that is evaluated can either be collected by the original experimenters, known as **primary data**, or can be collected indirectly from other experiments, referred to as **secondary data**.

During the earlier years in high school, a lot of the data obtained from experiments was based on observations made during reactions. Colour changes, odours and changes of physical state are all examples of the types of data that may have been collected. This type of data is classified as **qualitative data**. Qualitative data is based on observations and often describes (in words) what has occurred as a result of a reaction. **Qualitative analysis** allows us to gain information about what substances are made of.

An example of a technique that can gather qualitative data that will be studied in the course is acid-base titration. One component of acid-base titrations involves observing a colour change, which signifies the point at which the experiment should stop, as shown in figure 2. The advantage of collecting this type of data is that it can give a quick indication as to whether or not a reaction has taken place. Table 1 shows examples of qualitative data collected from different reactions.

USEFUL TIP

Primary data = 'first hand'

Secondary data = 'second hand'

Image: NamPu NiNu/shutterstock.com

Figure 2 Colour change is qualitative data

Table 1 Qualitative data collected from different reactions

Reaction	Example of qualitative data collected
Acid-base titrations	Colour change
Precipitation reactions	Formation of a solid
Esterification	Fruity odour

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Progress questions

Question 1

Qualitative analysis is optimal when

A. an experimenter requires information about colour differences.

B. an experimenter requires information about temperature changes.

Question 2

The image shown represents the collection of

A. qualitative data.

B. quantitative data.

Image: Ajamal/shutterstock.com

Question 3

The image in question 2 shows the collection of

A. primary data.

B. secondary data.

Quantitative data collection 2.3.2.2

Some experiments give numerical data.

How can numerical scientific data be collected?

Unlike qualitative data, **quantitative data** is expressed in numbers and can be used to calculate the amount of a substance. This type of data is important as it can help us determine the amount of a substance that is needed for a desired reaction, and also helps us to predict the products that can be formed. Quantitative data is obtained through **quantitative analysis**. Examples of quantitative analytical techniques that are studied in the course include solubility, gravimetric analysis, acid-base titrations and chromatography – see figure 3. These are shown in table 2.

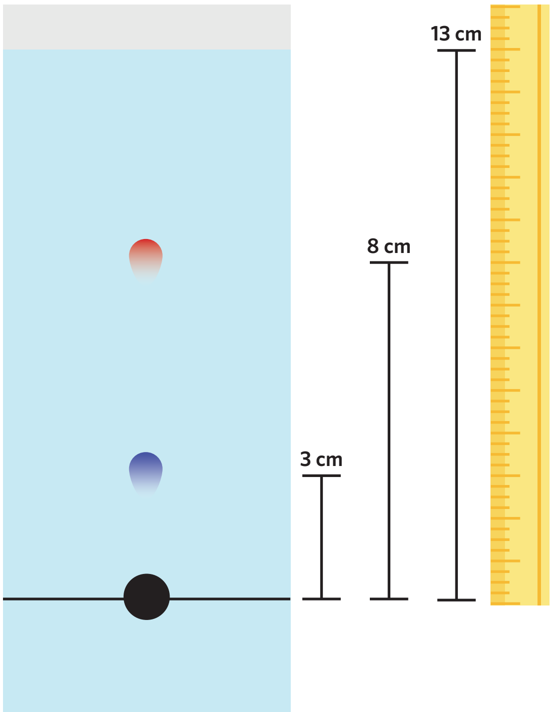

Figure 3 Numerical data (R_f) can be produced using paper chromatography

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itative data collected from different techniques

Technique	Example of quantitative data collected
Solubility	Mass of solute dissolved in a known volume of solvent
Gravimetric analysis	Mass of precipitate formed
Acid-base titrations	Concentration of unknown sample
Chromatography	Measuring the distance moved by different components

When selecting an appropriate technique for an experiment, it is important to understand the type of data that is required to answer the research question. If we pick a technique that is inappropriate for the investigation, the results may not address the research question which could jeopardise the validity of the experiment. We will learn about the concept of validity in greater depth in the following lessons.

USEFUL TIP

Quantitative data will normally have a unit of measurement e.g. mass in g, solubility in g mL^{-1}, and concentration in mol L^{-1}.

Progress questions

Question 4

This image shown represents the collection of

A. qualitative data.

B. quantitative data.

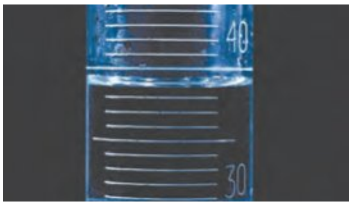

Image: ggw/shutterstock.com

Question 5

The image in question 4 shows the collection of

A. primary data.

B. secondary data.

Question 6

Secondary data can be

A. quantitative only.

B. both quantitative and qualitative.

Social, legal, political and economic influences 1.3.11.1

Scientific research has greater implications outside the scientific community.

How can scientific research affect society?

Regardless of the type of experiment conducted, when designing a research investigation, we need to make sure that we consider the wider implications of the study. Questions that could be asked during this process include:

- Who will be impacted by the results of the study?
- How will the results of the study contribute to the decision-making process of a governing political party?
- What are the implications of the study with respect to what is legal?

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Although it may seem like the experiment we conduct may not have a large impact on the wider community, the decisions that are made today are a result of the mountain of work that has been done by scientific researchers. For example, the different strategies and agreements that have been developed to reduce our impact on the environment come from the countless studies that have been conducted in all fields of science. When conducting research, we need to think about the:

- **Social considerations** – who is/are affected by the research?
- **Economic considerations** – what are the associated money and resourcing costs?
- **Legal considerations** – how is the research affected by the law?
- **Political considerations** – does the governing body have an influence?

The results from scientific research have flow-on effects that can impact many people and industries in different ways. Figure 4 shows an example of the potential effects of scientific research around dichlorodiphenyltrichloroethane (DDT), a chemical that was used as an insecticide.

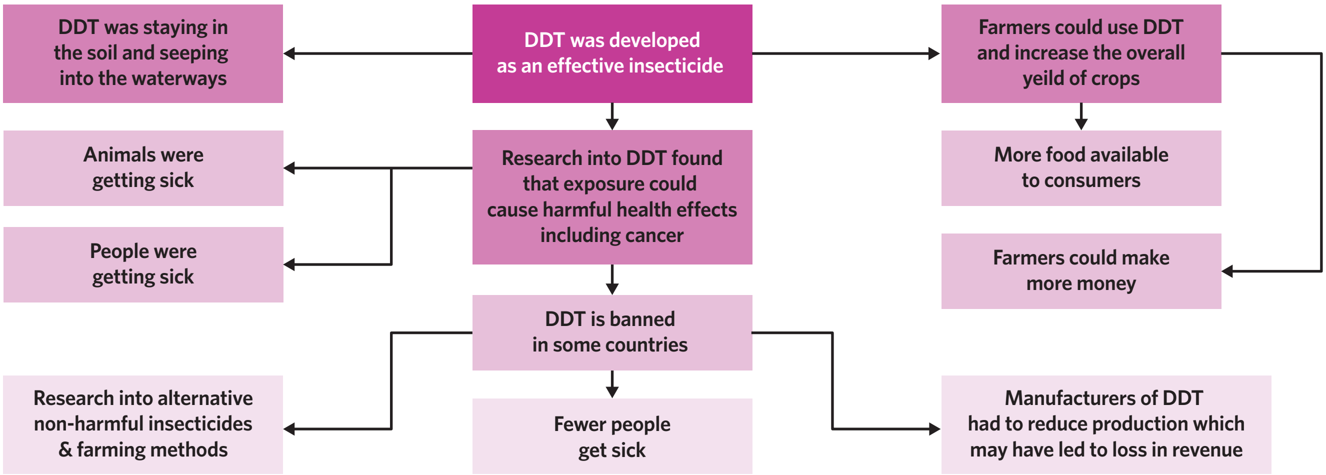

Figure 4 Illustration of potential impacts of scientific research

Progress questions

Question 7

All scientific research findings have positive implications for all areas of society.

A. True

B. False

Question 8

Scientists are responsible for all misuse of research findings.

A. True

B. False

Ethical considerations & safety guidelines 2.3.5.1 & 2.3.5.2

Ethics and safety guidelines should always be taken into account when conducting scientific research.

How do ethics influence scientific research?

Conducting scientific research is about more than just scientific theory. Due to the potential implications of the work, we are all responsible for the appropriate conduct and use of the knowledge obtained. As such, we need to make sure that we adhere to **ethical considerations** when conducting research. In other words, we need to make sure that we're trying our best to do the 'right' thing from a moral and principled perspective.

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about all of the ethical issues that may arise from chemistry-related studies, we can categorise them into two main areas: the internal and external domains. The internal domain consists mainly of the issues related to the experiment itself whereas the external domain consists of the wider societal implications of the research. Table 3 shows examples of the two types of ethical domains.

Table 3 Examples of ethical issues in internal and external domains

Internal domain	External domain
• Conducting experiments safely • Plagiarism • Data falsification • Acknowledgement of previous research (i.e. citations) • Animal/human testing	• Development of new substances • Environmental impacts • Societal impacts

Due to the importance of ethics in chemistry research, scientists met to develop a document known as the Hague Ethical Guidelines. This guideline outlines the main principles that we should try to abide by when conducting chemical research. This document is not directly related to the VCAA study design however is a good resource that gives further information about ethics in chemistry.

Why is it important to protect the safety of everyone involved in scientific investigations?

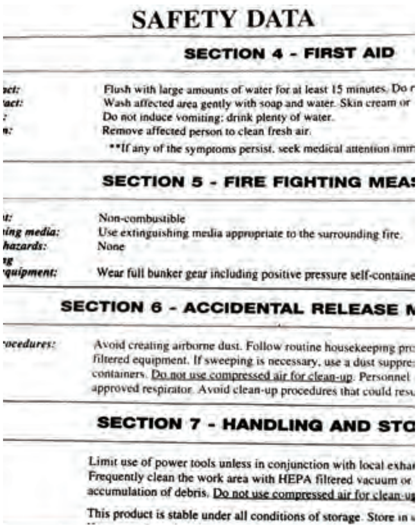
All scientific experiments come with their own safety risks. There are many different types of risks, which vary depending on the type of experiment conducted. These risks can arise through the use of certain substances as well as from the experimental procedure itself.

Prior to conducting any experiment, it is important that the experimenter undertakes a thorough risk assessment and refers to the **Safety Data Sheet (SDS)** of all of the substances that will be used as part of the experiment. The SDS contains all of the health and safety information relating to materials/substances that will be used, including strategies and guidelines to help mitigate the effects of each. Figure 5 shows sections that are often included in an SDS.

As part of the experimental process, scientists use an SDS to develop a risk assessment for the experiment to analyse the potential risks and outline specific strategies to minimise them. For example, experiments that result in the production of a potentially harmful gas would be required to be conducted in a fume hood. Other examples of safety measures that can be taken during an experiment include:

- Wearing protective equipment e.g. lab coat, safety glasses, etc.
- Knowing where and how to use the safety showers.
- Proper disposal of chemical waste.

For an experiment, we need to collate all of the information related to the substances that will be used in the experiment and develop a plan to minimise safety risks. This should be included as part of the 'Materials and Methods' section of a scientific investigation.



SAFETY DATA

SECTION 4 - FIRST AID

act: Flush with large amounts of water for at least 15 minutes. Do
tact: Wash affected area gently with soap and water. Skin cream or
: Do not induce vomiting; drink plenty of water.
n: Remove affected person to clean fresh air.
**If any of the symptoms persist, seek medical attention imm

SECTION 5 - FIRE FIGHTING MEA

t: Non-combustible
ing media: Use extinguishing media appropriate to the surrounding fire.
hazards: None
g
quipment: Wear full bunker gear including positive pressure self-containe

SECTION 6 - ACCIDENTAL RELEASE

ocedures: Avoid creating airborne dust. Follow routine housekeeping pr
filtered equipment. If sweeping is necessary, use a dust suppre
containers. Do not use compressed air for clean-up. Personnel
approved respirator. Avoid clean-up procedures that could res

SECTION 7 - HANDLING AND STO

Limit use of power tools unless in conjunction with local exha
Frequently clean the work area with HEPA filtered vacuum or
accumulation of debris. Do not use compressed air for clean-u
This product is stable under all conditions of storage. Store in

Image: JVinocur/wikimedia.org

Figure 5 Sample SDS of a chemical substance

Progress questions

Question 9

Scientists who conduct research in the same area are able to use each other's findings without needing to acknowledge the source.

A. True

B. False

Continues →

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Question 10

When a new substance or scientific technique is developed, it should be shared with the world immediately.

A. True

B. False

Question 11

Scientists need to conduct a primary test of all the substances they are going to use before performing their actual experiment in order to find out the potential risks of using the substances.

A. True

B. False

Question 12

Safety considerations for experiments only need to be documented after the experiment has been conducted, and only if issues arise during the experiments.

A. True

B. False

Theory summary

- Qualitative data is mainly observational.
- Quantitative data gives us numerical information about substances.
- Scientific findings can have social, legal and/or political, and economic implications.
- There are many ethical concerns related to chemical investigations that should be considered by the people conducting the research.
- It is very important to make sure that scientists are well informed about the nature of the experiment they are conducting so that they can put strategies in place to minimise the potential risks.

16B Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 13–15.

When increasing atmospheric concentrations of carbon dioxide (CO_2) are discussed in the context of climate change, a silver lining is often postulated: 'CO_2 is plant food.' That is, higher concentrations of CO_2 are generally acknowledged to stimulate plant photosynthesis and growth, with potential benefits for the productivity of the cereal crops that remain the world's most important sources of food. Cereal crops feed not just humans but also the animals that are important sources of protein for many. Since the mid-1960s, cereal production increased by approximately a billion tons, yet accelerated progress in agriculture is needed to keep pace with population growth anticipated to reach between 9–10 billion by 2050 [3], and to achieve Sustainable Development Goal (SDG) 2 to end hunger, achieve food security and improved nutrition, and promote sustainable agriculture.

But food security is about more than just production. While increases in CO_2 may make some crops grow more quickly, research shows that higher CO_2 concentrations can also reduce the nutritional quality of staple crops, from potatoes to barley, rice to wheat.

Source: https://journals.plos.org/plosmedicine/article?id=10.1371/journal.pmed.1002600

Citation: Ebi, K. L., & Ziska, L. H. (2018). *Increases in atmospheric carbon dioxide: Anticipated negative effects on food quality*. PLoS medicine, 15(7), e1002600. https://doi.org/10.1371/journal.pmed.1002600

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(1 MARK)

The effect of the CO_2 increase can be experienced by

A. direct consumers.

B. farmers.

C. food manufacturers.

D. all of the above

Question 14 (1 MARK)

What effect would an increase in CO_2 have on the health of consumers?

A. Consumers will be able to eat less to obtain the same amount of nutrients.

B. Consumers will have to eat more to obtain the same amount of nutrients.

C. Consumers will spend less on food.

D. Consumers will eat less cereal.

Question 15 (3 MARKS)

Describe an economic issue evident in the passage.

Exam-style

Question 16 (2 MARKS)

Cecelia and her team were conducting experiments to develop a new type of drug targeted at supporting patients with diabetes. As part of their work, they conducted extensive research on diabetes and the current available therapies and came across a study that was similar to theirs. The team used the results of the study to direct their research.

a. What type of data, primary or secondary, was Cecelia and her team collecting to do this research? (1 MARK)

b. In an effort to finish the research on time, Cecelia copied and pasted the research conducted by the other team and used it as her own. What is the main ethical issue around what she did? (1 MARK)

Question 17 (5 MARKS)

Luis was given two unknown samples of metallic substances and asked to determine the name of each. In his experience, he knew that he could use the flame test to differentiate between the two metallic substances.

Upon conducting the experiment, Luis noticed that each sample gave off a different coloured flame.

a. Identify whether this type of analysis is qualitative or quantitative. (1 MARK)

Luis wanted to use the results of the test to find how much of each substance was present.

b. Is it possible for Luis to calculate the amount of substance present based on his results? Explain. (4 MARKS)

Question 18 (6 MARKS)

One of the methods currently used to develop vaccines involves scientists intentionally infecting individuals with viruses, bacteria or parasites. Individuals can be given a vaccine against a particular disease and exposed to the pathogen responsible for causing the disease. In doing so, scientists are able to determine whether or not the vaccine is effective by measuring the extent to which the person shows symptoms of being ill. In comparison to traditional methods of vaccine development, this method is much quicker and requires only a few rounds of testing.

Alternatively, scientists can spend years developing vaccines that are rigorously tested in other contexts before being trialled on humans.

a. Identify the main concern discussed in the given passage. (1 MARK)

b. Identify one reason why scientists would choose to conduct experiments on humans. (2 MARKS)

c. Discuss the major disadvantages of both methods of vaccine development. (3 MARKS)

Back to contents (5 MARKS)

Elephant toothpaste can be made using reactants such as hydrogen peroxide.

A section from the SDS for hydrogen peroxide is shown:

Section 2: Hazard(s) identification		
2.1. Classification of the substance or mixture		
GHS–US classification		
Oxidising liguids category 1	H271	May cause fire or explosion; strong oxidiser
Acute toxicity (oral) category 4	H302	Harmful if swallowed
Acute toxicity (inhalation) category 4	H332	Harmful if inhaled
Skin corrosion/irritation category 1A	H314	Causes severe skin burns and eye damage

a. Identify two pieces of personal protective equipment that would allow you to safely conduct this experiment. (2 MARKS)

b. As part of the preparation for this experiment, a student asked the lab technician, who is not involved in the implementation of the method, to develop a risk assessment. Is the lab technician the appropriate person to develop the risk assessment? Explain. (3 MARKS)

Question 20 (4 MARKS)

Jo wanted to determine the concentration of different minerals in bottled mineral water. As part of the investigation, three different brands of mineral water were used.

a. Identify the independent variable. (1 MARK)

b. List three variables that should be kept constant throughout the experiment. (3 MARKS)

Question 21 (7 MARKS)

The following is an excerpt from a research article:

Background

Dietary sugar, especially in liquid form, increases risk of dental cavities, adiposity, and type 2 diabetes. The United Kingdom Soft Drinks Industry Levy (SDIL) was announced in March 2016 and implemented in April 2018 and charges manufacturers and importers at £0.24 per litre for drinks with over 8 g sugar per 100 mL (high levy category), £0.18 per litre for drinks with 5 to 8 g sugar per 100 mL (low levy category), and no charge for drinks with less than 5 g sugar per 100 mL (no levy category). Fruit juices and milk-based drinks are exempt. We measured the impact of the SDIL on price, product size, number of soft drinks on the marketplace, and the proportion of drinks over the lower levy threshold of 5 g sugar per 100 mL.

Source: https://journals.plos.org/plosmedicine/article?id=10.1371/journal.pmed.1003025

Citation: Scarborough, P., Adhikari, V., Harrington, R. A., Elhussein, A., Briggs, A., Rayner, M., Adams, J., Cummins, S., Penney, T., & White, M. (2020). *Impact of the announcement and implementation of the UK Soft Drinks Industry Levy on sugar content, price, product size and number of available soft drinks in the UK, 2015–19: A controlled interrupted time series analysis.* PLoS medicine, 17(2), e1003025. https://doi.org/10.1371/journal.pmed.1003025

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the information, write a hypothesis for what will happen to the proportion of drinks with over 5 g sugar per 100 mL as a result of the introduction of SDIL. (2 MARKS)

b. What is the primary type of ethical concern in this experiment? (1 MARK)

c. Describe a potential implication of the SDIL and how it would affect society. (4 MARKS)

Hints

16a. The data type depends on the source of the data.

16b. Scientific research reports should be written in the experimenter's own words.

17a. There are two main data types based on what the data measures.

17b. To calculate the amount of substance, numerical data needs to be collected.

18a. Special consideration needs to be taken into account when experimenting on humans.

18b. It is important to make medicine available in a timely manner.

18c. The disadvantages of each method depend on how we analyse the experiment.

19a. Safety precautions for experiments are selected based on the risks relating to substances used.

19b. As part of the planning process, researchers are required to analyse the potential risks of the experiment.

20a. Independent variables are changed by the experimenter.

20b. All conditions apart from the independent variable should be kept constant in an experiment.

21a. Generating a hypothesis requires us to understand the purpose of the experiment.

21b. In general, ethical concerns relate to the experiment or the implications of the research itself.

21c. Taxes are a way in which the government can help to shift behaviour that is more beneficial for society.

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16C Interpreting data

STUDY DESIGN DOT POINTS

- ways of organising, analysing and evaluating generated primary data and secondary data collected to identify patterns and relationships, and to identify sources of error
- the use of data representations, models and theories in organising and explaining observed phenomena and chemical concepts, and their limitations
- observations and investigations that are consistent with, or challenge, current scientific models or theories
- the nature of evidence and information: distinction between opinion, anecdote and evidence; and scientific and non-scientific ideas
- the key findings of a scientific investigation related to the production of gases, acid-base or redox reactions or the analysis of substances in water

16A 16B **16C** 16D 16E

2.3.9.1 & 1.3.10.1 Identifying patterns in data

2.3.7.1 & 1.3.2.1 Drawing conclusions from data

2.3.14.1 Linking chemical concepts to conclusions

ESSENTIAL PRIOR KNOWLEDGE

16B Data collected from experiments is analysed

16B Data is used to develop a deeper understanding

See questions 97–99.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: bigshot01/Shutterstock.com

Why is data the key finding in an investigation?

In this lesson, we will learn about the different ways in which we can interpret and represent data obtained from an experiment.

KEY TERMS AND DEFINITIONS

Anecdote the communication of an abstract idea about a person, place, or thing through the use of a story or narrative

Control group (in chemistry) sample that is subjected to the same conditions as all other samples without the independent variable

Opinion a view or judgement formed about something, not necessarily based on fact or knowledge

Identifying patterns in data 2.3.9.1 & 1.3.10.1

Experiments generally show a relationship between the variables being studied (figure 1).

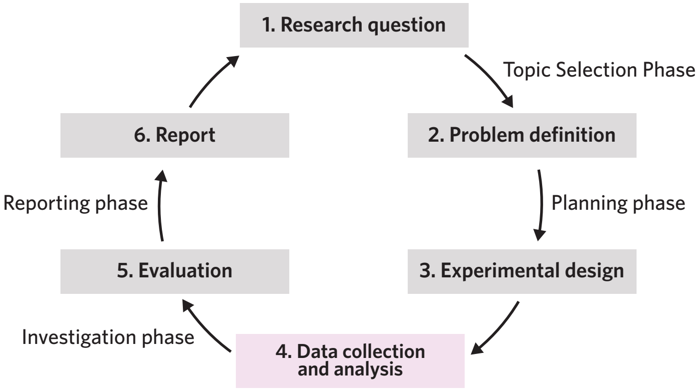

Figure 1 A scientific method

How can patterns in data be identified?

The main purpose of an experiment is to identify a possible relationship between the independent and dependent variables. Unless there are flaws in the experiment, we will generally be able to see a pattern emerge from the data collected.

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esent data can make trends easier to spot. For example, it is easier to identify trends in quantitative data when it is represented as a line graph than it would be if the same data was presented as a pie chart. Table 1 shows the common forms of data representation and the type of data they are mostly used for.

Table 1 Characteristics of common data representation formats

Type of data representation	Example	Type of data being represented	Advantage
Line graph	Calibration curves	Quantitative	Useful for displaying trends over time and for predicting trends over time
Bar graph	Water sampling in different areas	Both qualitative and quantitative	Easy to visualise and compare different categories of data
Pie chart	Composition of fuel	Both qualitative and quantitative	Used to show proportions (e.g. percentages)
Table	Gravimetric analysis	Both qualitative and quantitative	Helps to organise and compare multiple sets of values

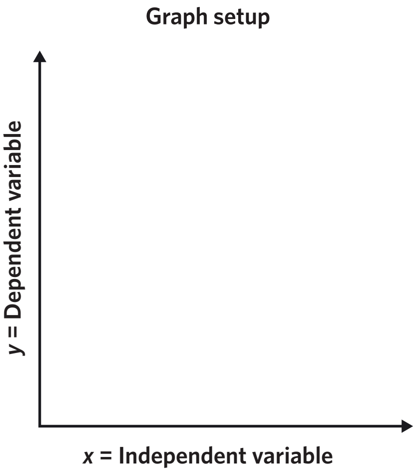

Figure 2 Representation of data in graphs

When constructing and analysing graphs, it is important to know the variable that is being represented on each axis. As seen in figure 2, the y-axis represents the dependent variable, whereas the x-axis represents the independent variable.

When interpreting graphs, we need to think about how the independent variable affects the dependent variable. For example, consider the graph shown in figure 3. In the graph, we can see that the dosage of medicine given to a dog (vertical axis) increases as the weight of the dog (horizontal axis) increases.

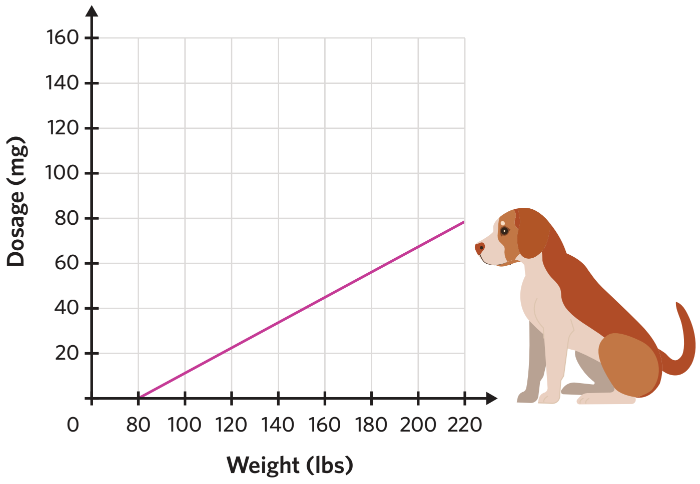

Figure 3 Graph showing information about the dosage of medicine for dogs

Regardless of whether we are interpreting our own data collected from our experiment, or we are interpreting experimental data from an experiment in a journal article, we are using the same strategy to identify trends. Being able to do so will help us determine the accuracy of the claims made by experimenters about their own research.

Figure 4 shows the calibration curve developed from the use of BPA standards. As the concentration of BPA increases, the area under the curve also increases We can use this graph to identify the concentration of BPA in the different samples in our experiment by cross referencing these values with the calibration curve.

Based on figure 5, it can be concluded that when bottles are left in a higher temperature, there is a greater increase in the concentration of BPA in the water.

BPA SCIENTIFIC INQUIRY

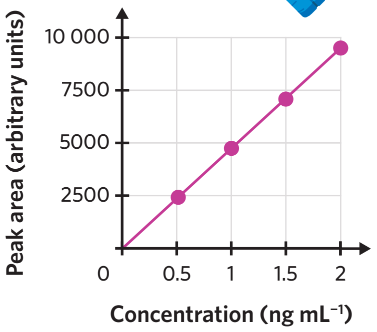

Figure 4 Calibration curve of known standards of BPA

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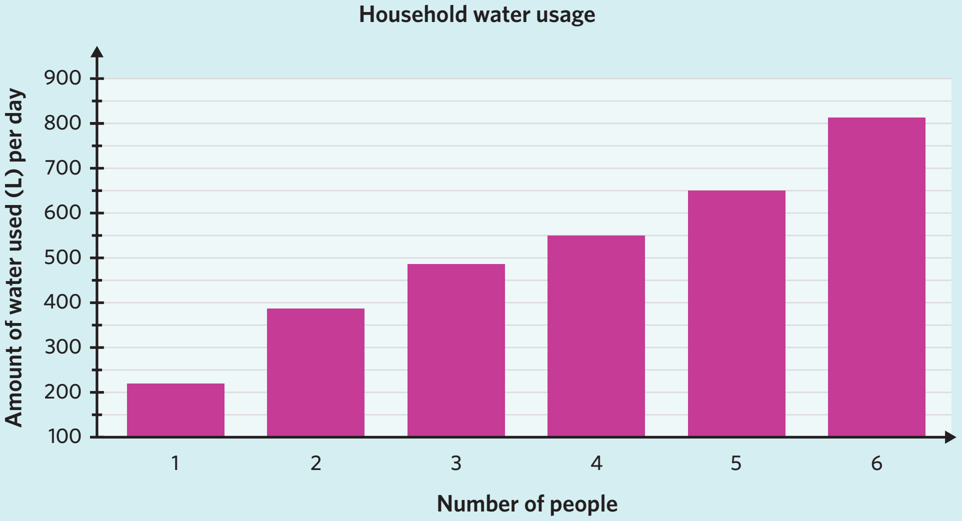
PLE 1

Water companies use data to estimate the water usage of different sized households. The graph provided shows the amount of water used based on the number of people in the house per day.



Predict the amount of water used in a household per day containing seven people.

What information is presented in the question?

A bar graph representing water usage.

What is the question asking us to do?

Predict the amount of water used in a house with seven people.

What strategies do we need in order to answer the question?

1. Identify the general trend in the data presented in the graph.
2. Find the average change in the amount of water used as the number of people increases.
3. Determine a possible range for the water used per day for seven people.

Answer

Overall, it can be seen that the more people there are in the household, the more water is used per day.

Upon closer analysis of the increase in water usage per person added, it can be seen that, on average, there is approximately a 130 L increase with each additional person.

Therefore, it is predicted that in a house of seven people, the water usage could range from approximately 860–990 L.

BPA SCIENTIFIC INQUIRY

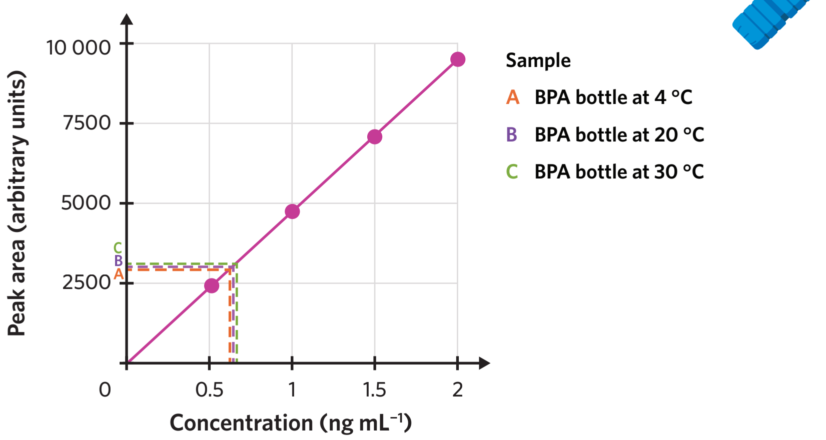


Figure 5 Determination of the concentration of BPA in samples A-C

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Progress questions

Question 1

Experimental data is always shown as a graph.

A. True

B. False

Question 2

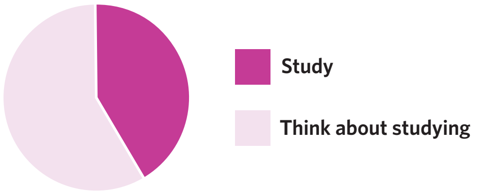
The type of data collected helps to determine the

A. way in which data is presented.

B. variables in the experiment.

Question 3

The figure shown would be best used to represent



A. qualitative data.

B. quantitative data.

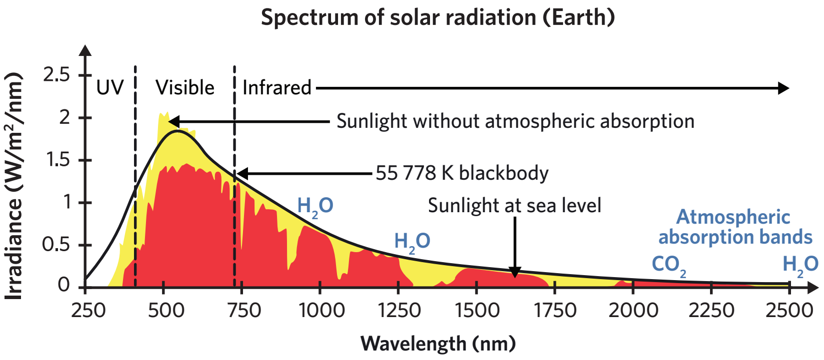
Question 4

Generally speaking, experimental data shows

A. trends between the controlled and independent variables.

B. trends between the independent and dependent variables.

Question 5



In the graph shown, wavelength represents the

A. independent variable.

B. dependent variable.

Question 6

In the graph provided in question 5, the ___________ is affected by the ___________.

A. wavelength, irradiance

B. irradiance, wavelength

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INQUIRY

Back to our experiment. Of the six samples, samples A-C involve BPA-containing bottles whereas D-F do not. The purpose of D-F is to act as a control to show that the changes in BPA concentration were due to the BPA present in the bottles that were subjected to different conditions, instead of just the change in temperature. To ensure the quality of our experiment, all other variables, such as volume of water and duration of the experiment were controlled (and therefore were our controlled variables).

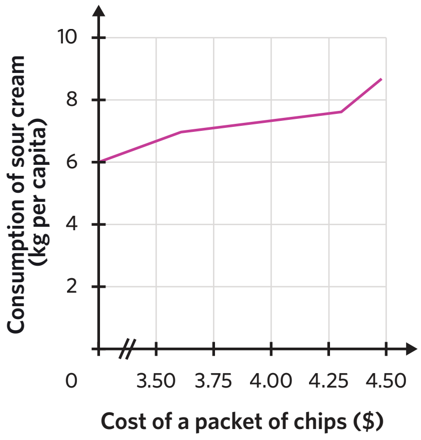

Figure 6 Line graph representing the relationship between sour cream consumption and cost of chip packets

Drawing conclusions from data 2.3.7.1 & 1.3.2.1

The trends represented by the data can help us develop a better understanding of the scientific theory related to the research.

How do we make conclusions about data?

The data collected through scientific research can help address the hypothesis, and the nature of this data is distinct from either non-scientific **opinion** or **anecdote**. The results of an experiment do not necessarily have to support the hypothesis; in fact, the greatest learnings come from experiments where the data challenges the hypothesis.

Firstly, to know that the trend in the experimental data collected was due to the independent variable, we need to make sure that a **control group** is present as a point of comparison to show that any change in the experiment was due to the impact of the independent variable. A control in an experiment is set up in a way that doesn't allow the independent variable to impact the results (i.e. the independent variable isn't applied or changed in the control sample). This is important as it allows us to attribute the changes in the experiment to the independent variable, and the effect of other variables.

One point to remember is that the trends seen in the data collected in an experiment do not necessarily illustrate a causal relationship; that is, it does not mean that the changes seen in the dependent variable are solely due to the independent variable. There may be a correlation, which is that the changes observed in the dependent variable are associated with the independent variable, but it is possible that other variables could also be responsible for this change. As such, we can't say that the result was caused by the independent variable; we can't establish causation. For example, based on the data presented in figure 6, it can be seen that as the cost of a packet of chips increases, the amount of sour cream consumed also increases. However, this does not necessarily mean the cost of chips is the reason why people eat more sour cream.

Furthermore, there may be factors that affect the validity and reliability of our conclusions – a concept that we will explore in the next lesson.

BPA SCIENTIFIC INQUIRY

For our water bottle experiment, we need to unpack the data to understand the implications of our observations. This will be the bulk of the discussion section of our experimental report.

Discussion

Compared to control samples D-F in which water was placed in a BPA-free bottle, the presence of BPA in the water bottles resulted in a greater concentration of BPA in the water samples. For samples A-C, it can be seen that an increase in temperature is correlated with an increase in the concentration of BPA measured. The samples placed in the 30 °C and 20 °C environments and in the fridge measured BPA concentrations in ng mL^{-1} of 0.63, 0.61 and 0.58 respectively.

BPA can be released into the contents of water bottles when the plastic begins to degrade. As shown in previous studies, an increase in temperature results in an increase in the degradation of plastic (Biedermann-Brem & Grob., 2008). In particular, an increase in temperature results in the increase in the rate at which the plastics are degraded. In the context of BPA-containing plastic water bottles, this could lead to an increase in the rate at which BPA is released into the water. This trend is shown in the results of this experiment. As can be seen in the experimental results, there was a greater concentration of BPA found in the sample subjected to the highest temperature (sample C) given that all samples were tested over the same period of time. It is possible then that at higher temperatures, there could be a greater release of BPA from plastic, resulting in a higher concentration of BPA recorded.

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Progress questions

Question 7

For an experiment measuring the effect of the amount of water used on the growth of a sunflower seed, a suitable control group would be

A. a sunflower seed receiving twice the amount of water compared to all of the other sunflower seeds.

B. a sunflower seed receiving no water.

Question 8

The following statement was made about an experiment. “As the temperature increases, it can be seen that students complete more homework.”

In this experiment, it is most likely that

A. the change in temperature causes students to do more homework.

B. the change in temperature is related to students completing homework.

BPA SCIENTIFIC INQUIRY

In our water bottle experiment, the main ideas related to the experiment involve

- concentration
- bonding
- temperature and chemical reactions

As a result, discussions about and explanations of the data need to be based on our current understanding of these key ideas. This thought process also applies when we are evaluating other research.

Linking chemical concepts to conclusions 2.3.14.1

The conclusions developed from research need to be explained by scientific theory.

How do we link the chemistry into the conclusions?

When evaluating research, the conclusions drawn need to be supported both by the experimental data collected during the experiment as well as existing theory. Regardless of whether or not the hypothesis is supported, it is important that we recognise the scientific theory that could explain the observations made.

For the research you will conduct this year, it will be mainly based around the concepts that you will be learning about. These are briefly outlined in table 2.

Table 2 Key research concepts in VCE chemistry units 1 and 2

Concept	Brief description
Redox reactions	Transfer of electrons between oxidation and reduction reactions
Production of gases	Combustion reactions
Solubility	Ability of substances to dissolve
Concentration	Amount of a substance in a given volume of water

Progress questions

Question 9

It is sometimes difficult to prove causation because

A. results are always affected by the same variable.

B. results can be affected by multiple factors.

Question 10

Conclusions in research will always prove a hypothesis correct.

A. True

B. False

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Theory summary

- Results generally show a relationship between the independent and dependent variables.
- A control group is important to show that the changes were due to the variable being tested.
- Anecdote and opinion are not considered to be scientific.
- The way in which we represent data depends on the type of data collected.
- Conclusions generated from data need to be supported by scientific theory.

16C Questions

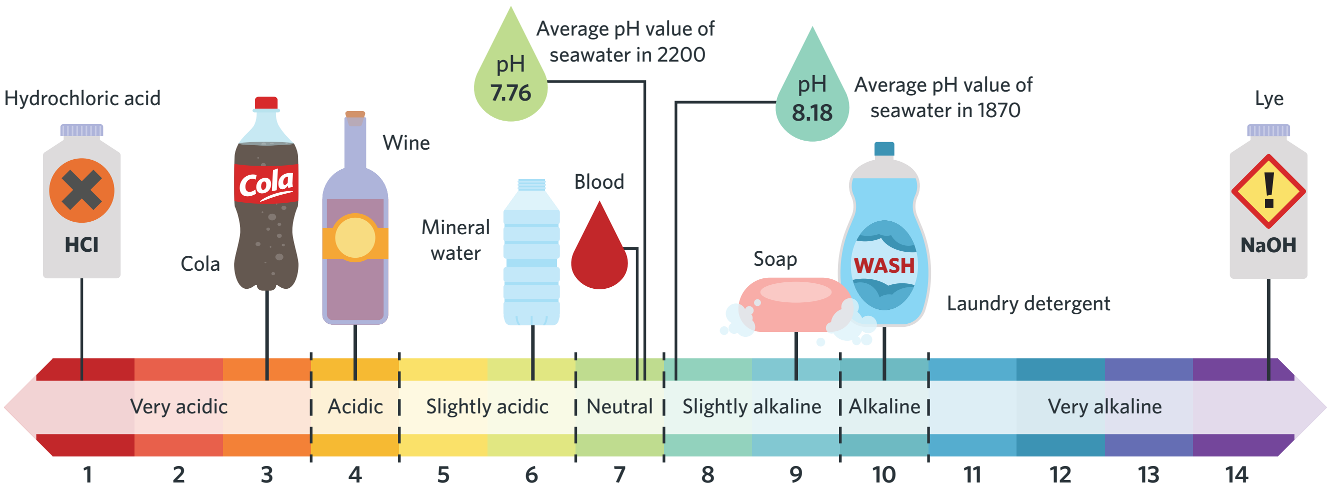
Mild Medium Spicy

Deconstructed

Use the following information to answer questions 11–13.

Scientists use pH to measure how acidic or basic substances are. For example, laundry detergents are considered basic whereas wine is considered acidic.

pH Scale: What is Acidic, What is Alkaline?



The difference may seem small, but the decline in the pH value from 1870 to 2100 would mean a 170 % increase in acidity. Much smaller changes already pose problems for many sea creatures.

Question 11 (1 MARK)

Substances that are neutral have a pH of approximately

A. 5.

B. 6.

C. 7.

D. 8.

Question 12 (1 MARK)

Generally speaking, when cola is added to water, the solution would become

A. more acidic.

B. more basic.

C. less acidic.

D. neutral.

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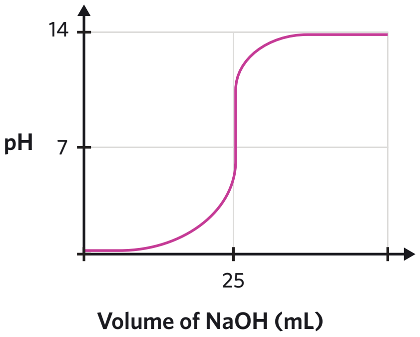
(2 MARKS)

The graph provided shows the pH levels in a solution to which lye (NaOH) was added over time.



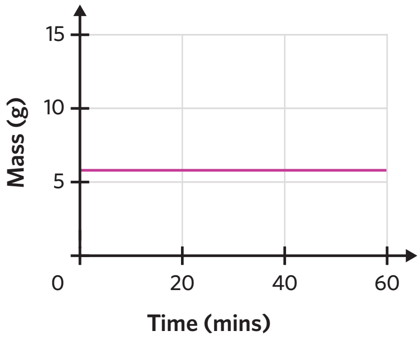
Would the original solution (before NaOH was added) be considered as acidic or basic? Explain.

Exam-style questions

Question 14 (3 MARKS)

Our current understanding of science is built from the knowledge developed through a large body of research. Consider the following graph, which represents data collected from an experiment, where a beaker containing a chemical reaction was left on a mass balance and weighed over time.

a. Identify the trend shown in the graph. (1 MARK)



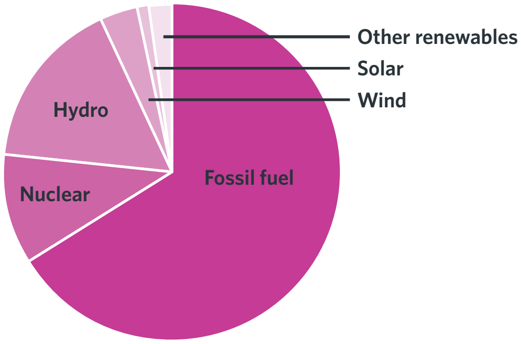
b. Use any of the following chemical concept(s) to justify the conclusion(s) drawn from the graph provided. (2 MARKS)

- **i.** The law of conservation of mass states that no mass is ever lost or gained overall.
- **ii.** Energy is never ‘lost’, it is merely transformed.
- **iii.** There are approximately 6.02×10^{23} atoms/molecules/ions in a mol.

Question 15 (4 MARKS)

The world requires different types of energy in order to sustain all of the work that occurs. The graph provided shows the world’s electricity sources in 2015.

a. Identify the overall breakdown of energy sources shown in the graph. (1 MARK)



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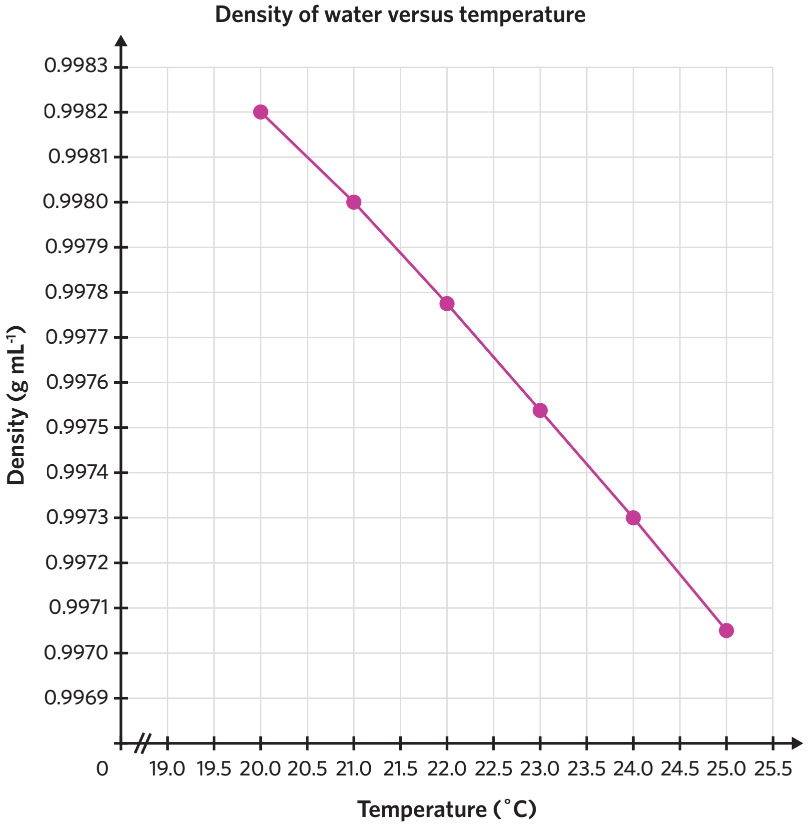
e formed over millions of years and generally extracted through deep drilling into the earth. The use of fossil fuels produces substances such as carbon dioxide, sulfur dioxide and nitrogen oxides. Given that the Earth's population continues to grow, what would be expected to occur over the next 10 years with respect to the usage and impact of fossil fuels? (3 MARKS)

Question 16 (2 MARKS)

The following graph shows the variation in the density of water with temperature.



a. Why do both axes have a break in the sequence of numbers? (1 MARK)

b. What conclusion can be drawn from the data shown in the graph? (1 MARK)

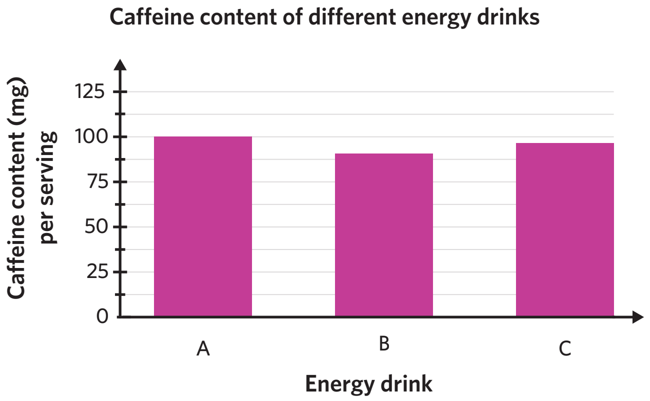
Adapted from VCAA (NHT) 2019 exam Short answer Q6

Question 17 (5 MARKS)

Energy drinks that contain a large amount of caffeine are becoming increasingly popular amongst teenage students. On average, each serving contains approximately 100 mg of caffeine, which is over double the amount of caffeine contained in a soft drink. A student wished to study the caffeine content in various brands of energy drinks (A, B and C).

a. What type of data would be collected in the experiment? (1 MARK)

The student presented the data in the graph shown.



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e general information presented in this graph? (1 MARK)

c. At high levels, caffeine is known to cause symptoms such as

- headaches,
- sleeping problems and
- increase in heart rate.

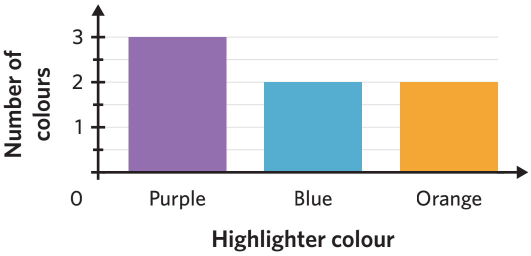
To further investigate the impact of caffeine on adolescents, the student proposed an experiment involving her friends. Identify a major ethical concern of this experiment. (2 MARKS)

d. Given that the student wanted to investigate the impact of caffeine-containing energy drinks on sleep, identify a suitable control for the experiment. (1 MARK)

Question 18 (3 MARKS)

A student wanted to conduct an experiment to determine the composition of different coloured highlighter pens. The graph provided shows the results of the experiment.

Highlighters and their colour composition



a. Identify the dependent variable. (1 MARK)

b. List two variables that need to be kept constant throughout the experiment. (2 MARKS)

Hints

14a. Trends in data relate to the response of the dependent variable as a result of manipulating the independent variable.

14b. Conclusions need to be supported by appropriate theory.

15a. The information presented by a pie chart depends on the detail shown in the graph.

15b. Data can be used to predict future outcomes.

16a. Appropriate use of space is important for data representation

16b. The labelled axes and title are very useful.

17a. The type of data depends on the nature of the data itself.

17b. Trends in data relate to the response of the dependent variable as a result of manipulating the independent variable.

17c. Ethical concerns can affect humans, animals, and society.

17d. Control groups are used to demonstrate the effect of the independent variable.

18a. Each axis on a graph represents a different variable.

18b. Outside of the independent and dependent variables, every other variable in an experiment should be controlled.

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16D Experimental factors affecting data

STUDY DESIGN DOT POINTS

- the nature of evidence and information: distinction between opinion, anecdote and evidence; and scientific and non-scientific ideas
- the quality of evidence, including validity and authority of data and sources of possible errors or bias
- accuracy, precision, repeatability, reproducibility, resolution, and validity of measurements in relation to the investigation
- the conventions of scientific report writing, including scientific terminology and representations, standard abbreviations and units of measurement
- the limitations of investigation methodologies and methods, and of data generation and/or analysis

16A 16B 16C **16D** 16E

1.3.2.1 & 1.3.3.1 Sources of error & bias

2.3.4.1 Precision and accuracy

2.3.13.1 Significant figures in reports

2.3.4.2 Validity and the three R's

2.3.11.1 Limitations of data

ESSENTIAL PRIOR KNOWLEDGE

16C Conclusions drawn from data can be affected by many different factors

See question 100.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

Image: Kiselev Andrey Valerevich/Shutterstock.com

How do we judge the quality of the data collected?

In this lesson, we will learn about the factors that affect the quality of the data collected during an experiment and how this can impact the research findings.

KEY TERMS AND DEFINITIONS

Accuracy how close measured values are to their true value

Bias difference between the average of a large set of measurements and the true value

Error difference between the measured value and the true value

Methodology different ways in which the planning and conducting of scientific investigations can be carried out

Mistake sometimes called personal errors, mistakes should not be included in reporting and analysis

Outlier data points or observations that differ significantly from other data points or observations

Precision how close measured values are to each other

Random error in measurement that differs in amount each time the experiment is conducted and is usually a 'one-off' error

Repeatability closeness of the agreement between the results of successive measurements of the same quantity being measured, carried out under the same conditions of measurement

Reproducibility closeness of the agreement between the results of measurements of the same quantity being measured, carried out under changed conditions of measurement

Resolution the smallest change in quantity that is measurable

Significant figures number of digits required to express a number to a certain level of accuracy

Systematic error in measurement by the same amount in the same direction every time which also includes errors that are inherent in the experiment

Uncertainty level of doubt regarding the data obtained (for VCE purposes, we are only looking at this from a qualitative perspective)

Validity whether or not the experiment and its components including the results address the aim and hypothesis of the research

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s of error & bias 1.3.2.1 & 1.3.3.1

Errors can affect the overall results collected during data analysis and are an important part of the scientific process (figure 1).

Why do we need to know about errors?

The **uncertainty** of the result of a measurement reflects the lack of exact knowledge of the value of the quantity being measured. VCE chemistry requires only a qualitative treatment of uncertainty. When conducting experiments, we always try to make sure that the conditions are well controlled. However, there is still a chance that errors can occur. These errors ultimately affect our results, creating a level of uncertainty in the data collected during the experiment. The main sources of error stem from two main types of errors: **systematic** and **random errors**.

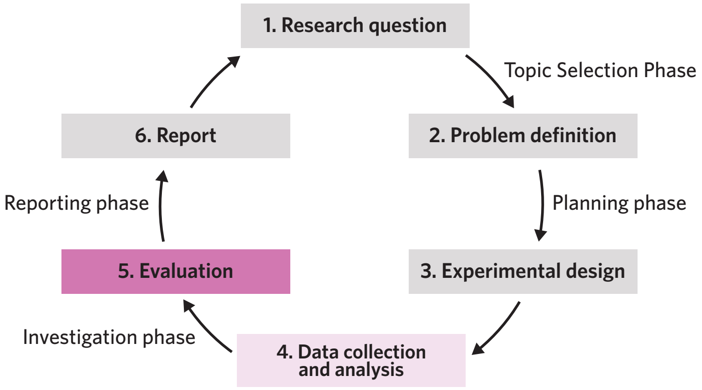

Figure 1 A scientific process

Systematic errors

Systematic errors cause measurements to differ from their true value in a systematic manner so that when a particular value is measured repeatedly, the error is always the same. These systematic errors can arise from limitations in the instruments used, incorrect calibration of instruments, or the use of inappropriate measuring methods (including the parallax error). A common example involves the reading of measuring cylinders, as shown in figure 2.

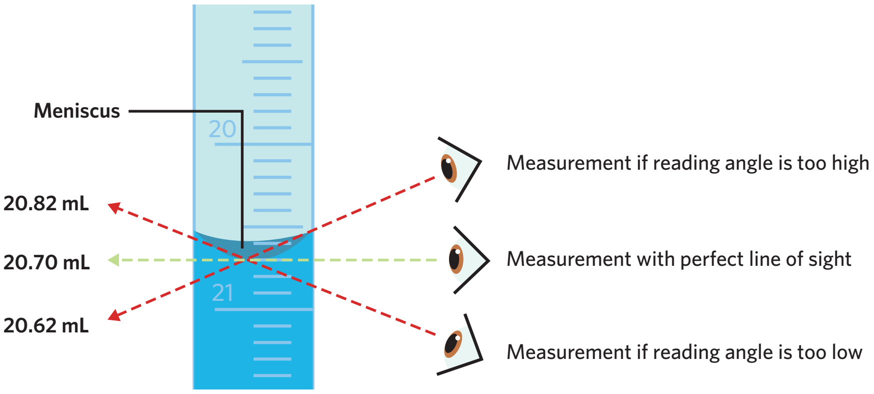

Figure 2 Errors that can arise during the use of measuring cylinders

To reduce the effect of systematic errors, we should calibrate all apparatus before conducting the experiment and analyse the method to identify steps that could introduce errors. Understanding correct measuring techniques is also beneficial.

BPA SCIENTIFIC INQUIRY

For our BPA experiment, a systematic error could occur in the volume of water in the water bottles used. Since we trusted the labelling of the unopened bottles, we assumed that they contained the correct volume of water. A way to rectify this error would be to measure the exact volume of water using scientific apparatus. Since this could also introduce a parallax error, we would have to be careful to ensure that we use the appropriate technique.

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Random errors

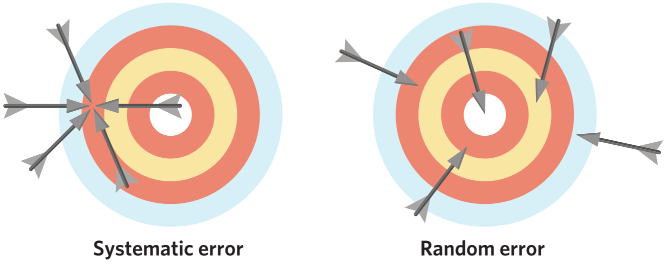

Figure 3 Effect of systematic and random errors on an experiment

Random errors affect the **precision** of a measurement and may be present in all measurements. These errors are unpredictable variations in the measurement process and result in a spread of readings. For example, a random error can occur when the volume of a liquid needs to be estimated as a consequence of using a measuring cylinder that is missing lines or numbers. We can minimise this type of error by repeating experiments multiple times to calculate the average of the results. Systematic and random errors affect an experiment in different ways as shown in figure 3.

From the perspective of a graph, systematic errors will cause the whole graph to be shifted by the same degree. The trend shown as a result of both errors is generally still the same, however the values will be different to the true value. With respect to random errors, we would have to draw a line of best fit. The effect of both systematic errors and random errors are depicted in figure 4.

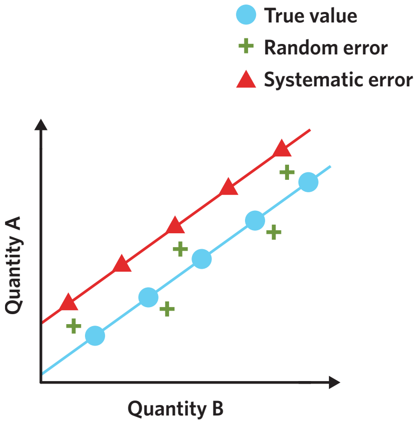

Figure 4 Graphical representation of the effect of systematic and random errors

We can use the measured values to determine the difference between the average of repeated experiments and the data (inclusive of systematic errors), which is referred to as **bias**. The best way to minimise bias would be to minimise both types of errors in an experiment.

Mistakes

Mistakes are sometimes called personal errors. Mistakes should not be included in reporting and analysis as part of the ethical consideration of data handling. Rather, the experiment should be repeated correctly. Fortunately, these errors can be minimised by repeating the experiment properly.

BPA SCIENTIFIC INQUIRY

In regards to our water bottle BPA experiment, a mistake that could occur may involve the extraction of a sample of pure BPA using a pipette that measures mL instead of µL for one of the samples.

Detecting errors in research

Errors can have a big impact on the quality of research. Therefore, it is important that we are able to identify errors in research that we read. This can range from theoretical misinterpretations, to analytical and measurement errors as mentioned previously. With respect to errors identified previously, some questions that will be useful in identifying errors could be:

- Are there **outliers** in the data? If so, why?
- Were the measuring instruments calibrated?
- Is the **methodology** prone to misinterpretation? (i.e. are any of the steps in the methodology ambiguous?)
- What is the expected value(s)? How far is the data collected in the experiment from the expected value(s)?

Being able to detect these types of errors will depend on:

- a deep understanding of the related theoretical concepts
- understanding the common methodology used for the type of experiment being conducted.

The best places to find these errors will be in the methods and results section of the research paper.

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Progress questions

Question 1

Which of the following is not a systematic error?

A. Spilling a solution

B. Using a 100 mL measuring cylinder to measure all solutions of less than 10 mL

Question 2

Which of the following is a random error?

A. Forgetting to add a reagent

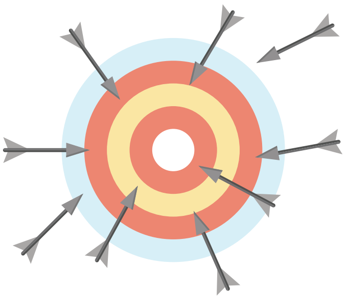
B. A parallax error

Question 3

The diagram shown represents

A. systematic errors.

B. random errors.



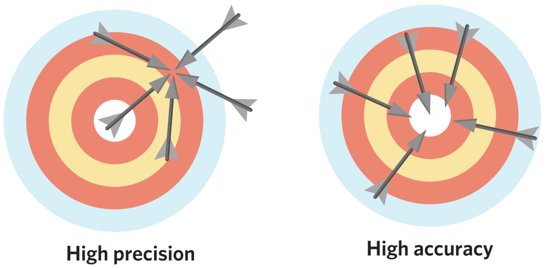
Precision and accuracy 2.3.4.1

Precision and **accuracy** affect experimental results in different ways.

How do precision and accuracy differ?

Precision is the measure of the **repeatability** or **reproducibility** of scientific measurements and refers to how close two or more measurements are to each other. A set of precise measurements will have values very close to the mean value of the measurements. Precision gives no indication of how close the measurements are to the true value and is therefore a separate consideration to accuracy. Precision is more impacted by random errors. For experiments that require estimation of results, the measurements taken may not be very precise. For example, this can occur during experiments where we are asked to identify changes in colour that are similar.

Accuracy refers to how close a measured value is to its true value. It is a qualitative term, meaning that it can describe a measurement value or result as being 'less accurate' or 'more accurate' when compared with a true value. As seen in figure 5, measurements that are precise aren't necessarily accurate; in fact the measurements can differ from the true value but can still be considered precise.



High precision

High accuracy

Figure 5 Comparison between precision and accuracy

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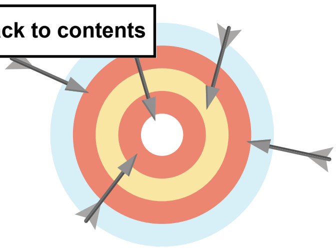
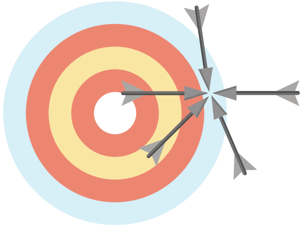
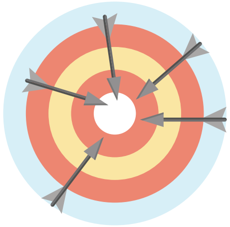


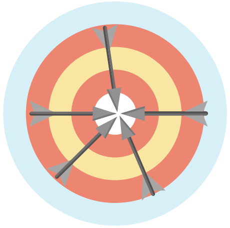
Low accuracy
Low precision



Low accuracy
High precision



High accuracy
Low precision



High accuracy
High precision

Figure 6 Illustration of the categorisation of data

It is also possible to collect data that is highly accurate yet not precise if the values are not close to each other. We can therefore categorise data based on both accuracy and precision. This concept is illustrated in figure 6.

The accuracy of data collected during an experiment is affected by the presence of systematic errors in the experimental design, some of which may include

- the purity of the substances used.
- use of poorly cleaned apparatus.
- uncalibrated measuring instruments.
- and unclear experimental method.

To increase the accuracy of data collected, we need to try to remove any systematic errors that may be present in the experiment. To test the accuracy of our data, we can determine the measurement error of the results by finding the difference between the measured values and the true value.

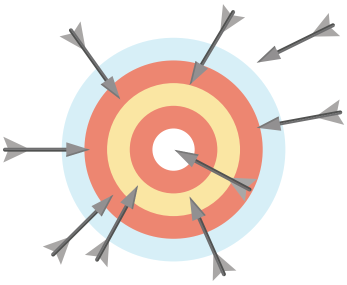
Validity refers to an experimental investigation that it sets out and/or claims to investigate. Both experimental design and the implementation should be considered when evaluating validity. An experiment and its associated data may not be valid, for example, if the investigation is flawed and controlled variables have been allowed to change. Data may not be valid, for example, if there is observer bias.

The collection of data that is both accurate and precise is important as it can affect the conclusions made from an experiment. It also affects the validity of an experiment, which refers to the ability of the results obtained through the particular experimental design to address the aim and hypothesis of a research study. To make sure that results are valid, we need to make sure that:

- The appropriate measuring instruments are used to collect accurate results.
- Controlled variables are kept constant.
- The experimental methodology is able to test the hypothesis.
- Data is collected under controlled and appropriate conditions.

Progress questions

Use the following diagram for questions 4–5.



Question 4

The diagram shows data that is

A. not precise.

B. very precise.

Question 5

For the data shown to be accurate, the data points need to be

A. as close as possible to the centre of the target.

B. clustered together anywhere on the target.

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ant figures in reports 2.3.13.1

Data is collected and expressed to various **significant figures**.

How do you work out significant figures?

Significant figures represent the precision of a measurement device. A measuring apparatus that expresses values in three significant figures is less precise than an electronic balance that expresses values in four significant figures.

Significant figures should be considered in all calculations. The following guidelines apply to VCE chemistry:

- all digits in numbers expressed in standard form are significant: for example, 5.320×10^{-6} has four significant figures
- all non-zero numbers are significant: for example, 53.2 has three significant figures
- zeros between two non-zero numbers are significant: for example, 5.320 has four significant figures
- leading zeros are not significant: for example, 0.0053 has two significant figures
- trailing zeros to the right of a decimal point are significant: for example, 53.00 has four significant figures
- for numbers less than one, 0.5 has one significant figure and 0.05 also has one significant figure, whereas 0.50 has two significant figures and 0.500 has three significant figures
- whole numbers written without a decimal point will have the same number of significant figures as the number of digits, with the assumption that the decimal point occurs at the end of the number: for example, 500 has three significant figures. Therefore, a stated volume of '500 mL' will be considered as having three significant figures.

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The appropriate use of significant figures is demonstrated in figure 7.

USEFUL TIP

Significant figures and decimal places are not the same thing.

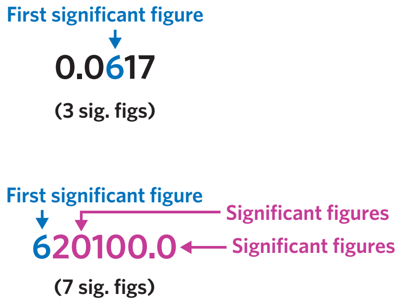

Figure 7 Consideration of significant figures

WORKED EXAMPLE 1

How many significant figures are found in the number 0.1104?

What information is presented in the question?

The number, 0.1104.

What is the question asking us to do?

Determine the number of significant figures.

What strategy do we need in order to answer the question?

Apply the rules to count the number of significant figures.

Answer

According to the rules, we start counting from the first non-zero number, which is the 1 immediately after the decimal place.

Since zeros that come after the first significant figure, that are followed by a non-zero number, are considered significant, the 0 after the 1 is considered significant.

Therefore, the number 0.1104 contains four significant figures.

When required to perform calculations, we need to express our answers to an appropriate number of significant figures. This depends on both the number of significant figures presented in the question, as well as the type of calculations that need to be performed:

- For addition/subtraction, answers are expressed with as many digits to the right of the decimal place according to the number with the least number of decimal places.
- For multiplication/division, answers are expressed with as many significant figures as the number with the least number of significant figures.

This is summarised in figure 8.

$$4.88 + 3.9 = 8.8$$

Answer is expressed to 1 decimal place (based on 3.9)

$$2.7 \times 1.11 = 3.0$$

Answer is expressed to 2 significant figures (based on 2.7)

Figure 8 The strategies involved in expressing quantitative values from a calculation

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PLE 2

As part of her calculations, Judee was to perform the following step:

$$= \frac{2.11}{1.3}$$

How many significant figures would her answer be expressed in?

What information is presented in the question?

The calculation to be performed.

What is the question asking us to do?

Determine the number of significant figures the answer should be expressed to.

What strategies do we need in order to answer the question?

1. Determine the type of calculation being performed.
2. Count the number of significant figures based on the appropriate rules.

Answer

This is a division calculation, therefore the answer will be expressed to the lowest number of significant figures given. For this question, the reference number will be 1.3, which contains two significant figures.

Therefore, the answer will be expressed to two significant figures.

USEFUL TIP

Sometimes the molar mass (M = 1.0 g mol^{-1}) in the data book is the limiting factor in calculations with respect to the number of significant figures in the final answer.

Sometimes numbers may be very large or very small which means that it is much easier to express them using scientific notation written as a decimal number multiplied by a power to the base of 10. The easiest way to do this is to:

- Move the decimal place in the number so that there is only one number to the left of the decimal point.
- Identify the number of significant figures required, and round the decimal number accordingly.
- Identify the number of times the decimal place has moved – this will be the exponent/power.

WORKED EXAMPLE 3

The diameter of the earth is approximately 12 742 km. Express this in scientific notation to three significant figures.

What information is presented in the question?

The diameter of the earth.

What is the question asking us to do?

Express the diameter of the earth to three significant figures using scientific notation.

What strategies do we need in order to answer the question?

1. Move the decimal place until there is only one number to the left of the decimal place.
2. Identify the number of significant figures required.
3. Write the answer in scientific notation.

Answer

12 742 can also be written as 12742.0. By moving our decimal, we can have a new number of:

$$1.2,742.0$$
$$= 1.27420$$

Considering we want the answer to three significant figures, our third significant number will be seven. Since the fourth number is four, according to the rounding rules, our third significant figure will stay at seven.

Third significant figure

$$1.27420$$
$$= 1.27$$

The number after 7 is less than 5, therefore our rounded number will be 1.27

Since we moved our decimal place four places to the left, our exponent/power will be four. Therefore the diameter can be expressed as:

$$1.27 \times 10^{4} \text{ km}$$

Number of times the decimal point was moved

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ng other people's research, we can check for precision and accuracy by comparing the results obtained with other pieces of research, but also by comparing the results to what is expected. These expected results will be detected by studying the data obtained and represented in the results section of research.

Progress questions

Question 6

The number 150050.0 contains

A. seven significant figures.

B. five significant figures.

Use the following information for questions 7 and 8.

The two following calculations were performed:

(1) 7.113 + 0.64 (2) 0.51 × 1.996

Question 7

Which of the following statements is true?

A. In calculation (1), the answer would be expressed to two decimal places.

B. In calculation (1), the answer would be expressed to two significant figures.

Question 8

Which of the following statements is true?

A. In calculation (2), the answer would be expressed to four significant figures.

B. In calculation (2), the answer would be expressed to two significant figures.

Validity and the three R's 2.3.4.2

Errors can affect the ability of an experiment to answer the research question.

How can an investigation be considered valid?

Unfortunately, not all experiments are considered valid. To ensure that experiments are able to test the aim and hypothesis, we need to make sure:

- We know what the independent variable is, and whether or not it relates to the question.
- We know the dependent variable, and whether or not it is what we need to measure to answer the question.
- We control all other variables.

There are three main ideas that are affected by the presence of errors and therefore determining the validity of an investigation.

Reproducibility is the closeness of the agreement between the results of measurements of the same quantity being measured, carried out under changed conditions of measurement. These changed conditions, involving replicate measurements on the same or similar objects, include a different observer, different method of measurement, different measuring instrument, different location, different conditions of use and different time. The purposes of reproducing experiments include checking of claimed precision and uncovering of any systematic errors that may affect accuracy from one or other experiments/groups. Experiments that use subjective human judgement(s) or that involve small sample sizes or insufficient measurements may also yield results that may not be reproducible. Reproducibility links closely to the accuracy of an experiment. Reproducibility can also be used to evaluate the quality of data in terms of the precision of measurement results.

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Repeatability is the closeness of the agreement between the results of successive measurements of the same quantity being measured, carried out under the same conditions of measurement. These conditions include the same observer, the same measurement procedure, the same measuring instrument used under the same conditions, the same location, and replicate measurements on the same or similar objects over a short period of time. Experiments that use subjective human judgement(s) or that involve small sample sizes may yield results that may not be repeatable. Repeatability can be used to evaluate the quality of data in terms of the precision of measurement results. Ideally, measurements should be repeated where possible to produce a measurement result.

Resolution is the smallest change in the quantity being measured that causes a perceptible change in the value indicated on the measuring instrument. This has implications for determining the number of decimal places to which a quantity may be quoted. For example, if the resolution of a burette is 0.05 mL, then the user must estimate the volume between the two marked intervals on the burette so that the value reported will be to two decimal places. For example, measurement readings of 10.50 mL or 10.55 mL are possible, but a measurement reading of 10.53 mL cannot be claimed. The meniscus of the liquid will either be on the burette line marking, in which case the reading would be 10.50, or it will lie between 10.50 and 10.60, in which case it is measured as 10.55 mL.

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Making sure that data is valid ensures that the evidence collected is strong. Having strong evidence means that the findings from the research study is more robust and believable. In contrast, weak evidence that is laden with errors is less likely to be accepted by the scientific community.

Progress questions

Question 9

Experiments with reproducible results

A. can be replicated by other experimenters under different conditions to yield similar results.

B. can be replicated by other experimenters under the same conditions to yield similar results.

Question 10

Experiments with repeatable results

A. can be replicated by other experimenters under the same conditions to yield similar results.

B. can be replicated by the same experimenter under different conditions to yield similar results.

Limitations of data 2.3.11.1

There are factors that can limit the ability to conduct experiments.

What are some of the limitations when collecting data?

It is difficult to control every variable in an experiment. This limits the validity, accuracy and reproducibility of experimental results, making it sometimes difficult to identify the relationship between the independent and dependent variable. These limitations can be a result of the design and methodology of the experiment, or even from the surrounding environment. Access to relevant apparatus and reagents as well as the depth of understanding of the related field can also act as limitations to the research.

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o what we learn in VCE chemistry, most of the data in the Data Book was developed from experiments that were conducted under specific conditions. Therefore, the information presented is only true under those conditions. However, when conducting our own experiments in the classroom, the conditions are not the same and therefore the data we use may not exactly align with the information stated in the data book.

Progress questions

Question 11

Experiments are valid if

A. all variables outside the independent (and dependent) variable are controlled.

B. all variables are slightly adjusted throughout the experiment.

Theory summary

- There are many different types of errors that can affect the quality of scientific research.
- Each error has a different effect on the overall research.
- Accuracy and precision measure different aspects of quantitative data.
- Significant figures allude to the precision of experimental data.
- The number of significant figures used to express answers depends on the mathematical operation being performed.
- Repeatability, resolution and reproducibility all refer to the ability of the results to be obtained by repeating the experiment under the same or different conditions.
- There will always be variables that limit research investigations.

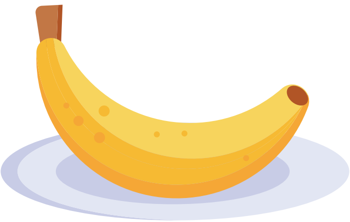
16D Questions

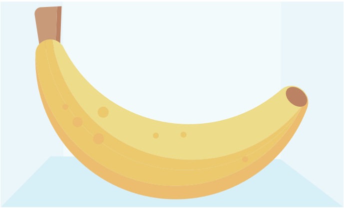
Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 12–14.

Farquaad developed an experiment to identify the rate at which bananas ripen in an open area compared to in a closed area. To do so, he set up the following experiment.





Farquaad measured the time it took, in hours, from peeling the banana to when the banana first started to turn brown. He checked the results every few hours during the day, but did not check it during the night.

Question 12 (1 MARK) 🌶

Which of the following steps will **not** improve the validity of the experiment?

A. Using bananas from the same bunch

B. Using the same type of bananas

C. Taking measurements at timed intervals

D. Doing the experiment on a different day

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(1 MARK)

Which of the following variables need to be kept constant?

A. The size of the bananas

B. The temperature of the experiment

C. The timer used to measure the time

D. All of the above

Question 14 (4 MARKS)

Based on the results of the experiment, Farquaad saw that the banana in the box began to ripen before the banana on the plate. Therefore, he concluded that bananas kept in a closed area ripen quicker than those left in an open environment. Comment on the validity of this conclusion.

Exam-style

Question 15 (4 MARKS)

For each of the descriptions, identify the type of error that has occurred.

Description	Type of error
Experimenter uses an uncalibrated thermometer to measure the heat of a reaction.	
Some reagent falls off a measuring glass as a student walks back to the bench.	
Student reads a measuring cylinder from the top of the meniscus.	
Student uses the wrong reagent in a single reaction.	

Question 16 (6 MARKS)

Ammonia is a strong cleaning agent found in many cleaning products. Depending on the purpose of the cleaner, different cleaners have different concentrations of ammonia. As part of an investigation to determine the concentration of ammonia, ammonia was reacted with different volumes of hydrochloric acid. It was expected that for every 20.00 mL of ammonia used there would be 15.00 mL of hydrochloric acid required.

The results from the test are provided.

Test number	Volume of HCl reacted (mL)
1	18.81
2	19.00
3	14.76
4	18.74

a. Which test results are most precise? (1 MARKS)

b. Are the results accurate? Explain. (4 MARKS)

c. How many decimal places should the concentration of ammonia, when calculated from the results obtained, be expressed to? (1 MARK)

Question 17 (8 MARKS)

For their research investigation, a student wanted to conduct an experiment using a substance that is not commonly available in schools. As a consequence, they had to source a small sample from a nearby university.

a. Identify two limitations of the experiment. (2 MARKS)

b. The student followed methodology presented by a research paper studying a similar question in the same field. Despite this, the results obtained in the experiment and the results reported by the researchers in the article were very different. Comment on the reproducibility of the results. (3 MARKS)

c. Suggest a way in which the reproducibility can be improved. (1 MARK)

d. What is the main difference between reproducibility and repeatability? (2 MARKS)

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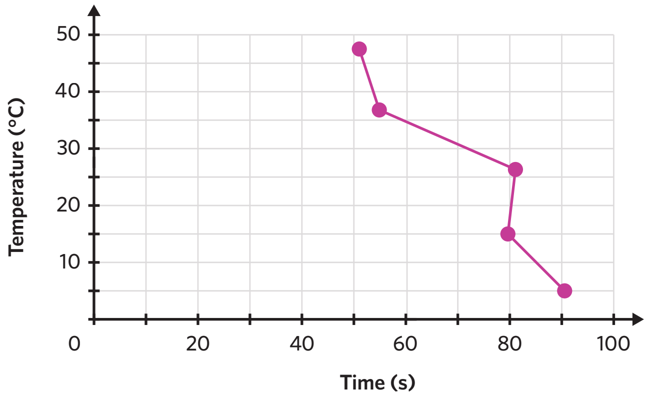
(1 MARK)

Consider the following graph.

Which of the following could not be used to describe the data point that was recorded at 25 °C?



A. Outlier

B. Systematic error

C. Random error

D. Mistake

VCAA 2020 exam Short answer Q9

Question 19 (1 MARK)

Which of the following statements about conducting an experiment is most correct?

A. Precise results may be biased.

B. Accuracy is assured if sensitive instruments are used.

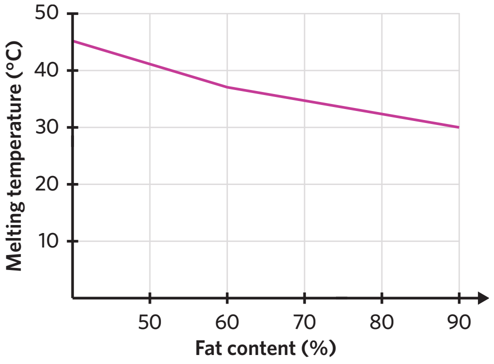
C. A method is valid if it identifies all controlled variables.

D. Repeating a procedure will remove the uncertainty of the results.

VCAA 2020 exam Multiple choice Q22

Question 20 (4 MARKS)

Balnea conducted an experiment to see how the fat content in chocolate affects the temperature at which the chocolate will melt.



a. Identify the independent variable. (1 MARK)

b. Identify an appropriate control group for the experiment. (1 MARK)

c. The results obtained were plotted on a graph as shown.
What conclusion can be drawn from the results? (1 MARK)

d. Balnea wanted to calculate the amount of fat present in the chocolate. To do so, 4.00 g samples of chocolate were required. What would be the maximum number of significant figures that the answer be expressed to? (1 MARK)

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(4 MARKS)

Type 1 diabetes affects the body's ability to bring glucose into cells as there is not enough insulin. Insulin is responsible for helping to move glucose from the bloodstream to cells. This means that most of the sugar, which contains glucose, isn't delivered to cells and stays in the blood. Therefore, those with type 1 diabetes need to closely monitor their blood glucose levels.

The different sugar content in foods can affect blood glucose in different ways. Thomas wanted to test different foods and their effect on his blood glucose. As part of this test, he ate a variety of different snacks with different glucose contents and measured his blood glucose at baseline, and then every hour for four hours.

a. Identify the dependent variable. (1 MARK)

b. Write an appropriate hypothesis for this research. (1 MARK)

c. At the end of the test, Thomas decided to calibrate his glucose metre and realised that his metre was measuring values 0.20 mmol L^{-1} less than the true value. Identify how this would affect his results. (2 MARKS)

Hints

15. Different actions lead to different errors based on their effect on the data.
16a. Precision describes how close values are to each other.
16b. Accuracy describes how close values are to the true value.
16c. The decimal places expressed depend on the data used.
17a. Limitations can affect the ability of an experiment to be performed.
17b. Reproducibility is measured in certain conditions.
17c. Improvements depend on the nature of the variable that needs to be optimised.
17d. Reproducibility and repeatability differ in their experimental conditions.
18. Some errors occur all the time.
19. The scientific process is very considered.
20a. The independent variable is deliberately manipulated.
20b. Graphs show the relationship between the independent and dependent variable.
20c. Expression of significant figures depends on the data used.
20d. The number of significant figures depends on the question and the calculation performed.
21a. The dependent variable is affected by the independent variable.
21b. The hypothesis predicts how the independent variable will influence the outcome of the dependent variable.
21c. Different errors affect results in different ways.

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16E Writing scientific material

Image: hydebrink/Shutterstock.com

How can the findings of scientific research papers be communicated?

Scientific investigations are often communicated using a large poster format. In this lesson, we will learn about how to share scientific information in a way that is easy to understand and that follows the conventions of scientific writing.

Effective scientific communication 1.3.9.1

Scientific communication is part of the scientific process (see figure 1) and needs to be concise.

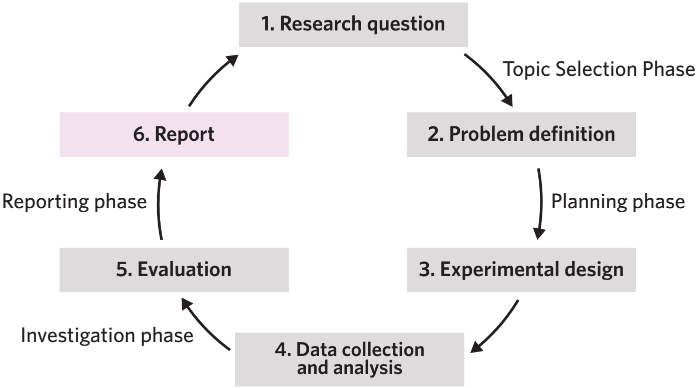

Figure 1 A scientific process

How can we communicate concisely in a poster?

The purpose of scientific articles and reports is to be able to convey a finding or message to the reader. Therefore, it is important to be concise and clear in the way we represent information without being verbose. Getting to the point is more important than sounding sophisticated. Some key techniques for effective communication in science include:

- use of short sentences that get to the point,
- consistent voice,
- maintaining a neutral stance and tone by avoiding subjective language,
- use of visuals to support writing where necessary and
- referencing of related theory.

STUDY DESIGN DOT POINTS

- characteristics of effective science communication: accuracy of chemical information; clarity of explanation of chemical concepts, ideas and models; contextual clarity with reference to importance and implications of findings; conciseness and coherence; and appropriateness for purpose and audience
- chemical concepts specific to the investigation: definitions of key terms; and use of appropriate chemical terminology, conventions and representations
- the conventions of scientific report writing, including scientific terminology and representations, standard abbreviations and units of measurement
- the use of data representations, models and theories in organising and explaining observed phenomena and chemical concepts, and their limitations

16A 16B 16C 16D 16E

ESSENTIAL PRIOR KNOWLEDGE

16A Hypotheses

See question 101.

ACTIVITIES

Log into your Edrolo account for activities that support this lesson.

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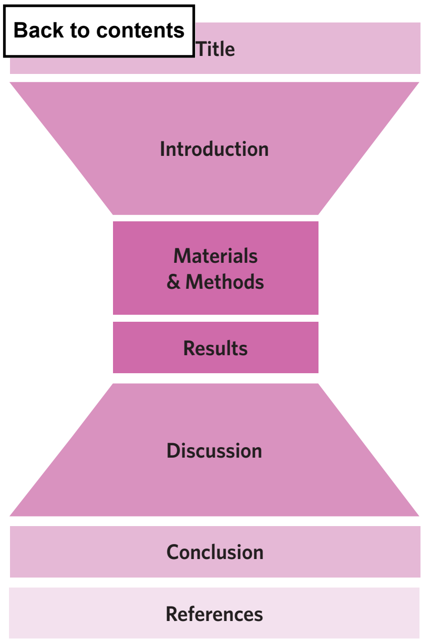

Figure 2 Outline of a scientific report

BPA SCIENTIFIC INQUIRY

When we put this into the context of the water bottle experiment, we can develop a scientific poster to share our experiment as seen in figure 4.

Another important factor to consider when communicating scientific ideas, especially in a research article, is the appropriate formatting of the main text as well as any references used. For example, scientific research is written in a particular order with predetermined sections as shown in figure 2.

The different sections of the poster convey different types of information, and together they provide a deeper understanding of an idea. Based on figure 2, we can see that a large portion of the poster is the discussion section. In fact, this is quite common in any scientific writing as this is the section where we bring all of the data and ideas together. The discussion section gives us an opportunity to elaborate on and explain the results, and to show how the research and its results contributes to the pool of knowledge relating to the same or similar ideas. Regardless of whether or not the results support the hypothesis, the analysis of the results and what they mean is the main purpose of the discussion. From this analysis, we are able to identify any future implications of the experiment.

We can now use everything that we've written about to provide a conclusion for the research. Similar to an essay, the conclusion is a very brief summary of what was found in the experiment.

Understanding that scientific ideas are built upon previous research, it's important to acknowledge the source of the information that formed the basis of the research. Depending on where scientific work is submitted to, the referencing requirements may vary. For example, some publications require references to be done in Harvard style, whereas others may require APA style. Information about the exact formatting required is available online.

Progress questions

Question 1

Scientific writing is supposed to be

A. verbose and sophisticated.

B. simple and clear.

Question 2

The following are examples of how water cycles through the environment. Which of the two statements would most likely be preferred when summarising this concept?

A. Water exists in many forms and can therefore be seen to move through different stages in the environment.

B. Due to the structural composition of water, which impacts its physical and chemical properties, it is able to interact with the environment in such a way that renders it capable of moving through different environments in different forms.

Question 3

A scientific report must include

A. numbers.

B. a methods section.

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al terminology, conventions and abbreviations 1.3.13.2

Chemistry has its own language when it comes to representing information.

How do we speak chemistry?

We know that chemistry has a particular language when presenting ideas. This includes the terms used, structural representations of chemicals, and symbolic representations of elements. The most recognisable example is the use of symbols to represent the names of the chemical elements, for example the use of the symbol O to represent the element oxygen. Other examples include:

- the use of subscripts and superscripts to represent multiples or charges of ions e.g. NH_3 and Mg^{2+},
- the use of brackets in chemical formulas and structural formulas e.g. $Mg(NO_3)_2$,
- the use of variables/symbols to represent chemical concepts e.g. N_A, n, M,
- key terms representing multi-faceted chemical concepts e.g. compound, and molecule.

Although there are certain terms and conventions that are used in different contexts, scientific concepts require specific conventions and terminology. Therefore, it is important that we are able to select these according to the context. At times, scientific research will include an appendix that outlines the conventions and abbreviations relevant to the research that are not commonly used elsewhere.

It is also important that anything included in the research is labelled appropriately. For example, tables and figures included in scientific writing all need to be accompanied by a table or figure legend. Furthermore, graphs need to be appropriately titled, with axis headings included with their appropriate units.

How are chemical reactions written?

An important part of chemistry is investigating chemical reactions. Chemical reactions involve the rearrangement of atoms found in the reactants to produce different products. This process is illustrated by a chemical equation as shown.

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

Reactants → Products

Chemical equations show a lot of information about a chemical reaction as shown in figure 3.

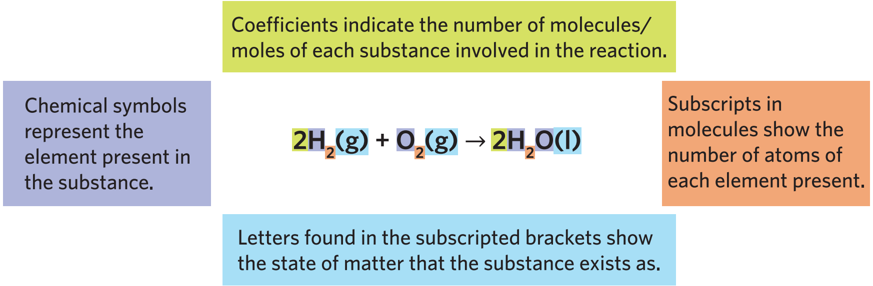

Figure 3 Key elements in a chemical equation

USEFUL TIP

If there isn't a coefficient present, imagine that there is a 1 in front of the formula.

As we know, substances can exist in different states, each of which have their own convention when represented in a chemical reaction as shown in table 1.

Table 1 States of matter conventions

Convention	Example
(g)	Gas
(s)	Solid
(l)	Pure liquid or pure molten liquid
(aq)	Aqueous (chemical species dissolved in water, also known as a solution)

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Is it dangerous to drink water from a plastic water bottle that has been left in a warm environment?

INTRODUCTION

Bisphenol A (BPA), as shown in figure 3.1, is a monomer, a small building block, used to produce a strong plastic polymer known as polycarbonate (Vogel, 2009). Due to its widespread uses in bottles and containers, there is growing concern over the health effects of using BPA in household plastic items.

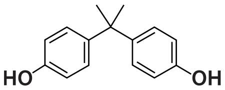

Figure 4.1 Structure of a bisphenol A (BPA) monomer

In water-containing products such as drink bottles, BPA is known to leach, when hydrolysed by heat, pH or microwaves, into the water over time (Le *et al*, 2008). When consumed, this can result in the disruption of important hormone-driven processes in the body (Diamanti-Kandarakis *et al*. 2009).

This experiment aims to test the effect of temperature on the release of BPA into water. It's expected that if the breakdown of BPA from plastic depends on the temperature of the surrounding environment, then an increase in temperature would result in an increase in the concentration of BPA found in the water.

MATERIALS

- 3 × 600 mL unopened BPA-containing bottles of water (from the brand Rolo Water)
- 3 × 600 mL unopened bottles of water in a BPA-free plastic bottle (from the brand Rolo Water)
- HPLC machine
- 0.50 ng mL^{-1}, 1.0 ng mL^{-1}, 1.5 ng mL^{-1} and 2.0 ng mL^{-1} BPA standards
- Micropipette
- Pipette tips
- Fridge
- Incubator

METHOD

Part A

Identification of the retention time of BPA.

1. Inject a 5 μL sample of a 2.0 ng mL^{-1} BPA standard into the HPLC machine.
2. Record the retention time of the most prevalent peak.
3. Repeat steps 1 and 2 with a fresh sample of the same standard.

Part B

Development of the BPA calibration curve.

1. Inject a 5 μL sample of 0.50 ng mL^{-1} BPA standard through the HPLC machine.
2. Record the area under the peak for the sample.
3. Repeat steps 1–2 for the 1.0ng mL^{-1}, 1.5 ng mL^{-1} and 2.0 ng mL^{-1} standards.
4. Construct a calibration curve using the data collected.

Part C

Determination of the concentration of BPA.

1. Label each BPA-containing bottle A, B and C, and each BPA-free bottle D, E and F.
2. Place bottles A and D in a fridge at 4 °C.
3. Place bottles B and E in an incubator set at 20 °C.
4. Place bottles C and F in an incubator set at 30 °C.
5. Leave the samples for 2 hours.
6. After the 2 hour period, inject a 5 μL sample of water from bottle A into the HPLC machine using the micropipette.
7. Record the area under the peak for sample A at the retention time determined for BPA in Part A.
8. Repeat steps 6 and 7 for bottles B–F.

Figure 4 Water bottle experiment scientific poster

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RESULTS

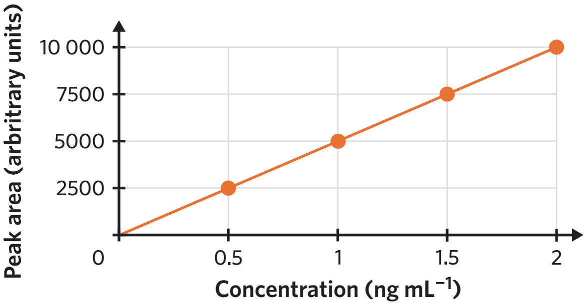

Figure 4.2 Calibration curve of known standards of BPA

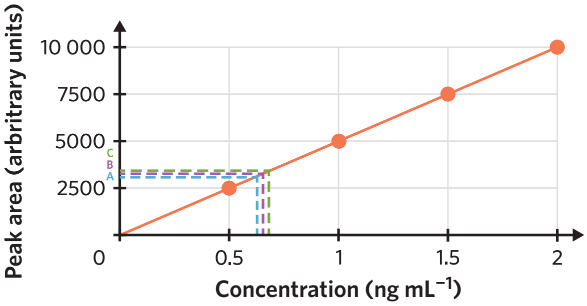

Figure 4.3 Determination of the concentration of BPA in samples A–F

Table Concentration of BPA in each sample

Sample	**Peak Area (arbitrary units)**	**Concentration (ng mL^{-1})**
A BPA bottle at 4 °C	2610	0.58
B BPA bottle at 20 °C	2720	0.61
C BPA bottle at 30 °C	2830	0.63
D BPA-free bottle at 4 °C	101	0.022
E BPA-free bottle at 20 °C	100	0.022
F BPA-free bottle at 30 °C	101	0.022

DISCUSSION

Compared to control samples D-F in which water was placed in a BPA-free bottle, the presence of BPA in the water bottles resulted in a greater concentration of BPA in the water samples. For samples A-C, it can be seen that an increase in temperature is correlated with an increase in the concentration of BPA measured (0.58, 0.61, and 0.63 respectively).

BPA can be released into the contents of water bottles when the plastic begins to degrade (Biedermann-Brem & Grob., 2008). In particular, an increase in temperature results in the increase in the rate at which the plastics are degraded. This trend is shown in the results of this experiment.

Future implications

As BPA is found in a number of different plastic materials, the results of this experiment can help consumers make more informed decisions about the way they currently store goods.

Limitations

As there are so many different types of plastics available, the results of the study are specific to only Rolo Water.

CONCLUSION

In the same amount of time, an increase in temperature results in a slight increase in the concentration of BPA found in water of BPA-containing water bottles.

REFERENCES

- Biedermann-Brem and K. Grob (2008). *Release of bisphenol A from polycarbonate baby bottles; water hardness as the most relevant factor*. Eur. Food res. Technol., 228(5):679–684.
- Diamanti-Kandarakis, E., Bourguignon, J. P., Giudice, L. C., Hauser, R., Prins, G. S., Soto, A. M., Zoeller, R. T., & Gore, A. C. (2009). *Endocrine-disrupting chemicals: an Endocrine Society scientific statement*. Endocrine reviews, 30(4), 293–342.
- Le, H. H., Carlson, E. M., Chua, J. P., & Belcher, S. M. (2008). *Bisphenol A is released from polycarbonate drinking bottles and mimics the neurotoxic actions of estrogen in developing cerebellar neurons*. Toxicology letters, 176(2), 149–156.
- Vogel S. A. (2009). *The politics of plastics: the making and unmaking of bisphenol a 'safety'*. American journal of public health, 99 Suppl 3(Suppl 3), S559–S566.

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Another important component of a chemical equation is the use of coefficients in front of the formula of each substance. These numbers represent the ratio relationship between each of the substances present, and links to the mole concept. This is an important concept that we will continue to revisit throughout our studies.

In the chemical reaction shown, we can see that hydrogen gas reacts with oxygen gas to produce liquid water. The product (H_2O) formed is made from atoms that were present in the reactant molecules (H_2 and O_2). If we look even closer, we will notice that not only are the elements present in the reactants and products the same, but the number of atoms of each element is also the same. When this is the case, we can say that the chemical equation is balanced. This highlights a concept known as the conservation of matter, whereby the total number and type of atoms in a system remains the same before and after a chemical change, where atoms are neither created nor destroyed.

To see if a chemical equation is balanced, we do a quick check as shown in figure 5 and table 2.

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

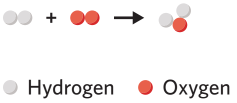
Figure 5 Process involved to check if a chemical equation is balanced

Table 2 Process of ensuring that reactants and products are balanced in a chemical equation

	Reactants	**Products**
Hydrogen	There are two hydrogen atoms per molecule, however we have a total of two hydrogen molecules as indicated by the coefficient 2. Therefore, we have a total of: $2 \times 2 = 4$ hydrogen atoms	Hydrogen is found as part of the water molecule. Each water molecule contains two hydrogen atoms. The coefficient in front of the H_2O in the equation shows that there are two water molecules in total. Therefore, we have a total of: $2 \times 2 = 4$ hydrogen atoms
Oxygen	There are two oxygen atoms per molecule. We only have one molecule of oxygen and therefore we have a total of: $1 \times 2 = 2$ oxygen atoms	Oxygen is found as part of the water molecule. Each water molecule contains one oxygen atom. The coefficient in front of the H_2O in the equation shows that there are two water molecules in total. Therefore, we have a total of: $2 \times 1 = 2$ oxygen atoms

The concept of a balanced chemical equation is extremely important as it is an underlying skill that is required for all of the chemical theory that we will be learning about. Therefore, it's important that we understand how to balance an equation.

Going back to the original chemical reaction presented earlier between hydrogen gas and oxygen gas, the initial equation would be as shown in figure 6.



	Present in the reactant	**Present in the product side**
Hydrogen	Two	Two
Oxygen	Two	One

Figure 6 Initial equation for the reaction between hydrogen and oxygen

The number of hydrogen atoms found on the reactant and product sides is the same, however there is an imbalance of oxygen atoms, with a greater number of oxygen atoms on the reactant side than the product side. To balance the oxygen atoms, we need to add another oxygen atom on the product side. However, we cannot add single atoms. Since oxygen is present as part of a water molecule, in order to add another oxygen atom, we must add a whole water molecule to the reaction. Therefore, the equation would change as shown in figure 7.

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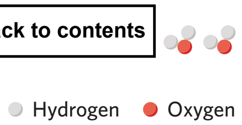
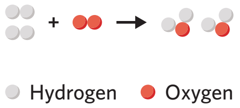


	Present in the reactant	Present in the product side
Hydrogen	Two	Four
Oxygen	Two	One

Figure 7 Balancing of oxygen atoms in a reaction between hydrogen and oxygen

By adding the additional water molecule, we have also added another two hydrogen atoms to the product side, which has caused an imbalance in the number of hydrogen atoms. To balance the number of hydrogen atoms, we need more hydrogen atoms on the reactant side. To do so, we can add a hydrogen molecule so that the equation changes as shown in figure 8.



	Present in the reactant	Present in the product side
Hydrogen	Four	Four
Oxygen	Two	Two

Figure 8 Balancing of hydrogen atoms in a reaction between hydrogen and oxygen

The equation is now balanced. As we can see, in the balanced equation, there are two hydrogen gas molecules, one oxygen gas molecule and two water molecules. Therefore the balanced equation is written as:

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

According to the balanced equation, for every two molecules of hydrogen that reacts with one molecule of oxygen, two molecules of water can be formed.

USEFUL TIP

Do not add single atoms or change the subscript of elements in formulas when balancing equations.
This changes the substance entirely!

WORKED EXAMPLE 1

Balance the following chemical equation which represents the process of photosynthesis. Interpret what the numbers in front of the molecular formulas mean.

$CO_2(g) + H_2O(l) \rightarrow C_6H_{12}O_6(aq) + O_2(g)$

What information is presented in the question?

Unbalanced chemical equation for photosynthesis:
$CO_2(g) + H_2O(l) \rightarrow C_6H_{12}O_6(aq) + O_2(g)$

What is the question asking us to do?

Balance the chemical equation.

What strategies do we need in order to answer the question?

1. Determine the number of carbon, hydrogen and oxygen atoms on either side of the equation.
2. Change the number of atoms on either side of the equation by adding molecules where appropriate, which changes the coefficients in the equation.

Answer

Initially, there are one carbon atom, two hydrogen atoms and three oxygen atoms on the reactant side of the equation and six carbon atoms, 12 hydrogen atoms and eight oxygen atoms on the product side of the equation.

In order to balance the carbon atoms, we need a total of six CO_2 molecules:
$6CO_2(g) + H_2O(l) \rightarrow C_6H_{12}O_6(aq) + O_2(g)$

There are now:

Reactant	:	Products
6 carbon atoms	:	6 carbon atoms
13 oxygen atoms	:	8 oxygen atoms
2 hydrogen atoms	:	12 hydrogen atoms

In order to balance the hydrogen atoms, we need six H_2O molecules:
$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + O_2(g)$

There are now:

Reactant	:	Product
6 carbon atoms	:	6 carbon atoms
18 oxygen atoms	:	8 oxygen atoms
12 hydrogen atoms	:	12 hydrogen atoms

In order to balance the oxygen atoms, we need six O_2 molecules:
$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$

There are now:

Reactant	:	Product
6 carbon atoms	:	6 carbon atoms
18 oxygen atoms	:	18 oxygen atoms
12 hydrogen atoms	:	12 hydrogen atoms

The equation is now balanced.

The numbers in front of the molecular formulas indicate the ratio of different molecules reacting. In this case, six molecules of CO_2 react with six molecules of H_2O to form one molecule of $C_6H_{12}O_6$ and six molecules of O_2.

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When balancing complicated equations, it may be easier to start by balancing key elements; those that are only found in one product.

Given that balanced chemical equations represent the number of molecules required in a reaction and the number of molecules produced in a reaction, they also represent mole ratios. If we recall lesson 6A, the mole is a unit which represents a fixed number of molecules/atoms (Avogadro's number). Therefore, using the example above, the mole ratio is:

$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

$$\begin{array}{ccccccc} CO_2 & : & H_2O & : & C_6H_{12}O_6 & : & O_2 \\ 6 & : & 6 & : & 1 & : & 6 \end{array}$$

This means that six mol of CO_2 reacts with six mol of H_2O to produce one mol of $C_6H_{12}O_6$ and six mol of O_2. Based on this ratio, we can calculate the amount required for a specific reaction to occur, or the amount of product formed.
The general formula to perform these calculations is:

$$n(unknown) = \frac{WANT}{KNOW} \times n(known)$$

Where WANT is the coefficient or number in front of the chemical species whose amount we want to determine and KNOW is the coefficient or number in front of the chemical species whose amount we know. Using the example above, we can now determine the amount of $C_6H_{12}O_6(aq)$ produced. For example, if one mol of $CO_2(g)$ reacts with sufficient $H_2O(l)$, 0.17 mol of $C_6H_{12}O_6$ will be produced as shown in figure 9.

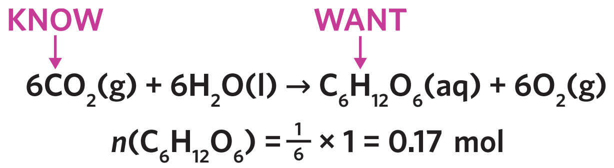

Figure 9 Steps to determine the amount of an unknown substance from a substance where the amount is known

Progress questions

Question 4

Chemical terminology is used to describe

A. a multi-process concept only.

B. both a multi-process concept and single ideas.

Units of measurement in reports 2.3.13.3

Different units are used to represent different measurements.

Why are units so important in chemistry?

In chemistry, we work with various different types of data ranging from amount, mass, volume and temperature, all of which are measured in different units. Table 3 shows some common units that can be found in the VCE Data Book.

Table 3 Common units used in chemistry as outlined by the VCE Data Book

Name	Symbol	Value
Avogadro constant	N_A or L	6.02×10^{23} mol^{-1}
Charge on one electron (elementary charge)	e	-1.60×10^{-19} C
Faraday constant	F	96 500 C mol^{-1}
Molar gas constant	R	8.31 J mol^{-1} K^{-1}
Molar volume of an ideal gas at SLC (25 °C and 100 kPa)	V_m	24.8 L mol^{-1}
Specific heat capacity of water	c	4.18 kJ kg^{-1} K^{-1} or 4.18 J g^{-1} K^{-1}
Density of water at 25 °C	d	997 kg m^{-3} or 0.997 g mL^{-1}

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mplete list of all of the units that we will use. The units used will depend on the type of information that we are collecting and trying to convey, and therefore, they vary contextually.

Progress questions

Question 5

Units

A. can be used to represent multiple different concepts.

B. are often specific to what is being measured.

Question 6

The most suitable unit to measure distance would be

A. minutes.

B. metres.

Representing ideas with models 1.3.10.2

Chemical concepts can also be represented visually.

Why are models useful?

Models are often used to represent chemical concepts in a way that helps to make it easier to understand conceptual theory. Visual models are the most commonly used and perhaps most useful when conveying conceptual information. For example, the ball-and-stick model in figure 10 is a common representation of molecules and their bonds, and is often used in resources to demonstrate the complexities of chemical bonds.

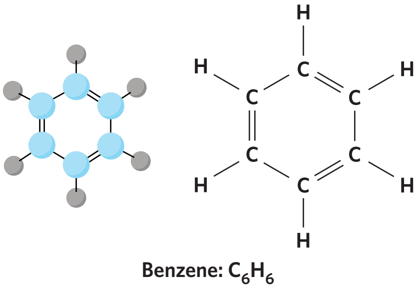

Figure 10 Chemical structures can be modelled in different ways

Other visual representations include flow charts and diagrams such as the example in figure 11, which functions to summarise the overall steps involved in a process.

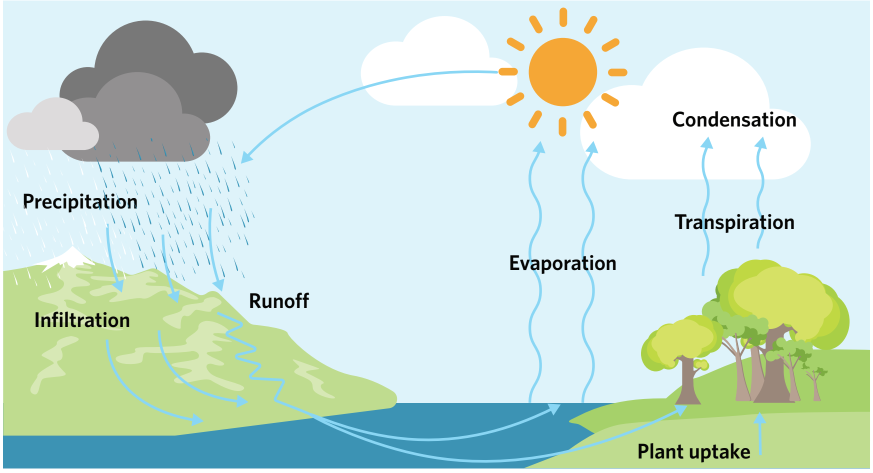

Figure 11 The water cycle

Other commonly used visual representations include graphs and tables.

Models do not just have to be visual, but can also be worded, or even formulaic. Chemical equations and formulas are also examples of models used in chemistry as they are different ways of representing concepts.

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Progress questions

Question 7

A chemical equation can be considered as a model.

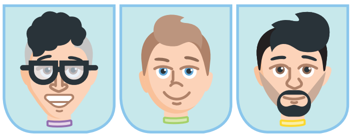
A. True

B. False

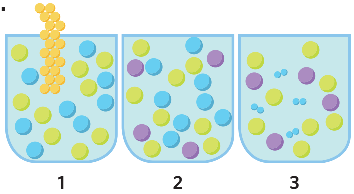
Question 8

Which of the following is an example of a visual model.

A.



B.



Limitations of representations 1.3.10.1

Representations often do not show the complete picture.

How can limitations influence our understanding?

Models and figures can make it easy to understand the general idea of chemical concepts, however, they are sometimes unable to illustrate the idea in detail. For example, with the water cycle shown in figure 11, we can identify the major components in the water cycle. However, we cannot see what actually happens at each stage. The addition of the smaller details would make the model too complicated and hard to understand. Therefore, we need to understand that models are often used to summarise or demonstrate a particular idea, not to show all of the details involved in a concept.

These concepts, along with all of the other concepts covered in the chapter, allow us to both develop our own research as well as critique other experiments during our research journey. The outcome of this can be shown in the poster related to our water bottle experiment, as shown in figure 3.

Progress questions

Question 9

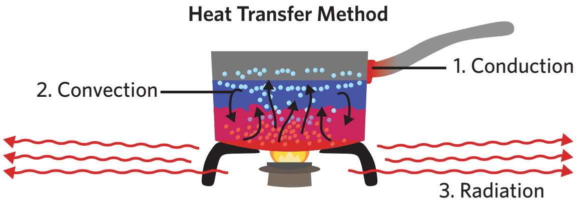
Visual models generally cannot

A. include words.

B. show all details related to a concept.

Question 10

The image provided shows the different ways in which heat can be transferred.



From this, we can see that

A. convection involves the transfer of heat through the movement of particles of different kinetic energy.

B. there are three main types of heat transfer.

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Theory summary

- Being clear and concise is important when communicating scientific ideas.
- Different concepts contain different terminology and conventions.
- The units of measurement used depend on the type of information that is being collected.
- Chemical concepts can be represented visually, however, these can be limited in their ability to show the idea in great detail.

16E Questions

Mild 🌶 Medium 🌶🌶 Spicy 🌶🌶🌶

Deconstructed

Use the following information to answer questions 11–13.

The following are excerpts from an article titled 'What is a virus? How do they spread?'

> **Are viruses alive?**
> Viruses rely on the cells of other organisms to survive and reproduce, because they can't capture or store energy themselves. In other words they cannot function outside a host organism, which is why they are often regarded as non-living.
>
> Outside a cell, a virus wraps itself up into an independent particle called a virion. The virion can "survive" in the environment for a certain period of time, which means it remains structurally intact and is capable of infecting a suitable organism if one comes into contact.
>
> **How do viruses spread?**
> Once a person is infected with a virus, their body becomes a reservoir of virus particles which can be released in bodily fluids – such as by coughing and sneezing – or by shedding skin or in some cases even touching surfaces.
>
> The virus particles may then either end up on a new potential host or an inanimate object. These contaminated objects are known as fomites, and can play an important role in the spread of disease.
>
> Source: Tajouri, L. (2020). *What is a virus? How do they spread? How do they make us sick?* The Conversation. *theconversation.com/what-is-a-virus-how-do-they-spread-how-do-they-make-us-sick-133437*

Question 11 (1 MARK) 🌶

The style of scientific writing above can be described as being

A. lacking in detail.
B. clear.
C. quantitative.
D. basic.

Question 12 (1 MARK) 🌶🌶

Which of the following terms used in the article could be considered as a science-specific term related to viruses?

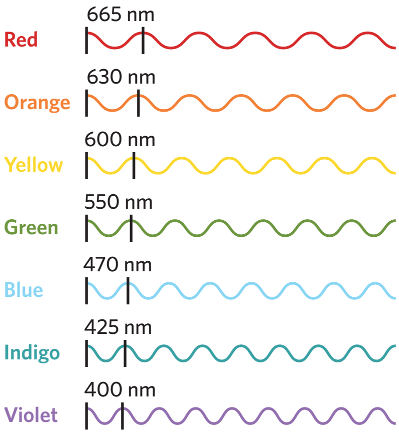
A. Non-living
B. Reservoir
C. Virion
D. Contaminated

Question 13 (3 MARKS) 🌶🌶🌶

Construct a visual model to show the process of how a handshake can help spread viruses.

Question 14 (3 MARKS)

The light that we see is a combination of different colours.



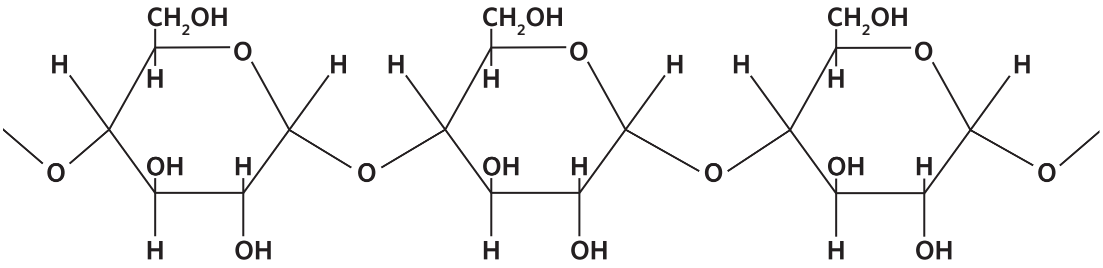
a. What is the main message conveyed by the visual model provided? (1 MARK)

b. What is the unit of measurement used to measure the wavelength of light? (1 MARK)

c. Identify the main limitation with this visual representation. (1 MARK)

Question 15 (1 MARK)

The chemical formula shown is a small section of a sugar molecule. Give one limitation of this model.



Question 16 (1 MARK)

Which of the following is not a unit of measurement commonly used in chemistry calculations?

A. L

B. M

C. N

D. K

Question 17 (3 MARKS)

Olaf wanted to conduct an experiment to see the effect of sunlight on the rate of growth of carrots. The results of the experiment are shown with the units omitted.

Test	Length of sun exposure	Length of carrot (cm)
A	0	0.15
B	1	15
C	2	24
D	3	29

a. Identify the most appropriate unit for the data relating to the length of sun exposure. (1 MARK)

b. Which of the tests would be considered as the control group? (1 MARK)

c. Identify the dependent variable. (1 MARK)

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(3 MARKS)

Medusa received a crystal growing kit as a gift and wanted to figure out how to grow the largest crystals possible. The instructions for the kit are as follows:

1. Add boiling water to a crystal growing tray.
2. Stir mixture until everything is dissolved.
3. Place rocks found in the kit at the bottom of the tray.

The kit contained multiple different sized rocks that could be used to make crystals. The three crystal mixtures included in the kit were identical.

a. Identify the independent variable. (1 MARK)

b. Write an appropriate aim for the experiment. (1 MARK)

c. Medusa wanted to write a poster to share the findings with other classmates. As part of this, outline the main sections that must be included in the poster. (1 MARK)

Hints

14a. Different colours are represented differently.
14b. Visual representations need to be labelled appropriately.
14c. Visual representations often do not include the reasoning for certain concepts.
15. Atoms are very small.
16. The databook contains all the units used in VCE chemistry.
17a. Units are appropriate for the data being measured.
17b. Controls help to demonstrate that the reason for change is due to the independent variable.
17c. The dependent variable is affected by the independent variable.
18a. The independent variable is the variable deliberately manipulated by the experimenter.
18b. An aim outlines the purpose of the experiment.
18c. Scientific writing has a predetermined format.

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Chapter 16 review

Multiple choice (10 MARKS)

Use the following information to answer questions 1–2.

Dean conducted an observational experiment to determine people's emotional response to the use of face masks.

Question 1 (1 MARK)

The data collected would be considered as

A. qualitative data.
B. quantitative data.
C. descriptive data.
D. concrete data.

Question 2 (1 MARK)

The following conclusion was drawn from the results obtained:

'As the use of face masks increased, the happiness of people decreased.'

The relationship between the use of face masks and happiness can be described as being

A. causal.
B. correlated.
C. unrelated.
D. captivated.

Question 3 (1 MARK)

The purpose of a control is to

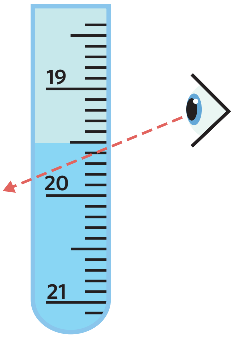
1. improve the accuracy of results.
2. improve the validity of results.
3. show that any changes were due to the independent variable.

A. 1 only
B. 3 only
C. 2 and 3 only
D. 2 only

Question 4 (1 MARK)

The error illustrated by the diagram would have the greatest effect on

A. the precision of results.
B. the accuracy of results.
C. the repeatability of an experiment.
D. the conclusion of an experiment.



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Question 5 (1 MARK)

Which of the following is **not** considered as a limitation?

A. Access to appropriate reagents
B. Misinterpretation of the methodology
C. Access to information
D. Relatively new field of research

Question 6 (1 MARK)

According to the SDS provided, the substance is dangerous as it is

A. considered 3 on the Richter scale.
B. really heavy.
C. able to freeze easily.
D. an irritant.

Hazards	
SDS	External SDS
R–phrase	R11, R36, R66, R67
S–phrase	S16, S26, S33
Main hazards	Flammable (F) Irritant (Xi)
NFPA 704	

Question 7 (1 MARK)

Forgetting to perform a step outlined in the methodology of an experiment would be considered as

A. a systematic error.
B. an obstructive error.
C. a random error.
D. bias.

Question 8 (1 MARK)

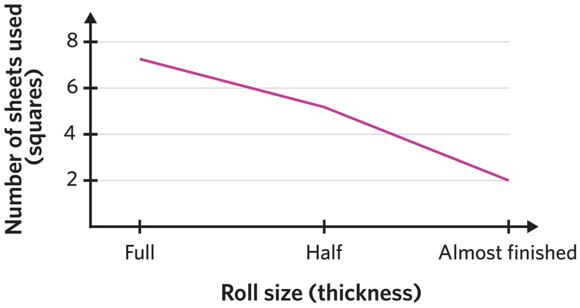
At a training session one day, Anthony measured the time it took him to run 400 m. The controlled variables in this experiment would include

A. the timer used, trial times and the shoes worn.
B. the time recorded, the running route and the apparel worn.
C. the distance, the running route and the rest period.
D. the water intake, the rest period and his pace.

Question 9 (1 MARK)

From the graph provided, it can be seen that

A. the number of sheets of paper used is constant.
B. the number of sheets used increases as the size of roll decreases.
C. the larger the roll, the fewer sheets used.
D. the smaller the roll, the fewer sheets used.



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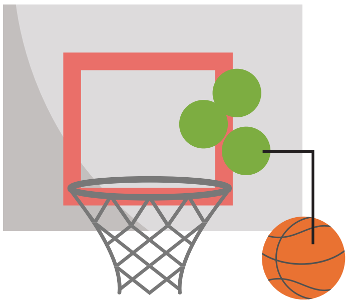
Question 10 (1 MARK) 🌶🌶



Kat was practising her basketball shots and used green paint to map each of her three shots. Based on the results, her shooting can be described as being

A. accurate but not precise.

B. precise but not accurate.

C. both precise and accurate.

D. neither accurate nor precise.

Short answer (35 MARKS)

Question 11 (4 MARKS) 🌶🌶

Oobleck is a substance that is unusual. As a liquid, it can be poured however when there is a force acting on it, it behaves like a solid. Oobleck is easy to make and consists of corn flour and water.

Shuka wanted to test the effect of the content of corn flour on the 'hardness' of the oobleck.

a. Identify the independent variable. (1 MARK)

b. What type of data is being collected in this experiment? (1 MARK)

c. As part of the experiment, Shuka tested the hypothesis using the following samples:

Sample 1	Sample 2	Sample 3
1.0 g cornflour + 100 mL of water	2.0 g of cornflour + 200 mL of water	3.0 g of cornflour and 300 mL of water

Comment on the validity of the experiment. (2 MARKS)

Question 12 (6 MARKS) 🌶🌶

To reduce the impact of diet-related diseases, researchers have tried to develop and suggest a taxation system that helps Australians to become healthier. The model that the researchers developed is shown.

Intervention	Tax or subsidy	Sources of assumptions
Saturated fat tax	\$1.37/100 g of saturated fat	Tax on saturated fat content in foods with > 2.3% saturated fat, excluding drinking milk
Excess salt tax	\$0.30/1 g of sodium	Tax on sodium in excess of Australian maximum recommended levels, excluding fresh fruits, vegetables, meats and dairy products
Sugar-sweetened beverage tax	\$0.47/L	Tax on sugar-sweetened soft drinks, energy drinks, cordials, and fruit drinks
Sugar tax	\$0.94/100 mL of ice cream; \$0.85/100 g of sugar	Tax on ice cream containing > 10 g of sugar per 100 g of ice cream; tax on sugar content in excess of 10 g per 100 g of all other products, excluding fresh fruits, vegetables and unflavoured dairy products

Table: doi: 10.1371/journal.pmed.1002232.t001

a. Write a hypothesis for this experiment. (1 MARK)

b. Suggest what effect this tax system would have on food consumption. (2 MARKS)

c. What are the main concerns related to this system? (3 MARKS)

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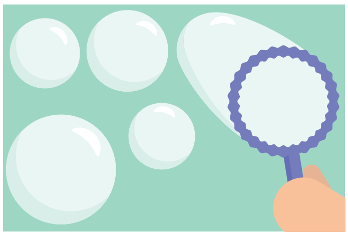
Question 13 (9 MARKS)



Jasper loves playing with bubbles. After running out of his bubble solution, he decided to develop a solution that would make the largest bubbles.

After conducting research, Jasper realised that bubble solutions are made from different substances and water. Therefore, he decided to test three different solutions containing dishwashing detergent, glycerin and corn syrup.

Method

1. Add 5 mL of dishwashing detergent to 200 mL of water.
2. Stir solution.
3. Place the bubble wand into the solution for 3 seconds.
4. Pull the bubble wand out of the solution and wave the hand once to make bubbles.
5. Take a photo of the bubbles formed.
6. Measure the diameter of the largest bubble in the photo.
7. Repeat for glycerin and corn syrup.

a. Identify an appropriate control for this experiment. (1 MARK)

b. The results of the experiment are shown.

	Diameter of the largest bubble (cm)			
Substance	**Test 1**	**Test 2**	**Test 3**	**Average**
Dishwashing detergent	5	4.8	3	4.3
Glycerin	7.1	7.2	7	7.1
Corn syrup	7.5	7.6	7.5	7.5

i. Would a bar graph, column graph or line graph be most suitable to represent the results shown in the table? (1 MARK)

ii. Which substance gave the most precise results? (1 MARK)

iii. Comment on the reproducibility of the results. (5 MARKS)

iv. Based on the data, what can be concluded about the experiment? (1 MARK)

Question 14 (8 MARKS)

Scientist Émilie du Châtelet proposed a law stating that energy cannot be created or destroyed; it can only be transformed or transferred. This principle is still believed to be true to this day.

a. Consider the scenario shown, involving the use of a coal barbecue to cook food. Based on this scenario, a student made the following statement about the law of conservation of energy. 'Energy can be lost because the chemical energy that was available in the coal was used to start a fire to cook food. However, some of the heat escaped the barbecue and was not able to be used to cook the food.' Comment on the accuracy of the statement made. (4 MARKS)

b. To test the theory, the student wanted to develop an experiment that could measure the energy in the coal, the barbecue, the food and the air to see if the total amount of energy would be the same.

i. Identify a limitation of the experiment. (1 MARK)

ii. List 3 variables that would need to be controlled. (3 MARKS)

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Question 15 (8 MARKS) 🌶🌶🌶

The taste buds on our tongue are sensitive to different tastes. The amount of food required to 'taste' something can vary between people, but can also be trained. For example, we can train our taste buds to increase the threshold of sourness that we can handle before we register the sour taste by consuming foods with a high sourness rating.

In an experiment to test this hypothesis, researchers wanted to find out whether or not having subjects consume varying concentrations of vinegar prior to eating a piece of lemon would reduce a person's reaction to consuming the lemon.

a. Identify the independent variable. (1 MARK)

b. To measure the response of a person to consuming the lemon, the experiment measured both facial expressions and a self-rating score of one to five, where five corresponded to 'very sour'. The results of the experiment are shown.

Concentration of vinegar (M)	1.0	2.0	3.0	4.0
Facial expression	😖	🙁	😢	😫
Self rating	3	2	4	4

i. Describe a key ethical concern in this experiment. (2 MARKS)

ii. Comment on the validity of the experiment. (4 MARKS)

iii. What conclusion can be drawn from the experiment? (1 MARK)

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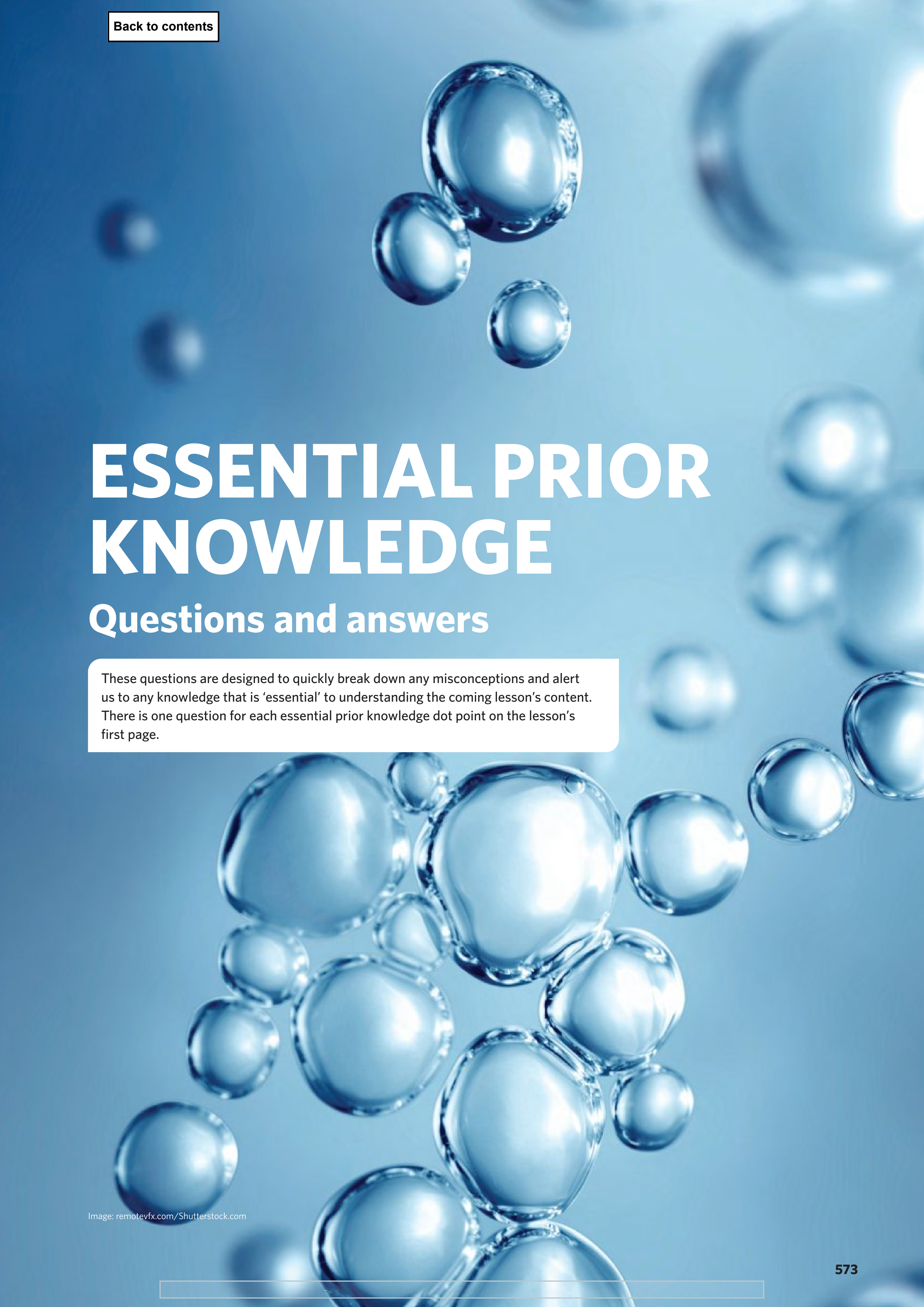

ESSENTIAL PRIOR KNOWLEDGE

Questions and answers

These questions are designed to quickly break down any misconceptions and alert us to any knowledge that is 'essential' to understanding the coming lesson's content. There is one question for each essential prior knowledge dot point on the lesson's first page.

Image: remotevfx.com/Shutterstock.com

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Essential prior knowledge questions

1A Atoms and elements

Question 1

How are the different types of atoms known in chemistry?

A. Substances

B. Elements

Question 2

Which of the following is not a subatomic particle?

A. Isotope

B. Electron

1B The periodic table - part 1

Question 3

Which of the following is not a subatomic particle?

A. Nucleus

B. Electron

Question 4

The atom of an element from the periodic table has equal numbers of both protons and electrons?

A. True

B. False

1C The periodic table - part 2

Question 5

Another term for the outermost electrons in the shell of an atom is

A. valence electrons.

B. bonding electrons.

Question 6

The size of an atom in chemistry is measured using the

A. diameter of the atom.

B. radius of the atom.

1D Recycling critical elements

Question 7

The periodic table of elements mainly consists of

A. non-metals.

B. metals.

Question 8

Blocks in the periodic table are determined by the number of electron shells.

A. True

B. False

2A Covalent bonding

Question 9

The valence shell of an atom is the

A. inner shell.

B. outer shell.

Question 10

According to the octet rule, atoms are most stable with how many electrons in their valence shell?

A. Two

B. Eight

Question 11

Which element is the most electronegative?

A. Fluorine

B. Oxygen

2B Intramolecular bonding and intermolecular forces

Question 12

A difference in electronegativity between two covalently bonded atoms results in

A. a non-polar covalent bond.

B. a polar covalent bond.

Question 13

A molecule with polar covalent bonds must be polar overall.

A. True

B. False

2C Macromolecules

Question 14

A single covalent bond is formed when

A. two atoms share a pair of electrons.

B. one atom donates an electron to another atom.

Question 15

The melting point of a substance is the temperature at which

A. the substance changes from a solid to a liquid.

B. the substance changes from a liquid to a gas.

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3A Metals

Question 16

Which word refers to the process of an atom losing an electron?

A. Electronegativity

B. Ionisation

3B Reactivity of metals - part 1

Question 17

Which of the following correctly describes ionisation energy?

A. The amount of energy required for a given amount of gaseous metal atoms to gain an outer electron.

B. The amount of energy required for a given amount of gaseous metal atoms to lose an outer electron.

Question 18

How does atomic radii affect the strength of attraction between the nucleus and the outer shell electrons?

A. The smaller the radius the greater the attraction

B. The smaller the radius the less the attraction

3C Metal recycling

Question 19

Metals cannot bond to other metals; they can only physically bond, forming mixtures. These mixtures are called

A. compounds.

B. alloys.

4A Ionic bonding

Question 20

According to the octet rule, how many electrons does each atom in an ionic bonding pair have in its valence shell?

A. 6

B. 8

4B Ionic compounds and equations

Question 21

In the Bohr model, which subatomic particles are contained in the nucleus of an atom?

A. Electrons

B. Protons

Question 22

An element with a low first ionisation energy ___________ lose electrons readily.

A. would

B. wouldn't

Question 23

Which groups do transition metals occupy on the periodic table?

A. 1–2

B. 3–12

Question 24

Crystal lattices demonstrate the 3D structure of ionic compounds.

A. True

B. False

4C Precipitation and ionic compounds

Question 25

What is the ionic formula of sodium hydroxide?

A. OHNa

B. NaOH

Question 26

Which of the following equations is balanced accurately?

A. $KCl + MgSO_4 \rightarrow K_2SO_4 + MgCl_2$

B. $3Ba(OH)_2 + 2(NH_4)_3PO_4 \rightarrow 6NH_4OH + Ba_3(PO_4)_2$

5A Investigating solubility

Question 27

Which of the following molecules is polar?

A. CH_4

B. CH_3OH

Question 28

Hydrogen bonds are the ___________ type of intermolecular force.

A. strongest

B. weakest

Question 29

Dipole-dipole forces exist between atoms of similar electronegativity.

A. True

B. False

6A Relative mass

Question 30

Which of the following best describes isotopes?

A. Atoms with the same number of neutrons but a different number of protons

B. Atoms with a different number of neutrons but the same number of protons

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Question 31

The term molecular formula refers to the formula of

A. covalent compounds.

B. ionic compounds.

Question 32

Ionic compounds exist as flat, 2D structures.

A. True

B. False

6B Avogadro's constant and the mole

Question 33

What is the molecular formula of carbon dioxide?

A. CO

B. CO_2

Question 34

The masses of all elements on the periodic table are given relative to

A. carbon–12.

B. oxygen–16.

6C Calculations involving mass

Question 35

Number of moles (n), mass (m) and molar mass (M) are related by the equation:

A. $n = \frac{m}{M}$

B. $n = mM$

Question 36

How many moles of elements are in TNT, $C_7H_5N_3O_6$?

A. 21

B. 1.26×10^{25}

7A Grouping hydrocarbons

Question 37

Covalent bonding involves the ___________ of electrons.

A. sharing

B. transfer

Question 38

Intermolecular forces occur

A. inside molecules.

B. between molecules.

7B Nomenclature

Question 39

Hydrocarbons are composed of ___________ and ___________, bonded ___________.

A. carbon, hydrogen, covalently

B. carbon, sodium, ionically

Question 40

A hydrocarbon with a double carbon to carbon bond is

A. saturated.

B. unsaturated.

7C Functional groups

Question 41

Electronegativity describes the ability of an atom to ___________ electrons.

A. attract

B. repel

Question 42

In hydrocarbons, carbon atoms are usually arranged in a ___________ sequence.

A. spherical

B. linear

Question 43

Parent carbon chains with only carbon to carbon single bonds present are given the suffix ___________.

A. '-ane'.

B. '-ene'.

7D Renewable sources

Question 44

Hydrocarbons are composed of ___________

A. hydrogen and carbon only.

B. hydrogen, carbon, and any other elements.

7E Organic compounds and society

Question 45

Examples of functional groups include

A. alkanes and alkenes.

B. halogens, hydroxyl, and carboxyl.

8A Polymers

Question 46

In structural formulas, a line represents

A. a covalent bond.

B. a hydrogen bond.

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Question 47

Much more energy is required to disrupt covalent bonds compared to intermolecular forces.

A. True

B. False

Question 48

A high melting point is indicative of

A. strong intermolecular forces.

B. weak intermolecular forces.

Question 49

A carboxylic acid contains the group of atoms __________, while an alcohol contains the group of atoms __________.

A. OH, COOH

B. COOH, OH

8B Plastics in our society

Question 50

Fossil fuel-based products are __________, while products from biomass are __________.

A. non-renewable, renewable

B. renewable, non-renewable

Question 51

Microplastics are

A. too small to be harmful to the environment.

B. extra harmful due to their small size.

Question 52

Thermoplastics __________ when heated, while thermosets __________ when heated.

A. melt, decompose

B. decompose, melt

8C Innovations in polymers

Question 53

Circular economies involve the recycling and reuse of resources.

A. True

B. False

Question 54

Alcohols are functional groups.

A. True

B. False

9A Physical and thermal properties of water

Question 55

Elements in the same group tend to have similar chemical and physical properties.

A. True

B. False

Question 56

Water is a __________ molecule.

A. polar

B. non-polar

Question 57

To undergo hydrogen bonding with itself, a substance must have

A. a hydrogen atom covalently bonded to an F, O, or N atom.

B. a hydrogen atom covalently bonded to an F, O, or N atom with a lone pair of electrons.

Question 58

Dispersion forces are usually __________ than hydrogen bonds.

A. stronger

B. weaker

10A Acids and bases

Question 59

What other name can be used to describe a hydrogen ion, H^+?

A. Proton

B. Neutron

10B Reactions of acids and bases

Question 60

Basic compounds consist of

A. non-metals only.

B. metals and non-metals.

Question 61

Acids can be defined as

A. proton donors.

B. proton acceptors.

Question 62

Which type of reaction is shown?

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

A. Combustion

B. Neutralisation

Question 63

If a solute does not dissolve in a solvent then the solute is said to be

A. soluble.

B. insoluble.

10C Calculating pH

Question 64

Water would be described as an __________ substance.

A. amphiprotic

B. diprotic

Question 65

A strong base would fully ionise in water.

A. True

B. False

10D Comparing pH measurements

Question 66

Which of the following is classified as a weak acid?

A. HCl

B. CH_3COOH

Question 67

Repeated trials in an experiment can __________ the accuracy of results.

A. increase

B. decrease

10E Applications of acid-base reactions in society

Question 68

When an element or a species gains oxygen is it said to have been

A. oxidised.

B. reduced.

Question 69

Weak acids are completely ionised.

A. True

B. False

Question 70

Basic and alkaline substances are proton acceptors.

A. True

B. False

Question 71

An acid reacting with an alkali can be classed as a

A. combustion reaction.

B. neutralisation reaction.

Question 72

A hydrogen ion consists of

A. a proton.

B. an electron and a proton.

11A Redox reactions and equations

Question 73

The smallest subatomic particles covered in VCE chemistry are

A. protons.

B. electrons.

Question 74

In ionic bonding, electrons are shared between atoms.

A. True

B. False

11B Reactivity of metals – part 2

Question 75

All metals react to the same extent with dilute acids.

A. True

B. False

Question 76

Which of the following gases are produced from the reaction between a reactive metal and hydrochloric acid, HCl?

A. H_2

B. O_2

11C Applications of redox reactions in society

Question 77

What type of reaction is occuring in the following half-equation?

$Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$

A. Reduction

B. Oxidation

Question 78

Which of the following metals is the strongest reducing agent?

A. $Mg(s)$

B. $Li(s)$

12A Measuring solutions

Question 79

Solutions are composed of a solvent and a solute.

A. True

B. False

Question 80

Which of the following equations can be used to calculate the amount of substance, in mol?

A. $E = mc^2$

B. $n = \frac{m}{M}$

12B Predicting solubility

Question 81

A precipitation reaction involves the formation of a(n) ___________ salt from two ___________ reactants.

A. soluble, gaseous

B. insoluble, aqueous

Question 82

Which of the following compounds is soluble in water?

A. KNO_3

B. AgCl

Question 83

For a substance to dissolve, which of the following is true?

A. The solvent-solute forces need to be overcome.

B. The solvent-solvent forces need to be overcome.

13A Volumetric analysis

Question 84

What is a mole a measure of?

A. Acidity

B. Quantity

Question 85

Indicators are used to determine the ___________ of a solution.

A. pH

B. temperature

Question 86

A solution that changes from 0.1 mol L^{-1} to 0.01 mol L^{-1} has undergone a ___________ dilution.

A. 10 ×

B. 100 ×

14A Gases

Question 87

A molecule always involves ___________ bonding.

A. ionic

B. covalent

Question 88

Stoichiometric ratios are always found in

A. a balanced equation.

B. the VCE Data Book.

15A Analysing ions using gravimetric analysis

Question 89

Moles are a measure of the

A. amount of particles in a sample.

B. the number of protons in the nucleus of an atom.

Question 90

Which of the following is the actual formula of a compound?

A. Empirical

B. Molecular

Question 91

Precipitation reactions produce an insoluble salt.

A. True

B. False

Question 92

Concentration of a solution is measured in

A. L mol^{-1}

B. mol L^{-1}

15B Analysing ions using light

Question 93

When atoms absorb energy they can absorb different coloured light.

A. True

B. False

16A Scientific research

Question 94

If it rains two Wednesdays in a row, the idea that it is more likely to rain on a Wednesday is a

A. scientific idea.

B. non-scientific idea

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16B Conducting an experiment

Question 95

Which of the following is a scientific consideration?

A. How much energy is used in a process

B. What you feel about an energy process

Question 96

Non-numerical data could be

A. how much something costs.

B. the colour of a person's eyes.

16C Interpreting data

Question 97

The ___________ variable is manipulated by the experimenter, whilst the ___________ variable is affected by the ___________ variable.

A. independent, dependent, independent

B. dependent, independent, dependent

Question 98

The mass of a sample is an example of

A. quantitative data.

B. qualitative data.

Question 99

The eye colour of an individual is an example of

A. quantitative data.

B. qualitative data.

16D Experimental factors affecting data

Question 100

Which of the following is a factor that can affect writing a conclusion?

A. The type of evidence used

B. How long the data took to analyse

16E Writing scientific material

Question 101

A hypothesis is

A. a prediction of how the independent variable will affect the dependent variable.

B. a prediction of what will happen in an experiment.

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Essential prior knowledge answers

1A	1. B	2. A			
1B	3. A	4. A			
1C	5. A	6. B			
1D	7. B	8. B			
2A	9. B	10. B	11. A		
2B	12. B	13. B			
2C	14. A	15. A			
3A	16. B				
3B	17. B	18. A			
3C	19. B				
4A	20. B				
4B	21. B	22. A	23. B	24. A	
4C	25. B	26. B			
5A	27. B	28. A	29. B		
6A	30. B	31. A	32. B		
6B	33. B	34. A			
6C	35. A	36. B			
7A	37. A	38. B			
7B	39. A	40. B			
7C	41. A	42. B	43. A		
7D	44. A				
7E	45. B				
8A	46. A	47. A	48. A	49. B	
8B	50. A	51. B	52. A		
8C	53. A	54. B			
9A	55. A	56. A	57. B	58. B	
10A	59. A				
10B	60. B	61. A	62. B	63. B	
10C	64. A	65. A			
10D	66. B	67. A			
10E	68. A	69. B	70. A	71. B	72. A
11A	73. B	74. B			
11B	75. B	76. A			
11C	77. A	78. B			
12A	79. A	80. B			
12B	81. B	82. A	83. B		
13A	84. B	85. A	86. A		
14A	87. B	88. A			
15A	89. A	90. B	91. A	92. B	
15B	93. A				
16A	94. B				
16B	95. A	96. B			
16C	97. A	98. A	99. B		
16D	100. A				
16E	101. A				

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ANSWERS

CONTENTS

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1A Atoms and elements

Progress questions

1. B. The atomic number is the number of protons in the nucleus of an atom.

2. A. Protons are positively charged, electrons are negatively charged, neutrons are neutral and have no charge.

3. A. Protons and electrons have equal and opposite charges so cancel each other.

4. B. Mass = number of protons + number of neutrons

5. B. In a neutral atom the number of protons equals the number of electrons.

6. B. The proton number is equal to the atomic number. Isotopes of the same element have the same number of protons, but a different number of neutrons.

7. A. Elements on the periodic table are ordered horizontally (from left to right) in order of increasing atomic (proton) number.

8. A. The electron has a significantly lower mass (about 1/1800) than both the proton and neutron.

9. A. number of neutrons = mass number − proton number

10. A. Mass number does not include the number of electrons and is found by only adding the number of protons and neutrons found in the nucleus.

11. B. Isotopes vary in the number of neutrons and mass number.

12. B. Some isotopes can be radioactive.

Deconstructed

13. B

14. C

15. [The mass number is the sum of protons and neutrons in an atom.[1]] [Based on the diagram, the image on the left contains a total of nine protons and 10 neutrons, therefore would have a mass of 19. The diagram on the right shows an atom with 10 protons and 10 neutrons, therefore would have a mass number of 20.[2]] [Although both contain the same number of neutrons, they both have a different mass number due to the number of protons in the nucleus.[3]]

✓ ✗ I have defined mass number.[1]

✓ ✗ I have determined the mass number of both atoms.[2]

✓ ✗ I have described the key difference in mass number of the atoms.[3]

Exam-style

16. Different number of neutrons in the nucleus

17. [The atom contains six protons and six neutrons in its nucleus.[1]] [If it were to undergo neutron release, the atom would lose a neutron to become a different isotope.[2]] [The atom would then only contain five neutrons instead of 6, compared to the original atom.[3]]

✓ ✗ I have identified the particles in the nucleus.[1]

✓ ✗ I have described the effect of neutron release.[2]

✓ ✗ I have compared the nucleus of the atom before and after neutron release.[3]

18. **a.** [R = 34.[1]] [I = 36.[2]] [T = 37.[3]] [A = 39.[4]]

✓ ✗ I have identified the mass number of R.[1]

✓ ✗ I have identified the mass number of I.[2]

✓ ✗ I have identified the mass number of T.[3]

✓ ✗ I have identified the mass number of A.[4]

b. I and T

c. [Considering elements I and T are isotopes of each other, they have the same atomic number.[1]] [As such, they should have the same elemental symbol, and therefore the table should be changed so that both elements share the same elemental symbol.[2]]

✓ ✗ I have identified the key characteristic of isotopes.[1]

✓ ✗ I have described the change that needs to be made to the table.[2]

d. [Atomic number = 19.[1]] [Mass number = 39.[2]] [Symbol = A.[3]]

✓ ✗ I have identified the atomic number for element A.[1]

✓ ✗ I have identified the mass number for element A.[2]

✓ ✗ I have identified the symbol for element A.[3]

e. [Both molecules and compounds require the bonding of two atoms.[1]] [If atoms of the elements R and A were to bond together, this would mean that there are two atoms involved in RA. As such, it could be defined as a molecule.[2]] [However, considering A is a metal and R is a non-metal it would be called a compound.[3]]

✓ ✗ I have identified the common characteristic of both molecules and compounds.[1]

✓ ✗ I have explained how RA can be considered a molecule.[2]

✓ ✗ I have described how RA is a compound.[3]

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Key science skills

19. a. By calibrating the scales, Matthew is minimising the potential for systematic errors.

b. [The mass number of an element is the total number of protons and neutrons in the atom.[1]][Since every element has a different atomic number, it would have a different number of protons compared to other elements.[2]][However, elements can also exist as isotopes, therefore the mass number of elements can still be the same given that the number of neutrons can be different.[3]][As a result, Matthew's statement is inaccurate.[4]]

✓ ✗ I have defined mass number.[1]

✓ ✗ I have described the difference in protons between elements.[2]

✓ ✗ I have described the impact of isotopes on mass number.[3]

✓ ✗ I have linked my answer to the question.[4]

FROM LESSON 16D

1B The periodic table – part 1

Progress questions

1. A. More than one electron can occupy an energy shell.

2. A. Orbitals are created by solving a mathematical equation and do not exist in reality.

3. B. Electrons do not completely fill shells before moving onto the next shell in this case the 4s subshell is filled first (as it is lower energy than the the 3d subshell).

4. B. Electrons fill by order of energy from lowest to highest.

5. B. Elements in the same period contain the same number of occupied electron shells.

6. A. Elements in the same group have the same number of valence electrons.

7. A. Groups and periods are seen as columns and rows respectively on the periodic table.

8. B. Elements across a period have an increasing number of protons.

9. B. Elements in the same period have the same number of inner shell electrons and therefore atomic radius decreases across the period.

10. B. Atomic radius decreases across a period.

11. B. An increase in the number of protons in the nucleus causes elements in the same period to have a smaller atomic radius.

12. B. Blocks in the periodic table highlight the outermost subshell.

13. B. Different blocks are found in the same period.

14. B. The p subshell can hold a maximum of six electrons.

15. B. Oxygen is found in group 16, which is located in the p block.

Deconstructed

16. D

17. D

18. $1s^22s^22p^63s^23p^4$ (2 MARKS)

Exam-style

19. B. All energy level subshells have a fixed maximum number of electrons that they can hold (s = 2, p = 6, d = 10 and f = 18).

20. a. Ca and V, P and Si

b. Be and Ca or P and Si

c. Calcium

21. a. Element A: 32 protons Element B: 20 protons

b. [Element A: $1s^22s^22p^63s^23p^64s^23d^{10}4p^2$.[1]]
[Element B: $1s^22s^22p^63s^23p^64s^2$.[2]]

✓ ✗ I have correctly written Element A's subshell electron configuration.[1]

✓ ✗ I have correctly written Element B's subshell electron configuration.[2]

c. [Both elements are found in the same period, meaning that they have the same number of occupied energy shells, however element A has more protons in its nucleus than element B.[1]][Therefore, the force of attraction between the positive nucleus and the valence electrons of element A is greater than that of element B.[2]][As a result, element A is able to pull the valence shell closer to the nucleus and would therefore have a smaller atomic radius compared to element B.[3]]

✓ ✗ I have compared the similarity and difference between both elements.[1]

✓ ✗ I have compared the ability of each element to attract the valence shell.[2]

✓ ✗ I have linked my answer to the question.[3]

Key science skills

22. a. Clarity of the image

b. i. Qualitative

ii. [Reproducibility is based on the closeness of data collected upon repeating an experiment under the same conditions.[1]][Although test 1 and test 2 were conducted under the same conditions, the data collected varies.[2]][Based on these results, the experiment seems to have low reproducibility.[3]]

✓ ✗ I have defined reproducibility.[1]

✓ ✗ I have identified the conditions of both experiments.[2]

✓ ✗ I have linked my answer to the question.[3]

FROM LESSONS 16A, 16B & 16D

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Questions from multiple lessons

23. **a.** 113; 113; 171

b. group 13 (or group 3) period 7

c. Tl (Thallium)

d. The number of protons in the nucleus of the atoms increases across the period.

FROM LESSON 1A

1C The periodic table – part 2

Progress questions

1. B. Electronegativity describes the ability of an atom to attract electrons to itself.

2. A. It increases due to the decrease in atomic radii of the elements.

3. B. It decreases due to the addition of extra shells of electrons.

4. A. Elements across a period have an increasing number of protons with the same number of inner shell electrons.

5. B. The core charge is the same for all elements in the same group regardless of the atomic size.

6. B. Atomic radius decreases across a period.

7. B. An increase in core charge causes elements in the same row to have a smaller atomic radius.

8. A. Elements towards the top of a group have a smaller atomic radius, and therefore the valence electrons are held closer to the nucleus.

9. A. The closer the valence electrons are to the nucleus, the more difficult it is to remove a valence electron.

10. B. Metals are known for their ability to easily lose electrons.

11. B. The most metallic elements are found in groups 1 and 2.

12. B. Metalloids possess both metallic and non-metallic characteristics and are therefore considered to be intermediates between metals and non-metals.

13. A. Highly reactive atoms are those that can gain or lose electrons easily.

14. A. Highly reactive atoms are those that can lose or gain electrons relatively easily.

15. B. It becomes harder for elements to lose electrons as we move between groups.

Deconstructed

16. D

17. B

18. [Barium and magnesium are both found in group 2.[1]] [Barium, found in period 6, has a greater atomic radius than magnesium, found in period 3, and therefore has a weaker attraction between its nucleus and valence electrons in comparison to magnesium.[2]][As a result, barium will lose electrons more readily than magnesium, and therefore would react more quickly with water than magnesium would.[3]]

✓ ✗ I have identified the groups of both elements on the periodic table.[1]

✓ ✗ I have compared the atomic radius of both elements.[2]

✓ ✗ I have compared the reactivity of both elements.[3]

Exam-style

19. **a.** Sodium: +1
Potassium: +1

b. [Sodium and potassium are found in the same group and also have the same core charge.[1]][However, potassium has one more energy shell than sodium, and would therefore have a greater atomic radius than sodium.[2]]

✓ ✗ I have identified the similarity between both elements.[1]

✓ ✗ I have described the differences in atomic radius.[2]

c. [Since potassium has a greater atomic radius than sodium, the attraction between the positive nucleus of potassium and its valence electrons is weaker than that of sodium given they have equal core charge.[1]][As a result, potassium would have a lower ionisation energy than sodium.[2]]

✓ ✗ I have compared the attraction of each element to its valence electrons.[1]

✓ ✗ I have identified the element with a lower ionisation energy.[2]

20. **a.** [Fluorine and bromine are both found in group 17 and would therefore have the same number of valence electrons and core charge.[1]][However bromine's atomic radius is larger than fluorine due to the increase in the number of energy shells.[2]] [As a result, it would be harder to remove a valence electron from fluorine than bromine.[3]]

✓ ✗ I have compared the properties of both elements.[1]

✓ ✗ I have compared the atomic radius of both elements.[2]

✓ ✗ I have linked my answer to the question.[3]

b. [Although both elements have the same core charge, fluorine has a smaller atomic radius compared to bromine.[1]] [As a result, fluorine is able to attract electrons to itself more strongly, and therefore would be more electronegative than bromine.[2]][Therefore, it is most likely that fluorine has an electronegativity score of 3.98 and bromine 2.96.[3]]

✓ ✗ I have compared the properties of both elements.[1]

✓ ✗ I have compared the ability of both elements to attract electrons.[2]

✓ ✗ I have linked my answer to the question.[3]

c. Fluorine

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21. [Row 1: Incorrect.[1]][Metals with high reactivity lose electrons easily and have low electronegativity. Non-metals with high reactivity gain electrons easily and have high electronegativity.[2]]
[Row 2: Incorrect.[3]][Depending on the atomic structure, elements can either gain or lose electrons during a reaction.Metals down the bottom of a group have electrons that are less attracted to the nucleus and can therefore lose electrons easily and are highly reactive. Non-metals at the top of a group will gain electrons easily and are highly reactive.[4]]
[Row 3: Correct.[5]][The elements on the left side of the periodic table have fewer valence electrons, and so prefer to lose electrons and therefore act as metals to satisfy the octet rule.[6]]
[Row 4: Correct.[7]][The amount of energy it takes to remove a valence electron from an atom is the first ionisation energy. Given that the noble gases have full valence shells, the amount of energy needed to remove a valence shell electron is much higher compared to the other elements.[8]]

✓ ✗ I have identified whether the statement is correct or incorrect.[1]

✓ ✗ I have given a justification for my choice.[2]

✓ ✗ I have identified whether the statement is correct or incorrect.[3]

✓ ✗ I have given a justification for my choice.[4]

✓ ✗ I have identified whether the statement is correct or incorrect.[5]

✓ ✗ I have given a justification for my choice.[6]

✓ ✗ I have identified whether the statement is correct or incorrect.[7]

✓ ✗ I have given a justification for my choice.[8]

Key science skills

22. a. i. Compound

ii. [Na: Period 3, group 1.[1]][Cl: Period 3, group 17.[2]]

✓ ✗ I have identified the period and group number for sodium.[1]

✓ ✗ I have identified the period and group number for chlorine.[2]

b. [The data indicates that the first ionisation energy increases left-to-right across a period.[1]][This is explained by the increase in core charge across the period in the data.[2]][Given that we are comparing elements in the same period, as the core charge increases, the energy required to remove an electron increases, as they are more strongly attracted to the nucleus.[3]]

✓ ✗ I have interpreted the data.[1]

✓ ✗ I have linked core charge to the data.[2]

✓ ✗ I have explained the link between core charge and first ionisation energy.[3]

c. First ionisation energy

d. [Systematic error.[1]][A systematic error affects all measurements to the same extent, in the same direction.[2]]

✓ ✗ I have identified the type of error.[1]

✓ ✗ I have explained how this error would affect the results.[2]

FROM LESSONS 16C, 16A & 16D

Questions from multiple lessons

23. a. Both are from period 2

b. [Both atoms have 2 electron shells but a different number of electrons and therefore a different number of protons.[1]][Element A has nine protons, whereas element B only has eight protons.[2]][As a result, A has a greater positive charge in comparison to B and is therefore able to attract valence electrons more strongly. Therefore, it would take more energy to remove a valence electron from A than from B, meaning A has a higher ionisation energy than B.[3]][Furthermore, due to the stronger attraction between A's nucleus and the valence shell, A would have a smaller atomic radius compared to B.[4]][As a result, Element A is fluorine and Element B is oxygen.[5]]

✓ ✗ I have identified the similarity and difference between both elements.[1]

✓ ✗ I have compared the ability of both elements to attract valence electrons.[2]

✓ ✗ I have compared the ionisation energy of both elements.[3]

✓ ✗ I have compared the atomic radius of both elements.[4]

✓ ✗ I have linked my answer to the question.[5]

FROM LESSON 1B

24. a. $1s^22s^22p^1$

b. p block

FROM LESSON 1B

1D Recycling critical elements

Progress questions

1. A. Without a critical element many important products or materials could not be created.

2. B. Sc, Ti, and V are all classified as transition metals.

3. B. The actinoids are below the lanthanoids.

4. A. Using an electromagnet would separate these three metals from all other non-magnetic metals.

5. A. Helium is not recovered as it was not used in the first instance.

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Deconstructed

6. D

7. A

8. [Phosphorus is found in protein-rich foods.^{1}][Most humans eat protein and therefore have phosphorus in their urine as a waste product.^{2}][If urine (and sewage) was collected it could be evaporated and condensed.^{3}][The phosphorus could then be used again by either animals or plants that are again consumed.^{4}]

✓ ✗ I have identified where the source of phosphorus in the diet comes from.^{1}

✓ ✗ I am aware that humans need to eat some form of protein.^{2}

✓ ✗ I understand the process involved in isolating the phosphorus.^{3}

✓ ✗ I have explained how the phosphorus can be recycled.^{4}

Exam-style

9. a. Transition metals or d-block

b. Lanthanoids (Lanthanides) or f-block

c. [The infrastructure already exists for recycling metals that are commonly used like aluminium and iron.^{1}][Recycling plants for relatively new rare-earth elements in electronic devices are very rare.^{2}]

✓ ✗ I have described that some metals are much more commonly used and recycled.^{1}

✓ ✗ I have explained that some metals are not recycled.^{2}

d. [Less than 1% of lithium is recycled.^{1}][Given that lithium is consumed more rapidly than it is produced, this is not sustainable as lithium reserves will not last forever.^{2}]

✓ ✗ I have determined the percentage of lithium that is recycled.^{1}

✓ ✗ I have concluded on the sustainability of lithium giving a reason.^{2}

10. a. Zirconium, and Yttrium

b. [France cannot produce their own rare-earth metals.^{1}] [Recycling their current supply of rare-earth elements is advantageous as it reduces their dependence on other countries to supply them.^{2}]

✓ ✗ I have understood that France imports rare-earth metals.^{1}

✓ ✗ I have justified with a reason why it is an advantage.^{2}

c. [Products containing rare-earth elements are difficult to recycle.^{1}][The recycling process is very expensive or not yet possible for particular elements.^{2}]

✓ ✗ I have given a reason.^{1}

✓ ✗ I have explained the reason.^{2}

Key science skills

11. a. [In some cultures it is accepted that children work to provide an income for the family.^{1}][Even if acceptable in some countries, it is still unethical to make children work and deprive them of an education.^{2}]

✓ ✗ I have discussed an ethical issue in regards to age.^{1}

✓ ✗ I have discussed an ethical issue in regards to education.^{2}

b. [In some poorer countries, it is necessary for all able family members to work in order to survive.^{1}][Minimum wages should be in place.^{2}]

✓ ✗ I have discussed an economic issue.^{1}

✓ ✗ I have discussed an economic issue about wages.^{2}

c. [Most countries have laws regarding minimum legal age and wages.^{1}][Many countries also have laws and regulations regarding safe working conditions.^{2}]

✓ ✗ I have discussed the importance of a legal minimum wage.^{1}

✓ ✗ I have discussed working conditions and basic laws.^{2}

FROM LESSON 16B

Questions from multiple lessons

12. a. Fe

b. Ar

c. Na

d. F

e. Li

FROM LESSONS 1A & 1B

Chapter 1 review

Multiple choice

1. A. Period 3 indicates that there are three electron shells and group II indicates that there are two valence electrons.

2. C. The atomic number represents the number of protons in the nucleus of an atom which may be different to the number of neutrons.

3. C. Calcium is an alkali earth metal.

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4. D. Metallic character decreases from left to right (more difficult to remove electrons) and increases down a group (easier to remove electrons).

5. C. Do not forget to include helium which has only 2 electrons.

6. B. Electronegativity increases across the period from left to right and up a group.

7. D. Atomic radius increases initially down the group 7 and then continues to increase from right to left across the third period from Cl to Mg and then to finally increase further as we go from group 2 to group 1, period four.

8. A. The first ionisation energy increases from left to right across the third period and then increases further as we go up the inert gas group from period 3 to period 2.

9. C. Concepts of atomic number and mass number were unknown at the time of Mendeleev.

10. D. The mass number is equal to the sum of the number of protons (atomic number) and the number of neutrons.

Short answer

11. a. [Name: Titanium and chemical symbol: Ti.[1]] [Atomic number: 22.[2]] [Relative atomic mass: 47.9.[3]]

✓ ✗ I have identified the element and symbol.[1]

✓ ✗ I have referenced the atomic number.[2]

✓ ✗ I have referenced the relative atomic mass.[3]

b. [Titanium is found in the d-block.[1]] [This is because its highest energy electrons are in the d-subshell.[2]]

✓ ✗ I have identified which block on the periodic table the element is found in.[1]

✓ ✗ I have justified my answer with reference to its electron configuration.[2]

c. [Electronegativity is a measure of how strongly electrons are attracted to the nucleus of an atom. [1]] [Electronegativity increases across a period from left to right as the core charge increases and therefore, titanium will be less electronegative than copper.[2]]

✓ ✗ I have defined electronegativity.[1]

✓ ✗ I have compared the electronegativities of copper and the element.[2]

d. [Metallic character is based on an element's ability to lose electrons easily.[1]] [Since zirconium is further down in the group, it has the same core charge but its valence electrons are further away, resulting in greater metallic character as it can lose electrons more easily than titanium.[2]]

✓ ✗ I have defined metallic character.[1]

✓ ✗ I have compared the metallic character of titanium and zirconium.[2]

e. [Titanium is a metal. [1]] [Therefore, it is more likely to react by losing electrons as opposed to gaining electrons.[2]]

✓ ✗ I have identified titanium as a metal.[1]

✓ ✗ I have identified its preferred mechanism of reactivity.[2]

12. a. [Decreases.[1]] [Increasing the core charge (nuclear charge), outer electrons are pulled closer to nucleus (emphasis on the effect of the core charge on size).[2]] [Increases.[3]] [Increasing the core charge (nuclear charge), greater pull/attraction on outer shell electrons (emphasis on core charge on electron attracting power).[4]]

✓ ✗ I have identified the change in atomic size.[1]

✓ ✗ I have identified the change in electronegativity.[2]

✓ ✗ I have explained the trend in atomic size.[3]

✓ ✗ I have explained the trend in electronegativity.[4]

b. The energy required to remove an electron (not electrons).

c. [The outermost electron in potassium is further from the nucleus than the outer shell electron in sodium.[1]] [Therefore potassium has weaker attraction (the electron is more easily removed).[2]]

✓ ✗ I have described the effect of the different atomic size (this does not affect core charge).[1]

✓ ✗ I have explained the effect of the size on the attraction.[2]

13. a. [They have similar chemical properties.[1]] [They have the same number of valence electrons.[2]]

✓ ✗ I have stated a reason for their grouping.[1]

✓ ✗ I have stated a second reason for their grouping.[2]

b. They all have the same number of electron shells.

c. [The size of the calcium atom is greater than that of magnesium.[1]] [Therefore electrons are less shielded by the nucleus, and hence less strongly attracted.[2]]

✓ ✗ I have compared the sizes of calcium and magnesium.[1]

✓ ✗ I have explained how this affects the attraction between the nucleus and valence electrons.[2]

d. [Ca – $1s^22s^22p^63s^23p^64s^2$.[1]] [Ni – $1s^22s^22p^63s^23p^64s^23d^8$.[2]]

✓ ✗ I have used the appropriate conventions for writing configurations.[1]

✓ ✗ I have used the appropriate conventions for writing configurations where the 4s subshell fills before the 3d subshell.[2]

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e. [Due to systematic filling of a d-subshell,^{1}][the d-subshell can only have a maximum of 10 electrons.^{2}][After the d subshell is filled the next elements become p-block elements.^{3}]

✓ ✗ I have identified the d-subshell.^{1}

✓ ✗ I have described the capacity of the d-subshell.^{2}

✓ ✗ I have explained what happens after the d-subshell is complete.^{3}

f. [It is considered in short supply and high demand.^{1}] [Europium can be recycled, or alternative sources can be used.^{2}]

✓ ✗ I have defined what a critical element is.^{1}

✓ ✗ I have given an example of how its use can be made more sustainable.^{2}

Key science skills

14. a. The aim of this experiment is to prove the law of conservation of mass by carrying out a reaction between mercury and oxygen in a sealed vessel.

b. [If the law of conservation of mass is true, there should be no change in the mass of the sealed vessel.^{1}][This should be true both before and after the chemical reaction has occurred. ^{2}]

✓ ✗ I have stated no change in mass.^{1}

✓ ✗ I have explained that it is independent of the progress of the reaction.^{2}

c. [The mass of the sealed vessel before and after the experiment would need to be collected.^{1}][This is quantitative data.^{2}] [The reason this data needs to be collected is to prove that even though a chemical reaction has occurred, the mass of the vessel has not changed which means that atoms have only been rearranged.^{3}]

✓ ✗ I have identified what data needs to be collected.^{1}

✓ ✗ I have specified what type of data it is.^{2}

✓ ✗ I have justified why this data needs to be collected.^{3}

d. [If the jar is not sealed properly and the reactant (oxygen) escapes, the recorded mass of the vessel would be lower after the reaction has occurred compared to the initial mass, apparently disproving the law of conservation of mass.^{1}] [This is a random error.^{2}]

✓ ✗ I have identified the effect on the results the error has.^{1}

✓ ✗ I have classified the type of error.^{2}

e. Personal protective equipment including gloves, glasses and a lab coat should be worn.

FROM LESSONS 16A, 16B & 16D

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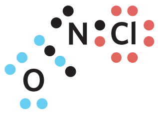
2A Covalent bonding

Progress questions

1. B. When atoms are covalently bonded together, they form molecules.
2. B. Covalent bonding relies on the sharing of electrons.
3. B. Some atoms, such as sulfur, do not obey the octet rule.
4. B. Double covalent bonds have two bond pairs of electrons.
5. A. The shape depicted has four atoms/pairs of electrons spaced equally around a central atom.
6. A. When determining molecular shape, double bonds are considered as one pair of electrons.
7. A. A non-polar covalent bond can be formed because the difference in electronegativity is negligible.
8. A. Not all compounds containing polar covalent bonds are considered polar.

Deconstructed

9. A
10. A
11. [Nitrogen-chlorine single covalent bond drawn correctly.[1]]
[Nitrogen-oxygen double covalent bond drawn correctly.[2]]



✓ ✗ I have identified the nitrogen-chlorine single covalent bond.[1]

✓ ✗ I have identified the nitrogen-oxygen double covalent bond.[2]

Exam-style

12. a. [Oxygen has six valence electrons and therefore, it needs another atom or atoms to share two electrons with it.[1]] [In this case, since both oxygen atoms need another atom to share two electrons with it, a double covalent bond would form, as each oxygen atom shares two electrons with the other oxygen atom.[2]]

✓ ✗ I have identified the number of electrons needed for oxygen to satisfy the octet rule.[1]

✓ ✗ I have described the type of covalent bond that would form.[2]

b. [Oxygen has six valence electrons and therefore, it needs another atom or atoms to share two electrons with it.[1]] [In this case, since both oxygen atoms need to be bonded to one hydrogen atom and the other oxygen atom, two single covalent bonds will form (O–H and O–O). This means the oxygen-oxygen bond will be a single covalent bond.[2]]

✓ ✗ I have identified the number of electrons needed for oxygen to satisfy the octet rule.[1]

✓ ✗ I have described the type of covalent bond that would form.[2]

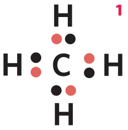
c. [Double covalent bonds are stronger than single covalent bonds because there are more electrons shared.[1]] [This means that the oxygen atoms are bonded to each other more strongly in O_2 than in H_2O_2.[2]] [This would explain why oxygen gas (O_2) is more stable than hydrogen peroxide (H_2O_2), as shown by the fact that H_2O_2 decomposes explosively while O_2 is more stable in nature.[3]]

✓ ✗ I have explained the relative strengths of double and single covalent bonds.[1]

✓ ✗ I have compared the strength of the covalent bonds present in O_2 and H_2O_2.[2]

✓ ✗ I have compared the stability of O_2 and H_2O_2.[3]

13. a. [The shape is tetrahedral.[2]] [The molecule is non-polar.[3]]



✓ ✗ I have drawn the Lewis structure of the molecule.[1]

✓ ✗ I have used this diagram to determine the shape of the molecule.[2]

✓ ✗ I have used shape and symmetry to infer the molecule's polarity.[3]

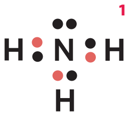
b. [The shape is linear.[2]] [The molecule is non-polar.[3]]

✓ ✗ I have drawn the Lewis structure of the molecule.[1]

✓ ✗ I have used this diagram to determine the shape of the molecule.[2]

✓ ✗ I have used shape and symmetry to infer the molecule's polarity.[3]

c. [The shape is trigonal pyramidal.[2]] [The molecule is polar.[3]]



✓ ✗ I have drawn the Lewis structure of the molecule.[1]

✓ ✗ I have used this diagram to determine the shape of the molecule.[2]

✓ ✗ I have used shape and symmetry to infer the molecule's polarity.[3]

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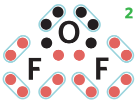
14. a.

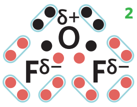
b. [The lone pairs repel each other, giving the molecule a V-shape.[1]]



✓ ✗ I have described the effect of the lone pairs of electrons on the shape of the molecule.[1]

✓ ✗ I have circled all the lone pairs of electrons.[2]

c. [OF_2 is polar.[1]]



✓ ✗ I have stated the polarity of the molecule.[1]

✓ ✗ I have used δ+ and δ− symbols to indicate the polarity of the molecule.[2]

15. [The electron clouds represent the density of electrons around each atom.[1]][When comparing molecules A and B, we can see that molecule B has an even density of electrons around both atoms whereas molecule A has an uneven electron density around each atom, where the atom found on the right is shown to have a greater density than the atom on the left.[2]][This suggests that the atom on the right of molecule A is more electronegative than the atom on the left, resulting in an uneven distribution of the shared pair of electrons.[3]][Therefore, molecule A has two permanent dipoles of differing charge and hence is polar.[4]]

✓ ✗ I have identified the function of electron clouds.[1]

✓ ✗ I have described the electron distribution of both molecules.[2]

✓ ✗ I have described the electron distribution within molecule A.[3]

✓ ✗ I have linked my answer to the question.[4]

Key science skills

16. a. The independent variable is the identity of the chemical substance A or B.

b. The aim of this experiment is to determine the type of substance present by performing a boiling point determination.

c. Quantitative

d. [Substance A is more likely to be a diatomic molecule than substance B.[1]][This is because its boiling point, a measure of the force of attraction between molecules, is much lower than substance B. Diatomic molecules have much lower boiling points compared to most covalent compounds because it generally takes more energy to disrupt the electrostatic attraction in diatomic molecules which are all gases at room temperature compared to the intermolecular forces between liquid and solid covalent compounds.[2]]

✓ ✗ I have identified which substance is more likely to be a diatomic molecule.[1]

✓ ✗ I have justified my choice by comparing the energy required to break the different forces.[2]

FROM LESSONS 16A & 16B

Questions from multiple lessons

17. a. Ca and vanadium, phosphorus and Si

b. Be and Ca

c. Calcium

d. Calcium

e. Calcium

FROM LESSONS 1B & 1C

2B Intramolecular bonding and intermolecular forces

Progress questions

1. A. Intramolecular bonds refer to the bonds within a molecule (covalent).

2. B. Increased bond length reduces the electrostatic force of attraction between two covalently bonded atoms, while increasing the number of bonding electrons has the opposite effect.

3. A. The number of electrons can affect an atom's ability to form instantaneous dipole moments, leading to stronger intermolecular forces.

4. B. Dispersion forces occur due to electrostatic attraction between instantaneous dipoles.

5. A. Permanent dipoles only occur in polar molecules. Permanent dipole-dipole attractions are much stronger than temporary dipole-dipole attractions.

6. B. Permanent dipoles can only form in polar molecules.

7. A. Hydrogen bonding is stronger than all other molecular forces due to the difference in electronegativity between hydrogen and nitrogen, oxygen or fluorine.

8. A. The greater the difference in electronegativity between bonded atoms, the more polar the compound. This results in a stronger dipole-dipole interaction. Compared to HCl, HF is more polar and therefore has stronger intermolecular forces.

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9. B. In addition to dispersion forces, water molecules also form hydrogen bonds, which requires more energy to break than dipole-dipole interactions and dispersion forces alone.

10. B. A larger surface area allows for more interactions between molecules.

Deconstructed

11. C

12. C

13. [Dispersion forces are the only intermolecular force affecting these molecules.[1]][Larger molecules have stronger dispersion forces and thus, higher boiling points.[2]][I_2 is a bigger molecule than Br_2, Cl_2 and F_2. Hence I_2 is a solid at room temperature, whilst Br_2 is a liquid and Cl_2 and F_2 are gaseous.[3]]

✓ ✗ I have identified the similarities between the intermolecular forces of all the molecules.[1]

✓ ✗ I have related atomic size to the boiling points of the molecules.[2]

✓ ✗ I have established which molecules are the largest and related this to their state at room temperature.[3]

Exam-style

14. a. CF_4

b. CBr_4

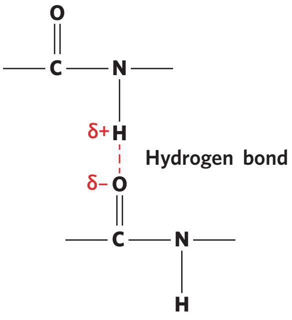
c. CBr_4

15. a. [Hydrogen bonds are a much stronger form of intermolecular forces than dispersion forces.[1]][These strong intermolecular forces will stabilise protein structure, as they are less easily disrupted than dispersion forces.[2]]

✓ ✗ I have compared the strength of hydrogen bonds to dispersion forces.[1]

✓ ✗ I have linked this comparison to the importance of hydrogen bonding in protein structures.[2]

b.



c. [Molecules drawn correctly.[1]][Hydrogen bond correctly labeled.[2]]

✓ ✗ I have drawn the two molecules.[1]

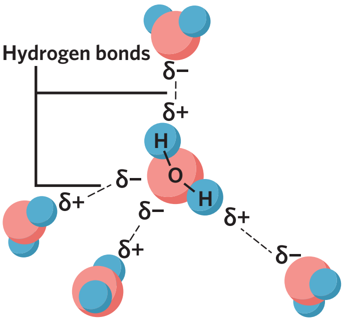
✓ ✗ I have labeled the hydrogen bond between the two molecules.[2]

d. [An increase in temperature would increase the average kinetic energy of molecules, disrupting the hydrogen bonding present.[1]][Since hydrogen bonding helps to stabilise protein structure, a disruption in hydrogen bonding would disrupt the structure and therefore function of the protein.[2]]

✓ ✗ I have identified one way to disrupt intermolecular forces.[1]

✓ ✗ I have linked this effect to protein function.[2]

16. a.



[Molecules dipoles correctly annotated.[1]][Hydrogen bonds drawn and labeled correctly.[2]]

✓ ✗ I have labeled the polarity of each atom in the molecule.[1]

✓ ✗ I have labeled the hydrogen bonds.[2]

b. [Hydrogen bonding occurs between water molecules, whereas only dispersion forces operate between CH_4 molecules.[1]] [Since hydrogen bonds are a much stronger intermolecular force than dispersion forces, more energy is required to break them. Hence, water has a much higher melting point.[2]]

✓ ✗ I have identified the type of intermolecular forces for each molecule.[1]

✓ ✗ I have linked the difference in intermolecular forces to melting point.[2]

c. [Tetradecane is a very large molecule with a large surface area and has 14 carbon atoms and 30 hydrogen atoms which means there are 114 electrons in total. As a result, the dispersion forces between molecules are very strong.[1]][Tetradecane therefore has a higher melting point than water because the dispersion forces between tetradecane molecules are stronger than the combination of dispersion forces and hydrogen bonding amongst water molecules.[2]]

✓ ✗ I have described the size of tetradecane and its effect on the strength of the dispersion forces present.[1]

✓ ✗ I have explained why tetradecane has a higher melting point than water.[2]

Key science skills

17. a. The substance used

b. To determine the effect of the structure of a molecule on its boiling point.

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c. [The results suggest that substance A has a higher boiling point than substance B.[1]][The fact that two different masses of substance were used (5 g of substance A and 4 g of substance B) would have decreased the validity of the results and may have lead to an incorrect conclusion being drawn.[2]][This is due to the fact that there was more than one variable that was changed during the experiment (multiple independent variables).[3]]

✓ ✗ I have identified the conclusion of the experiment.[1]

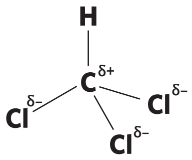
✓ ✗ I have identified the key error affecting the validity of the experiment.[2]

✓ ✗ I have described the effect of multiple uncontrolled variables on the validity of the experiment.[3]

FROM LESSONS 16A & 16D

Questions from multiple lessons

18. a.



[Molecule shape correct.[1]][Dipoles of molecule correctly annotated.[2]]

✓ ✗ I have used lone and bonded pairs of electrons to determine the shape of the molecule.[1]

✓ ✗ I have used the electronegativity of elements to determine the permanent dipoles of the molecule as well as the overall dipole.[2]

b. [Dispersion forces are present in trichloromethane as they are found between all molecules.[1]][Dispersion forces arise from instantaneous dipole-dipole moments which are created from the constantly changing distribution of electrons.[2]][Permanent dipole-dipole forces are also present.[3]][This is because the molecule has an overall permanent dipole moment as shown in part a, which means that there is the electrostatic attraction between the positive and negative ends of $CHCl_3$ molecules.[4]]

✓ ✗ I have identified the presence of dispersion forces.[1]

✓ ✗ I have explained the origin of dispersion forces.[2]

✓ ✗ I have identified the presence of permanent dipole-dipole forces.[3]

✓ ✗ I have explained the origin of permanent dipole-dipole forces.[4]

c. [A substance's state of matter at room temperature depends on the strength of intermolecular forces present.[1]][Chloroform has both dispersion forces (a relatively weak intermolecular force) and permanent dipole-dipole forces (a stronger intermolecular force).[2]][There are sufficiently strong intermolecular forces for the molecules to be held together as a liquid at room temperature but not enough for it to be solid.[3]]

✓ ✗ I have related the molecules state of matter to the strength of intermolecular forces.[1]

✓ ✗ I have described the intermolecular forces operating between the molecules.[2]

✓ ✗ I have related this to the question.[3]

FROM LESSON 2A

2C Macromolecules

Progress questions

1. A. 'Tetra' means four.

2. B. Carbon allotropes all contain carbon atoms bonded in a different way.

3. B. Diamond consists of a tetrahedral carbon lattice.

4. A. Graphite is layers of carbon lattice held together by dispersion forces.

5. B. The density of graphite is lower than diamond as there is space between the layers.

6. B. Graphite is a good conductor because it has delocalised electrons. Diamond however, does not.

7. A. In both instances, the intermolecular forces between water molecules are not strong enough to break carbon-carbon covalent bonds.

8. A. Covalent bonds require a lot of energy to break, which results in a high melting point.

Deconstructed

9. A

10. C

11. [The structure of diamond is a covalent network lattice, where each carbon atom is covalently bonded to four other carbon atoms.[1]] [Due to the strength of these covalent bonds, the structure is rigid, causing it to maintain its structure while making a cut into the wood.[2]][On the other hand, graphite is a covalent layer lattice, where each carbon atom is bonded to three others. The layers are held together by weak dispersion forces.[3]][The weak intermolecular forces between layers of graphite cause them to be rubbed off onto the wood.[4]]

✓ ✗ I have described the structure of diamond.[1]

✓ ✗ I have described the properties of diamond and its effect on the wood.[2]

✓ ✗ I have described the properties of graphite and its effect on the wood.[4]

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✓ ✗ I have described the structure of graphite.^{3}

Exam-style

12. a. [Due to the fact that carbon atoms in graphite bond to only three other carbons, the single unpaired electron from every carbon becomes a delocalised electron.^{1}][The delocalised electrons are able to move through the layers of graphite when electricity is applied.^{2}][Therefore, graphite is able to conduct electricity through the movement of charged particles.^{3}]

✓ ✗ I have described the structure of graphite.^{1}

✓ ✗ I have identified the effect of electricity on the structure of graphite.^{2}

✓ ✗ I have linked my answer to the question.^{3}

b. [Diamonds do not conduct electricity.^{1}][Diamonds are abrasive and may damage the machinery.^{2}]

✓ ✗ I have mentioned diamond's non-conductivity.^{1}

✓ ✗ I have mentioned diamond's abrasiveness.^{2}

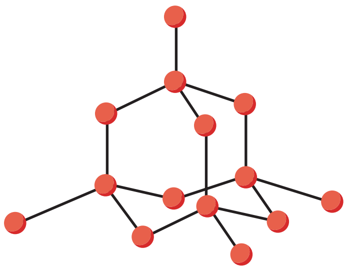
c. [Graphite has an extremely high melting point.^{1}][The thermal energy generated in the electric motor is unlikely to interfere with the functionality of the carbon brushes.^{2}][This is because the thermal energy generated would be insufficient to overcome the strength of the strong network of covalent bonds within the layers of graphite.^{3}]

✓ ✗ I have identified the high melting point of graphite.^{1}

✓ ✗ I have linked my answer to the question.^{2}

✓ ✗ I have described the effect of heat on the structure of graphite.^{3}

13. a.



✓ ✗ I have shown each carbon atom bonded to four other carbon atoms.^{1}

✓ ✗ I have illustrated the structure of diamond as a lattice structure.^{2}

b. [Diamond has a rigid three-dimensional lattice structure.^{1}][This makes it very hard and resistant to wear and thus, an effective abrasive.^{2}]

✓ ✗ I have described the structure of diamond.^{1}

✓ ✗ I have identified the effectiveness of diamond as an abrasive.^{2}

c. [While diamond is hard and rigid, graphite is relatively soft.^{1}][The layer lattices in graphite are held together by weak dispersion forces.^{2}][These intermolecular forces make the overall structure quite weak, causing graphite to rub off when pressed into most surfaces which is why is is commonly used as a industrial lubricant.^{3}][Thus, diamond is a more appropriate material for use as a drilling tool.^{4}]

✓ ✗ I have compared the hardness of graphite to diamond.^{1}

✓ ✗ I have identified the relevant structural property of graphite.^{2}

✓ ✗ I have related this to graphite's hardness.^{3}

✓ ✗ I have linked my answer to the question.^{4}

Key science skills

14. a. It is likely a systematic error.

b. Changing the method or replacing equipment or recalibration of equipment.

c. The results are accurate, however they are imprecise.

FROM LESSON 16D

Questions from multiple lessons

15. a. [A critical element is a vital element with an extremely limited natural supply.^{1}][Diamonds are made entirely of carbon, which is an abundant element on earth.^{2}][Although diamonds are rare, they could not be considered a critical element, and hence the student's proposal is not supported.^{3}]

✓ ✗ I have defined a critical element.^{1}

✓ ✗ I have identified why diamond is not a critical element.^{2}

✓ ✗ I have linked my answer to the question.^{3}

b. Reduce the use of artificial fertilisers or increase the use of natural fertilisers or remove phosphorus from sewage for reuse or use algae to remove the phosphate from polluted waterways.

FROM LESSON 1D

Chapter 2 review

Multiple choice

1. D. The more electrons that are shared in a covalent bond, the stronger the force of attraction between the atoms and hence, the greater the bond strength.

2. C. Molecular formulas tell us which elements are present in a molecule and how many atoms of each element there are.

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3. C. C_2H_2 has a triple covalent bond meaning that both carbon atoms have a linear arrangement of electron pairs.

4. D. Molecules with a permanent dipole are polar, which means they have one or more polar covalent bonds present and are asymmetrical.

5. C. Ethane is the largest molecule and hence, has the strongest dispersion forces.

6. C. Some molecules with polar covalent bonds are non-polar.

7. D. Hydrogen bonds occur when a hydrogen atom is 'sandwiched' between two electronegative atoms (either fluorine, oxygen or nitrogen).

8. B. The strength of intermolecular forces contributes to the boiling point of compounds.

9. B. Diamond consists of a tetrahedral carbon lattice structure.

10. B. Diamond is extremely hard, making it perfect for use in cutting tools. It is however sometimes described as brittle, meaning it does not deform before breaking.

Short answer

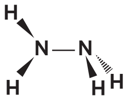
11. a.

✓ ✗ I have drawn the electron dot diagram of nitrogen.[1]

✓ ✗ I have drawn the electron dot diagram of hydrogen.[2]

b.

c. [Each nitrogen atom has one lone pair of electrons, and three pairs of bonding electrons.[1]][As a result of this combination, according to VSEPR, each nitrogen atom will have a trigonal pyramidal shape.[2]][Hence, the overall shape of the molecule is trigonal pyramidal as shown.[3]]



✓ ✗ I have identified the number of lone pairs and bonding electrons of each nitrogen atom.[1]

✓ ✗ I have identified the shape of each nitrogen atom according to VSEPR.[2]

✓ ✗ I have described the three-dimensional shape of hydrazine.[3]

d. [Hydrazine has polar N-H covalent bonds present.[1]] [Since hydrazine is not symmetrical, these permanent dipole moments do not cancel out and therefore, it is a polar molecule.[2]]

✓ ✗ I have identified the polar covalent bonds present in hydrazine.[1]

✓ ✗ I have determined the polarity of the molecule.[2]

e. [The molecule with the greatest boiling point would have the strongest intermolecular forces.[1]][Since C_2H_4 is non-polar, there will be only dispersion forces present whereas N_2H_4 is polar with dispersion forces and hydrogen bonding.[2]] [Hydrogen bonds are stronger than dispersion forces and therefore would require more energy to break, resulting in N_2H_4 having a higher boiling point.[3]]

✓ ✗ I have identified the relationship between boiling point and intermolecular forces.[1]

✓ ✗ I have identified the intermolecular forces present between adjacent molecules in C_2H_4 and N_2H_4.[2]

✓ ✗ I have described the effect of intermolecular forces on the boiling point and have identified N_2H_4 as having the highest boiling point.[3]

12. a.

✓ ✗ I have drawn the Lewis structure of H_2.[1]

✓ ✗ I have drawn the Lewis structure of O_2.[2]

✓ ✗ I have drawn the Lewis structure of N_2.[3]

b. All of these molecules are non-polar.

c. [The more electrons shared in a covalent bond, the stronger the force of attraction between the atoms.[1]][Therefore, H_2 will require the least amount of energy due to its single covalent bond, O_2 will require a greater amount of energy due to its double covalent bond, and N_2 will require the greatest amount of energy due to its triple covalent bond.[2]]

✓ ✗ I have identified the relationship between the number of electrons in a covalent bond and the strength of the covalent bond.[1]

✓ ✗ I have ordered the molecules based on increasing energy required to break the intramolecular bonds.[2]

d. [These molecules are all non-polar.[1]][Therefore, the only intermolecular forces present between these molecules are dispersion forces.[2]]

✓ ✗ I have identified that these molecules are non-polar.[1]

✓ ✗ I have identified the intermolecular forces present.[2]

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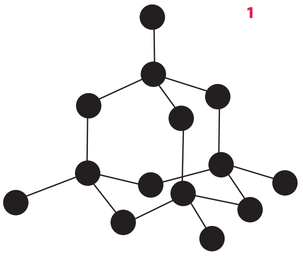
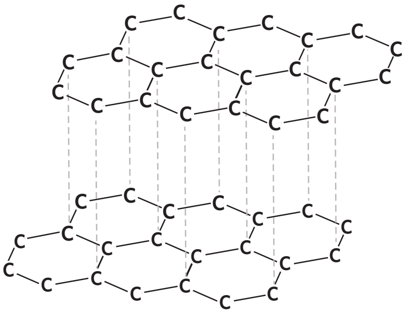
13. a. [1]



[2]



✓ ✗ I have drawn the correct number of bonds for each carbon atom.[1]

✓ ✗ I have shown the three-dimensional structure of the molecule.[2]

b. i. [Graphite is a covalent layer lattice, with layers held together by dispersion forces.[1]][Hence, the layers can easily slide over each other, meaning graphite can reduce friction between two surfaces (lubrication).[2]]

✓ ✗ I have identified the aspect of graphite's structure that makes it appropriate for this use.[1]

✓ ✗ I have linked this structural property to real-world application.[2]

ii. [In graphite, each carbon atom only bonds to three other carbon atoms, leaving one electron delocalised.[1]] [These delocalised electrons allow graphite to conduct electricity when a current is applied, making it suitable for use in a battery-powered brush.[2]]

✓ ✗ I have identified the aspect of graphite's structure that makes it appropriate for this use.[1]

✓ ✗ I have linked this structural property to real-world application.[2]

c. [As shown in the diagrams, diamond forms a tetrahedral lattice while graphite forms a layer lattice, making it a very hard substance.[1]][Furthermore, as each carbon atom is bonded to four other carbon atoms in diamond, leaving no delocalised electrons, diamond cannot conduct electricity.[2]] [The properties of hardness and non-conductivity make diamond inappropriate for use as a lubricant or battery-powered brush.[3]]

✓ ✗ I have identified diamond's hardness with reference to its structure.[1]

✓ ✗ I have identified diamond's non-conductivity with reference to its structure.[2]

✓ ✗ I have linked my answer to the question.[3]

Key science skills

14. a. [In a fumehood, set up a bunsen burner, pour a 10 mL sample of each compound into a test tube with a thermometer and hold the sample over the bunsen burner with a retort stand.[1]] [Once bubbles are observed, record the temperature reached by the sample.[2]][Repeat for the other samples. [3]]

✓ ✗ I have considered the materials required.[1]

✓ ✗ I have identified what data to record.[2]

✓ ✗ I have ensured that this method is a fair test (appropriate variables are controlled) and will produce a complete set of results.[3]

b. [The independent variable is the compound being tested (e.g., $C_3H_8(l)$, $C_3H_7OH(l)$ and $C_3H_7Cl(l)$).[1]][The dependent variable is the boiling point (°C) of the sample.[2]] [Possible controlled variables include the type of thermometer used, the volume of liquid in a sample, and the distance between the test tube and the bunsen burner's flame.[3]]

✓ ✗ I have identified the independent variable.[1]

✓ ✗ I have identified the dependent variable.[2]

✓ ✗ I have identified a possible controlled variable.[3]

c. [The data collected by the thermometer will reveal the boiling points of the different compounds.[1]][The stronger the intermolecular forces between the molecules, the greater the amount of energy that is required to break the bonds and thus, the higher the boiling point.[2]][By comparing the boiling points of each substance, we can determine the relative strengths of dispersion forces (C_3H_8), permanent dipole-dipole forces (C_3H_7Cl) and hydrogen bonds (C_3H_7OH).[3]]

✓ ✗ I have explained that the thermometer has collected the data.[1]

✓ ✗ I have explained the link between boiling point and the strength of intermolecular forces.[2]

✓ ✗ I have linked my answer to the question.[3]

d. Hazard: open flame. Solution: wear protective gloves and use tongs or Hazard: inhalation of volatile organic compounds. Solution: use a fume hood or Hazard: contact of chemicals with skin or eyes. Solution: wear personal protective equipment or Hazard: flammability of compounds. Solution: wear personal protective equipment.

FROM LESSONS 16B, 16D & 16E

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3A Metals

Progress questions

1. A. Metals are extracted from ores, which are the deposits of minerals containing metals.

2. B. The names of electrons indicate whether they can move freely or they are fixed.

3. B. Delocalised electrons are removed from metal atoms and are therefore attracted to metal cations.

4. A. Malleability, ductility, lustre, heat, and electrical conductivity can be explained by the metallic bonding model.

5. A. Electrostatic force of attraction in the metallic bonding model is the non-directional attractive force between the sea of delocalised electrons and metal cations.

6. A. Delocalised electrons are removed from metal atoms and therefore can absorb heat and transfer kinetic energy to other delocalised electrons.

7. B. Delocalised electrons have negative charge, so they are pushed by the negative electrode and attracted by the positive electrode.

8. A. Metallic bonds are strong, so they can withstand high temperatures.

Deconstructed

9. A

10. C

11. [Based on the data given, iron has the highest boiling point of all the metals.[1]][This suggests that the metallic bonds involving iron are the strongest.[2]][Therefore, it would be able to withstand more weight and heat compared to the other metals.[3]][As a result, iron would be the most suitable metal for the company to use.[4]]

✓ ✗ I have identified the metal with the highest melting point.[1]

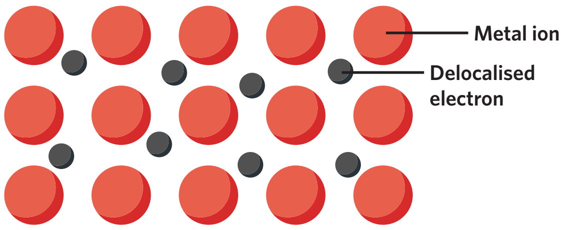
✓ ✗ I have identified the metal with the strongest bonds.[2]

✓ ✗ I have described the effect of bond strength on different properties.[3]

✓ ✗ I have linked my answer to the question.[4]

Exam-style

12. a.



b. [Metals are held together by the force of attraction that exists between the positive metal cation and the sea of negatively charged delocalised electrons.[1]][When a force is applied to the lattice structure, the layers of positive ions can move, however due to the delocalised electrons, the 3-D non-directional force of attraction between the electrons and positive ions remains,[2]][holding the structure together and therefore allowing it to be bent without breaking.[3]]

✓ ✗ I have described metallic bonding.[1]

✓ ✗ I have explained what happens when a force is applied to a metallic structure.[2]

✓ ✗ I have linked my answer to the question.[3]

13. a. Nick can conclude that metals can conduct heat.

b. [According to the metallic bonding model, there are many delocalised electrons between metal cations.[1]][When a piece of metal is heated, the rise in temperature causes the increase in kinetic energy of delocalised electrons and metal cations. Therefore, the metallic ions can vibrate more rapidly and delocalised electrons can move faster, transferring the kinetic energy to nearby ions and electrons.[2]][Hence, when Nick touched the spoon, the kinetic energy carried by the moving electrons was transferred to his hand in the form of heat energy and he therefore had a ‘burning’ feeling.[3]]

✓ ✗ I have described delocalised electrons in the metallic bonding model.[1]

✓ ✗ I have used the metallic bonding to explain how metals can conduct heat.[2]

✓ ✗ I have linked my explanation with Nick’s experience.[3]

14. [The student’s explanation is not correct.[1]][For a material to be able to conduct electricity, it needs to allow for the movement of electrons.[2]][Metals can conduct electricity because according to the metallic bonding model, delocalised electrons in a solid piece of metal can move freely. Therefore, when one end of the metal is connected to the positive electrode and the other end is connected to the negative electrode of an electrical source, there is a movement of delocalised electrons away from the negative electrode and towards the positive electrode.[3]]

✓ ✗ I have determined if the student’s explanation is correct.[1]

✓ ✗ I have described metal conductivity.[2]

✓ ✗ I have explained how metals can conduct electricity.[3]

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15. **a.** [According to the results, the structures of both metals are unaffected at 200 °C. At 400 °C the lead rail began to lose its structure whereas copper remained unaffected.[1]] [This suggests that the energy supplied at 400 °C was enough to break the metallic bonds within the lead rail, implying that the melting temperature of lead is within the range of 400 °C.[2]] [In contrast, the structure of copper's rail remained unchanged, suggesting that the melting point for copper is greater than 600 °C.[3]] [Through these results, it can be seen that lead has a lower melting point than copper.[4]]

✓ ✗ I have identified the key results in the experiment.[1]

✓ ✗ I have explained the effect of heat on the bonding of lead.[2]

✓ ✗ I have explained the effect of heat on the bonding of copper.[3]

✓ ✗ I have linked my answer to the question.[4]

b. [Although the metallic bonding model is able to explain most of the properties of metals, it isn't able to explain why different metals exhibit slightly different variations of certain properties.[1]] [For example, for metals to conduct heat, the kinetic energy is passed on through the movement of electrons.[2]] [However since all metallic structures contain delocalised electrons, we are unable to use the model to explain why different metals are able to conduct heat more or less efficiently.[3]]

✓ ✗ I have identified the main limitation of the metallic bonding model.[1]

✓ ✗ I have described how heat is conducted through metals.[2]

✓ ✗ I have described the inability of the metallic bonding model to explain heat conductivity.[3]

Key science skills

16. **a.** [Gabriella can conclude that her ring is not entirely made of silver.[1]] [There was dull light from the bulb when the ring was implemented in the circuit, while there was strong and dazzling light from the bulb when the piece of pure silver was implemented in the circuit.[2]]

✓ ✗ I have given a conclusion.[1]

✓ ✗ I have used my observation in the experiment to explain my conclusion.[2]

b. Gabriella can use different sets of wires which are made of different metals to confirm that the same results will be obtained.

FROM LESSONS 16C & 16D

Questions from multiple lessons

17. **a.** Titanium

b. [Titanium is a transition metal while plastic is made of non-metal atoms.[1]] [Titanium is more durable than the plastic which could wear down or even break.[2]] [The plastic might react with body fluids whereas titanium is inert (chemically inactive).[3]]

✓ ✗ I have determined the most appropriate material to be used.[1]

✓ ✗ I have compared the strength and potential reactivity of the two materials.[2]

✓ ✗ I have used my explanation to justify my answer.[3]

FROM LESSONS 1A & 2B

3B Reactivity of metals – part 1

Progress questions

1. A. Not all but the majority of metallic elements in the periodic table can react with acid, water or oxygen.

2. A. Metals have low ionisation energies, so electrons are easily taken from them.

3. B. Across the periods, ionisation energies of metals increase, and therefore their reactivity with acid decreases.

4. B. Since a positive metal ion bonds to a negative ion from an acid to form a salt, free hydrogen ions in the solution bond to each other to produce hydrogen gas

5. B. Metals with lower ionisation energies react more quickly with water, while metals with higher ionisation energies react more slowly with water.

6. A. A "pop" sound is heard when hydrogen gas reacts with oxygen in the atmosphere in the presence of a flame.

7. A. A reaction between a metal and water will produce a metal hydroxide.

8. A. When ignited a mixture of hydrogen and oxygen is explosive. Pure hydrogen burns gently.

9. B. Rust is a result of a reaction between iron and oxygen.

10. B. Some metals, like gold and platinum, cannot react with oxygen under normal conditions.

11. A. Zinc can react with oxygen to produce zinc oxides under normal conditions, while silver cannot.

Deconstructed

12. A

13. B

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14. [Due to the humid condition in the laboratory, there was water in the room atmosphere.[1]][Group 1 metals are highly reactive due to their low ionisation energies.[2]][The small explosion was the result of the reaction between group 1 metals with water in the atmosphere at room temperature with hydrogen gas produced.[3]]

- ✓ ✗ I have described the effect of the humidity on the room atmosphere.[1]
- ✓ ✗ I have explained the reactivity of group 1 metals.[2]
- ✓ ✗ I have explained the underlying reason for the explosion observed.[3]

Exam-style

15. C

16. B

17. [Box A contains calcium and box B contains lithium.[1]][Calcium has higher ionisation energy than lithium, meaning it is more difficult for calcium to lose electrons. Therefore, calcium reacts slower with water and hydrogen gas bubbles are produced more slowly. Lithium has lower ionisation energy than calcium, meaning it is easier for lithium to lose electrons. Hence, lithium reacts more vigorously with dilute sulfuric acid and more bubbles are produced at a faster rate.[2]]

- ✓ ✗ I have determined the metal contained in each box.[1]
- ✓ ✗ I have explained the observation of the reaction between the metals and the acid.[2]

18. [The engineer should choose iron.[1]][Barium reacts more strongly with water than iron so using iron will help maintain better quality of agrimotors during the rain season.[2]]

- ✓ ✗ I have determined which metal should be chosen.[1]
- ✓ ✗ I have explained my answer with respect to the reactivity of the metals.[2]

19. a. [The student's method is incorrect.[1]][Both lithium and zinc can react with hydrochloric acid to produce colourless solutions of salts and hydrogen gas.[2]]

- ✓ ✗ I have identified the accuracy of the student's method.[1]
- ✓ ✗ I have described the reaction of the metals with hydrochloric acid.[2]

b. [Lithium and zinc can be distinguished by placing the same amount of each metal into water at room temperature.[1]][Lithium can react rapidly with water at room temperature to produce hydrogen gas, while zinc cannot react with water at room temperature.[2]]

[$2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$
$Zn(s) + H_2O(l) \rightarrow$ no reaction[3]]

- ✓ ✗ I have suggested one method to distinguish the two metals.[1]
- ✓ ✗ I have described the outcome of the method chosen.[2]
- ✓ ✗ I have included the possible balanced chemical equation(s) to support my answer.[3]

Key science skills

20. a. [The remaining solid mixture contains Mg(s) and Cu(s).[1]][This is because calcium will react rapidly with water to produce a hydroxide solution, whereas magnesium and copper will not react rapidly and therefore remain in the mixture.[2]]

$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

- ✓ ✗ I have determined the metals remained in the solid mixture.[1]
- ✓ ✗ I have justified my answer by comparing the ability to react with cold water of the metals.[2]
- ✓ ✗ I have included the occurred chemical reaction.

b. [The student's suggestion is not correct and is very dangerous.[1]][If concentrated sulfuric acid is placed into the mixture first, magnesium, calcium and copper will all react with the acid to produce a solution, meaning that we cannot collect pure solid copper.[2]]

- ✓ ✗ I have evaluated the student's suggestion.[1]
- ✓ ✗ I have explained my answer describing what would occur.[2]

c. [The students separated more than 5.0 g of solid copper because a small amount of either calcium or magnesium did not react completely with the dilute sulfuric acid.[1]][To collect only pure copper, they can increase the temperature of the acid or increase its concentration.[2]][This would ensure that only copper remained with a mass of 5.0 g.[3]]

- ✓ ✗ I have explained the reason the students collected more copper than expected.[1]
- ✓ ✗ I have suggested one change that could be made to collect just copper.[2]
- ✓ ✗ I have described a reason to support my suggestion.[3]

FROM LESSON 16B

Questions from multiple lessons

21. a. P and N

b. Fe

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c. [P.[1]][Recover phosphorus from human and animal waste, or recycle industrial waste, eating less meat and dairy to reduce the number of livestock eating plants.[2]]

✓ ✗ I have identified the critical element.[1]

✓ ✗ I have described a way in which phosphorus supplies can be preserved.[2]

FROM LESSONS 1D & 3A

3C Metal recycling

Progress questions

1. B. A linear economy life cycle ends up in the disposal of resources and therefore the same amount of resources need to be added to create the product again. A circular economy does not require the addition of the same amount of resources to create the product again as it is not disposed of.

2. B. Economies that do not use as many resources are more sustainable.

3. A. Glass can be infinitely recycled, whereas a single use plastic bag is disposed of in landfill.

4. B. Radioactive (U, Pu) or very reactive metals (Na) are not suitable for recycling as they are very dangerous and need to be managed by special recovery agencies.

5. A. It is better for all if iron is used as it has a reduced impact.

6. B. Recycling does not reduce the number of cans that are produced but it does reduce the amount of new metal that needs to be mined and produced from its ore.

Deconstructed

7. C

8. C

9. [Collecting the scrap metal using trucks.[1]][Using a specialised electromagnetic crane to separate out the magnetic metals.[2]] [The melting of the metals to obtain pure iron.[3]]

✓ ✗ I have planned a sequence of events starting with collection.[1]

✓ ✗ I have identified how the iron is separated.[2]

✓ ✗ I have described how the iron is purified ready for resale.[3]

Exam-style

10. D. In most cases, new chemicals that are produced can persist in the environment for very long periods of time.

11. **a.** [Can be produced at a greater rate than it is consumed.[1]] [Consequences of the use of this chemical product are minimal for society and the environment.[2]]

✓ ✗ I have defined sustainable.[1]

✓ ✗ I have identified possible consequences.[2]

b. [Ferrous means that it contains some degree of iron.[1]] [Non-ferrous does not contain any iron.[2]]

✓ ✗ I have understood what ferrous means.[1]

✓ ✗ I have explained the difference between the two terms.[2]

c. [When mixtures of metals melt this separates them.[1]] [This then allows them to be purified, recycled and used again and again.[2]][This is part of a closed cycle which is indicative of a circular economy.[3]]

✓ ✗ I have identified the property.[1]

✓ ✗ I have explained how they can be separated.[2]

✓ ✗ I have identified that recycling is part of a closed cycle.[3]

12. [Fewer fossil fuels will need to be combusted which will reduce greenhouse gas emissions.[1]][Fewer fossil fuels will need to be extracted which will help conserve Earth's natural resources and protect communities, wildlife and the environment.[2]]

✓ ✗ I have mentioned less energy used means less greenhouse gases.[1]

✓ ✗ I have addressed the issue surrounding Earth's natural resources and the environment.[2]

Key science skills

13. **a.** The amount of e-waste produced is forecast to continue increasing (per capita).

b. Small

c. Screens and monitors

d. [Precious metals are extremely difficult to find.[1]][They are expensive to mine.[2]][We only have a finite supply.[3]]

✓ ✗ I have described that they are difficult to find.[1]

✓ ✗ I have described that they are expensive to mine from the ground.[2]

✓ ✗ I have described that not a lot of some metals (e.g. gold) are left in the ground.[3]

FROM LESSONS 16B & 16C

14. [Although recycling may be costly,[1]] [extracting a metal from its ore will be a lot more expensive.[2]] [Recycling metals is much more economical.[3]] [It is also more environmentally friendly than extracting new metals.[4]] [Mining and extraction are still required due to demand.[5]]

✓ ✗ I have addressed the issue that recycling does cost a lot of money.[1]

✓ ✗ I have addressed the issue that extracting metals from the ore is much more expensive.[2]

✓ ✗ I have used economy to compare extraction and recycling.[3]

✓ ✗ I have addressed the issue of landfill and environmental damage.[4]

✓ ✗ I have mentioned that extraction is still necessary to meet demand.[5]

FROM LESSON 16E

Questions from multiple lessons

15. **a.** Oxidation

b. [Hydrogen pop test.[1]] [Involves carefully capturing hydrogen gas and igniting it and a controlled explosion indicates the presence of hydrogen gas.[2]]

✓ ✗ I described the name of the test.[1]

✓ ✗ I described the procedure for carrying out the test.[2]

c. As in nature it is found as a pair of covalently bonded atoms, $F_2(g)$.

d. [Reactive metals like sodium readily give away electrons.[1]] [On the other hand reactive non-metals like fluorine gas readily accept electrons.[2]]

✓ ✗ I understand that reactive metals give away electrons.[1]

✓ ✗ I understand that reactive non-metals accept electrons.[2]

FROM LESSONS 1B, 2A & 3B

Chapter 3 review

Multiple choice

1. D. The reactivity of metals depends on their ionisation energy.

2. D. The reactivity of metals increases going down a group.

3. B. Rust forms when iron is exposed to oxygen and water over a certain period of time.

4. C. A high melting point is an indicator of the strength of the metallic bonds between metal atoms.

5. B. It is important that the complete life cycle of the chemical product is mapped out.

6. A. Metals are able to conduct electricity due to the movement of delocalised electrons.

7. B. When metals and an acid react, hydrogen gas is produced.

8. A. Linear is an open cycle where products are disposed of.

9. C. Reactivity depends on the ionisation energy of a metal.

10. C. The delocalised electrons in metals are responsible for many of their unique properties including lustre and heat conductivity.

Short answer

11. **a.** [Copper → Used in electrical wiring due to high conductance.[1]]
[Gold → Found in the Earth as a pure metal.[2]]
[Aluminium → Resistant to corrosion due to a strong oxide layer.[3]]
[Iron → Most recycled material on Earth.[4]]
[Sodium → Reacts rapidly with water producing hydrogen gas.[5]]

✓ ✗ I have identified the electrical properties of copper.[1]

✓ ✗ I have identified the unreactive nature of gold.[2]

✓ ✗ I have identified that aluminium is a reactive metal but its surface is protected.[3]

✓ ✗ I have identified that iron is the most commonly used metal on Earth.[4]

✓ ✗ I have identified that sodium is very reactive.[5]

b. **i.** [Finite resources of bauxite.[1]] [Extracted at 950 °C – large amount of energy required.[2]] [More expensive.[3]]
[Conserves resources.[4]] [Cheaper to recycle.[5]]
[Less electricity needs to be used than original extraction of aluminium from bauxite.[6]]

✓ ✗ I have referred to one reason based on the extraction processes.[1]

✓ ✗ I have referred to a second reason based on the extraction processes.[2]

✓ ✗ I have referred to a third reason based on the extraction processes.[3]

✓ ✗ I have referred to one reason based on the recycling processes.[4]

✓ ✗ I have referred to a second reason based on the recycling processes.[5]

✓ ✗ I have referred to a third reason based on the recycling processes.[6]

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ii. [Cryolite reduces the melting point of impure aluminium oxide by around 1300 °C which saves a lot of energy.[1]] [Melting scrap aluminium requires around 250 °C less to melt it which saves energy.[2]]

✓ ✗ I have quoted the temperature values when cryolite is used and its effect on energy use.[1]

✓ ✗ I have quoted the temperature values when scrap aluminium is used and its effect on energy use.[2]

iii. [Renewable resources (e.g. solar energy) to provide the electrical energy supply.[1]] [Increase the amount of aluminium that is recycled by doing a regional or national campaign.[2]] [Use renewable fuels in the machinery that is used to extract and recycle the aluminium.[3]]

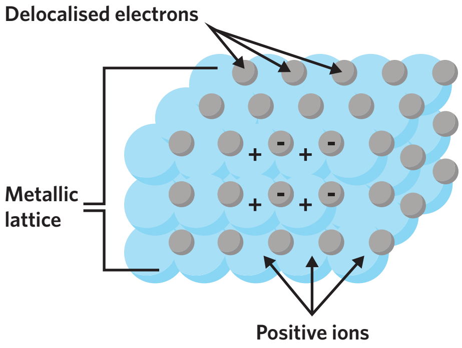
✓ ✗ I have described how energy can be supplied by alternative renewable fuels.[1]

✓ ✗ I have suggested how more aluminium could be recycled.[2]

✓ ✗ I have described how machinery could be supplied by alternative renewable fuels.[3]

12. a. An alloy is a mixture consisting of a metal physically combined with another element which may be another metal or even a non-metal (normally carbon).

b. [Metallic bonding consists of metal cations in a lattice structure surrounded by a 'sea' of delocalised electrons as shown in the diagram.[1]] [In order to conduct electricity, there needs to be the flow of charged particles within the substance.[2]] [In the case of metals, they can conduct electricity because of the 'sea' of delocalised electrons, which are able to flow throughout the metallic lattice, carrying charge and, therefore, electricity.[3]]



✓ ✗ I have described the model of metallic bonding.[1]

✓ ✗ I have described the requirements for a substance to conduct electricity.[2]

✓ ✗ I have justified why metals are able to conduct electricity.[3]

✓ ✗ I have included a diagram.

c. The wires of a pacemaker need to be malleable to be inserted through the irregularly shaped veins and arteries of the body.

13. a. [Barium would have the greatest reactivity.[1]] [Although all three elements are found in the same group, barium contains more occupied energy shells than beryllium and magnesium. As a result, the valence electrons are further away from the nucleus of the barium atom and are therefore held less strongly to the atom. Subsequently, valence electrons of barium are lost more easily, making it highly reactive.[2]]

✓ ✗ I have identified the element with the greatest reactivity.[1]

✓ ✗ I have explained the difference in reactivity based on the number of energy shells of the atom.[2]

b. [$Ba(s) + 2H_2O(l) \rightarrow Ba(OH)_2(aq) + H_2(g)$.[1]] [If a flame were placed inside the test tube, a 'pop' would be heard due to the ignition of hydrogen gas.[2]]

✓ ✗ I have included a balanced equation.[1]

✓ ✗ I have identified the effect of placing a flame close to the reaction.[2]

c. [Over time, metals exposed to the atmosphere will start to react with oxygen and start to form metal oxides and, therefore, the metals will start to corrode.[1]] [This could be used to determine the reactivity of metals by comparing the times at which they react and form metal oxides.[2]] [Barium is predicted to react first with oxygen according to the equation: $2Ba(s) + O_2(g) \rightarrow 2BaO(s)$.[3]] [Magnesium will be predicted to react next with oxygen according to the equation: $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$.[4]] [Finally, beryllium will be the last to react with oxygen according to the equation: $2Be(s) + O_2(g) \rightarrow 2BeO(s)$.[5]]

✓ ✗ I have identified what will happen when metals are left outside.[1]

✓ ✗ I have explained how this could be used to determine the reactivity of metals.[2]

✓ ✗ I have given the reaction between barium and oxygen.[3]

✓ ✗ I have given the reaction between magnesium and oxygen.[4]

✓ ✗ I have given the reaction between beryllium and oxygen.[5]

Key science skills

14. a. [Validity means whether the experimental technique can support the aim or, in this case, whether practising with sodium is helpful before handling caesium.[1]] [This is a valid technique because sodium is another highly reactive metal in group 1, but it is less reactive than caesium and therefore is safer to use.[2]]

✓ ✗ I have defined validity.[1]

✓ ✗ I have identified whether this technique is valid.[2]

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b. [Accuracy relates to how close a value is to the true value, whereas precision relates to how close results are to each other.[1]][In this case, atomic clocks could be accurate but they don't have to be precise because different clocks could be accurately timekeeping but not showing the exact same reading (i.e. they may be in different locations).[2]]

✓ ✗ I have defined accuracy and precision.[1]

✓ ✗ I have explained why accuracy does not necessarily result in precision.[2]

15. a. Atmospheric pressure in the lab.

b. [The different samples of metals.[1]][The relative amount of hydrogen gas bubbles produced.[2]]

✓ ✗ I have identified the independent variable.[1]

✓ ✗ I have identified the dependent variable.[2]

c. [Surface area is not the same as mass and this will affect the rate at which bubbles are produced.[1]][It is very difficult to control or measure the surface area of each different sample.[2]]

✓ ✗ I have identified another variable.[1]

✓ ✗ I have explained why it cannot be controlled.[2]

d. Measure the volume of gas produced quantitatively rather than qualitatively using some sort of gas syringe to measure the volume of gas in mL.

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4A Ionic bonding

Progress questions

1. A. Bohr models can show how ionic bonds are formed and the ratio of cations and anions but cannot demonstrate how ions are arranged in space.

2. B. In a sample of sodium chloride, each metal ion is bonded to adjacent negative non-metal ions if they are close enough.

3. B. Ionic bonds can be formed between a single cation and multiple neighbouring anions and vice versa, which forms a 3D structure of a sample of an ionic compound.

4. B. The bond between each sodium ion and its 6 neighbouring chloride atoms are all ionic bonds.

5. A. The ionic bonds between cations and anions determine ionic compounds' properties.

6. B. Ionic bonds between cations and anions are strong and require a lot of energy to disrupt.

7. A. Ionic bonds between cations and anions are strong but when a strong force is applied, like-charged ions are pushed close to each other and the repulsive forces between them can break the ionic lattice.

8. B. Ionic bonds between cations and anions are strong so it requires a large amount of heat to disrupt them.

9. B. Ionic compounds can conduct electricity only when cations and anions move freely.

10. B. In aqueous state, cations and anions can move freely within the solution.

Deconstructed

11. B

12. C

13. [The campers can dissolve the table salt in water and then connect the salt solution to the wires.^{1}][Table salt is primarily made of sodium chloride which is an ionic compound. Therefore, it cannot conduct electricity in solid state due to the fact that all cations and anions are held in fixed positions via ionic bonds.^{2}][However, when sodium chloride is dissolved in water, sodium cations and chloride anions can move freely within the solution and therefore allow for the electric charge to be carried through the circuit. As a result, sodium chloride in aqueous state can conduct electricity and hence allow the light bulb to shine.^{3}]

✓ ✗ I have suggested one thing the students can do.^{1}

✓ ✗ I have explained my suggestion by explaining the compound's inability to conduct electricity in solid state.^{2}

✓ ✗ I have explained my suggestion by explaining the compound's ability to conduct electricity in aqueous state.^{3}

Exam-style

14. [The student's statement is incorrect.^{1}][The bond between a potassium ion and the oxide ion it donates its electrons to and the bonds between that potassium ion and the other 3 neighbouring oxide ions are all ionic bonds. This is because ionic bonds can be formed between a single cation and multiple anions if they are close enough to each other.^{2}]

✓ ✗ I have determined the accuracy of the student's statement.^{1}

✓ ✗ I have justified my answer based on the condition of the formation of ionic bonds.^{2}

15. [Limestone is made of $CaCO_3$ which is an ionic compound.^{1}] [It was not broken in the first three attempts because ionic bonds between Ca^{2+} ions and CO_3^{2-} ions are strong, meaning that the initial force applied was not strong enough to break the ionic bonds.^{2}][However, when more force was applied, the lattice became disrupted in a way that Ca^{2+} ions were pushed close to one other, and CO_3^{2-} ions were pushed close to one another. The repulsive forces between the like-charged ions disrupted the lattice structure of $CaCO_3$, causing the limestone to break.^{3}]

✓ ✗ I have identified the class of compound to which limestone belongs.^{1}

✓ ✗ I have described the strength of ionic bonding.^{2}

✓ ✗ I have explained the effect of the disruption of the ionic lattice on the brittleness of limestone.^{3}

16. [The student's explanation is incorrect.^{1}][BaO has a high melting point due to strong ionic bonds between Ba^{2+} cations and O^{2-} anions, not as a result of the repulsive forces between similarly charged ions.^{2}]

✓ ✗ I have identified the accuracy of the student's explanation.^{1}

✓ ✗ I have explained my answer with reference to the nature of ionic bonds.^{2}

17. [Tap water contains freely moving ions.^{1}][When a hair dryer is dropped into a bathtub filled with tap water, the ions found in the water can move freely and thereby transfer the electric charge from the hair dryer to the human body. This can result in an electrocution.^{2}]

✓ ✗ I have identified the properties of tap water based on given data.^{1}

✓ ✗ I have explained how water can cause electrocution based on the conduction of charge.^{2}

Key science skills

18. **a.** [The independent variable is the concentration of dissolved NaCl.^{1}][The dependent variable is the current recorded by the ammeter.^{2}]

✓ ✗ I have correctly identified the independent variable.^{1}

✓ ✗ I have correctly identified the dependent variable.^{2}

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b. [Solid NaCl cannot conduct electricity.[1]][NaCl solutions in both beaker I and beaker II can conduct electricity. However, the NaCl solution in beaker II has a higher electrical conductivity than the NaCl solution in beaker I.[2]]

✓ ✗ I have identified the electrical conductivity of an ionic compound in solid state.[1]

✓ ✗ I have identified the electrical conductivities of an ionic compound in aqueous state.[2]

c. [NaCl solutions in both beaker I and beaker II can conduct electricity as both solutions contain Na$^+$ cations and Cl$^-$ anions that are able to move freely in the solution.[1]][The NaCl solution in beaker II has a higher electrical conductivity than the NaCl solution in beaker I because more solid NaCl was dissolved into water in beaker II, meaning that beaker II has a greater concentration of Na$^+$ cations and Cl$^-$ anions available to carry the electric current.[2]]

✓ ✗ I have explained my conclusion using the movement of ions in aqueous solutions of an ionic compound.[1]

✓ ✗ I have explained my conclusion by comparing the concentration of moving ions in each case.[2]

FROM LESSONS 16A & 16C

Questions from multiple lessons

19. [Sodium is a metallic substance and therefore contains delocalised electrons in its structure.[1]][The delocalised electrons can move freely within the sodium lattice structure, resulting in the movement of electrons when the lattice is connected to an electrical circuit.[2]][In contrast, in the sodium chloride lattice structure, ions are fixed in position and therefore are unable to move to carry electric charge.[3]][Therefore, a sodium lattice has a high value of electrical conductivity whilst a sodium chloride lattice cannot conduct electricity.[4]]

✓ ✗ I have identified the structure of a piece of sodium metal.[1]

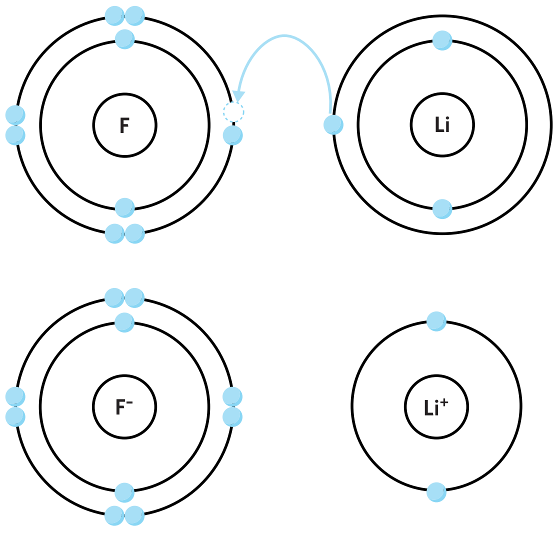
✓ ✗ I have described the ability of sodium to conduct electricity.[2]

✓ ✗ I have compared the structure of sodium to that of sodium chloride and its ability to conduct electricity.[3]

✓ ✗ I have linked the difference in structure to the difference in electrical conductivity.[4]

FROM LESSON 3A

20. a.



(1 MARK) (1 MARK)

b. [Lithium fluoride has a high melting point due to the fact that the ionic bonds between Li$^+$ cations and F$^-$ anions are strong.[1]][Therefore, a large amount of heat energy is required to disrupt the lattice to melt lithium fluoride.[2]]

✓ ✗ I have described the ionic bonds between lithium and fluoride ions.[1]

✓ ✗ I have identified the effect of the bonds between lithium and fluoride ions on the melting point.[2]

FROM LESSONS 1B & 4A

4B Ionic compounds and equations

Progress questions

1. A. The cation is placed before the anion in the formula of an ionic compound.
2. B. Ionic compounds can only form from a cation-anion pair.
3. B. The sulfate ion is an anion, and therefore has a negative charge.
4. A. NH$_4^+$ is a polyatomic cation.
5. A. Brackets are not required around monatomic ions.
6. A. Oxide ions have a charge of −2, so 2 vanadium (III) ions with a charge of +3 are needed to balance the ionic formula.
7. B. When naming an ionic compound, the cation is written before the anion.
8. B. NO$_3^-$ is the polyatomic ion and is named nitrate.
9. A. The opposite charges between positive cations and negative anions create strong forces of attraction.
10. B. The magnitude of an ion's charge is written as a superscript after the chemical symbol.
11. B. Charged particles can form ionic bonds with multiple particles of opposite charge.

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12. B. Balanced chemical equations have the same number of atoms of each element on either side of the equation.

13. A. The charge of transition metal ions is given in brackets as roman numerals.

Deconstructed

14. A

15. D

16. [$5VO + 2VCl_5 \rightarrow$^{1}][$V_2O_5 + 5VCl_2$^{2}]

✓ ✗ I have accurately identified the ionic formulae of the reactants.^{1}

✓ ✗ I have accurately identified the ionic formulae of the products.^{2}

Exam-style

17. **a.** Ag^+ **b.** Copper (II)

c. OH^- **d.** Fluoride

e. NO_3^- **f.** Ammonium

g. SO_4^{2-} **h.** Iron (III)

i. PO_4^{3-}

18. **a.** Name: Sodium nitrate
Formula: $NaNO_3$

b. Name: Silver sulfate
Formula: Ag_2SO_4

c. Name: Ammonium fluoride
Formula: NH_4F

d. Name: Strontium hydrogen carbonate
Formula: $Sr(HCO_3)_2$

e. Name: Potassium oxide
Formula: K_2O

f. Name: Calcium carbonate
Formula: $CaCO_3$

g. Name: Aluminium hydroxide
Formula: $Al(OH)_3$

h. Name: Iron (III) sulfide
Formula: Fe_2S_3

i. Name: Zinc phosphate
Formula: $Zn_3(PO_4)_2$

19. **a.** $Cu_2SO_4 + NiI_2 \rightarrow 2CuI + NiSO_4$ (2 MARKS)

b. $K_2O + FeSO_4 \rightarrow K_2SO_4 + FeO$ (2 MARKS)

c. $Mg_3N_2 + 6NaHCO_3 \rightarrow 3Mg(HCO_3)_2 + 2Na_3N$ (2 MARKS)

d. $Ba_3(PO_4)_2 + 6NH_4F \rightarrow 2(NH_4)_3PO_4 + 3BaF_2$ (2 MARKS)

Key science skills

20. **a.** Quantitative

b. [Independent variable: identity of element (X, Y or Z).^{1}]
[Dependent variable: charge of element.^{2}]

✓ ✗ I have identified the independent variable.^{1}

✓ ✗ I have identified the dependent variable.^{2}

c. [This is a systematic error as it affects all results to the same degree.^{1}][It could be rectified by using an object of known exact mass to determine the difference between true mass and mass recorded by the balance, then adding or subtracting this amount to all the data collected.^{2}]

✓ ✗ I have identified the type of error.^{1}

✓ ✗ I have suggested a way to correct this error.^{2}

FROM LESSONS 16A, 16B & 16D

Questions from multiple lessons

21. **a.** [Ca^{2+}, calcium ion.^{1}][PO_4^{3-}, phosphate ion.^{2}]
[F^-, fluoride ion.^{3}]

✓ ✗ I have identified that calcium ions are present in fluorapatite.^{1}

✓ ✗ I have identified that phosphate ions are present in fluorapatite.^{2}

✓ ✗ I have identified that fluoride ions are present in fluorapatite.^{3}

b. [The strong electrostatic forces of attraction between ions in an ionic compound take a significant amount of energy to disrupt, making ionic compounds hard.^{1}][Fluorapatite is a hard ionic compound, making it suitable to function as tooth enamel to protect the sensitive nerve structures in a tooth.^{2}]

✓ ✗ I have explained the hardness of ionic compounds.^{1}

✓ ✗ I have linked my answer to the suitability of fluorapatite as tooth enamel.^{2}

FROM LESSON 4A

4C Precipitation and ionic equations

Progress questions

1. A. Precipitates are insoluble in the solutions they form in.

2. A. Precipitation reactions occur between ionic compounds.

3. B. Magnesium hydroxide is insoluble in water.

4. A. AgCl is insoluble in water, creating a precipitate to be removed.

5. A. A solid is formed from a precipitation reaction.

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6. A. The full equation includes all ions present in the solution.

7. B. The ionic equation does not contain spectator ions.

8. B. The spectator ions do not change states during the reaction.

Deconstructed

9. C

10. A

11. Full equation:
$Co(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Co(OH)_2(s) + 2NaNO_3(aq)$
(1 MARK)

Ionic equation: $Co^{2+}(aq) + 2OH^-(aq) \rightarrow Co(OH)_2(s)$ (1 MARK)

Spectator ions: $Na^+(aq)$, $NO_3^-(aq)$ (1 MARK)

Exam-style

12. [$HgCl_2$ is insoluble.[1]][Na_2CO_3 is soluble.[2]][K_3PO_4 is soluble.[3]][$Ca(OH)_2$ is soluble.[4]]

✓ ✗ I have identified the solubility of $HgCl_2$.[1]

✓ ✗ I have identified the solubility of Na_2CO_3.[2]

✓ ✗ I have identified the solubility of K_3PO_4.[3]

✓ ✗ I have identified the solubility of $Ca(OH)_2$.[4]

13. **a.** [The chemist would observe the test tube becoming cloudy,[1]][or the formation of a sediment at the bottom of a test tube, as a solid precipitate is formed.[2]]

✓ ✗ I have identified cloudiness as evidence of precipitation formation.[1]

✓ ✗ I have identified sediment formation as evidence of precipitation formation.[2]

b. [The test tube becoming cloudy is evidence of an insoluble copper precipitate forming.[1]][According to solubility rules, of the three compounds that could be formed – $CuCl_2$, $Cu(NO_3)_2$ and $Cu(OH)_2$ – only $Cu(OH)_2$ is insoluble.[2]][Therefore, $Cu(OH)_2$ is the precipitate responsible for the cloudy solution, and solution B contains OH^- ions.[3]]

✓ ✗ I have explained the chemist's observation.[1]

✓ ✗ I have determined which of the possible compounds is insoluble.[2]

✓ ✗ I have linked my answer to the question.[3]

c. $Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$

d. [The formation of a sediment is evidence of an insoluble silver precipitate forming.[1]][According to solubility rules, of the two compounds that could be formed – AgCl and $AgNO_3$ – only AgCl is insoluble.[2]][Therefore, AgCl is the precipitate responsible for the sediment formed, and solution A contains Cl^- ions.[3]][By process of elimination, solution C must contain NO_3^- ions.[4]]

✓ ✗ I have explained the chemist's observation.[1]

✓ ✗ I have determined which of the possible compounds is insoluble.[2]

✓ ✗ I have determined the ions in solution A.[3]

✓ ✗ I have determined the ions in solution C.[4]

e. $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

14. **a.** $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$

b. [A higher concentration of Ca^{2+} ions in the water means there is more opportunity for $CaCO_3(s)$ to form.[1]][Since $CaCO_3(s)$ is insoluble, it will remain in the solid state and block the pipes more frequently.[2]]

✓ ✗ I have explained the effect of higher Ca^{2+} concentration.[1]

✓ ✗ I have linked my answer to the question.[2]

Key science skills

15. **a.** Qualitative

b. [According to the table, precipitates formed from both cobalt and manganese would appear pink.[1]][Since this error is embedded into the experimental design, this is a systematic error.[2]]

✓ ✗ I have described the similarity between the metals.[1]

✓ ✗ I have identified the type of error present.[2]

c. [Accuracy is a measure of how close values are to the true value.[1]][Since concentration based on one scientist's observation of cloudiness is subjective, the results are not likely to be very accurate.[2]]

✓ ✗ I have outlined the meaning of precision.[1]

✓ ✗ I have commented on the effect of the experimental design on precision.[2]

FROM LESSONS 16B & 16D

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Questions from multiple lessons

16. **a.** $MgSO_4(aq) + BaCl_2(aq) \rightarrow MgCl_2(aq) + BaSO_4(s)$

b. [According to the solubility table, magnesium sulfate compounds are soluble in water.[1]] [Thus, epsom salts, which are composed of magnesium sulfate, would be soluble in water.[2]]

✓ ✗ I have determined the solubility of magnesium sulfate.[1]

✓ ✗ I have linked my answer to the question.[2]

c. Group 2

FROM LESSONS 1B & 4B

Chapter 4 review

Multiple choice

1. B. An ion can form ionic bonds with multiple oppositely charged ions.

2. D. Ionic compounds are made up of cations and anions. Cations are usually formed from metals, and anions are usually formed from non-metals.

3. C. Electrical conductivity requires the movement of charged particles. Ions in a molten salt or in solution have sufficient mobility to conduct electricity, whilst ions in a solid ionic compound are fixed.

4. B. Ionic bonds are strong, making ionic compounds hard. However, a force strong enough to push like-charged ions next to each other can shatter the lattice, making ionic compounds brittle.

5. A. The high melting point of lithium fluoride enables it to be used as a component in engines.

6. C. Magnesium is the cation and is, therefore, written first, and magnesium ions have a charge of +2 while fluoride has a charge of −1, so two fluoride ions are needed to balance the charge of the compound.

7. D. Iron (III) has a charge of +3, and hydroxide ions are polyatomic ions with an overall charge of −1, so three hydroxide ions are required to balance the charge of the compound.

8. A. Ions with like charges will repel each other, while ions with opposite charges will attract each other.

9. C. For an equation to be balanced, there must be the same number of each element on each side of the equation.

10. B. Precipitation reactions involve the formation of a solid from two aqueous reactants.

Short answer

11. **a.** [$Ba_3(PO_4)_2$ [1]] [Barium phosphate [2]]

✓ ✗ I have identified the ionic formula of the compound formed from the two ions.[1]

✓ ✗ I have identified the name of the compound formed from the two ions.[2]

b. **i.** $CaBr_2 + Cu_2SO_4 \rightarrow CaSO_4 + 2CuBr$

ii. $Ag_2CO_3 + 2KF \rightarrow 2AgF + K_2CO_3$

iii. $Cu_3N_2 + 3Na_2S \rightarrow 3CuS + 2Na_3N$

c. [Calcium sulfate and copper (I) bromide.[1]] [Silver fluoride and potassium carbonate.[2]] [Copper (II) sulfide and sodium nitride.[3]]

✓ ✗ I have identified the names of the products formed from the reaction between calcium bromide and copper (I) sulfate.[1]

✓ ✗ I have identified the names of the products formed from the reaction between silver carbonate and potassium fluoride.[2]

✓ ✗ I have identified the names of the products formed from the reaction between copper (II) nitride and sodium sulfide.[3]

12. **a.** **i.** +1

ii. Metal

b. JX

c. Yes, all nitrate salts are soluble in water.

13. **a.** [For a medium to be electrically conductive, there mustbe adequate movement of charged particles.[1]] [In the solid state, the ions in magnesium chloride are in fixed positions, and therefore cannot conduct electricity.[2]] [In molten or dissolved states, the ions in magnesium chloride have sufficient mobility to conduct electricity.[3]]

✓ ✗ I have identified the conditions necessary for electrical conductivity.[1]

✓ ✗ I have described the conductivity of magnesium chloride in the solid state.[2]

✓ ✗ I have described the conductivity of magnesium chloride in the molten and dissolved states.[3]

b. $AgCl(s)$

c. [Spectator ions: $Mg^{2+}(aq)$ and $NO_3^-(aq)$.[1]]
[Ionic equation: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$ [2]]

✓ ✗ I have identified the spectator ions in the given reaction.[1]

✓ ✗ I have determined the ionic equation for the given reaction.[2]

d. [The solution may begin to turn cloudy.[1]] [A sediment of the precipitate may begin to form at the bottom of the beaker.[2]]

✓ ✗ I have identified one observation of precipitate formation.[1]

✓ ✗ I have identified another observation of precipitate formation.[2]

Back to contents

14. **a.** NaCl

b. [The scientist’s hypothesis is likely incorrect.[1]] [Ionic compounds have high melting points because of the high strength of ionic bonds.[2]] [Therefore, exposing food to moderate heat would be insufficient to break the compound down, and it would, therefore, contain the same amount of sodium and have the same health effects.[3]]

- ✓ ✗ I have stated whether the scientist’s hypothesis is correct.[1]
- ✓ ✗ I have explained that strong ionic bonds give ionic compounds high melting points.[2]
- ✓ ✗ I have linked my answer to the question.[3]

c. [Like-charged particles repel one another.[1]] [When sufficient force is applied to the block of sodium chloride, the sodium ions would be pushed next to each other, and the chloride ions would be pushed next to each other.[2]] [The strength of the repulsive forces between the two sodium ions, or between two chloride ions, would cause the block to shatter.[3]]

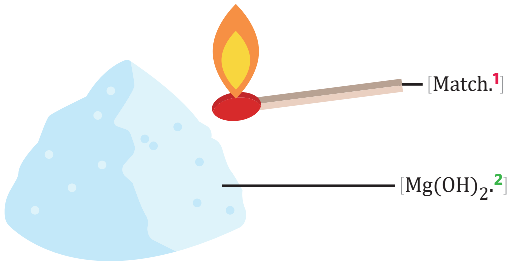
- ✓ ✗ I have described the force between like-charged particles.[1]
- ✓ ✗ I have identified the effect of a force on the structure of the compound.[2]
- ✓ ✗ I have linked my answer to the question.[3]

d. [High melting point. [1]] [Hardness. [2]] [Electrical conductivity in dissolved form. [3]]

- ✓ ✗ I have identified the relevant property of magnesium oxide.[1]
- ✓ ✗ I have identified the relevant property of calcium phosphate.[2]
- ✓ ✗ I have identified the relevant property of ammonium chloride.[3]

Key science skills

15. **a.**



- ✓ ✗ I have drawn the key elements of the experimental setup.[1]
- ✓ ✗ I have accurately labelled the key elements of the experimental setup.[2]

b. [The independent variable is the identity of the compound.[1]] [The dependent variable is the interaction observed in the presence of a flame.[2]]

- ✓ ✗ I have identified the independent variable.[1]
- ✓ ✗ I have identified the dependent variable.[2]

c. [Temperature of the flame or distance of sample from the flame.[1]] [Amount of time the sample is exposed to the flame or amount of sample.[2]]

- ✓ ✗ I have identified a possible variable.[1]
- ✓ ✗ I have identified a second possible variable.[2]

d. **i.** [This is not a repeatable experimental method.[1]] [A repeatable experiment is one that yields similar results when carried out under the same conditions.[2]] [As each student obtained different results despite carrying out the experiment under the same conditions, this method is not repeatable.[3]]

- ✓ ✗ I have correctly identified whether the experiment is repeatable.[1]
- ✓ ✗ I have defined repeatability.[2]
- ✓ ✗ I have linked my answer to the question.[3]

ii. Systematic error

FROM LESSONS 16B, 16D & 16E

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5A Investigating solubility

Progress questions

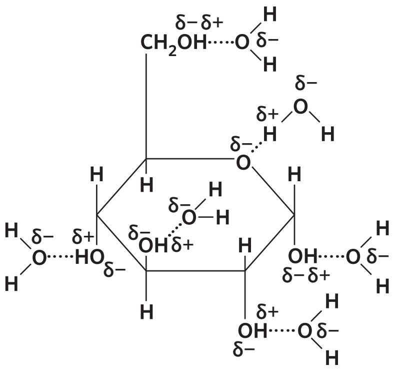
1. A. All solutions are homogeneous mixtures where solute particles are distributed evenly.
2. B. Substances dissolve when the solute–solute and solvent–solvent attractive forces are overcome by the solute–solvent attractive forces.
3. B. Hydrogen bonds are able to form between water molecules and the polar hydroxyl group on CH_3OH.
4. B. Ionisation produces charged particles which can be held in solution by polar water molecules.
5. A. Substances which are largely non–polar will be unable to dissolve in water.
6. B. When a sample is run through chromatography, components are continuously adsorbing to and desorbing from the stationary phase.
7. B. Components will adhere more strongly to the phase that exhibits similar polarity, resulting in less movement from the origin.
8. B. The R_f values obtained from chromatography trials run at different temperatures will differ.
9. A. A chromatogram shows the retention times of different components in a sample as well as their absorbance peaks.
10. A. The mobile phase is the liquid that passes through the HPLC column packed with a solid stationary phase.
11. B. The absorbance of a component reflects the concentration at which it is present in the sample.

Deconstructed

12. B

13. C

14. **a.**



b. [New hydrogen bonds form between glucose and water molecules.[1]][The hydrogen bonds formed are stronger than the solute–solute and solvent–solvent forces, and glucose is therefore able to dissolve.[2]][This process is represented by:

$C_6H_{12}O_6(s) \xrightarrow{H_2O(l)} C_6H_{12}O_6(aq)$.[3]]

✓ ✗ I have identified the formation of hydrogen bonds.[1]

✓ ✗ I have explained how solute–solvent hydrogen bonds are stronger than the existing solute–solute and solvent–solvent forces.[2]

✓ ✗ I have given the appropriate equation to support my answer.[3]

Exam-style

15. D

16. C

17. A

18. **a.** [HPLC is based upon the repeated adsorption and desorption of components between the stationary and mobile phase.[1]][Components which are more strongly adsorbed to the stationary phase will spend more time stationary compared to those which are more strongly desorbed into the mobile phase.[2]][Therefore, the components experience different rates of motion through the column, leading to varying retention times.[3]]

✓ ✗ I have identified the basic principle of HPLC.[1]

✓ ✗ I have explained how components have different retention times.[2]

✓ ✗ I have linked this idea to separation in a HPLC column.[3]

b. [The scientific researcher would have to run a set of known concentrations of the specialised immune cell in the HPLC column under the same conditions.[1]][A calibration curve can then be constructed to relate the area under the peak to the concentration of the immune cell in the patient's blood.[2]]

✓ ✗ I have outlined the process of creating a calibration curve.[1]

✓ ✗ I have explained how a calibration curve can be used to determine the concentration of the immune cell in the patient's blood.[2]

c. [A new calibration curve will have to be generated.[1]][This is because a calibration curve is only accurate if it is generated with conditions identical to when the experiment with the sample of unknown concentration is carried out.[2]]

✓ ✗ I have stated whether a new calibration has to be generated.[1]

✓ ✗ I have explained the factors affecting the accuracy of a calibration curve.[2]

19. **a.** 4

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b. [When a polar stationary phase is used, the component with the longest retention time will be the most polar as it is most strongly adsorbed to the stationary phase.[1]][Therefore, P will be the most polar component.[2]]

✓ ✗ I have linked the length of retention time to the level of polarity of a component.[1]

✓ ✗ I have identified the most polar component.[2]

c. The retention times of S, P, Q, and R could be matched with known compounds with the same retention times when the HPLC is carried out under identical conditions.

d. [The coroner would have to run a set of known concentrations of ethanol in the HPLC column under the same conditions.[1]] [A calibration curve can then be constructed to relate the area under the peak to the concentration of alcohol in the patient's blood.[2]]

✓ ✗ I have outlined the process of creating a calibration curve.[1]

✓ ✗ I have explained the use of a calibration curve to determine the concentration of alcohol.[2]

Key science skills

20. a. Qualitative

b. Identity of solid compound

c. [Validity relates to whether an experimental design accurately tests a hypothesis.[1]][In this case, the experiment is invalid because there is more than one independent variable (both the presence/absence of stirring and the identity of the compound), which means that Margaret cannot compare the solubility differences between compounds.[2]]

✓ ✗ I have defined validity.[1]

✓ ✗ I have explained why the experiment is invalid.[2]

d. Either stir both beakers to the same extent or stir neither beaker to ensure that it remains a controlled variable.

FROM LESSONS 16A, 16B & 16C

Questions from multiple lessons

21. a. i. Each chemical produces a distinct peak on the chromatogram, meaning the identity of each chemical can still be determined.

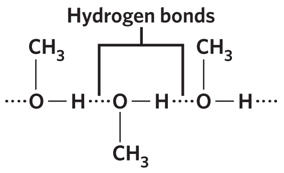
ii. [The analysis is to determine the amount of ethanol in the blood, which will require calculation of only the corresponding peak's area.[1]][0.9 minutes is the retention time of ethanol, the alcohol in blood that is analysed.[2]]

✓ ✗ I have identified that only the peak produced at the retention time of ethanol will be analysed.[1]

✓ ✗ I have identified that the retention time of ethanol is 0.9 minutes.[2]

b. Hydrogen bonds

c.



FROM LESSON 2B

Chapter 5 review

Multiple choice

1. A. Butane is a non-polar compound, and is therefore insoluble in water.

2. B. $R_f = \frac{16 - 2}{20 - 2}$

$R_f = \frac{14}{18}$

$R_f = 0.78$

3. A. The most strongly adsorbed dye moves the least distance from the origin:

$R_f = \frac{6 - 2}{18 - 2}$

$R_f = \frac{4.0}{16}$

$R_f = 0.25$

4. A. Compounds have a unique retention time under given conditions.

5. D. The peak area can be compared to a calibration curve of substances of known concentration.

6. D. Octanol has the largest non-polar region, and is therefore the most soluble in non-polar solvents.

7. A. The most polar molecule is most attracted to the polar stationary phase, and therefore moves the smallest distance from the origin.

8. C. Ethanol, C_2H_5OH, dissolves in water by forming hydrogen bonds.

9. D. Polar molecules will be more attracted to the polar solvent particles by dipole-dipole attraction and will therefore move more quickly as they are dissolved in the free moving mobile phase. Non-polar molecules will be more attracted to the non-polar stationary phase by dispersion phases, and will adsorb more readily than they desorb.

10. B. Any change in HPLC conditions will cause a change in retention times. Given the use of a polar solvent, the molecules with shortest retention times are the most polar. Concentration is indicated by peak area; compound W has the highest peak area, and is therefore present in the highest concentration.

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Short answer

11. [A compound of like polarity to a given solvent will be soluble in the solvent.[1]] [Cholesterol has only one polar hydroxyl (—OH) group, and a large non-polar region, making it largely non-polar.[2]] [Given water is a polar solvent, cholesterol would be relatively insoluble in water.[3]]

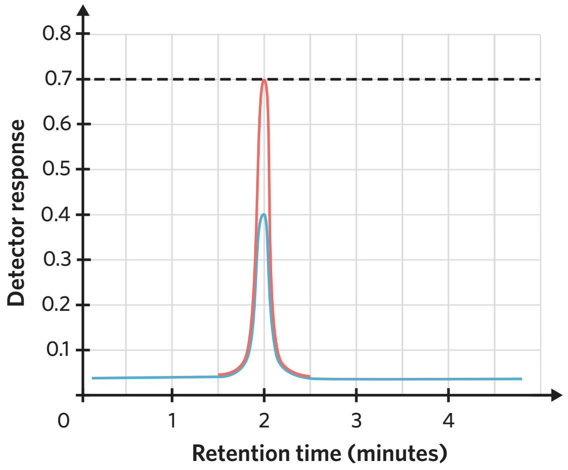
✓ ✗ I have identified the general relationship between polarity and solubility.[1]

✓ ✗ I have described the non-polar nature of cholesterol.[2]

✓ ✗ I have used the polar nature of water to describe the solubility of cholesterol in water.[3]

12. a. 2 minutes

b.



13. a. [The substance responsible for peak Q is glucose.[1]] [The substance responsible for peak R is glycerol.[2]]

✓ ✗ I have identified the substance responsible for peak Q.[1]

✓ ✗ I have identified the substance responsible for peak R.[2]

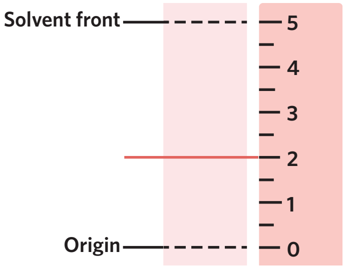
b. [Retention time depends on the relative attraction of a compound to the stationary and mobile phases.[1]] [The more soluble a compound is in the mobile phase, the more it will desorb into the mobile phase, and thus, the shorter its retention time will be.[2]] [The less soluble a molecule is in the mobile phase, the more it will adsorb onto the stationary phase, and thus, the longer its retention time will be.[3]]

✓ ✗ I have identified the general factors affecting retention time.[1]

✓ ✗ I have explained the retention time of compounds that are more soluble in the mobile phase.[2]

✓ ✗ I have explained the retention time of compounds that are less soluble in the mobile phase.[3]

14. a.



b. [Compounds that adsorb less strongly to the stationary phase in thin layer chromatography travel further up the plate from the origin.[1]] [Therefore, since the R_f value for a given compound is the ratio of the distance moved by the compound to the distance moved by the solvent front,[2]] [4-aminophenol would have a greater R_f value than paracetamol.[3]]

✓ ✗ I have explained that compounds that adsorb less strongly will migrate further from the origin.[1]

✓ ✗ I have explained that an R_f value is the ratio of the distance moved by a compound to the distance moved by the solvent front.[2]

✓ ✗ I have identified whether 4-aminophenol would have a greater or smaller R_f value than paracetamol.[3]

15. a. [Band A.[1]] [The components in band A migrated the least down the column, hence spending the most time adsorbed to the stationary phase.[2]]

✓ ✗ I have identified the band with the components that most strongly adsorbed to the stationary phase.[1]

✓ ✗ I have provided a justification for my answer.[2]

b. [The stationary phase could be modified, either by making the column longer or using more finely divided Al_2O_3.[1]] [The mobile phase could be changed to a solvent with a higher polarity.[2]]

✓ ✗ I have identified one possible modification.[1]

✓ ✗ I have identified a second possible modification.[2]

c. [Band C.[1]] [Band C has passed through the column fastest, and therefore contains the component that would be least strongly attracted to the stationary phase in an equivalent paper chromatography set-up.[2]] [As such, the component in band C would move furthest from the origin, producing the largest R_f value.[3]]

✓ ✗ I have identified the band with the component that would have the largest R_f value.[1]

✓ ✗ I have identified the attraction of the component to the stationary phase.[2]

✓ ✗ I have described how the component in band C would move furthest from the solvent front.[3]

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d. [Band A.[1]][Band A has adsorbed most strongly to the stationary phase, and would therefore be most attracted to the alumina stationary phase in the HPLC column.[2]][As such, the components in band A would move most slowly through the column, producing the largest retention time.[3]]

✓ ✗ I have identified the band with the component that would have the largest retention time.[1]

✓ ✗ I have identified the attraction of the component to the stationary phase.[2]

✓ ✗ I have described how the component in band A would move most slowly through the column.[3]

16. a. [Polar molecules are more soluble than non-polar molecules in polar solvents.[1]][Water is a polar solvent, and therefore, threonine will be more soluble than alanine in water.[2]]

✓ ✗ I have described the solubility of polar and non-polar molecules.[1]

✓ ✗ I have linked my answer to the question.[2]

b. Tyrosine

c. Y

d. [In chromatogram I, which used solvent G, alanine and threonine appeared as the same spot/were not separated since both had the same R_f value of 0.51, and hence with arginine ($R_f = 0.16$) and tyrosine ($R_f = 0.68$), there were 3 spots present on the chromatogram.[1]][In chromatogram II, which used solvent F, alanine and threonine had different R_f values and were therefore separated, leading to 4 spots on the second chromatogram.[2]]

✓ ✗ I have identified that the spots for alanine and threonine were merged in chromatogram I.[1]

✓ ✗ I have identified that the spots were separated due to differing R_f values in chromatogram II.[2]

Key science skills

17. Drink type

18. Quantitative

19. D

20. 124 ppm

21. [Caffeine has a retention time of 96 seconds under the given conditions.[1]][The chromatogram for soft drink B has no discernible peak at 96 seconds, and soft drink B is therefore unlikely to contain caffeine.[2]]

✓ ✗ I have given the retention time of caffeine under the given conditions.[1]

✓ ✗ I have noted the absence of a peak at 96 seconds on the chromatogram for soft drink B.[2]

22. [The caffeine peak area is beyond the range of the calibration graph provided.[1]][Extrapolation outside of the range of the calibration graph may provide inaccurate values of caffeine content for the espresso coffee.[2]]

✓ ✗ I have identified that the caffeine peak area for espresso coffee is outside the range of solutions.[1]

✓ ✗ I have explained the inaccuracy of results obtained from extrapolation.[2]

23. [The espresso coffee sample can be diluted.[1]][Diluting the solution of espresso coffee can bring the caffeine concentration of the sample within the range of the calibration curve.[2]]

✓ ✗ I have identified dilution as a possible solution.[1]

✓ ✗ I have explained the effect of dilution.[2]

FROM LESSONS 16B & 16C

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6A Relative mass

Progress questions

1. A. The masses on the periodic table are given relative to the mass of the carbon-12 isotope.

2. B. Relative atomic mass of an element takes into consideration the different isotopes in which the element can exist.

3. A. ^{14}Z is the most abundant of the two isotopes, therefore the relative atomic mass is likely to be closest to 14.

4. B. The relative molecular mass depends on the type and number of each atom in the molecule.

5. B. The mass of a molecule depends on the relative atomic mass of all atoms found in the molecule.

6. A. Relative masses are represented by M_r.

7. B. Cl_2 has two chlorine atoms in the molecule therefore, the value of M_r would be twice as much as the relative atomic mass of chlorine.

8. B. Each peak in a mass spectrum represents an isotope.

9. B. The least abundant isotope has the smallest peak.

10. A. The ^{90}Zr isotope is the most abundant isotope. Therefore, it is likely that the relative atomic mass of Zr is closer to 90.

Deconstructed

11. D

12. C

13. $A_r = \frac{\text{(relative abundance \% × relative isotopic mass) + (relative abundance \% × relative isotopic mass) + ...}}{100\%}$

$A_r = \frac{(75\% \times 35) + (25\% \times 37)}{100\%}$

$A_r = 35.5$ (no units) (1 MARK)

According to the periodic table, the element with a relative atomic mass of 35.5 is chlorine. Therefore, the element in the sample is likely to be chlorine. (1 MARK)

Exam-style

14. **a.** Relative abundance $= 100\% - (90.92 + 8.82)\%$

Relative abundance $= 0.26\%$

b. $A_r = \frac{\text{(relative abundance \% × relative isotopic mass) + (relative abundance \% × relative isotopic mass) + ...}}{100\%}$

$A_r = \frac{(90.92\% \times 19.99) + (0.26\% \times 20.99) + (8.82\% \times 21.99)}{100\%}$

(1 MARK)

$A_r = 20.17$ (no units) (1 MARK)

15. **a.** $A_r = \frac{\text{(relative abundance \% × relative isotopic mass) + (relative abundance \% × relative isotopic mass) + ...}}{100\%}$

$A_r = \frac{(79\% \times 24.0) + (10\% \times 25.0) + (11\% \times 26.0)}{100\%}$ (1 MARK)

$A_r = 24.3$ (no units) (1 MARK)

b. According to the periodic table, magnesium has a relative atomic mass of 24.3. Therefore, the element in the sample is likely to be magnesium.

16. **a.** Three

b. Total peak height = 90.9 mm + 0.3 mm + 8.8 mm = 100mm

Relative abundance of peak with m/z ratio of 20

$= \frac{90.9 \text{ mm}}{100 \text{ mm}} \times 100\%$

Relative abundance of peak with m/z ratio of 20

$= 90.9\%$ (1 MARK)

Relative abundance of peak with m/z ratio of 21

$= \frac{0.30 \text{ mm}}{100 \text{ mm}} \times 100\%$

Relative abundance of peak with m/z ratio of 21

$= 0.30\%$ (1 MARK)

Relative abundance of peak with m/z ratio of 22

$= \frac{8.9 \text{ mm}}{100 \text{ mm}} \times 100\%$

Relative abundance of peak with m/z ratio of 22

$= 8.8\%$ (1 MARK)

c. $A_r = \frac{\text{(relative abundance \% × relative isotopic mass) + (relative abundance \% × relative isotopic mass) + ...}}{100\%}$

$A_r = \frac{(90.9\% \times 20) + (0.30\% \times 21) + (8.8\% \times 22)}{100\%}$

$A_r = 20$ (no units) (1 MARK)

According to the periodic table, neon has a relative atomic mass of 20.2. Therefore, the element in the sample is likely to be neon. (1 MARK)

17. **a.** ^{175}Lu, given it has a relative isotopic mass closest to the relative atomic mass of lutetium.

b. Relative abundance of ^{175}Lu:

$A_r = \frac{\text{(relative abundance \% × relative isotopic mass) + (relative abundance \% × relative isotopic mass) + ...}}{100\%}$

$174.97 = \frac{(x\% \times 174.94) + ((100 - x)\% \times 175.94)}{100\%}$ (1 MARK)

$174.97 = \frac{(174.94x \times 17594 - 175.94x)\%}{100\%}$

$174.97 \times 100\% = (174.94x + 17594 - 175.94x)\%$

$17497 = -x + 17594$

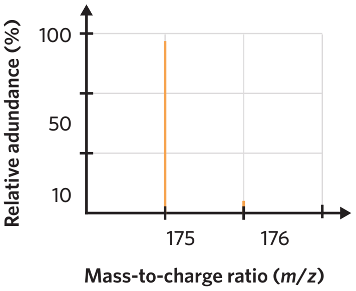
$-97.000 = -x$

$x = 97.000$ (no units) (1 MARK)

Therefore, the relative abundance of ^{175}Lu is 97.000%, and the relative abundance of ^{176}Lu is:

$100 - 97.000 = 3.0000\%$ (1 MARK)

c.



✓ ✗ I have labelled both axes.[1]

✓ ✗ I have drawn two peaks, one at 175 m/z and one at 176 m/z.[2]

✓ ✗ I have drawn the height of each peak relative to the abundance of each isotope.[3]

d. $M_r(\text{LuBr}_3) = (1 \times 174.97) + (3 \times 79.9)$

$M_r = 415$ (no units)

Key science skills

18. a. [Repeatability of an experiment depends on the closeness of results measured from an experiment in identical conditions.[1]] [Tests 1 and 2 were conducted by the same experimenter using the same method, one after the other.[2]] [The large variability in the data between tests 1 and 2 indicates a low repeatability.[3]]

✓ ✗ I have explained the term repeatability in relation to scientific investigation.[1]

✓ ✗ I have identified that tests 1 and 2 were conducted under identical conditions.[2]

✓ ✗ I have evaluated the repeatability of the experiment.[3]

b. Exclude outlying 'Test 2' results.

$$= \frac{24.9\% + 24.5\%}{2}$$

Relative abundance of isotope with m/z ratio of 71 $= 24.7\%$ (1 MARK)

Relative abundance of isotope with m/z ratio of 73

$$= \frac{75.1\% + 75.5\%}{2}$$

Relative abundance of isotope with m/z ratio of 73 $= 75.3\%$ (1 MARK)

FROM LESSON 16D

Questions from multiple lessons

19. a. s block

b. $5s^1$

c. $$A_r = \frac{(\text{relative abundance \%} \times \text{relative isotopic mass}) + (\text{relative abundance \%} \times \text{relative isotopic mass}) + \ldots}{100\%}$$

$$A_r = \frac{(72\% \times 85) + (28\% \times 87)}{100\%}$$ (1 MARK)

$A_r = 86$ (1 MARK)

FROM LESSON 1B

20. a. 29 protons

b. $35 + 29 = 64$

c. $$A_r = \frac{(\text{relative abundance \%} \times \text{relative isotopic mass}) + (\text{relative abundance \%} \times \text{relative isotopic mass}) + \ldots}{100\%}$$

$$63.5 = \frac{(x\% \times 63) + ((100 - x)\% \times 65)}{100\%}$$ (1 MARK)

$$63.5 = \frac{(63x + 6500 - 65x)\%}{100\%}$$

$63.5 \times 100\% = (63x + 6500 - 65x)\%$

$6350 = -2x + 6500$

$-150 = -2x$

$\frac{-150}{-2} = x$

$x = 75.0$ (no units) (1 MARK)

Therefore, the relative abundance of ^{63}Cu is 75%, and the relative abundance of ^{65}Cu is:

$100 - 75.0 = 25.0\%$ (1 MARK)

d. $M(\text{CuO}) = 63.5 + 16.0$

$M(\text{CuO}) = 79.5$ (no units)

FROM LESSON 1A & 4B

6B Avogadro's constant and the mole

Progress questions

1. A. Avogadro's constant, 6.02×10^{23} mol^{-1}, gives the number of particles in one mole of substance.

2. B. $N = n \times N_A$

$N = 2 \text{ mol} \times (6.02 \times 10^{23} \text{ mol}^{-1})$

$N = 1.20 \times 10^{24}$ particles

3. B. One mole of CO_2 and one mole of NH_3 both contain 6.02×10^{23} molecules. One mole of CO_2 contains 1.81×10^{24} atoms, whilst one mole of NH_3 contains 2.41×10^{24} atoms.

4. B. In 1 mol of fluorine gas, F_2, there is 2 mol of fluorine atoms.

5. A. The molar mass is the mass of one mole of substance.

6. B. The number that represents the relative atomic mass of an element on the periodic table is the same as the element's molar mass.

7. B. The mass of 1 mol of BeO is 25 g, however, the units for molar mass are g mol^{-1}. Therefore, the molar mass of BeO is 25.0 g mol^{-1}.

8. B. The formula for the amount, in mol, is $n = \frac{m}{M}$, meaning that both the mass and molar mass must be known to calculate mol.

9. B. The formula for the mass of a substance, in grams, is $m = n \times M$. Therefore, as the amount of a substance, in mol, decreases, the mass of the substance will also decrease.

10. B. The formula for the amount, in mol, is $n = \frac{m}{M}$.

11. B. Both equations are used to calculate the amount of substance, n.

Deconstructed

12. D

13. D

14. $$n(\text{NH}_3) = \frac{N(\text{NH}_3)}{N_A}$$

$$n(\text{NH}_3) = \frac{6.11 \times 10^{26}}{6.02 \times 10^{23} \text{ mol}^{-1}}$$

$n(\text{NH}_3) = 1014.95$ mol (1 MARK)

$m(\text{NH}_3) = n(\text{NH}_3) \times M(\text{NH}_3)$

$m(\text{NH}_3) = 1014.95 \text{ mol} \times (14.0 + (3 \times 1.0)) \text{ g mol}^{-1}$

$m(\text{NH}_3) = 1.73 \times 10^4$ g (1 MARK)

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Exam-style

15. **a.** $M(\text{HI}) = M(\text{H}) + M(\text{I})$

$M(\text{HI}) = 1.0 \text{ g mol}^{-1} + 126.9 \text{ g mol}^{-1}$

$M(\text{HI}) = 127.9 \text{ g mol}^{-1}$

b. $M(\text{MgCl}_2) = M(\text{Mg}) + (2 \times M(\text{Cl}))$

$M(\text{MgCl}_2) = 24.3 \text{ g mol}^{-1} + (2 \times 35.5) \text{ g mol}^{-1}$

$M(\text{MgCl}_2) = 95.3 \text{ g mol}^{-1}$

c. $M(\text{C}_3\text{H}_6\text{O}_3) = (3 \times M(\text{C})) + (6 \times M(\text{H})) + (3 \times M(\text{O}))$

$M(\text{C}_3\text{H}_6\text{O}_3) = (3 \times 12.0) \text{ g mol}^{-1} + (6 \times 1.0) \text{ g mol}^{-1} + (3 \times 16.0) \text{ g mol}^{-1}$

$M(\text{C}_3\text{H}_6\text{O}_3) = 90.0 \text{ g mol}^{-1}$

16. **a.** $n(\text{Cl}_2) = \frac{N(\text{Cl}_2)}{N_A}$

$n(\text{Cl}_2) = \frac{3.33 \times 10^{33}}{6.02 \times 10^{23} \text{ mol}^{-1}}$

$n(\text{Cl}_2) = 5.53 \times 10^{9} \text{ mol}$ (1 MARK)

Therefore, 2.11 mol of iron (Fe) atoms $< 3.33 \times 10^{33}$ molecules of chlorine gas (Cl_2). (1 MARK)

b. $n(\text{NO}_2) = \frac{N(\text{NO}_2)}{N_A}$

$n(\text{NO}_2) = \frac{5.17 \times 10^{11}}{6.02 \times 10^{23} \text{ mol}^{-1}}$

$n(\text{NO}_2) = 8.59 \times 10^{-13} \text{ mol}$ (1 MARK)

Therefore, 5.17×10^{11} molecules of nitrogen dioxide < 2.98 mol of carbon tetrachloride (CCl). (1 MARK)

c. $n(\text{CS}_2) = \frac{N(\text{CS}_2)}{N_A}$

$n(\text{CS}_2) = \frac{7.01 \times 10^{9}}{6.02 \times 10^{23} \text{ mol}^{-1}}$

$n(\text{CS}_2) = 1.16 \times 10^{-14} \text{ mol}$ (1 MARK)

$n(\text{PH}_3) = \frac{N(\text{PH}_3)}{N_A}$

$n(\text{PH}_3) = \frac{2.70 \times 10^{9}}{6.02 \times 10^{23} \text{ mol}^{-1}}$

$n(\text{PH}_3) = 4.49 \times 10^{-15} \text{ mol}$ (1 MARK)

Therefore, $n(\text{H})$ in $\text{PH}_3 = 3 \times n(\text{PH}_3)$

$n(\text{H})$ in $\text{PH}_3 = 3 \times 4.49 \times 10^{-15} \text{ mol}$

$n(\text{H})$ in $\text{PH}_3 = 1.35 \times 10^{-14} \text{ mol}$ (1 MARK)

Therefore, 7.01×10^{9} molecules of carbon disulfide (CS_2) $<$ number of hydrogen atoms in 2.70×10^{9} molecules of phosphine (PH_3). (1 MARK)

17. **a.** $m(\text{CO}) = n(\text{CO}) \times M(\text{CO})$

$m(\text{CO}) = 201.3 \text{ mol} \times (12.0 + 16.0) \text{ g mol}^{-1}$

$m(\text{CO}) = 5636.4 \text{ g}$

$m(\text{CO}) = 5.64 \text{ kg}$

b. Car A: $n(\text{CO}) = \frac{m(\text{CO})}{M(\text{CO})}$

$n(\text{CO}) = \frac{114 \text{ g}}{(12.0 + 16.0) \text{ g mol}^{-1}}$

$n(\text{CO}) = 4.07 \text{ mol}$ (1 MARK)

Car B: $n(\text{CO}) = \frac{N(\text{CO})}{N_A(\text{CO})}$

$n(\text{CO}) = \frac{12.04 \times 10^{33}}{6.02 \times 10^{23} \text{ mol}^{-1}}$

$n(\text{CO}) = 2.00 \times 10^{10} \text{ mol}$ (1 MARK)

Therefore, Car A releases significantly less carbon monoxide per 40 L tank of petrol than Car B. (1 MARK)

18. **a.** $M(\text{C}_6\text{H}_{12}\text{O}_6) = (6 \times M(\text{C})) + (12 \times M(\text{H})) + (6 \times M(\text{O}))$

$M(\text{C}_6\text{H}_{12}\text{O}_6) = (6 \times 12.0) \text{ g mol}^{-1} + (12 \times 1.0) \text{ g mol}^{-1} + (6 \times 16.0) \text{ g mol}^{-1}$

$M(\text{C}_6\text{H}_{12}\text{O}_6) = 180 \text{ g mol}^{-1}$ (1 MARK)

$M(\text{C}_2\text{H}_5\text{OH}) = (2 \times M(\text{C})) + (6 \times M(\text{H})) + M(\text{O})$

$M(\text{C}_2\text{H}_5\text{OH}) = (2 \times 12.0) \text{ g mol}^{-1} + (6 \times 1.0) \text{ g mol}^{-1} + 16.0 \text{ g mol}^{-1}$

$M(\text{C}_2\text{H}_5\text{OH}) = 46.0 \text{ g mol}^{-1}$ (1 MARK)

Therefore, ethanol has a lower molar mass. (1 MARK)

b. $n(\text{C}_6\text{H}_{12}\text{O}_6) = \frac{m(\text{C}_6\text{H}_{12}\text{O}_6)}{M(\text{C}_6\text{H}_{12}\text{O}_6)}$

$n(\text{C}_6\text{H}_{12}\text{O}_6) = \frac{4.11 \text{ g}}{180 \text{ g mol}^{-1}}$

$n(\text{C}_6\text{H}_{12}\text{O}_6) = 2.28 \times 10^{-2} \text{ mol}$

c. $n(\text{C}_6\text{H}_{12}\text{O}_6)_{required} = 3.00 \text{ mol} - 2.28 \times 10^{-2} \text{ mol}$

$n(\text{C}_6\text{H}_{12}\text{O}_6)_{required} = 2.977 \text{ mol}$ (1 MARK)

$m(\text{C}_6\text{H}_{12}\text{O}_6)_{required} = n(\text{C}_6\text{H}_{12}\text{O}_6)_{required} \times M(\text{C}_6\text{H}_{12}\text{O}_6)$

$m(\text{C}_6\text{H}_{12}\text{O}_6)_{required} = 2.977 \text{ mol} \times 180 \text{ g mol}^{-1}$

$m(\text{C}_6\text{H}_{12}\text{O}_6)_{required} = 536 \text{ g}$ (1 MARK)

19. B. $n(\text{C}) = \frac{m(\text{C})}{M(\text{C})}$

$n(\text{C}) = \frac{0.157 \text{ g}}{12.0 \text{ g mol}^{-1}}$

$n(\text{C}) = 0.0131 \text{ mol}$

$n(\text{Taxol}^{®}) = \frac{n(\text{C})}{47}$

$n(\text{O}) = 14 \times n(\text{Taxol}^{®})$

$n(\text{O}) = 14 \times \frac{n(\text{C})}{47}$

$n(\text{O}) = 14 \times \frac{0.0131 \text{ mol}}{47}$

$n(\text{O}) = 0.00390 \text{ mol}$

$m(\text{O}) = n(\text{O}) \times M(\text{O})$

$m(\text{O}) = 0.00390 \text{ mol} \times 16.0 \text{ g mol}^{-1}$

$m(\text{O}) = 0.0624 \text{ g}$

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Key science skills

20. a. The amount of methane collected

b. $N(CH_4) = n(CH_4) \times N_A$

$N(CH_4) = 15.8 \text{ mol} \times (6.02 \times 10^{23} \text{ mol}^{-1})$

$N(CH_4) = 9.51 \times 10^{24}$ molecules

c. [In the test run by the miners, large holes are drilled deep into the earth.[1]][This could destroy the land and potentially destroy the habitats of the organisms living in the area.[2]]

✓ ✗ I have identified the key environmental concern.[1]

✓ ✗ I have explained the effect that drilling can have on the environment.[2]

FROM LESSONS 16A & 16B

Questions from multiple lessons

21. a. $m(C_6H_{12}O_6) = n(C_6H_{12}O_6) \times M(C_6H_{12}O_6)$

$m(C_6H_{12}O_6) = 0.98 \text{ mol} \times ((6 \times 12.0) + (12.0 \times 1.0)$
$+ (6 \times 16.0)) \text{ g mol}^{-1}$

$m(C_6H_{12}O_6) = 1.8 \times 10^2 \text{ g}$

b. $m(NaHCO_3) = n(NaHCO_3) \times M(NaHCO_3)$

$m(NaHCO_3) = 0.11 \text{ mol} \times (23.0 + 1.0 + 12.0$
$+ (3 \times 16.0)) \text{ g mol}^{-1}$

$m(NaHCO_3) = 9.24 \text{ g}$ (1 MARK)

Mass of both $= 9.24 \text{ g} + 176.4 \text{ g}$

Mass of both $= 186 \text{ g}$ (1 MARK)

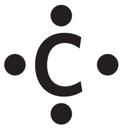
c. [Sodium.[1]][Period 3.[2]]

✓ ✗ I have identified the name of the metallic element in bicarbonate.[1]

✓ ✗ I have identified the period in which the metallic element would be found on the periodic table.[2]

FROM LESSONS 1B & 3A

22. a.



b. $1s^2 2s^2 2p^2$

c. $m(C)_{formed} = 17.41 \text{ g} - (10 + 5) \text{ g}$

$m(C)_{formed} = 2.41 \text{ g}$ (1 MARK)

$n(C)_{formed} = \frac{m(C)}{M(C)}$

$n(C)_{formed} = \frac{2.41 \text{ g}}{12.0 \text{ g mol}^{-1}}$

$n(C)_{formed} = 0.201 \text{ mol}$ (1 MARK)

FROM LESSONS 1B & 2A

23. a. 18

b. $1s^2 2s^2 2p^6 3s^2 3p^6$

c. [The relative atomic mass is derived from the average mass of all isotopes, and is most heavily weighted towards the most abundant isotope.[1]][According to the periodic table, the relative atomic mass of argon is 39.9.[2]][This value is closest to the relative isotopic mass of ^{40}Ar which is, therefore, the most abundant isotope.[3]]

✓ ✗ I have described how the relative atomic mass is calculated.[1]

✓ ✗ I have identified the relative atomic mass of Ar.[2]

✓ ✗ I have identified the most abundant isotope.[3]

d. $N(Ar) = n(Ar) \times N_A$

$N(Ar) = 2.44 \text{ mol} \times (6.02 \times 10^{23}) \text{ mol}^{-1}$

$N(Ar) = 1.47 \times 10^{24}$

Therefore there are 1.47×10^{24} atoms of argon present in a 2.44 mol sample.

FROM LESSONS 1B & 6A

6C Calculations involving mass

Progress questions

1. B. $\%$ by mass of nitrogen in $C_5H_{10}N_2O_2 = \frac{28.0 \text{ g mol}^{-1}}{130 \text{ g mol}^{-1}} \times 100\%$

$\%$ by mass of nitrogen in $C_5H_{10}N_2O_2 = 21.5\%$

$\%$ by mass of nitrogen in $C_5H_{10}N_2O_2 \approx 22\%$

2. B. $\%$ by mass of carbon in $C_2H_5OH = \frac{(2 \times 12.0) \text{ g mol}^{-1}}{46.0 \text{ g mol}^{-1}} \times 100\%$

$\%$ by mass of carbon in $C_2H_5OH = 52.2\%$

$\%$ by mass of carbon in $C_2H_5OH \approx 52\%$

3. B. The mass of an element present in a molecule is calculated by multiplying the number of moles of an element present in one mole of the molecule by the molar mass of the element.

4. B. Empirical formulas illustrate the lowest whole number ratio of elements in a compound.

5. B. $C_6H_{12}O_6$ can be divided by 6 to get the lowest whole number ratio of CH_2O, whilst $C_{12}H_{22}O_{11}$ is only divisible by 1, and is, therefore, an empirical formula. Empirical formulas illustrate the lowest whole number ratio of elements in a compound.

6. A. Molecular formulas depict the actual number of atoms of each element present.

7. B. Sometimes the molecular formula is the same as the empirical formula.

8. B. The empirical formula can be determined from the molecular formula by identifying a common factor to simplify the ratio.

9. A. The molar mass of the molecule is required to determine the number of empirical formula units in the molecular formula.

10. A. The molecular formula can be determined from the empirical formula and molecular mass of a molecule through the calculation of an empirical formula multiplier value.

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Deconstructed

11. C

12. D

13. Fe : O

0.00760 mol : 0.0114 mol

$\frac{0.00760}{0.00760} : \frac{0.0114}{0.00760}$

1.00 : 1.50 (1 MARK)

Multiply by 2 to get the lowest whole number ratio:

2.00 : 3.00

Therefore, the empirical formula is Fe_2O_3. (1 MARK)

[The molar mass of the empirical formula is $(2 \times 55.8\text{ g mol}^{-1}) + (3 \times 16.0\text{ g mol}^{-1}) = 159.6\text{ g mol}^{-1}$.[3]]

[Since the molar mass of the empirical formula is the same as the molar mass of the compound's ionic formula, the empirical formula and ionic formula are the same.[4]]

✓ ✗ I have determined the lowest ratio of iron to oxygen in the compound.[1]

✓ ✗ I have determined the empirical formula of the compound.[2]

✓ ✗ I have calculated the molar mass of the empirical formula.[3]

✓ ✗ I have determined the molecular formula.[4]

Exam-style

14. a. Mass of H in 1 mol of HCl = 1.0 g

$M(\text{HCl}) = 1.0\text{ g mol}^{-1} + 35.5\text{ g mol}^{-1}$

$M(\text{HCl}) = 36.5\text{ g mol}^{-1}$ (1 MARK)

$\%\text{ mass by composition} = \frac{1.0\text{ g mol}^{-1}}{36.5\text{ g mol}^{-1}} \times 100\%$

% mass by composition = 2.7% (1 MARK)

b. Mass of Mn in 1 mol of $KMnO_4$ = 54.9 g

$M(\text{KMnO}_4) = 39.1\text{ g mol}^{-1} + 54.9\text{ g mol}^{-1} + (4 \times 16.0\text{ g mol}^{-1})$

$M(\text{KMnO}_4) = 158.0\text{ g mol}^{-1}$ (1 MARK)

$\%\text{ mass by composition} = \frac{54.9\text{ g mol}^{-1}}{158.0\text{ g mol}^{-1}} \times 100\%$

% mass by composition = 34.8% (1 MARK)

c. Mass of H in 1 mol of $(NH_4)_3PO_4 = 12\text{ mol} \times 1.0\text{ g mol}^{-1}$

Mass of H in 1 mol of $(NH_4)_3PO_4$ = 12.0 g

$M((\text{NH}_4)_3\text{PO}_4) = (3 \times 14.0\text{ g mol}^{-1}) + (12 \times 1.0\text{ g mol}^{-1})$

$+ 31.0\text{ g mol}^{-1} + (4 \times 16.0\text{ g mol}^{-1})$

$M((\text{NH}_4)_3\text{PO}_4) = 149.0\text{ g mol}^{-1}$ (1 MARK)

$\%\text{ mass by composition} = \frac{12.0\text{ g mol}^{-1}}{149.0\text{ g mol}^{-1}} \times 100\%$

% mass by composition = 8.05% (1 MARK)

d. Mass of Fe in 1 mol of $Fe_2O_3 = 2\text{ mol} \times 55.8\text{ g mol}^{-1}$

Mass of Fe in 1 mol of Fe_2O_3 = 111.6 g

$M(\text{Fe}_2\text{O}_3) = (2 \times 55.8)\text{ g mol}^{-1} + (3 \times 16.0)\text{ g mol}^{-1}$

$M(\text{Fe}_2\text{O}_3) = 159.6\text{ g mol}^{-1}$ (1 MARK)

$\%\text{ mass by composition} = \frac{111.6\text{ g mol}^{-1}}{159.6\text{ g mol}^{-1}} \times 100\%$

% mass by composition = 69.9% (1 MARK)

e. $m(\text{Fe}) = 0.6992 \times 36.2\text{ g} = 25.3\text{ g}$

15. a. H : Cl

2.8% : 97.2%

2.8 g : 97.2 g

$\frac{2.8\text{ g}}{1.0\text{ g mol}^{-1}} : \frac{97.2\text{ g}}{35.5\text{ g mol}^{-1}}$

2.8 mol : 2.74 mol (1 MARK)

$\frac{2.8\text{ mol}}{2.74\text{ mol}} : \frac{2.74\text{ mol}}{2.74\text{ mol}}$

1.02 : 1.00

Empirical formula: HCl (1 MARK)

b. Pb : C : H

38.43 g : 17.83 g : 3.74 g

$\frac{38.43\text{ g}}{207.2\text{ g mol}^{-1}} : \frac{17.83\text{ g}}{12.0\text{ g mol}^{-1}} : \frac{3.74\text{ g}}{1.0\text{ g mol}^{-1}}$

0.185 mol : 1.49 mol : 3.74 mol (1 MARK)

$\frac{0.185\text{ mol}}{0.185\text{ mol}} : \frac{1.49\text{ mol}}{0.185\text{ mol}} : \frac{3.74\text{ mol}}{0.185\text{ mol}}$

1.00 : 8.01 : 20.2

Empirical formula: PbC_8H_{20} (1 MARK)

c. Al : S : O

15.8% : 28.1% : 56.1%

15.8 g : 28.1 g : 56.1 g

$\frac{15.8\text{ g}}{27.0\text{ g mol}^{-1}} : \frac{28.1\text{ g}}{32.1\text{ g mol}^{-1}} : \frac{56.1\text{ g}}{16.0\text{ g mol}^{-1}}$

0.585 mol : 0.875 mol : 3.51 mol (1 MARK)

$\frac{0.585\text{ mol}}{0.585\text{ mol}} : \frac{0.875\text{ mol}}{0.585\text{ mol}} : \frac{3.51\text{ mol}}{0.585\text{ mol}}$

1.00 : 1.50 : 6.00

2.00 : 3.00 : 12.00

Empirical formula: $Al_2S_3O_{12}$ (1 MARK)

d. C : H

4.8 g : 6.4 g – 4.8 g

4.8 g : 1.6 g

$\frac{4.8\text{ g}}{12.0\text{ g mol}^{-1}} : \frac{1.6\text{ g}}{1.0\text{ g mol}^{-1}}$

0.40 mol : 1.6 mol (1 MARK)

$\frac{0.40\text{ mol}}{0.40\text{ mol}} : \frac{1.6\text{ mol}}{0.40\text{ mol}}$

1.0 : 4.0

Empirical formula: CH_4 (1 MARK)

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e. N : O

$0.608\ \text{g} : 1.996\ \text{g} - 0.608\ \text{g}$

$0.608\ \text{g} : 1.388\ \text{g}$

$\dfrac{0.608\ \text{g}}{14.0\ \text{g mol}^{-1}} : \dfrac{1.388\ \text{g}}{16.0\ \text{g mol}^{-1}}$

$0.0434\ \text{mol} : 0.0868\ \text{mol}$ (1 MARK)

$\dfrac{0.0434\ \text{mol}}{0.0434\ \text{mol}} : \dfrac{0.0868\ \text{mol}}{0.0434\ \text{mol}}$

$1.00 : 2.00$

Empirical formula: NO_2 (1 MARK)

16. a. Molar mass of empirical formula $= 12.0\ \text{g mol}^{-1}$

$+ (2 \times 1.0\ \text{g mol}^{-1})$

Molar mass of empirical formula $= 14.0\ \text{g mol}^{-1}$

Empirical formula multiplier: $\dfrac{70.0\ \text{g mol}^{-1}}{14.0\ \text{g mol}^{-1}} = 5.0$ (1 MARK)

Molecular formula is $C_{(1 \times 5)}H_{(2 \times 5)} = C_5H_{10}$ (1 MARK)

b. Molar mass of empirical formula $= 12.0\ \text{g mol}^{-1}$

$+ (2 \times 1.0\ \text{g mol}^{-1}) + 16.0\ \text{g mol}^{-1}$

Molar mass of empirical formula $= 30.0\ \text{g mol}^{-1}$

Empirical formula multiplier: $\dfrac{180\ \text{g mol}^{-1}}{30\ \text{g mol}} = 6.0$ (1 MARK)

Molecular formula is $C_{(1 \times 6)}H_{(2 \times 6)}O_{(1 \times 6)}$

$= C_6H_{12}O_6$ (1 MARK)

c. Molar mass of empirical formula $= (6 \times 12.0\ \text{g mol}^{-1})$

$+ (10 \times 1.0\ \text{g mol}^{-1}) + (2 \times 32.1\ \text{g mol}^{-1}) + 16.0\ \text{g mol}^{-1}$

Molar mass of empirical formula $= 162.2\ \text{g mol}^{-1}$

Empirical formula multiplier: $\dfrac{162.2\ \text{g mol}^{-1}}{162.2\ \text{g mol}^{-1}} = 1.00$ (1 MARK)

Molecular formula is $C_{(6 \times 1)}H_{(10 \times 1)}S_{(2 \times 1)}O_{(1 \times 1)}$

$= C_6H_{10}S_2O$ (1 MARK)

d. Molar mass of empirical formula $= (1 \times 12.0\ \text{g mol}^{-1})$

$+ (4 \times 1.0\ \text{g mol}^{-1}) + (1 \times 14.0\ \text{g mol}^{-1})$

Molar mass of empirical formula $= 30.0\ \text{g mol}^{-1}$

Empirical formula multiplier: $\dfrac{60.0\ \text{g mol}^{-1}}{30.0\ \text{g mol}^{-1}} = 2.00$ (1 MARK)

Molecular formula is $C_{(1 \times 2)}H_{(4 \times 2)}N_{(1 \times 2)}$

$= C_2H_8N_2$ (1 MARK)

e. Molar mass of empirical formula $= (3 \times 12.0\ \text{g mol}^{-1})$

$+ (8 \times 1.0\ \text{g mol}^{-1}) + 14.0\ \text{g mol}^{-1}$

Molar mass of empirical formula $= 58.0\ \text{g mol}^{-1}$

Empirical formula multiplier: $\dfrac{116.0\ \text{g mol}^{-1}}{58.0\ \text{g mol}^{-1}} = 2.00$ (1 MARK)

Molecular formula is $C_{(3 \times 2)}H_{(8 \times 2)}N_{(1 \times 2)}$

$= C_6H_{16}N_2$ (1 MARK)

Key science skills

17. a. [The carbon dioxide measurement will be accurate, but not the water measurement.[1]][This is because accuracy is a measure of how close the measured value is to the true value, and since the water absorber was exposed to the atmosphere, atmospheric water would have been absorbed, reducing the accuracy of the results.[2]]

✓ ✗ I have commented on the accuracy of the results.[1]

✓ ✗ I have justified my answer with reference to the definition of accuracy.[2]

b. [It is a systematic error.[1]][This is because the uncalibrated scale will provide measurements which are always incorrect by the same degree.[2]]

✓ ✗ I have identified the type of error.[1]

✓ ✗ I have justified my answer with reference to the definition of a systematic error.[2]

c. [The experimenter is incorrect.[1]][This is because the results will still be precise, since the values are all affected to the same extent and therefore, will be close to each other. However, the results will be inaccurate because the measurements are not the true values as the scale is uncalibrated.[2]]

✓ ✗ I have commented on the experimenter's statement.[1]

✓ ✗ I have justified my answer with reference to the definition of precision and accuracy.[2]

FROM LESSON 16D

Questions from multiple lessons

18. a. $Li(s) + H_2O(l) \rightarrow LiOH(aq) + H_2(g)$

b. [$2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$.[1]][The above equation is now balanced because there is the same number of lithium, hydrogen and oxygen atoms on both sides of the chemical equation.[2]]

✓ ✗ I have written the balanced chemical equation.[1]

✓ ✗ I have justified why the chemical equation is balanced.[2]

c. $1s^2 2s^1$

d. [Lithium and fluorine are in the same period of the periodic table.[1]][However, fluorine is more electronegative (able to attract electrons) because it has a greater core charge with an increased number of protons and the same number of inner shell electrons compared to lithium.[2]][This allows fluorine to attract electrons more strongly than lithium.[3]]

✓ ✗ I have identified the commonality between lithium and fluorine.[1]

✓ ✗ I have justified my answer with reference to the trends in the periodic table.[2]

✓ ✗ I have compared the ability of both elements to attract electrons.[3]

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e. [Lithium and potassium are in the same group of the periodic table and also have the same core charge.[1]] [However, potassium has three occupied electron shells compared to lithium's two, resulting in the valence electron of potassium being further from the nucleus.[2]] [Therefore, the attraction between the valence electron and the nucleus of potassium is weaker, which results in potassium being able to lose its valence electron easier than lithium.[3]] [As a result, potassium is more reactive than lithium.[4]]

- ✓ ✗ I have identified the commonalities between lithium and potassium.[1]
- ✓ ✗ I have compared the structure of lithium and potassium.[2]
- ✓ ✗ I have described the effect of the number of occupied electron shells on the ability of the atoms to lose electrons.[3]
- ✓ ✗ I have linked my answer to the question.[4]

FROM LESSONS 1C & 3B

Chapter 6 review

Multiple choice

1. C. Butane has a molecular formula of C_4H_{10} according to the structure shown, which corresponds to a relative molecular mass of 58.0. The molecular formula can be divided by 2 to give the empirical formula C_2H_5.

2. D. $n(CaCO_3) = \frac{m}{M}$

$$n(CaCO_3) = \frac{25.1\ g}{(40.1 + 12.0 + (3 \times 16.0))\ g\ mol^{-1}}$$

$n(CaCO_3) = 0.2507$ mol

There are three moles of oxygen per mole of $CaCO_3$.

Therefore, $n(O) = 3 \times n(CaCO_3)$

$n(O) = 3 \times 0.2507$ mol

$n(O) = 0.752$ mol

3. C. $n(Ba(NO_3)_2) = \frac{N(Ba(NO_3)_2)}{N_A}$

$$n(Ba(NO_3)_2) = \frac{2.020 \times 10^{23}}{6.02 \times 10^{23}\ mol^{-1}}$$

$n(Ba(NO_3)_2) = 0.3355$ mol

There are six moles of oxygen per mole of $Ba(NO_3)_2$.

Therefore, $n(O) = 6 \times n(Ba(NO_3)_2)$

$n(O) = 6 \times 0.3355$ mol

$n(O) = 2.01$ mol

4. B. As a relative measurement, relative atomic mass does not have units.

5. C. The empirical formula CH_3 gives a molar mass of 15 g mol^{-1}, so the molecular formula must be C_2H_6.

6. B. ^{50}X has the lowest relative abundance of all the isotopes.

7. A. The relative atomic mass of an element is most heavily weighted towards the most abundant isotope.

8. D. Titanium has a relative atomic mass of 47.9.

9. B. There are 6 moles of oxygen atoms in 1 mole of glucose, so there are 12 moles of oxygen atoms in 2 moles of glucose.

10. B. $M = (55.8 + (5 \times 12.0) + (5 \times 16.0))\ g\ mol^{-1}$

$M = 195.8\ g\ mol^{-1}$

$$\%\ by\ mass\ of\ Fe = \frac{55.8\ g\ mol^{-1}}{195.8\ g\ mol^{-1}}$$

% by mass of Fe = 28.5%

$$\%\ by\ mass\ of\ C = \frac{(5 \times 12.0)\ g\ mol^{-1}}{195.8\ g\ mol^{-1}}$$

% by mass of C = 30.6%

$$\%\ by\ mass\ of\ O = \frac{(5 \times 16.0)\ g\ mol^{-1}}{195.8\ g\ mol^{-1}}$$

% by mass of O = 40.9%

Short answer

11. **a.** $n(CH_3COOH) = \frac{m}{M}$

$$n(CH_3COOH) = \frac{30.0\ g}{((2 \times 12.0) + (4 \times 1.0) + (2 \times 16.0))\ g\ mol^{-1}}$$

$n(CH_3COOH) = 0.500$ mol

b. $n(H) = 4 \times n(CH_3COOH)$

$n(H) = 4 \times 0.500$ mol

$n(H) = 2.00$ mol (1 MARK)

$N(H) = n(H) \times N_A$

$N(H) = 2.00\ mol \times (6.02 \times 10^{23}\ mol^{-1})$

$N(H) = 1.20 \times 10^{24}$ (1 MARK)

c. Mass of oxygen in 1 mol of CH_3COOH = 32.0 g

$M(CH_3COOH) = (2 \times 12.0\ g\ mol^{-1}) + (2 \times 16.0\ g\ mol^{-1}) + (4 \times 1.0\ g\ mol^{-1})$

$M(CH_3COOH) = 60.0\ g\ mol^{-1}$ (1 MARK)

$$Oxygen\ \%\ mass\ by\ composition = \frac{32.0\ g\ mol^{-1}}{60.0\ g\ mol^{-1}} \times 100\%$$

Oxygen % mass by composition = 53.3% (1 MARK)

12. **a.** Mass of K in 1 mol of $KMnO_4$ = 39.1 g

$M(KMnO_4) = 39.1\ g\ mol^{-1} + 54.9\ g\ mol^{-1} + (4 \times 16.0)\ g\ mol^{-1}$

$M(KMnO_4) = 158.0\ g\ mol^{-1}$ (1 MARK)

$$\%\ mass\ by\ composition = \frac{39.1\ g\ mol^{-1}}{158.0\ g\ mol^{-1}} \times 100\%$$

% mass by composition = 24.7% (1 MARK)

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b. Mass of N in 1 mol of $(NH_4)_3PO_4$ = 3 mol × 14.0 g mol^{-1}

Mass of N in 1 mol of $(NH_4)_3PO_4$ = 42.0 g

$M((NH_4)_3PO_4) = (3 \times 14.0)\text{ g mol}^{-1} + (12 \times 1.0)\text{ g mol}^{-1}$

$+ 31.0\text{ g mol}^{-1} + (4 \times 16.0)\text{ g mol}^{-1}$

$M((NH_4)_3PO_4) = 149.0\text{ g mol}^{-1}$ (1 MARK)

$\%\text{ mass by composition} = \dfrac{42.0\text{ g mol}^{-1}}{149.0\text{ g mol}^{-1}} \times 100\%$

% mass by composition = 28.2% (1 MARK)

13. $n(CO_2) = \dfrac{m}{M}$

$n(CO_2) = \dfrac{5.68\text{ g}}{44.0\text{ g mol}^{-1}}$

$n(CO_2) = 0.1291$ mol (1 MARK)

$n(C_2H_5OH) = n(CO_2)$ (1 MARK)

$m(C_2H_5OH) = n \times M$

$m(C_2H_5OH) = 0.1291\text{ mol} \times ((2 \times 12.0) + (6 \times 1.0) + 16.0)\text{ g mol}^{-1}$

$m(C_2H_5OH) = 5.94$ g (1 MARK)

14. $n(C) = \dfrac{m}{M}$

$n(C) = \dfrac{0.120\text{ g}}{12.0\text{ g mol}^{-1}}$

$n(C) = 1.00 \times 10^{-2}$ mol (1 MARK)

$n(C_{14}H_9Cl_5) = \dfrac{n(C)}{14}$

$n(C_{14}H_9Cl_5) = \dfrac{1.00 \times 10^{-2}\text{ mol}}{14}$

$n(C_{14}H_9Cl_5) = 7.14 \times 10^{-4}$ mol (1 MARK)

$n(Cl) = 5 \times n(C_{14}H_9Cl_5)$

$n(Cl) = 5 \times (7.14 \times 10^{-4}\text{ mol})$

$n(Cl) = 3.57 \times 10^{-3}$ mol (1 MARK)

$m(Cl) = n \times M$

$m(Cl) = (3.57 \times 10^{-3}\text{ mol}) \times 35.5\text{ g mol}^{-1}$

$m(Cl) = 0.127$ g (1 MARK)

15. **a.** Mass spectrometer

b. ^{25}Mg

c. $A_r = \dfrac{(\text{relative abundance }\% \times \text{relative isotopic mass}) + (\text{relative abundance }\% \times \text{relative isotopic mass}) + \ldots}{100\%}$

$A_r = \dfrac{(78.99\% \times 23.985) + (10.00\% \times 24.986) + (11.01\% \times 25.983)}{100\%}$

(1 MARK)

$A_r = 24.31$ (no units) (1 MARK)

16. **a.** $n(H_2O) = \dfrac{m}{M}$

$n(H_2O) = \dfrac{3.21\text{ g}}{18.0\text{ mol}^{-1}}$

$n(H_2O) = 0.1783$ mol

$n(H) = 2 \times n(H_2O)$

$n(H) = 0.3566$ mol (1 MARK)

$m(H)$ in compound = $n(H) \times M(H)$

$m(H)$ in compound = 0.3566 mol × 1.0 g mol^{-1}

$m(H)$ in compound = 0.36 g

$n(CO_2) = \dfrac{n}{M}$

$n(CO_2) = \dfrac{7.85\text{ g}}{44.0\text{ g mol}^{-1}}$

$n(CO_2) = 0.1784$ mol

$n(C) = 1 \times n(CO_2)$

$n(C) = 0.1784$ mol (1 MARK)

$m(C)$ in compound = $n \times M$

$m(C)$ in compound = 0.1784 mol × 12.0 g mol^{-1}

$m(C)$ in compound = 2.141 g

$m(O)$ in compound = total mass of compound – $m(C)$ – $m(H)$

$m(O)$ in compound = 3.07 g – 2.141 g – 0.36 g

$m(O)$ in compound = 0.572 g

$n(O) = \dfrac{m}{M}$

$n(O) = \dfrac{0.572\text{ g}}{16.0\text{ mol}^{-1}}$

$n(O) = 0.0357$ mol (1 MARK)

$$\begin{array}{ccccc} C & : & H & : & O \\ \end{array}$$

0.1784 mol : 0.357 mol : 0.0357 mol

$$\frac{0.1784}{0.0357} : \frac{0.357}{0.0357} : \frac{0.0357}{0.0357}$$

$$4.99 : 10.0 : 1.00$$

Therefore, the empirical formula is $C_5H_{10}O$. (1 MARK)

b. Molar mass of empirical formula = (5 × 12.0) g mol^{-1} + (10 × 1.0) g mol^{-1} + (1 × 16.0) g mol^{-1}

Molar mass of empirical formula = 86.0 g mol^{-1} (1 MARK)

$\text{Empirical formula multiplier} = \dfrac{172.0\text{ g mol}^{-1}}{86.0\text{ g mol}^{-1}}$

Empirical formula multiplier = 2 (1 MARK)

Molecular formula is $2 \times (C_5H_{10}O) = C_{(5 \times 2)}H_{(10 \times 2)}O_{(1 \times 2)}$

Molecular formula is = $C_{10}H_{20}O_2$ (1 MARK)

c. An isotope is an atom of the same element (same number of protons) with a different number of neutrons.

d. [The one small peak most likely corresponds to a molecule with one ^{13}C atom instead of all ^{12}C atoms.[1]][This is because the relative molecular mass is one greater than what is expected and the ^{13}C isotope is the most common compared to ^{2}H or ^{17}O.[2]]

✓ ✗ I have identified what causes the peak at m/z 173.[1]

✓ ✗ I have justified my answer with reference to the relative abundances of isotopes.[2]

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Key science skills

17. a. [^{63}Cu is more abundant in nature.[1]][This is because the relative atomic mass of copper is closer to 63 than it is to 65, indicating that there is a greater percentage of ^{63}Cu isotopes present than ^{65}Cu isotopes.[2]]

✓ ✗ I have specified which isotope is more abundant in nature.[1]

✓ ✗ I have justified my answer.[2]

b. $2Cu_2O(s) \rightarrow O_2(g) + 4Cu(s)$

c. According to the balanced chemical equation, two moles of Cu_2O decompose to produce one mole of $O_2(g)$ and four moles of $Cu(s)$.

$n(O_2) = \frac{1}{2} \times 1.00\ \text{mol}$ (1 MARK)

$n(O_2) = 0.500\ \text{mol}$ (1 MARK)

$N = 6.02 \times 10^{23}\ \text{mol}^{-1} \times 0.500\ \text{mol}$

$N = 3.01 \times 10^{23}$ molecules (1 MARK)

d. Quantitative

e. Mass spectrometry requires the use of a mass spectrometer, which is very expensive and requires a technician to operate, limiting the accessibility of the technique.

FROM LESSONS 16B & 16D

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7A Grouping hydrocarbons

Progress questions

1. B. Hydrocarbons are carbon-based molecules, making them organic.

2. B. Organic molecules can be naturally occurring or synthetically produced.

3. A. Hydrocarbons are usually arranged in a linear, chain-like fashion.

4. B. The size of a hydrocarbon chain affects its boiling point.

5. B. Unsaturated hydrocarbons have carbon to carbon double bonds.

6. A. Alkenes readily undergo addition reactions.

7. B. The molecule is non-polar, and therefore adjacent molecules will be held together by dispersion forces.

8. B. According to the general formula C_nH_{2n+2}, the molecule will contain 6 carbon atoms.

9. A. As the smaller molecule, the dispersion forces will be weaker in C_2H_6 and therefore the melting point will be lower.

10. B. Alkanes and alkenes have relatively similar physical properties, but can be differentiated by their different chemical properties.

Deconstructed

11. D

12. D

13. [Both structures are non-polar compounds. As a result, dispersion forces will be the strongest force of attraction between both methane and octane molecules.[1]] [However, octane is a much larger molecule than methane, and would therefore have stronger dispersion forces between adjacent molecules than methane.[2]] [As a result, it would have a higher boiling point than methane, and therefore would be able to withstand higher temperatures.[3]] [Therefore, fuel A is methane and fuel B is octane.[4]]

✓ ✗ I have identified the strongest intermolecular force present for the two molecules.[1]

✓ ✗ I have identified the larger molecule, and stated the effect of this on the strength of dispersion forces.[2]

✓ ✗ I have related the strength of dispersion forces to physical properties.[3]

✓ ✗ I have identified and named fuel A and fuel B.[4]

Exam-style

14. **a.** [Molecule A is considered saturated.[1]] [The term 'saturated' refers to the number of hydrogens bonded to the carbon atoms in the molecule.[2]]

✓ ✗ I have identified the saturated molecule.[1]

✓ ✗ I have explained what the term 'saturated' refers to.[2]

b. [Molecule B would take part in an addition reaction.[1]] [The reactive, electron-rich double bond allows the molecule to undergo such a reaction.[2]]

✓ ✗ I have identified the molecule that would take part in an addition reaction.[1]

✓ ✗ I have identified a property that allows the molecule to undergo such a reaction.[2]

c. Molecule A, as alkanes combust to a greater extent than alkenes.

Key science skills

15. **a.** Alkane/saturated hydrocarbon

b. [Most hydrocarbons are not able to interact with water due to the differences in their polarities,[1]] [therefore the two substances are unable to mix as they are insoluble in one another.[2]]

✓ ✗ I have identified the differences in polarities between water and hydrocarbons.[1]

✓ ✗ I have linked my answer to the question.[2]

c. Personal error/mistake

FROM LESSON 16D

Questions from multiple lessons

16. **a.** **i.** Yes

ii. Only single bonds between carbons in the carbon chain

iii. No

b. **i.** Yes

ii. At least one double (or triple) bond between carbons in the carbon chain

iii. Yes

FROM LESSON 2A

17. **a.** [Each covalent bond represents a shared pair of electrons.[1]] [In a carbon to carbon double bond, there are a total of two covalent bonds.[2]] [As a result, there are a total of four electrons that are being shared.[3]]

✓ ✗ I have identified the nature of covalent bonds.[1]

✓ ✗ I have identified the number of bonds in a double bond.[2]

✓ ✗ I have identified the total number of shared electrons.[3]

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b. [Diamonds have a stronger structure than hydrocarbons.[1]] [The structure of hydrocarbons consists of a linear chain of carbon atoms held together by dispersion forces, whereas diamond consists of a 3D carbon lattice held together by covalent bonds.[2]] [Consequently, it is more difficult to disrupt the structure of diamond compared to a hydrocarbon, since covalent bonds are significantly stronger than dispersion forces.[3]]

✓ ✗ I have identified that diamonds are stronger than hydrocarbons.[1]

✓ ✗ I have described the structures of hydrocarbons and diamonds.[2]

✓ ✗ I have explained why the structure of diamond is stronger than the structure of hydrocarbons.[3]

FROM LESSON 2C

18. **a.** 40.0 L = 40 000 mL

Mass of octane = $40\ 000 \text{ mL} \times 0.703 \text{ g mL}^{-1}$ (1 MARK)

Mass of octane = 28 120 g

Mass of octane = 2.81×10^4 g (1 MARK)

b. $$n(\text{octane}) = \frac{2.81 \times 10^4 \text{ g}}{(8 \times 12.0 \text{ g mol}^{-1}) + (18 \times 1.0 \text{ g mol}^{-1})}$$

(1 MARK)

$n(\text{octane}) = 246.67$ mol

$n(\text{octane}) = 2.47 \times 10^2$ mol (1 MARK)

c. Energy = $2.47 \times 10^2 \text{ mol} \times 5\ 460 \text{ kJ mol}^{-1}$ (1 MARK)

Energy = 1 346 800 kJ

Energy = 1.35×10^6 kJ (1 MARK)

FROM LESSON 6C

7B Nomenclature

Progress questions

1. A. The parent name 'prop-' describes three carbon atoms.

2. B. The name 'octane' describes an eight-carbon chain, with only carbon to carbon single bonds.

3. A. The suffix used for carbon to carbon double bonds is '-ene'.

4. A. Ensuring that the double bond occurs at the lower numbered carbon, the name is but-1-ene.

5. B. The parent chain in an organic molecule does not have to be written linearly.

6. A. Both have the same molecular formula of C_5H_{12}, but have a different name, due to the different spatial arrangement of the atoms.

7. B. The parent chain is numbered based on the position of carbon to carbon double bonds first.

8. A. Both molecules have the same parent chain, and a methyl group located at the same numbered carbon.

9. B. Electron dot formulas show the valence electrons of each atom.

10. A. There are a total of five carbons in the molecule.

Deconstructed

11. D

12. D

13. [The longest carbon chain in this molecule is eight carbon atoms long, and therefore the parent name should be 'oct-' rather than 'hept-'.[1]] [There is a methyl group on the 3rd carbon in the carbon chain rather than an ethyl group on the 2nd carbon, and thus the prefix 3-methyl should be used.[2]] [Therefore, the correct name for this molecule is 3-methyloctane.[3]]

✓ ✗ I have identified the appropriate parent name.[1]

✓ ✗ I have identified the appropriate notation for the alkyl group.[2]

✓ ✗ I have provided the IUPAC systematic name for this molecule.[3]

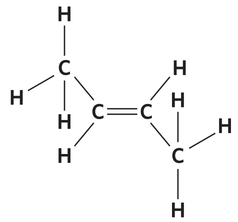
Exam-style

14. D. The longest carbon chain is 7 carbons long. There are methyl groups coming off the 3rd and 4th carbons.

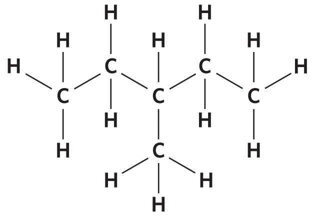
15. C. The longest carbon chain is 6 carbons long. There are methyl groups coming off the 3rd and 4th carbons.

16. C. The longest carbon chains in structures I and IV are 6 carbons long, meaning the names of the compounds would have to end in -hexane. In both structures II and III, the longest carbon chains are 5 carbons long. There are also 3 methyl groups in each structure, and in both cases, 2 of the methyl groups project from the 2nd carbon, and one methyl group projects from the 4th carbon.

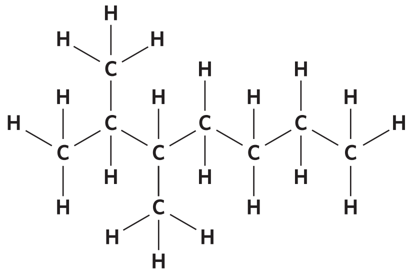
17. **a.** **i.**



ii.



iii.



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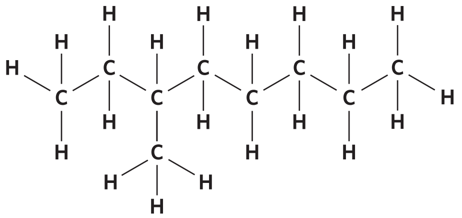
b. [When naming a molecule, side groups must be numbered such that they are attached to the lowest numbered carbon atom possible. This molecule should therefore be 3-methyloctane not 6-methyloctane.^{1}][Since the naming of the molecule is incorrect, this molecule cannot exist.^{2}]

✓ ✗ I have explained why this molecule's naming is incorrect.^{1}

✓ ✗ I have linked my answer to the question.^{2}

c. 3-methyloctane

d.



18. **a.** 2-methylpent-2-ene

b. 3,3-dimethylbut-1-ene

c. 3,3-dimethylpent-1-ene

d. 2-methyloct-3-ene

Key science skills

19. **a.** **i.** [The results are inaccurate, as they are not close to the true value.^{1}][However, they are precise as the measured values are close together.^{2}]

✓ ✗ I have commented on the accuracy of the results.^{1}

✓ ✗ I have commented on the precision of the results.^{2}

ii. [The results indicate a systematic error.^{1}][This is because all of the results have been affected to the same degree, which indicates that there may be an error with the procedure or the instruments being used, rather than a personal error.^{2}]

✓ ✗ I have identified that the most likely type of error is a systematic error.^{1}

✓ ✗ I have explained why the results are likely the result of a systematic error.^{2}

b. [Use of instruments to maintain distance from flame e.g. crucible tongs or use a fume hood to protect from hazardous fumes.^{1}][Use a safety flame when flame is not in use or tie up hair.^{2}]

✓ ✗ I have identified one precaution that may help to prevent injury.^{1}

✓ ✗ I have identified a second precaution that may help to prevent injury.^{2}

FROM LESSONS 16B & 16D

Questions from multiple lessons

20. **a.** Molecular formula of the compound: C_8H_{16}

$M(C_8H_{16}) = ((12.0 \times 8) + (1.0 \times 16))\ \text{g mol}^{-1}$

$M(C_8H_{16}) = 112.0\ \text{g mol}^{-1}$

b. 3-ethylhex-3-ene

c. [Unsaturated hydrocarbons contain at least one carbon to carbon double bond in the parent chain, whereas saturated hydrocarbons contain only single carbon to carbon bonds in the parent chain.^{1}][Due to the carbon to carbon double bond located at carbon 3 in the parent chain, this molecule is considered an unsaturated hydrocarbon.^{2}]

✓ ✗ I have compared the definitions of saturated and unsaturated hydrocarbons.^{1}

✓ ✗ I have identified the molecule as either saturated or unsaturated.^{2}

d. This molecule is branched, which results in the molecules having weaker dispersion forces between adjacent molecules.

FROM LESSONS 2B, 6A & 7A

21. **a.** [Due to the small difference in electronegativity between carbon and hydrogen atoms, the bonds occurring between carbon and hydrogen in the hydrocarbon are non-polar.^{1}] [As a result, the molecule shown is a non-polar compound, and therefore the forces responsible for holding these molecules together are weak dispersion forces.^{2}][As a result, to change phase of the substance made up of these compounds, the only intermolecular forces that need to be disrupted are dispersion forces.^{3}]

✓ ✗ I have described the bonding between carbon and hydrogen atoms.^{1}

✓ ✗ I have described the non-polar nature of the compound.^{2}

✓ ✗ I have linked my answer to the question.^{3}

b. $CH_2\,CHCH_2CH(CH_3)CH_2CH_3$

c. $m = n \times M$

$m(C_7H_{14}) = 0.122\ \text{mol} \times ((12.0 \times 7) + (1.0 \times 14))\ \text{g mol}^{-1}$

$m(C_7H_{14}) = 12.0\ \text{g}$

FROM LESSONS 2B & 6B

7C Functional groups

Progress questions

1. A. The reactivity of haloalkanes means they are useful as intermediates in chemical processes.
2. B. Haloalkanes have higher boiling points than alkanes due to the presence of dipole-dipole interactions.
3. B. When the parent chain is 2 carbons long or less, no number designation is required.

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4. B. In 2-bromobutane, the halogen is attached to the second carbon in the parent chain.

5. A. A hydroxyl group is a functional group consisting of an oxygen and a hydrogen atom (–OH).

6. B. The influence of functional groups decreases as the carbon chain length increases.

7. A. The prefix 'ethan-' denotes a two-carbon saturated hydrocarbon, and the suffix '-ol' denotes an alcohol.

8. A. In pentan-3-ol, the hydroxyl group is attached to the third carbon in the parent chain.

9. B. The carboxyl group is able to take part in hydrogen bonding.

10. B. Carboxylic acids have a higher boiling point than alcohols as the carbonyl group also contributes to the polarity of the molecule.

11. A. The parent name 'propan-' is used to indicate three- carbon atoms are in the saturated carbon chain, and the suffix '-oic acid' indicates this is a carboxylic acid.

12. A. The carbon in the COOH group should be included when numbering the parent chain.

Deconstructed

13. C

14. A

15. [Butan-2-ol is named correctly,[1]] [however the molecule labelled butan-4-ol should be named butan-1-ol instead.[2]] [This is because functional groups should always be named on the lowest numbered carbon atom.[3]]

- ✓ ✗ I have identified which molecules are named correctly.[1]
- ✓ ✗ I have explained why any incorrectly named molecules are incorrect.[2]
- ✓ ✗ I have identified the correct name of any incorrectly named molecules.[3]

Exam-style

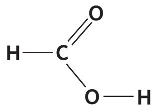
16. A. Ethane is the least polar compound, and is hence the least soluble in water at room temperature.

17. D. Of the options listed, only carboxylic acids contain 2 oxygen atoms.

18. a. [Methanoic acid is a carboxylic acid,[1]] [and therefore possesses a carboxyl functional group.[2]]

- ✓ ✗ I have identified the type of molecule.[1]
- ✓ ✗ I have identified the functional group contained.[2]

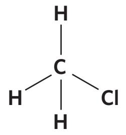
b.



c. i. Haloalkane

ii. Chloromethane

iii.



19. a. [The addition of a hydroxyl group makes the hydrocarbon polar.[1]] [This is due to a difference in electronegativity between the oxygen and hydrogen atoms in the hydroxyl functional group, making it a polar bond and therefore a polar functional group.[2]] [Consequently, hydrocarbons with hydroxyl groups are able to form stronger dipole-dipole interactions or potentially hydrogen bonds with other molecules.[3]]

- ✓ ✗ I have identified the effect of a hydroxyl group on the polarity of a compound.[1]
- ✓ ✗ I have explained why hydroxyl groups are polar functional groups.[2]
- ✓ ✗ I have described the intermolecular forces that will be seen in hydrocarbons with hydroxyl groups.[3]

b. [The stronger intermolecular forces between alcohols will result in a higher boiling point than in hydrocarbons without functional groups.[1]] [This is because more energy is needed to overcome the hydrogen bonds between alcohol molecules in comparison to the dispersion forces between alkane molecules.[2]]

- ✓ ✗ I have identified the compound with a higher boiling point.[1]
- ✓ ✗ I have explained that the increased boiling point relates to the increased amount of energy needed to overcome intermolecular forces.[2]

c. i. Carboxyl group

ii. [A dimer is a molecular complex consisting of two carboxylic acids with intermolecular interactions between them.[1]] [The carboxylic acids take part in hydrogen bonds between the –C=O and –OH groups of each of the molecules.[2]]

- ✓ ✗ I have identified what a dimer is.[1]
- ✓ ✗ I have described the arrangement of intermolecular forces in a dimer.[2]

Key science skills

20. a. [Keep test tube in a fume hood to protect from fumes or tie hair back.[1]] [Use gloves when handling chemical substances or use goggles to protect eyes.[2]]

- ✓ ✗ I have identified one factor.[1]
- ✓ ✗ I have identified another factor.[2]

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b. [The independent variable in this experiment is the type of organic substance used.[1]][This is because the independent variable is the variable changed by the experimenter, and the experiment involves exposing alcohols and carboxylic acids to a flame for comparison.[2]]

✓ ✗ I have identified the independent variable in this experiment.[1]

✓ ✗ I have explained why the type of substance is the independent variable.[2]

c. Boiling point

d. [Amount of the substances used or amount of heat from the bunsen burner.[1]][Position relative to flame.[2]][Type of test tube or equipment used to measure boiling point.[3]]

✓ ✗ I have identified one variable.[1]

✓ ✗ I have identified a second variable.[2]

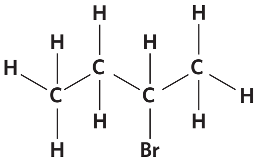
✓ ✗ I have identified a third variable.[3]

FROM LESSONS 16A & 16B

Questions from multiple lessons

21. a. [Pentan-1-ol.[1]][$CH_3CH_2CH_2CH_2CH_2OH$.[2]]

[2-bromobutane.[3]]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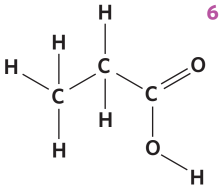
[Propanoic acid.[5]]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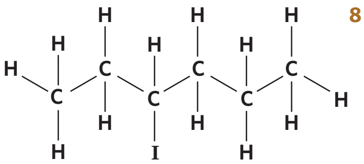
[$CH_3CH_2CH(I)CH_2CH_2CH_3$.[7]]


8

✓ ✗ I have named the alcohol.[1]

✓ ✗ I have drawn a semi-structural formula for the alcohol.[2]

✓ ✗ I have named the haloalkane.[3]

✓ ✗ I have drawn the structural formula (showing all bonds) for the haloalkane.[4]

✓ ✗ I have named the carboxylic acid.[5]

✓ ✗ I have drawn the structural formula (showing all bonds) for the carboxylic acid.[6]

✓ ✗ I have drawn the semi-structural formula for the haloalkane.[7]

✓ ✗ I have drawn the structural formula (showing all bonds) for the haloalkane.[8]

b. 3-iodohexane

c. i. [These molecules all have a higher boiling point than their alkane counterpart due to the presence of a functional group.[1]][These functional groups all have greater polarity than the non-polar bonds in the alkanes, making the overall molecule more polar and thus able to form stronger dipole-dipole interactions than the weak dispersion forces in alkanes,[2]][which produces a higher boiling point as more energy is required to break the intermolecular forces.[3]]

✓ ✗ I have identified the feature which provides these molecules with a higher boiling point than their alkane counterparts.[1]

✓ ✗ I have explained how functional groups contribute to the intermolecular interactions of molecules.[2]

✓ ✗ I have identified that these intermolecular forces increase the boiling point.[3]

ii. [The difference between the boiling points of an alcohol and its alkane counterpart decreases as the carbon chain length increases because of the reduced influence of the hydroxyl group.[1]][As the length of the carbon chain increases, the dispersion forces in both molecules become stronger[2]][and, since the non-polar portion of the alcohol molecule increases, its polarity decreases, thus decreasing the ability of the alcohols to have strong dipole-dipole interactions.[3]][As boiling point is proportional to the strength of intermolecular forces, these factors together result in the boiling points becoming increasingly similar as the carbon chain length increases.[4]]

✓ ✗ I have identified that the hydroxyl group gradually exerts less influence on intermolecular forces as the carbon chain increases.[1]

✓ ✗ I have explained that dispersion forces increase in both molecules.[2]

✓ ✗ I have explained that the capacity for dipole-dipole interactions is reduced.[3]

✓ ✗ I have described how these factors result in the boiling points coming closer together.[4]

FROM LESSONS 6B & 7B

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7D Renewable sources

Progress questions

1. A. Crude oil is the result of millions of years of heat, pressure, and sediment build-up on ancient marine life.

2. B. Currently, crude oil remains the major global source of organic chemicals.

3. B. Biomass is a generic term for material from plants used by humans for fuel or chemical production.

4. A. 95% of global ethanol production occurs through fermenting crops, such as corn and sugarcane.

5. B. Although plant-based biomass is renewable, the high demand for organic chemicals means that they often cannot be produced from plants at a rate faster than the rate at which they are consumed.

Deconstructed

6. C

7. A

8. [Dextrin glue is a low-strength adhesive that would be ideal for cardboard coffee cups.[1]][It is produced by dry-roasting starch, a component of all green plants. Since green plants replenish naturally, this is a renewable source.[2]]

✓ ✗ I have identified dextrin glue as a suitable alternative adhesive.[1]

✓ ✗ I have explained why dextrin glue is renewable.[2]

Exam-style

9. [PVA glue is an organic chemical synthesised from hydrocarbons, which are derived from crude oil.[1]][Crude oil is formed from the remains of ancient sea creatures subjected to millions of years of heat and pressure[2]][therefore making the student's statement valid.[3]]

✓ ✗ I have explained that PVA glue is derived from crude oil.[1]

✓ ✗ I have explained that crude oil is formed from ancient marine life.[2]

✓ ✗ I have evaluated the validity of the student's statement.[3]

10. a. [Currently, organic chemicals are sourced primarily from crude oil, which is non-renewable and thus will not be replenished for millions of years.[1]][Crude oil will become more and more expensive as it becomes scarcer, and could in fact be used up completely.[2]][As a result, the renewable alternative biomass must be used to ensure society has continuing access to organic chemicals in the future.[3]]

✓ ✗ I have explained that the current primary source of organic chemicals is non-renewable.[1]

✓ ✗ I have linked crude oil's renewability to its likely increasing scarcity and cost.[2]

✓ ✗ I have explained that biomass is renewable and hence provides longer-term access to organic chemicals.[3]

b. [Many organic chemicals have no plant-based alternative source/plant-based sources often each produce a specific chemical, compared with crude oil's widespread uses.[1]] [Many plant-based chemicals require additives from crude oil to be used industrially.[2]][Sourcing chemicals from plant-based biomass can be highly expensive.[3]][Large land areas are needed to grow sufficient crops to fulfil high demand for organic chemicals/high demand means plant crops are sometimes grown in an unethical or environmentally harmful way.[4]]

✓ ✗ I have given a first reason.[1]

✓ ✗ I have given a second reason.[2]

✓ ✗ I have given a third reason.[3]

✓ ✗ I have given a fourth reason.[4]

11. a. [Ethanol can be synthesised from ethene, a hydrocarbon derived from the separation of crude oil, a non-renewable resource.[1]] [Crude oil itself is extracted by drilling underground deposits.[2]]

✓ ✗ I have explained how ethanol is produced from crude oil.[1]

✓ ✗ I have explained how crude oil is extracted by humans.[2]

b. [Ethanol can be produced from the fermentation of plant crops, a renewable resource.[1]][Large fields of sugarcane or corn are the major plant crops used in ethanol production.[2]]

✓ ✗ I have explained how ethanol is produced from fermentation of plant crops.[1]

✓ ✗ I have identified the major source of these plant crops.[2]

c. Any one of the following: as a solvent (for example, in hand sanitiser), in alcoholic beverages, as a fuel.

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Key science skills

12. a. [Onshore and offshore drilling for fossil fuels such as crude oil is a process often very harmful to land and marine environments, disrupting natural ecosystems and destroying habitats of plant and animal life.[1]][Plant biomass can be a more environmentally-friendly source, but rapid harvesting and replanting crops can cause soil degradation, while clearing land for crops can also destroy habitats.[2]]

✓ ✗ I have discussed the environmental issues associated with using fossil fuels as a source of organic chemicals.[1]

✓ ✗ I have discussed the environmental advantages and issues with biomass as a source of organic chemicals.[2]

b. [Fossil fuels are not a sustainable source of organic chemicals, as they are replenished at a far slower rate than they are currently consumed.[1]][On the other hand, plant-based biomass is quickly replenished by replanting, making it a much more sustainable source.[2]]

✓ ✗ I have discussed the sustainability of fossil fuels as a source of organic chemicals.[1]

✓ ✗ I have discussed the sustainability of biomass as a source of organic chemicals.[2]

c. [Currently, fossil fuels are often a much cheaper source of organic chemicals, but as they are used up and become scarcer, their cost will increase significantly.[1]][The cost of producing chemicals from plant-based biomass can vary significantly depending on plant availability and the efficiency of the process, but as research continues, this cost will likely decrease.[2]]

✓ ✗ I have discussed the economic cost of fossil fuels as a source of organic chemicals.[1]

✓ ✗ I have discussed the economic cost of biomass as a source of organic chemicals.[2]

d. [Crude oil drilling can often destroy sites of cultural and aesthetic significance, and has faced numerous ethical issues for this reason.[1]][Similarly, plant-based sources have struggled to keep up with the high demand for organic chemicals, leading to worker exploitation of farmers in order to meet this supply expectation.[2]]

✓ ✗ I have discussed the ethical issues associated with using fossil fuels as a source of organic chemicals.[1]

✓ ✗ I have discussed the ethical issues associated with using biomass as a source of organic chemicals.[2]

FROM LESSON 16B

Questions from multiple lessons

13. a. Dispersion forces

b. [As the size of an alkane increases, the number of electrons present in the molecule increases.[1]][As a result, the strength of the instantaneous dipoles increases, meaning larger alkanes have stronger dispersion forces between them.[2]][More energy is therefore required to disrupt the stronger intermolecular forces between larger alkane molecules than those of smaller alkanes.[3]][Hence larger alkanes need higher temperatures to boil, and thus have higher boiling points than smaller alkanes.[4]]

✓ ✗ I have linked larger alkane size with higher electron count.[1]

✓ ✗ I have linked higher electron count with stronger dispersion forces.[2]

✓ ✗ I have linked stronger dispersion forces with higher energy required to disrupt them.[3]

✓ ✗ I have linked my explanation to larger alkanes having higher boiling points.[4]

c. [Fractional distillation is not usually the final step in the production of organic chemicals.[1]][After being separated by fractional distillation, hydrocarbons must be combined with other compounds to form most everyday organic chemicals.[2]]

✓ ✗ I have observed that fractional distillation is not usually the final step in organic chemical production.[1]

✓ ✗ I have explained that synthesis with other compounds is usually necessary after fractional distillation.[2]

FROM LESSONS 2B & 7A

7E Organic compounds and society

Progress

1. A. Biomagnification is the phenomenon by which contaminant concentration is higher in animals higher on the food chain.

2. B. Sugars, fats, proteins, and food additives are all organic compounds.

3. B. Certain organic chemicals diminish the ozone layer, which is a safeguard from harmful UV rays.

4. A. Plastics tend to be long-lasting and damage-resistant.

5. A. Microplastics have no proven negative health effects on humans, but have been shown to harm certain animals.

Deconstructed

6. D

7. C

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8. [Microbeads are tiny organic compounds, traditionally used in exfoliating products and toothpastes, that are small enough to pass through sewerage systems and into aquatic ecosystems.[1]][They then absorb harmful chemicals in the water and are mistaken for food by small aquatic animals.[2]][Due to bioaccumulation and biomagnification, animals higher on the food chain thus consume large quantities of the harmful chemicals absorbed by microbeads, harming these species.[3]]

✓ ✗ I have explained how microbeads enter water ecosystems from cosmetic products.[1]

✓ ✗ I have explained how microbeads act as 'vessels' for harmful chemicals.[2]

✓ ✗ I have explained how bioaccumulation means animals higher on the food chain are worst affected.[3]

Exam-style

9. **a.** [Fire resistance/softness.[1]][This property reduces risk of fire inside cars/decreases severity of collisions between car occupants and dashboards.[2]][It thus reduces harm and saves lives in car-crash situations.[3]]

✓ ✗ I have identified a property that makes the plastics used in car dashboards advantageous.[1]

✓ ✗ I have explained that this property reduces dangerousness of dashboards.[2]

✓ ✗ I have explained that safer dashboards reduce the severity of casualties from car crashes.[3]

b. [Flexibility/adhesivity/durability/inertness.[1]][This property enables total artificial hearts to replace the function of the heart for several years.[2]][It thus extends lifespan/improves quality of life/reduces urgency of heart transplants for patients with damaged hearts.[3]]

✓ ✗ I have identified a property that makes the plastics used in total artificial hearts advantageous.[1]

✓ ✗ I have explained that this property enables total artificial hearts to be a long-term replacement for the heart.[2]

✓ ✗ I have explained that total artificial hearts extend and improve lives of patients with reduced heart function.[3]

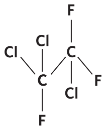
c. [High strength/lightweight/durability.[1]][This property increases the weight of loads pulleys can lift/means pulley ropes need to be replaced less often.[2]][It thus enhances efficiency of machinery used in manufacturing and construction.[3]]

✓ ✗ I have identified a property that makes the plastic used in nylon pulleys advantageous.[1]

✓ ✗ I have explained that this property enables pulleys to be stronger for longer.[2]

✓ ✗ I have explained that nylon pulley ropes improve effectiveness and longevity of machinery.[3]

10. **a.**



b. [Certain organic solvents enter the stratosphere and react with ozone.[1]][This process depletes this layer of ozone, which reflects UV rays away from Earth.[2]][As a result of this thinning, more UV rays from the Sun pass through the ozone layer and reach Earth.[3]]

✓ ✗ I have explained that certain organic solvents react with ozone in the stratosphere.[1]

✓ ✗ I have explained the role ozone plays in protecting the Earth.[2]

✓ ✗ I have linked ozone depletion to an increase in UV rays reaching Earth.[3]

Key science skills

11. **a.** [Measurement equipment or personnel must not impede the driver's ability to safely drive the car.[1]][Car occupants should not be forcibly exposed to unsafe levels of toxic gases, for example by allowing passengers to open windows if their health is threatened.[2]]

✓ ✗ I have mentioned the necessity of preserving the driver's ability to drive.[1]

✓ ✗ I have mentioned a safety consideration necessary when dealing with toxic gases.[2]

b. [This would not be a valid conclusion to draw.[1]][This is because this conclusion is based on a singular anecdote, which is insufficient to draw a conclusion.[2]][Furthermore, the driver's lethargy can reasonably be attributed to the mental strain of long car trips, so the cause is not necessarily gases in new cars.[3]]

✓ ✗ I have evaluated the validity of the student's conclusion.[1]

✓ ✗ I have noted that conclusions cannot be drawn from a singular anecdote.[2]

✓ ✗ I have explained that other factors may have caused the observed effect.[3]

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c. [This excerpt of the study relies on anecdotal reports, rather than quantitative evidence, making it difficult to compare results or draw conclusions.[1]][Effects were reported retrospectively by civilians, rather than directly measured by scientists, introducing potential bias and unreliability of 'results'.[2]][Only three types of car were studied, which is an insufficient sample size from which to draw conclusions about new cars on the whole.[3]]

✓ ✗ I have explained that anecdotal evidence is standardly inferior to quantitative evidence.[1]

✓ ✗ I have observed the potential bias and unreliability of relying on unobserved reports.[2]

✓ ✗ I have explained that the sample size of car varieties is not large enough to draw broad conclusions.[3]

FROM LESSONS 16B & 16D

Questions from multiple lessons

12. a. [Proteins/sugars/fats are a class of organic compound naturally present in many foods.[1]][Artificial colours/artificial flavours/any other organic food additives are a class of organic chemical often added to foods.[2]]

✓ ✗ I have identified one class of organic compound naturally present in many foods.[1]

✓ ✗ I have identified one class of organic compound often added to foods.[2]

b. Xylitol acts as a low-energy, 'healthier' alternative to sugar for sweetening foods.

c. [Pentane[1]][-1,2,3,4,5-[2]][pentol.[3]]

Pentane-1,2,3,4,5-pentol

✓ ✗ I have correctly counted five carbons in the longest carbon chain.[1]

✓ ✗ I have correctly numbered the five hydroxyl functional groups.[2]

✓ ✗ I have correctly named the molecule according to its five hydroxyl groups.[3]

FROM LESSON 7C

13. a. $M(C_6H_5Cl) = 6 \times M(C) + 5 \times M(H) + M(Cl)$

$M(C_6H_5Cl) = ((6 \times 12.0) + (5 \times 1.0) + 35.5)\ \text{g mol}^{-1}$

$M(C_6H_5Cl) = 112.5\ \text{g mol}^{-1}$ (1 MARK)

$n(C_6H_5Cl) = \frac{m}{M}$

$n(C_6H_5Cl) = \frac{450\ \text{g}}{112.5\ \text{g mol}^{-1}}$

$n(C_6H_5Cl) = 4.00\ \text{mol}$ (1 MARK)

b. According to the reaction equation given:

$n(C_{14}H_9Cl_5) : n(C_6H_5Cl)$

1 : 2 (1 MARK)

$n(C_{14}H_9Cl_5) = \frac{1}{2} \times n(C_6H_5Cl)$

$n(C_{14}H_9Cl_5) = \frac{1}{2} \times 4.00\ \text{mol}$

$n(C_{14}H_9Cl_5) = 2.00\ \text{mol}$ (1 MARK)

c. $m(C_{14}H_9Cl_5) = n(C_{14}H_9Cl_5) \times M(C_{14}H_9Cl_5)$

$m(C_{14}H_9Cl_5) = 2.00\ \text{mol} \times ((14 \times 12.0) + (9 \times 1.0) + (5 \times 35.5))\ \text{g mol}^{-1}$

$m(C_{14}H_9Cl_5) = 2.00\ \text{mol} \times 354.5\ \text{g mol}^{-1}$ (1 MARK)

$m(C_{14}H_9Cl_5) = 709\ \text{g}$ (1 MARK)

d. [DDT was used as a pesticide in agriculture, killing insects that ate crops and thus allowing for greater harvests.[1]] [DDT was also sprayed on humans to deter mosquitoes and other insects, temporarily eradicating the spread of diseases like malaria.[2]][However, DDT entered the water supply and due to biomagnification, predator birds like peregrine falcons were poisoned, causing eggshell thinning and decreasing populations.[3]]

✓ ✗ I have identified the benefit DDT provided to the agriculture industry.[1]

✓ ✗ I have identified the benefit DDT provided in reducing spread of disease.[2]

✓ ✗ I have identified the negative impact DDT had on animals high up on the food chain.[3]

FROM LESSON 6B

Chapter 7 review

Multiple choice

1. D. Carboxyl is the functional group, carboxylic acid is a molecule containing a carboxyl group.
2. B. Alkenes have the general formula C_nH_{2n}.
3. C. Crude oil is non-renewable, found in underground deposits, and composed of hydrocarbons.
4. D. Carboxylic acids have higher boiling points than alcohols, which in turn have higher boiling points than alkanes. Larger alkanes have higher boiling points than smaller ones.
5. D. Propane has molecular formula C_3H_8.
6. B. Carboxyl groups are not numbered.
7. C. The structural isomers of C_4H_9Cl are 1-chlorobutane, 2-chlorobutane, 1-chloro-2-methylpropane, and 2-chloro-2-methylpropane.
8. A. The largest chain is a five-carbon alkene, with a methyl group attached.
9. A. As there is only one possible form of chloroethene, numbering is not required.

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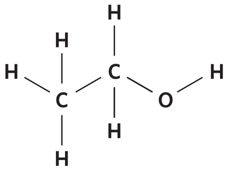
10. C. There are three chloro and three fluoro groups which can be arranged in different ways, so the prefixes 'tri-' as well as numbering are required.

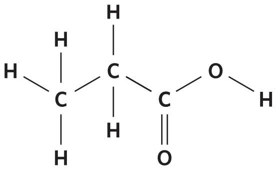
Short answer

11. a.



Ethanol

(1 MARK)



Propanoic acid

(1 MARK)

b. i. [The propanoic acid sample would have a higher boiling point.[1]][This is because carboxylic molecules form a dimer, stronger than the hydrogen bonding interactions between ethanol molecules.[2]][As a result, propanoic acid would require more energy to convert from a liquid to a gaseous state, leading to a higher boiling point.[3]]

✓ ✗ I have identified that the sample that would be expected to have a higher boiling point.[1]

✓ ✗ I have compared the intermolecular forces of both compounds.[2]

✓ ✗ I have described the impact of intermolecular forces on the boiling point.[3]

ii. [The boiling points of ethanol and propanoic acid would both be greater than ethane.[1]][This is because these are both polar molecules, while ethane is non-polar, and this leads to dipole-dipole interactions between ethanol molecules and between propanoic acid molecules, which are stronger than the dispersion forces found between ethane molecules.[2]]

✓ ✗ I have compared the boiling points of both compounds in comparison to ethane.[1]

✓ ✗ I have identified the polarity of each molecule and compared the relative strength of dipole-dipole forces and dispersion forces.[2]

12. a. $C_7H_6O_3$

b. [Salicylic acid has the molecular formula $C_7H_6O_3$, while 7-hydroxyheptanoic acid has molecular formula $C_7H_{14}O_3$.[1]][Since these molecules have different molecular formulas, they are not isomers.[2]]

✓ ✗ I have determined the molecular formula of 7-hydroxyheptanoic acid.[1]

✓ ✗ I have used the respective molecular formulas to determine isomerism.[2]

c. Skincare products or cosmetic products or exfoliating products or aspirin or medication

13. a. [Hydroxyl (top left, middle right).[1]][Halogen (fluoro) (top left, top right).[2]][Carboxyl (middle right, bottom middle).[3]]

✓ ✗ I have identified an hydroxyl group.[1]

✓ ✗ I have identified a halogen group.[2]

✓ ✗ I have identified a carboxyl group.[3]

b. [The production of synthetic chemicals begins by drilling crude oil, a mixture of hydrocarbons, from underground deposits.[1]][These hydrocarbons are then separated according to size, and hydrocarbons of an appropriate size for the chemical are extracted.[2]][These extracted hydrocarbons are reacted with other compounds to synthesise more complex organic chemicals, such as 'Oregon Green'.[3]]

✓ ✗ I have mentioned that hydrocarbons are predominantly accessed by drilling crude oil.[1]

✓ ✗ I have explained that crude oil must be separated into hydrocarbons of different sizes.[2]

✓ ✗ I have explained that these hydrocarbons must usually be reacted with other compounds.[3]

c. [Since 'Oregon Green' is sourced from crude oil, which takes millions of years to reform from dead organisms under heat and pressure, it does not replenish by natural means in a short timeframe. Hence, it is non-renewable.[1]][On the other hand, biomass replenishes naturally in a short timeframe, and is therefore renewable.[2]][Whilst non-renewable sources will become scarcer, more expensive, and eventually run out, using renewable sources for organic chemicals ensures society will have continuing access to these compounds into the future.[3]]

✓ ✗ I have explained that crude oil does not naturally replenish naturally.[1]

✓ ✗ I have explained that biomass replenishes naturally and is hence renewable.[2]

✓ ✗ I have explained that renewable sources ensure continuing access to organic compounds.[3]

14. a. i. Alkene

ii. Alkane

iii. Haloalkane

iv. Alkene

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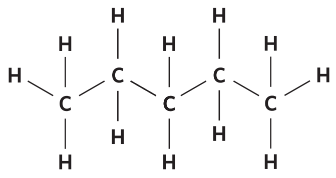
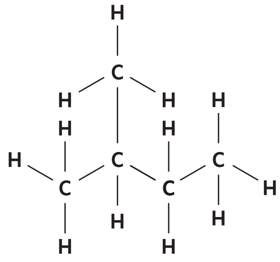
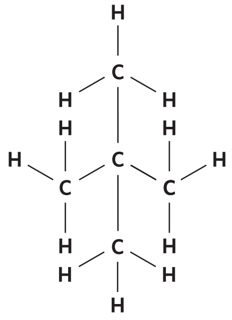
b.


^{1}

Pentane ^{2}


^{3}

Methylbutane ^{4}


^{5}

Dimethylpropane ^{6}

✓ ✗ I have correctly drawn pentane.^{1}

✓ ✗ I have correctly named pentane.^{2}

✓ ✗ I have correctly drawn methylbutane.^{3}

✓ ✗ I have correctly named methylbutane (or 2-methylbutane).^{4}

✓ ✗ I have correctly drawn dimethylpropane.^{5}

✓ ✗ I have correctly named dimethylpropane (or 2,2-dimethylpropane).^{6}

Key science skills

15. a. [The three carboxylic acids could be placed in separate, labelled beakers with a thermometer in them, then heated over a bunsen burner until they boiled.^{1}] [The temperature at which they each boiled could then be recorded and the boiling point of each carboxylic acid compared.^{2}] [If the experiment is performed without major mistakes or errors, the highest boiling point will correspond with propanoic acid, the lowest boiling point will correspond with methanoic acid, and the other compound must be ethanoic acid.^{3}]

✓ ✗ I have selected an appropriate quantitative test which could be performed on all three acids.^{1}

✓ ✗ I have mentioned that the experimental results of the three unknown acids could be compared.^{2}

✓ ✗ I have explained that these experimental results could be aligned with expected results to determine the identity of the unknown carboxylic acids.^{3}

b. Wearing gloves or lab coat or safety goggles

FROM LESSON 16B

16. a. [Many organic solvents have been banned worldwide due to their impact on the ozone layer.^{1}] [There are also legal restrictions on the sale of products containing certain solvents to children, due to the potential for abuse.^{2}]

✓ ✗ I have identified an environment-related legal factor surrounding organic solvents.^{1}

✓ ✗ I have identified a health-related legal factor surrounding organic solvents.^{2}

b. [The company should ensure the solvents used in the markers are non-toxic, as it is unethical to sell products to children that could harm their health.^{1}] [Any animal or human testing used to determine the toxicity of these solvents should also be performed according to ethical standards.^{2}]

✓ ✗ I have identified potential harm to children as a factor for ethical consideration.^{1}

✓ ✗ I have identified animal or human testing as a source of ethical discussion.^{2}

c. [The impact of production of the solvent, including oil drilling and manufacturing, should be considered.^{1}] [Additionally, the potential for water and air pollution posed by disposal of the solvent must also be taken into account.^{2}]

✓ ✗ I have identified an environmental issue with production of solvents.^{1}

✓ ✗ I have identified an environmental issue with use or disposal of solvents.^{2}

FROM LESSON 16B

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8A Polymers

Progress Questions

1. A. Monomers are the individual units of polymer chains.
2. B. Only condensation polymerisation produces water as a by-product.
3. A. Polypeptides (proteins) form naturally in living organisms.
4. A. One representation of a polymer is square brackets with *n* in subscript after it.
5. A. Polar monomers create strong permanent dipole-dipole forces between polymer chains.
6. A. Longer polymer chains have a larger surface area over which dispersion forces can act.
7. B. Branching means that polymers cannot pack tightly together, decreasing the strength of intermolecular forces.
8. B. HDPE is produced under low pressure conditions, creating linear polymers.
9. A. Linear polymers only have weak intermolecular forces which enables remoulding.
10. A. Cross-links are covalent bonds between polymer chains.

Deconstructed

11. A **12.** C

13. a. [A thermoplastic polymer should be used to make a squeezable shampoo bottle.[1]][Thermoplastics tend to be softer and less rigid than thermosets, since there are no cross-links.[2]]

✓ ✗ I have identified the type of polymer to use.[1]

✓ ✗ I have identified the relevant property to this application.[2]

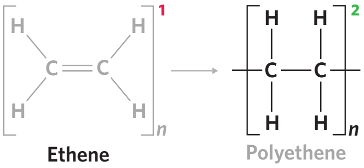
b. [A thermoset polymer should be used to make the knob of a saucepan lid.[1]][This is because thermosets are heat resistant polymers.[2]]

✓ ✗ I have identified the type of polymer to use.[1]

✓ ✗ I have identified the relevant property to this application.[2]

Exam-style

14. a.



Ethene Polyethene

✓ ✗ I have identified the name of the monomer.[1]

✓ ✗ I have drawn the structure of the polymer, using appropriate conventions.[2]

b. Addition polymerisation

c. Condensation polymerisation

d. [Similarities: Both produce a polymer **or** both involve the joining of monomers.[1]] [Differences: condensation polymerisation produces water, while addition polymerisation does not **or** condensation polymerisation involves monomers with functional groups such as carboxylic acid, while addition polymerisation involves alkene monomers.[2]]

✓ ✗ I have identified a similarity.[1]

✓ ✗ I have identified a difference.[2]

15. a. Addition polymerisation

b. [High temperature and pressure generates lots of branching in polymer molecules.[1]][Low temperature and pressure minimises branching of polymer molecules.[2]]

✓ ✗ I have identified the effect of high temperature and pressure on polyethene's structure.[1]

✓ ✗ I have identified the effect of low temperature and pressure on polyethene's structure.[2]

c. [Leo should create this polymer under low temperature and pressure.[1]][With minimal branching, the molecules can pack tightly together, increasing the strength of the intermolecular forces.[2]][This will increase the toughness of the material, making it appropriate for use as a ballistics material.[3]]

✓ ✗ I have identified the conditions under which the polymer should be made.[1]

✓ ✗ I have explained the decision with reference to intermolecular forces.[2]

✓ ✗ I have further explained the decision with reference to a physical property.[3]

d. High-density polyethene

Key science skills

16. a. The 'bendability' of the plastic

b. Qualitative data

c. [The bendability of the plastic is a subjective judgement made by the experimenter. [1]][As a result, the classification of 'bendability' will be different between different experimenters.[2]][Therefore, the results from the experiment would not be very reproducible as the results would vary depending on the experimenter.[3]]

✓ ✗ I have identified that 'bendability' is subjective.[1]

✓ ✗ I have described the impact of this factor on the results of the experiment.[2]

✓ ✗ I have linked my answer to the question.[3]

FROM LESSONS 16B & 16D

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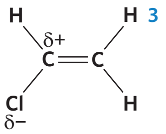
Multiple lessons

17. **a.** [Ethene is non-polar and hence has no permanent dipoles.[1]] [Chloroethene is polar.[2]]



✓ ✗ I have identified the polarity of ethene.[1]

✓ ✗ I have identified the polarity of chloroethene.[2]

✓ ✗ I have labelled the permanent dipoles of chloroethene.[3]

b. [There are permanent dipole-dipole forces and dispersion forces between polychloroethene chains but only dispersion forces between polyethene chains.[1]] [Permanent dipole-dipole forces are stronger than dispersion forces.[2]] [Thus, polychloroethene would have a higher melting point.[3]]

✓ ✗ I have identified the intermolecular forces present for each polymer.[1]

✓ ✗ I have compared the strengths of these intermolecular forces.[2]

✓ ✗ I have identified which polymer would have the highest melting point.[3]

c. [The polyethene chains may be longer than the polychloroethene chains or the polychloroethene chains may have more branching than the polyethene chains.[1]] [Increased chain length strengthens intermolecular forces between chains or increased branching weakens intermolecular forces between chains.[2]] [These stronger intermolecular forces would increase the melting point of polyethene or these weakened intermolecular forces would decrease the melting point of polychloroethene.[3]]

✓ ✗ I have identified a variable that could affect polymer melting point.[1]

✓ ✗ I have explained how that variable affects the intermolecular forces between polymer chains.[2]

✓ ✗ I have related this change in intermolecular forces to the relative melting points of the two polymers.[3]

FROM LESSONS 2A & 2B

8B Plastics in our society

Progress questions

1. A. Currently, 99% of plastics used are fossil fuel-based.

2. B. Bio-PP is a drop-in bioplastic, and so it functions identically to regular PP.

3. B. PLA is produced from fermenting starch from corn and other crops.

4. A. Thermoplastics are all recyclable in theory, though in practice, recyclability varies.

5. B. Through chemical recycling, plastics can be recycled into raw materials, which can be used to make different plastics or even fuels.

6. A. ‘Composting’ and ‘organic recycling’ are essentially synonymous.

7. B. Some fossil fuel-based plastics are compostable.

8. B. Only 9% of plastic produced has been recycled.

9. A. Renewable plastics and recycling enable materials to be reused.

Deconstructed

10. C

11. B

12. [Bio-PE is a drop-in bioplastic, meaning it can be used with existing HDPE and LDPE manufacturing equipment without change.[1]] [Additionally, Bio-PE is produced from renewable biomass, meaning humans will have continued access to this resource into the future, unlike its fossil fuel-based counterpart.[2]] [Finally, the crops grown to produce Bio-PE absorb carbon dioxide, meaning Bio-PE production can be carbon neutral or carbon negative, unlike fossil fuel-based PE production.[3]]

✓ ✗ I have explained how Bio-PE’s drop-in property is advantageous.[1]

✓ ✗ I have explained how renewability is advantageous.[2]

✓ ✗ I have explained how Bio-PE production is carbon negative.[3]

Exam-style

13. [Household plastics are collected in curbside recycling bins, before being sorted and cleaned.[1]] [The plastics are then shredded and melted down into pellets, which can be reformed to make new products like furniture.[2]]

✓ ✗ I have explained how plastics are collected and sorted.[1]

✓ ✗ I have explained how plastics are melted and reformed to make new products.[2]

14. **a.** [Due to the risk of contamination, products cannot be mechanically recycled into the original product, preventing true circularity.[1]] [Furthermore, attempts to organically recycle plastics into organic nutrients often lead to microplastics and other toxic residues being left behind, meaning plastics still remain.[2]]

✓ ✗ I have identified a barrier to the circular plastic economy.[1]

✓ ✗ I have identified another barrier to the circular plastic economy.[2]

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b. [Almost all plastic products can be mechanically recycled to produce lower-grade plastic products, as contamination is less of an issue for objects like furniture compared to plastics in contact with food.[1]] [Chemical recycling methods can also be used to convert plastics into raw materials which can then be used as a fuel, offsetting some of the energy used in these processes and completing the 'spiral'.[2]]

✓ ✗ I have explained why mechanical recycling usually produces lower-grade products.[1]

✓ ✗ I have explained that chemical recycling can convert plastics into energy.[2]

Key science skills

15. a. [The green leaf on the label suggests that the bottle is 'eco-friendly', yet not all biodegradable products are as eco-friendly as many people believe.[1]] [This is because biodegradable plastics can take many years to decompose, and may leave toxic residues like microplastics, which are in fact often more harmful to the environment than the plastic itself.[2]]

✓ ✗ I have explained that 'biodegradable' does not mean eco-friendly.[1]

✓ ✗ I have explained the environmental limitations of the term 'biodegradable'.[2]

b. [A number of plastic bottles with this label could be marked and weighed.[1]] [These could then be placed in an industrial composter and left for 180 days.[2]] [After this, the remnants of the bottles could be weighed, and if their mass is less than 10% of their starting mass, the bottles are compostable.[3]]

✓ ✗ I have suggested that a sample of bottles should be taken and weighed at the start of the experiment.[1]

✓ ✗ I have suggested that the sample of bottles should be placed in a composting environment for 180 days.[2]

✓ ✗ I have explained how the final mass of bottles can be used to determine compostability.[3]

FROM LESSON 16A

Questions from multiple lessons

16. a. Formula of propene: C_3H_6 (1 MARK)

$M(C_3H_6) = ((3 \times 12.0) + (6 \times 1.0))\ \text{g mol}^{-1}$

$M(C_3H_6) = 42.0\ \text{g mol}^{-1}$ (1 MARK)

$m(C_3H_6) = n \times M$

$m(C_3H_6) = 0.153\ \text{mol} \times 42.0\ \text{g mol}^{-1}$

$m(C_3H_6) = 6.43\ \text{g}$ (1 MARK)

b. $\%\ \text{converted} = \dfrac{m(\text{propene})}{m(\text{polypropene})} \times 100\%$

$\%\ \text{converted} = \dfrac{6.43\ \text{g}}{35\ \text{g}} \times 100\%$ (1 MARK)

$\%\ \text{converted} = 18\%$ (1 MARK)

c. [This process is not particularly efficient, as only 18% of the plastic was actually recycled, reducing its industrial appeal.[1]] [Moreover, the high temperatures required are energy-intensive, increasing the cost and environmental impact of the process.[2]]

✓ ✗ I have identified conversion efficiency as a barrier.[1]

✓ ✗ I have identified energy cost as a barrier.[2]

FROM LESSONS 6C & 7A

8C Innovations in polymers

Progress questions

1. A. Water is produced as a byproduct of condensation reactions.

2. B. Lactic acid contains both a carboxylic acid group and a hydroxyl group, so it can condensation polymerise with itself to produce PLA.

3. B. Two functional groups on each molecule are required to undergo condensation polymerisation – there is only one carboxyl group present in the carboxylic acid.

4. A. Water is a reactant in hydrolysis reactions.

5. A. Hydrolysis is the breaking down of a polymer using water while condensation is the formation of water.

6. A. Circular economies involve the recycling of resources.

7. B. Humans must transition from a linear to a circular economy to preserve Earth's natural resources.

Deconstructed

8. C

9. C

10. [The hydrolysis of renewably-sourced polymers allows for the development of biodegradable polymer products.[1]] [These products, when processed correctly following disposal, can be used in the renewal of reactants for subsequent polymer synthesis.[2]] [This allows for a circular economy to be implemented, as the product is recycled and repurposed for the synthesis of a future product.[3]]

✓ ✗ I have related hydrolysis to biodegradability.[1]

✓ ✗ I have related biodegradability to the recycling and renewal of polymer products.[2]

✓ ✗ I have related the recycling and renewal of polymer products to the implementation of a circular economy.[3]

Exam-style

11. a. Condensation

b. Hydrolysis

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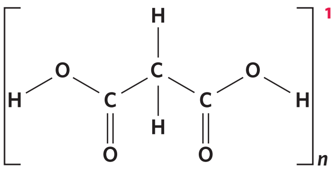
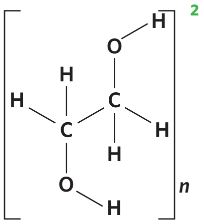
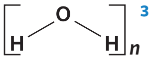
c.







✓ ✗ I have drawn one product of the hydrolysis of the polymer.[1]

✓ ✗ I have drawn a second product of the hydrolysis of the polymer.[2]

✓ ✗ I have drawn the other reactant (water) of the hydrolysis of the polymer.[3]

Key science skills

12. a. Material of coffee cup

b. Mass of the coffee cups

c. By leaving the cups in different places, Imogen has exposed them to different environments which may affect the rates of degradation or as different types of cups, they are likely to have different mass, introducing another independent variable that will affect the time of degradation.

d. [This is a systematic error.[1]][Imogen could add 6 g to all of her results.[2]]

✓ ✗ I have identified the type of error.[1]

✓ ✗ I have identified how this error could be rectified.[2]

FROM LESSONS 16A & 16D

Questions from multiple lessons

13. a. [The functional groups present in glycolic acid are a carboxyl group,[1]][and a hydroxyl group.[2]]

✓ ✗ I have identified one functional group present in glycolic acid.[1]

✓ ✗ I have identified a second functional group present in glycolic acid.[2]

b. $M(\text{glycolic acid}) = (2 \times M(\text{C})) + (4 \times M(\text{H})) + (3 \times M(\text{O}))$

$M(\text{glycolic acid}) = (2 \times 12.0 \text{ g mol}^{-1}) + (4 \times 1.0 \text{ g mol}^{-1}) + (3 \times 16.0 \text{ g mol}^{-1})$ (1 MARK)

$M(\text{glycolic acid}) = 76.0 \text{ g mol}^{-1}$ (1 MARK)

c. [As a condensation polymer, polyglycolic acid is able to undergo hydrolysis.[1]][This results in the relatively fast breakdown of the molecule.[2]][Polyglycolic acid would hence be expected to be biodegradable.[3]]

✓ ✗ I have related its nature as a condensation polymer to its ability to undergo hydrolysis.[1]

✓ ✗ I have related the ability to undergo hydrolysis to the rate of breakdown.[2]

✓ ✗ I have stated the expected biodegradability.[3]

FROM LESSONS 6C & 7C

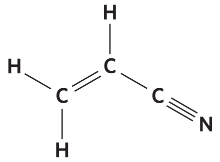
Chapter 8 review

Multiple choice

1. B. Non-polar polymers, such as Teflon, possess properties such as being non-stick, heat resistant, and low friction.
2. A. Low-density polyethene is produced at a higher pressure than high-density polyethene.
3. D. Bio-PP, PLA, Bio-PE, and PHB are all bioplastics.
4. B. Polar monomers result in strong intermolecular forces between polymer chains, thus increasing the hardness of the polymers.
5. C. There are only weak intermolecular forces present in thermoplastics.
6. A. Thermoplastics have no cross-links between chains.
7. B. Biodegradable plastics can leave toxic residues and may take many years to decompose.
8. B. Although almost all plastics can be recycled, only 10–20% of plastic produced is actually recycled.
9. C. Condensation reactions may have an alcohol as a reactant, but some may not. However, water is always a product.
10. B. The properties of condensation polymers allow them to undergo hydrolysis, which in turn makes them biodegradable.

Short answer

11. a.



b. 3

c. Permanent dipole-dipole forces

d. [They would increase the hardness/rigidity of the polymer or they would increase the melting point of the polymer or they would increase the density of the polymer.[1]] [This is because the chains are held more tightly together by intermolecular forces.[2]]

✓ ✗ I have identified a property influenced by strong intermolecular forces.[1]

✓ ✗ I have explained why intermolecular forces affect that property.[2]

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12. [Kettle handles are made from thermosetting polymers, meaning they have lots of cross-links.^{1}][This makes kettle handles very hard and heat resistant.^{2}]

✓ ✗ I have described the level of cross-linking in the item.^{1}

✓ ✗ I have related this to a desired property of the item.^{2}

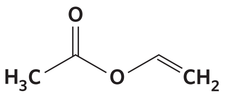
13. [The chains of thermosets are joined together by lots of covalent bonds (cross-links).^{1}][The cross-links require the same amount of energy to disrupt as the intramolecular bonds of the chains.^{2}] [Thus, upon exposure to heat, chains do not separate and mobilise. Instead, the whole polymer degrades as bonds are randomly broken.^{3}]

✓ ✗ I have identified the presence of cross-links in thermosets.^{1}

✓ ✗ I have compared the strengths of cross-links and intramolecular bonds.^{2}

✓ ✗ I have used this information to explain why thermosets decompose when heated.^{3}

14. a.

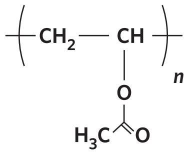


b. [The forces holding the polymer chain together are dipole-dipole^{1}][and dispersion forces.^{2}]

✓ ✗ I have identified an intermolecular force holding the polymer chains together.^{1}

✓ ✗ I have identified another intermolecular force holding the polymer chains together.^{2}

c.



15. a. [Chain length would affect the strength and hardness of these polymers^{1}][as it influences the surface area over which chains interact.^{2}][Polarity also influences the physical properties of polymers^{3}][as it creates permanent dipole-dipole attractions between chains.^{4}]

✓ ✗ I have identified a factor affecting polymer strength and hardness.^{1}

✓ ✗ I have described this factor.^{2}

✓ ✗ I have identified a factor affecting polymer strength and hardness.^{3}

✓ ✗ I have described this factor.^{4}

b. [This difference in hardness is caused by the increased branching that is characteristic of LDPE.^{1}][Increased branching pushes chains apart, reducing the strength of intermolecular forces and hence, the hardness of the polymer.^{2}]

✓ ✗ I have identified the cause of the difference in hardness between HDPE and LDPE.^{1}

✓ ✗ I have explained how this structural property influences hardness.^{2}

16. a. PLA is a bioplastic as it is produced from crop starch, a form of biomass.

b. [PLA plastic can be mechanically recycled, which involves shredding the plastic into tiny pieces which are then melted and formed into reusable pellets.^{1}][It can also be chemically recycled using a number of chemical processes that break the plastic down into either its polymer, constituent monomers, or raw chemical species.^{2}][Since PLA is compostable, it can also be organically recycled, whereby it is placed in an industrial composter and decomposed by microorganisms into water, carbon dioxide, and humus.^{3}]

✓ ✗ I have outlined the mechanical recycling of PLA.^{1}

✓ ✗ I have outlined the chemical recycling of PLA.^{2}

✓ ✗ I have outlined the organic recycling of PLA.^{3}

c. Since PLA can be produced from renewable sources and can be recycled to produce new PLA, it is considered a circular product.

17. a. Hydrolysis

b. [As these polymers are formed by condensation, they can also undergo hydrolysis.^{1}][This allows for them to be biodegradable,^{2}][and hence, when sourced from renewable sources, allows for the implementation of a circular economy.^{3}] [This is innovative, as polymers have traditionally been non-renewably sourced and have taken many years to be broken down.^{4}]

✓ ✗ I have understood that many reactions are reversible.^{1}

✓ ✗ I have used this information with regard to breaking down substances.^{2}

✓ ✗ I have applied this scenario to a circular economy.^{3}

✓ ✗ I have understood when to apply innovation to a scenario.^{4}

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Key science skills

18. a. [This is a bar graph.[1]][It represents both quantitative [2]] [and qualitative data.[3]]

✓ ✗ I have identified the type of graph.[1]

✓ ✗ I have identified the first type of data represented.[2]

✓ ✗ I have identified the second type of data represented.[3]

b. The x-axis

c. [A bar graph is an appropriate representation of the data.[1]] [Bar graphs effectively compare different categories of qualitative data (in this case, different types of polymers).[2]]

✓ ✗ I have commented on the effectiveness of the graph.[1]

✓ ✗ I have justified my answer.[2]

d. [No.[1]][Line graphs have quantitative data on both the x and y-axis.[2]]

✓ ✗ I have commented on whether a line graph could be used.[1]

✓ ✗ I have justified my answer.[2]

e. [Teflon would be the most appropriate.[1]][This is because it has the highest melting point at around 330 °C, making it the most heat resistant.[2]]

✓ ✗ I have identified the polymer that would be most appropriate for use in cookware.[1]

✓ ✗ I have justified my answer.[2]

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9A Physical and thermal properties of water

Progress questions

1. A. Gaseous water makes up 0.25% of the air.
2. B. The human body is inefficient at removing salt from water.
3. B. Most freshwater needs to be treated before consumption.
4. B. The vast majority of water on Earth is sea water with high salt content.
5. A. Sulfur is a group 16 element.
6. B. Water has a higher boiling point than other group 16 hydrides due to its ability to form hydrogen bonds.
7. B. Hydrogen bonds are the strongest intermolecular force in water.
8. A. Different substances have different heat capacities, so they experience different temperature changes given the same amount of heat energy.
9. B. Specific heat capacity is the heat energy required to raise the temperature of 1 g of a substance by 1 °C.
10. B. When performing calculations with specific heat capacity, we use the change in temperature (ΔT).
11. A. Kilojoules per kilogram is equivalent to joules per gram, and a change of one degree celsius is equal to a change of one Kelvin.
12. B. Water molecules are packed less tightly together as a solid than as a liquid.
13. A. Latent heat of vaporisation tells us the quantity of energy needed to boil a given amount of a substance.
14. A. Water has a high latent heat of vaporisation relative to other molecular substances.
15. B. Latent heat of vaporisation has a given value for water.
16. B. Much energy is needed to heat water, keeping ocean temperature fairly constant.
17. A. As much energy is absorbed by water as it boils, much energy is released from water as water vapour is released.
18. A. Many aquatic life forms cannot thrive in higher ambient temperatures.

Deconstructed

19. C
20. B
21. C
22. **Step 1:** heat water from 15 °C to 100 °C

$q = m \times c \times \Delta T$

The temperature change of water is calculated by $\Delta T = T_{final} - T_{initial}$

$\Delta T = 100\ °\text{C} - 15\ °\text{C}$

$\Delta T = 85\ °\text{C}$ (1 MARK)

The specific heat capacity of water is $4.18\ \text{J g}^{-1}\ °\text{C}^{-1}$

$q = 150\ \text{g} \times 4.18\ \text{J g}^{-1}\ °\text{C}^{-1} \times 85\ °\text{C}$

$q = 53.3 \times 10^3\ \text{J}$

$q = 53.3\ \text{kJ}$ (1 MARK)

Therefore, 53.3 kJ of heat energy is required to increase the temperature of water from 15 °C to 100 °C.

Step 2: boil water

The latent heat of vaporisation of water is $44.0\ \text{kJ mol}^{-1}$.

$q = n \times 44.0\ \text{kJ mol}^{-1}$

$$n(\text{H}_2\text{O}) = \frac{150\ \text{g}}{((2 \times 1.0) + 16.0)\ \text{g mol}^{-1}}$$

$n(\text{H}_2\text{O}) = 8.33\ \text{mol}$

Latent heat of vaporisation of water $= 44.0\ \text{kJ mol}^{-1}$

$q = 8.33\ \text{mol} \times 44.0\ \text{kJ mol}^{-1}$

$q = 367\ \text{kJ}$ (1 MARK)

Therefore, 367 kJ of heat energy is required to boil 150 g of water at 100 °C.

Total energy required = 53.3 kJ + 367 kJ

Total energy required = 4.2×10^2 kJ (1 MARK)

Exam-style

23. a. [Freshwater: ponds/ lakes/rivers/streams/aquifers.[1]] [Brackish water: swamps/estuaries/brackish seas/ lakes/swamps.[2]] [Water vapour: air.[3]]

✓ ✗ I have identified a source of freshwater.[1]

✓ ✗ I have identified a source of brackish water.[2]

✓ ✗ I have identified a source of water vapour.[3]

b. [There is truth to the sailor's claim, for water makes up more than 70% of the Earth's surface,[1]] [but despite this abundance, the vast majority of water is seawater and only a small percentage of freshwater is accessible and drinkable.[2]]

✓ ✗ I have justified the sailor's statement with reference to water's abundance on Earth.[1]

✓ ✗ I have explained that only a small portion of water on Earth is potable.[2]

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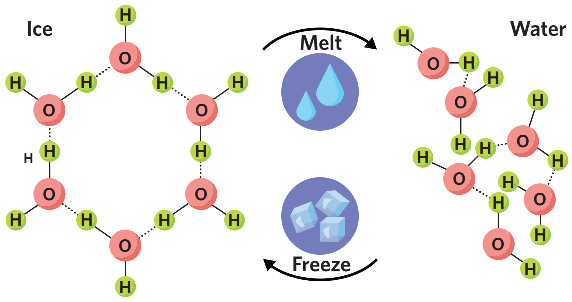
c. [To be potable, water must have a sufficiently low salt content in order for the human body to be able to process it.[1]][It is also important to ensure that water is free from contaminants.[2]]

✓ ✗ I have identified salt content as an important factor for potability.[1]

✓ ✗ I have identified contamination as an important factor for potability.[2]

d.



[When water freezes and becomes ice, the water molecules adopt a crystal-like structure with a regular hydrogen bonding arrangement.[1]][This hydrogen bonding arrangement is shown in the diagram.[2]][Since the water molecules are spaced further apart in this arrangement as opposed to in their liquid state, the volume of ice is greater for the same number of molecules and, therefore, ice has a lower density than water.[3]] [Consequently, icebergs are able to float on water as they are less dense than water.[4]]

✓ ✗ I have described the regular hydrogen bonding arrangement in ice.[1]

✓ ✗ I have included a diagram in my response.[2]

✓ ✗ I have described the density and volume of ice compared to water.[3]

✓ ✗ I have justified why ice floats on water.[4]

24. B

25. a. [Specific heat capacity is the amount of energy required to raise the temperature of 1 gram of a substance by 1 °C.[1]] [Latent heat of vaporisation is the energy that needs to be absorbed by 1 mole of a substance at its boiling point to convert it from a liquid to a gas.[2]]

✓ ✗ I have defined specific heat capacity.[1]

✓ ✗ I have defined latent heat of vaporisation.[2]

b. [Water molecules are held together by dispersion forces, permanent dipole-dipole forces and hydrogen bonds.[1]] [As a result of the strength of these intermolecular forces, predominantly the hydrogen bonds, water absorbs a large amount of energy in its bonds before its temperature increases or before it changes state, resulting in large values for the specific heat capacity and latent heat respectively.[2]]

✓ ✗ I have identified the intermolecular forces present in water.[1]

✓ ✗ I have explained how the intermolecular forces affect specific heat capacity and latent heat.[2]

c. [Water's high specific heat capacity enables it to store large amounts of energy in the form of heat before its temperature increases.[1]][This property of water means that excess heat can be removed more efficiently because more energy is transferred to the water molecules per degree of temperature increase.[2]]

✓ ✗ I have identified the property that makes water suitable to remove excess heat.[1]

✓ ✗ I have described how specific heat capacity relates to energy absorbance.[2]

26. a. [Ethanol has a lower specific heat capacity than water. As a result, less energy is needed to heat a volume of ethanol to a certain temperature than to heat the same volume of water to this temperature.[1]][Accordingly, ethanol will likely take less time to heat up before cooking and less energy will be used in the cooking process.[2]]

✓ ✗ I have explained a key thermal property of ethanol in contrast with water.[1]

✓ ✗ I have linked this property to the comparative suitability of cooking food in ethanol.[2]

b. [Water has a higher boiling point than ethanol. As a result, water would still be a liquid above 78.4 °C, whereas ethanol would have evaporated.[1]][This means that food can be cooked at higher temperatures in water (up to 100 °C) than in ethanol, without the water being boiled away.[2]]

✓ ✗ I have explained a key thermal property of water in contrast with ethanol.[1]

✓ ✗ I have linked this property to the comparative suitability of cooking food in water.[2]

c. Ethanol has a lower boiling point, lower specific heat capacity, and thus weaker intermolecular forces than those of water.

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27. $m(H_2O) = 250\ \text{mL} \times 0.997\ \text{g mL}^{-1}$

$m(H_2O) = 249\ \text{g}$

Step 1: heat water from 25 °C to 100 °C

$\Delta T = 100\ °\text{C} - 25\ °\text{C}$

$\Delta T = 75\ °\text{C}$ (1 MARK)

$q = m \times c \times \Delta T$

$q = 249\ \text{g} \times 4.18\ \text{J g}^{-1}\ °\text{C}^{-1} \times 75\ °\text{C}$

$q = 78\ 140\ \text{J}$

$q = 78.14\ \text{kJ}$ (1 MARK)

Step 2: boil water

$n(H_2O) = \frac{m}{M}$

$n(H_2O) = \frac{249\ \text{g}}{((2 \times 1.0) + 16.0)\ \text{g mol}^{-1}}$

$n(H_2O) = 13.85\ \text{mol}$

$q = n \times 44.0\ \text{kJ mol}^{-1}$

$q = 13.85\ \text{mol} \times 44.0\ \text{kJ mol}^{-1}$

$q = 609.3\ \text{kJ}$ (1 MARK)

Total: $78.14\ \text{kJ} + 609.3\ \text{kJ} = 687\ \text{kJ}$ (1 MARK)

Key science skills

28. a. Quantitative

b. Not all energy from combustion will be used to heat the water; some will be lost to the surroundings.

c. [Add a lid on the beaker to minimise heat loss or move the beaker closer to the flame. [1]][Surround the setup with insulating material e.g. polystyrene or use a known standard to determine heat loss, and factor this in.[2]]

✓ ✗ I have suggested one improvement.[1]

✓ ✗ I have suggested a second improvement.[2]

d. Repeating experiments enables us to assess the precision of results and minimise the impact of random error through averaging.

e. Initial temperature

f. [In trial 5, water's temperature appeared to stop increasing upon reaching 100 °C – the boiling point of water.[1]][Due to water's latent heat of vaporisation, much energy was being used to convert liquid water to water vapour, rather than to increase the water's temperature.[2]][As a result, a much lower ΔT value was obtained for this trial.[3]]

✓ ✗ I have identified the correlation between water's boiling point and the outlying result.[1]

✓ ✗ I have explained how latent heat of vaporisation causes energy to be absorbed without temperature increase.[2]

✓ ✗ I have linked this to the resulting ΔT value for this trial.[3]

FROM LESSONS 16B, 16C, 16D & 16E

Questions from multiple lessons

29. a. Group 16 hydrides consist of a group 16 element covalently bonded to hydrogen.

b. [Molecule A is water, which has the formula H_2O.[1]][Water has an exceptionally high boiling point because boiling point is a reflection of intermolecular force strength, and water is the only group 16 hydride able to form strong hydrogen bonds with itself.[2]][This is because of the highly polar O-H bonds in water, which – combined with the lone electron pairs on the oxygen atom – enable these very strong hydrogen bonds to form between water molecules.[3]]

✓ ✗ I have identified the name and formula of molecule A.[1]

✓ ✗ I have identified the intermolecular force responsible for molecule A having a high boiling point.[2]

✓ ✗ I have described how this intermolecular force arises.[3]

c. [The strength of intermolecular forces determines the boiling point of a substance.[1]][Therefore, boiling points increase from Substance B to E as the strength of dispersion forces increases.[2]][The strength of intermolecular dispersion forces increases because as one goes down a group, more electrons are added, which leads to stronger, and more frequent dipole-dipole interactions.[3]]

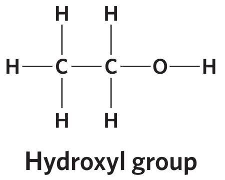
✓ ✗ I have explained how boiling points are determined by the strength of intermolecular forces present.[1]

✓ ✗ I have identified the intermolecular force responsible for this increase.[2]

✓ ✗ I have explained how the strength of intermolecular forces increases down a group.[3]

FROM LESSON 2B

30. a.



b. [The presence of the polar hydroxyl group in ethanol molecules means that dispersion forces, permanent dipole-dipole forces and hydrogen bonds form between water molecules.[1]][These strong hydrogen bonds can absorb large amounts of energy, so ethanol molecules have a high specific heat capacity.[2]][However, the specific heat capacity of ethanol is less than that of water because of the presence of the non-polar carbon chain which diminishes the effect of the polar hydroxyl group.[3]]

✓ ✗ I have identified the intermolecular forces present in ethanol.[1]

✓ ✗ I have explained ethanol's relatively high specific heat capacity.[2]

✓ ✗ I have compared the specific heat capacity of ethanol to water.[3]

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c. $m(H_2O) = 300\text{ mL} \times 0.997\text{ g mL}^{-1}$

$m(H_2O) = 299\text{ g}$ (1 MARK)

Energy released by ethanol $= 3.0\text{ g} \times 29.6\text{ kJ g}^{-1}$

Energy released by ethanol $= 88.8\text{ kJ}$

Energy released by ethanol $= 88\,800\text{ J}$ (1 MARK)

$q = m \times c \times \Delta T$

$88\,800\text{ J} = 299\text{ g} \times 4.18\text{ J g}^{-1}\,°\text{C}^{-1} \times \Delta T$

$\Delta T = \dfrac{88\,800\text{ J}}{299\text{ g} \times 4.18\text{ J g}^{-1}\,°\text{C}^{-1}}$

$\Delta T = 71\,°\text{C}$ (1 MARK)

$\Delta T = T_{final} - T_{initial}$

$71\,°\text{C} = T_{final} - 20\,°\text{C}$

$T_{final} = 91\,°\text{C}$

Therefore, the final temperature of the water is 91 °C. (1 MARK)

FROM LESSON 7C

Chapter 9 review

Multiple choice

1. D. Water makes up ~0.25% of the air everywhere in the world.
2. C. All group 16 hydrides are v-shaped, or bent.
3. D. Water's boiling point of 100 °C is high relative to other group 16 hydrides.
4. C. H_2O is the only group 16 hydride which has hydrogen bonds between adjacent molecules, an intermolecular force which is much stronger than the dipole-dipole interactions and dispersion forces between other group 16 hydride molecules.
5. C. Water has a high specific heat capacity relative to many other liquids, but not the highest of all.
6. B. Substances with identical mass but lower specific heat capacity heat up more with the same energy input.
7. D. Using $q = mc\Delta T$,

   $\Delta T = \dfrac{q}{m \times c}$

   $\Delta T = \dfrac{4300\text{ J}}{100\text{ g} \times 3.97\text{ J g}^{-1}\,°\text{C}^{-1}}$

   $\Delta T = 10.8\,°\text{C}$

   $T_{\text{final}} = T_{\text{initial}} + \Delta T$

   $T_{\text{final}} = 25\,°\text{C} + 10.8\,°\text{C}$

   $T_{\text{final}} = 35.8\,°\text{C}$
8. B. Pure water's boiling point is 100 °C, but other substances, like milk, have different boiling points.
9. A. The amount of water is $\dfrac{500\text{ kJ}}{44.0\text{ kJ mol}^{-1}} = 11.4$ mol,

   which converts to $11.4\text{ mol} \times (16.0 + 2 \times 1.0)\text{ g mol}^{-1} = 205\text{ g}$.
10. C. To heat water to 100 °C: $q = mc\Delta T$

    $q = 100\text{ g} \times 4.18\text{ J g}^{-1}\,°\text{C}^{-1} \times (100 - 20)\,°\text{C}$

    $q = 33\,400\text{ J}$ or 33.4 kJ

    To evaporate water at 100 °C: $q = 44.0\text{ kJ mol}^{-1} \times n(H_2O)$

    $q = 44.0\text{ kJ mol}^{-1} \times \left(\dfrac{100\text{ g}}{(16.0 + (2 \times 1.0))\text{ g mol}^{-1}}\right)$

    $q = 244\text{ kJ}$

    Total energy needed: $(244 + 33.4)\text{ kJ} = 277\text{ kJ}$

Short answer

11. a. [Water molecules contain highly polar covalent bonds between an oxygen atom and two hydrogen atoms.[1]][The two lone pairs of electrons on the oxygen atom form two sites of negative charge, while the two hydrogen atoms form two sites of partial positive charge.[2]][Hydrogen bonds form between oppositely charged sites on different water molecules, hence one water molecule can form hydrogen bonds with 4 other water molecules.[3]]

✓ ✗ I have recognised that water molecules contain two polar O-H bonds.[1]

✓ ✗ I have explained that lone pairs and partially positive hydrogens form a total of 4 charged sites per molecule.[2]

✓ ✗ I have explained that a total of 4 charged sites allows for 4 hydrogen bonds from each molecule.[3]

b. [H_2S, H_2Se, H_2Te, H_2Po,[1]][H_2O.[2]][Water contradicts the expected trend of increasing boiling point down the group since it has strong hydrogen bonds between adjacent molecules while the other group 16 hydrides do not.[3]]

✓ ✗ I have correctly ordered the group 16 hydrides (excluding H_2O) down the group.[1]

✓ ✗ I have correctly identified H_2O as having the highest boiling point of the group 16 hydrides.[2]

✓ ✗ I have identified the presence of hydrogen bonds as the reason for H_2O's abnormally high boiling point.[3]

12. a. High latent heat of vaporisation

b. Latent heat of vaporisation of water $= 44.0\text{ kJ mol}^{-1}$

$100\text{ kJ} = n(H_2O) \times 44.0\text{ kJ mol}^{-1}$

$n(H_2O) = \dfrac{100\text{ kJ}}{44.0\text{ kJ mol}^{-1}}$

$n(H_2O) = 2.27\text{ mol}$ (1 MARK)

$m(H_2O) = 2.27\text{ mol} \times (16.0 + (2 \times 1.0))\text{ g mol}^{-1}$

$m(H_2O) = 40.9\text{ g}$ (1 MARK)

Density of water $= 0.997\text{ g mL}^{-1}$

$V(H_2O) = \dfrac{40.9\text{ g}}{0.997\text{ g mL}^{-1}}$

$V(H_2O) = 41.0\text{ mL}$ (1 MARK)

Therefore, 41.0 mL of water must be perspired.

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c. [Ice cubes float on water because ice is less dense than liquid water.[1]][This is because the lattice structure of ice involves hydrogen bonds that hold water molecules at a further distance from each other than in their random arrangement as a liquid, which is tightly packed.[2]] [As a result, there is more empty space between water molecules as a solid than as a liquid, leading to ice's lower density.[3]]

✓ ✗ I have identified the density property of water that enables ice to float on liquid water.[1]

✓ ✗ I have compared the effect of intermolecular forces on the lattice structure of ice and the random arrangement of molecules in liquid water.[2]

✓ ✗ I have explained the link between empty space and density of the two states.[3]

d. [Liquid water is denser than whiskey.[1]][This is evident because ice sinks in whiskey, therefore whiskey is less dense than ice and thus less dense than liquid water.[2]]

✓ ✗ I have identified liquid water as the denser liquid.[1]

✓ ✗ I have justified this by comparing ice's behaviour in the two liquids.[2]

13. a. [The property that water is less dense as a solid than a liquid means that icebergs float on top of the ocean.[1]] [By reflecting solar energy away from the ocean, icebergs prevent the ocean from absorbing as much heat, keeping the surface temperature down.[2]]

✓ ✗ I have linked ice's lower density as a solid than as a liquid to icebergs' tendency to float.[1]

✓ ✗ I have explained how ice's reflective property contributes to ocean temperature regulation.[2]

b. [Water has a high latent heat of vaporisation because strong hydrogen bonds form between water molecules.[1]][These strong intermolecular forces can absorb a large quantity of energy before being overcome and forming a gas.[2]]

✓ ✗ I have identified hydrogen bonds as the key to water's high latent heat of vaporisation.[1]

✓ ✗ I have linked hydrogen bonds' strength to their ability to absorb a large quantity of heat energy.[2]

c. [Water's high latent heat of vaporisation means that a large amount of energy is removed from the ocean when water evaporates and rises from the ocean surface as vapour.[1]] [This reduces the amount of heat energy stored by the water, preventing the ocean from heating up significantly and thus maintaining its temperature.[2]][As a result, water temperatures remain within a narrow range, vital for the angelfish, which can only survive a narrow range of temperatures.[3]]

✓ ✗ I have linked water's high latent heat of vaporisation to the large amount of energy removed in evaporation.[1]

✓ ✗ I have explained that energy removed from water reduces its temperature.[2]

✓ ✗ I have explained the importance of regulated temperature to the angelfish's survival.[3]

14. a. $\Delta T = 120\ °\text{C} - 20\ °\text{C}$

$\Delta T = 100\ °\text{C}$ (1 MARK)

$\Delta q = 1000\ \text{J} - 0\ \text{J}$

$q = 1000\ \text{J}$

$q = m \times c \times \Delta T$

$1000\ \text{J} = 4.90\ \text{g} \times c \times 100\ °\text{C}$ (1 MARK)

$$c = \frac{1000\ \text{J}}{4.90\ \text{g} \times 100\ °\text{C}}$$

$c = 2.04\ \text{J g}^{-1}\ °\text{C}^{-1}$ (1 MARK)

Therefore, the unknown substance is ethanoic acid. (1 MARK)

b. $q_{\text{vaporisation}} = 3000\ \text{J} - 1000\ \text{J}$

$q_{\text{vaporisation}} = 2000\ \text{J}$ or $2.00\ \text{kJ}$ (1 MARK)

$M(\text{ethanoic acid}) = 60.0\ \text{g mol}^{-1}$ (from table)

$$n(\text{ethanoic acid}) = \frac{m(\text{ethanoic acid})}{M(\text{ethanoic acid})}$$

$$n(\text{ethanoic acid}) = \frac{4.90\ \text{g}}{60.0\ \text{g mol}^{-1}}$$

$n(\text{ethanoic acid}) = 0.0817\ \text{mol}$ (1 MARK)

$q_{\text{vaporisation}} =$ latent heat of vaporisation of ethanoic acid $\times\ n(\text{ethanoic acid})$

$2.00\ \text{kJ} =$ latent heat of vaporisation of ethanoic acid $\times\ 0.0817\ \text{mol}$ (1 MARK)

$$\text{latent heat of vaporisation of ethanoic acid} = \frac{2.00\ \text{kJ}}{0.0817\ \text{mol}}$$

latent heat of vaporisation of ethanoic acid $= 24.5\ \text{kJ mol}^{-1}$ (1 MARK)

Key science skills

15. a. [The student's conclusion is flawed,[1]][because it is not possible to make a valid conclusion about global ocean temperatures based on measurements from only Anglesea Beach.[2]]

✓ ✗ I have assessed the conclusion as flawed.[1]

✓ ✗ I have explained that the scope of the conclusion is too broad for the experiment.[2]

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b. [Research and analyse temperature measurements from other oceans around the world.[1]][Take multiple measurements throughout each year and average them or research temperature measurements at Anglesea Beach from previous years to broaden timeframe.[2]]

✓ ✗ I have identified the need for measurements from many worldwide locations to make a global conclusion.[1]

✓ ✗ I have identified the need for more measurements from this location to see a trend.[2]

c. [The major issue with the graph is that the vertical axis scale is too large to distinguish between the different temperature values, making the graph appear flat.[1]][To improve this, the range of vertical axis values could be narrowed (for example, to a range of 16 – 18 °C) to more clearly show the trend in data.[2]]

✓ ✗ I have identified the vertical axis scale as the main issue with the graph.[1]

✓ ✗ I have provided a valid suggestion to more clearly display the trend in the data.[2]

d. [It is unreasonable to say that a temperature change of 0.4 °C in 6 years is negligible.[1]][A change of this magnitude would in fact have devastating consequences for aquatic life forms, which rely on ocean temperatures remaining relatively constant.[2]][This is because many such creatures, like the clownfish, can only thrive in a narrow habitat temperature range.[3]]

✓ ✗ I have evaluated the claim.[1]

✓ ✗ I have justified my evaluation by noting that a small temperature change can significantly affect aquatic life.[2]

✓ ✗ I have explained that many aquatic species can only survive a narrow range of temperatures.[3]

e. ... “anyone will tell you that the water does not feel any warmer than it used to” is a non-scientific idea.

FROM LESSONS 16C, 16D & 16E

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10A Acids and bases

Progress questions

1. B. According to the Brønsted-Lowry theory of acids and bases, an acid-base reaction is defined as a reaction involving the transfer of a proton, or H^+ ion.

2. A. During an acid-base reaction, a H^+ ion is transferred from a donating acid to an accepting base.

3. B. HCl(aq) donates a H^+ ion to form Cl^-(aq), therefore acting as an acid. H_2O(l) accepts a H^+ ion to form H_3O^+(aq), therefore acting as a base.

4. B. HF has one proton available to donate, making it a monoprotic acid. HF cannot also accept a H^+ ion and so is not amphiprotic as well.

5. B. $H_2PO_4^-$ has two available H^+ ions to donate and so is, therefore, a diprotic acid. It can also accept a H^+ ion to form H_3PO_4, making it amphiprotic as well.

6. A. If an acid is polyprotic, it is capable of donating more than one H^+ ion.

7. B. CH_3COOH, H_2CO_3 and NH_4^+ (ammonium) are common weak acids.

8. A. HCl, as a strong acid, will ionise (almost) completely in water, readily donates its proton and breaks up into H^+(aq) and Cl^-(aq). As a result, almost zero molecules of HCl will remain in the solution.

9. D. The given solution has only a few molecules present, meaning that it is most likely a dilute solution. Complete ionisation is characteristic of strong acids.

Deconstructed

10. A

11. B

12. **a.** $HNC(NH_2)_2(aq) + H_2O(l) \rightarrow C(NH_2)_3^+(aq) + OH^-(aq)$

b. [The species acting as an acid is H_2O(l).[1]][This is because H_2O(l) donates a H^+ ion to form OH^-(aq), which is consistent with the Brønsted-Lowry theory of acids and bases that classifies an acid as a proton donor.[2]]

✓ ✗ I have identified the chemical species that is acting as an acid.[1]

✓ ✗ I have justified my answer with reference to the Brønsted-Lowry theory of acids and bases.[2]

c. [H_2O(l) is classified as amphiprotic.[1]][As H_2O(l) is capable of both accepting a H^+ ion to form H_3O^+(aq) and donating a H^+ ion to form OH^-(aq), it is classified as an amphiprotic species.[2]]

✓ ✗ I have stated how the chemical species is classified.[1]

✓ ✗ I have explained my answer with reference to the relevant theory.[2]

Exam-style

13. D. H_2SO_4 is a strong acid, however 0.03 M is a dilute concentration.

14. C. The article suggests a molecule with the formula $H_9O_4^+$ best represents the true structure of a solvated proton. In the structure, the oxygen atoms would orient themselves towards a central H^+ ion.

15. D. '…the hydronium ion, which remains the simplest and most practical way to indicate the covalent attachment of H^+ ions to water.'

16. **a.** Monoprotic acid

b. $HClO_4(aq) + NH_3(aq) \rightarrow ClO_4^-(aq) + NH_4^+(aq)$

17. C. Of the 20 mol of citric acid in the vessel, only a small portion of the molecules will actually deprotonate, and only a small portion of those molecules will deprotonate more than once.

18. B. The net negative charge of a compound that has deprotonated pulls the remaining protons in.

Key science skills

19. **a.** **i.** [Independent variable: Identity of the acid (HF/HCl).[1]] [Dependent variable: Final concentration of H_3O^+ in the solution.[2]]

✓ ✗ I have identified the independent variable.[1]

✓ ✗ I have identified the dependent variable.[2]

ii. [Consistent environmental temperature or make-up of H_2O solution (no contaminants or foreign species).[1]][Time for which the pH probe is recording or volume of H_2O used.[2]] [Concentration of each acid (same number of mol of each species in the solution) or proticity of each acid.[3]]

✓ ✗ I have identified one variable.[1]

✓ ✗ I have identified a second variable.[2]

✓ ✗ I have identified third variable.[3]

b. **i.** $n(\text{HF}) = \frac{m}{M}$

$$n(\text{HF}) = \frac{5.00 \text{ g}}{(19.0 + 1.0) \text{ g mol}^{-1}}$$

$n(\text{HF}) = 0.25$ mol

$n(\text{HCl}) = \frac{m}{M}$

$$n(\text{HCl}) = \frac{9.125 \text{ g}}{(35.5 + 1.0) \text{ g mol}^{-1}}$$

$n(\text{HCl}) = 0.25$ mol

10A ANSWERS

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[Since this experiment is only investigating the strength of each acid, it is important that as many other variables as possible are controlled for, including concentration.^{1}] [Therefore, the two masses of HF(s) and HCl(s) are different, but the number of mol is equal (0.250 mol), leading to solutions of equal concentration (1.00 M).^{2}]

- ✓ ✗ I have included a calculation in my answer.
- ✓ ✗ I have stated the importance of controlling for concentration.^{1}
- ✓ ✗ I have justified my answer with reference to my calculations.^{2}

ii. [Repeating an experiment reduces the effects of random fluctuations in experimental results (variance).^{1}] [Averaging multiple trials leads to a final result that is closer to the true value (more accurate) as long as there are no systematic errors present in the method.^{2}]

- ✓ ✗ I have described how repetition reduces the effect of variance in results.^{1}
- ✓ ✗ I have described the impact on the accuracy of results obtained.^{2}

c. [Caoimhe has collected data on the concentration of H_3O^+(aq) ions present in each acid's solution over time. The point at which the curve plateaus (is constant) will give the final concentration of H_3O^+(aq) ions in the solution.^{1}] [Since both HF and HCl are monoprotic acids, this final concentration of H_3O^+(aq) can be used to determine what percentage of HCl or HF ionised into H^+(aq) (H_3O^+(aq)) ions, which is an indicator of each acid's strength.^{2}] [The final concentrations of H_3O^+(aq) were 0.025 M (HF) and 1.0 M (HCl), making hydrochloric acid significantly stronger than hydrofluoric acid given that they both started at the same concentration (1.00 M).^{3}]

- ✓ ✗ I have described how to analyse the data collected.^{1}
- ✓ ✗ I have explained how the data collected can be used to determine each acid's strength.^{2}
- ✓ ✗ I have identified which acid is stronger with justification from the data collected.^{3}

d. [Acid-resistant personal protective equipment (gloves, protective eyewear, aprons) should be worn.^{1}] [The HF beaker should be made of a HF-resistant material other than glass.^{2}]

- ✓ ✗ I have identified a safety issue about personal protection.^{1}
- ✓ ✗ I have identified a safety issue about the materials used.^{2}

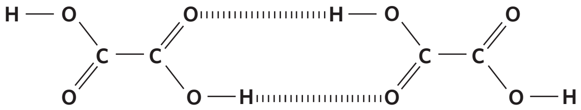
e. [Unlike HF and HCl, phosphoric acid is a triprotic acid and so will deprotonate in three stages, each of which will give off a proton to form H_3O^+(aq) ions.^{1}] [However, the current experimental technique only measures the total hydronium concentration, not from where those ions have come.^{2}] [This means that a H_3PO_4 molecule that has donated two or three protons will contribute twice or three times as much respectively as one that has only lost one proton. This will skew the data and give an incorrect estimation of the acid's strength.^{3}] [Since there is an extra variable here (the proticity of the acid) affecting the results (dependent variable), the experiment is no longer considered valid.^{4}]

- ✓ ✗ I have identified the key difference between H_3PO_4, HF and HCl.^{1}
- ✓ ✗ I have identified the issue with the experimental technique.^{2}
- ✓ ✗ I have explained why this key difference will result in inaccurate data being collected.^{3}
- ✓ ✗ I have explained how this will affect the validity of the experiment.^{4}

FROM LESSONS 16A, 16C & 16D

Questions from multiple lessons

20. a.



b. i. The table of values given does not include the presence of H_2O(l), which will be a significant proportion of the mass of the solution.

ii. [Since the compounds $(COOH)_2$, $(COOH)(COO)^-$ and $(COO)_2{}^{2-}$ are all very similar in molar mass, their relative % composition by mass in the solution gives an indication of their concentrations.^{1}] [The most highly concentrated species is $(COOH)_2$ at 25.7%, with the deprotonated forms of oxalic acid only occupying 4.30% and 0.912% of the solution's mass.^{2}] [This means that only a small proportion of oxalic acid is deprotonated at all and as a result it can be classified as a weak acid.^{3}]

- ✓ ✗ I have described how % composition by mass can indicate concentration.^{1}
- ✓ ✗ I have made reference to the data present in the table.^{2}
- ✓ ✗ I have used these data to justify my answer with reference to appropriate theoretical concepts.^{3}

c. i. $N((COOH)_2) = 11.5 \text{ mol} \times (6.02 \times 10^{23} \text{ mol}^{-1})$

$N((COOH)_2) = 6.92 \times 10^{24}$ molecules

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ii. $n(\text{OH}^-) : n((\text{COOH})_2)$

$2 : 1$

$n(\text{OH}^-) = \frac{2}{1} \times n((\text{COOH})_2)$

$n(\text{OH}^-) = 2 \times 11.5$ mol

$n(\text{OH}^-) = 23.0$ mol (1 MARK)

$m(\text{OH}^-) = n \times M$

$m(\text{OH}^-) = 23.0 \text{ mol} \times (16.0 + 1.0) \text{ g mol}^{-1}$

$m(\text{OH}^-) = 391$ g (1 MARK)

d. Diprotic acid

e. (1) $(\text{COOH})(\text{COO})^-$(aq)

(2) H_2O(l)

(3) OH^-(aq)

FROM LESSONS 2B & 6B

10B Reactions of acids and bases

Progress questions

1. B. Spectator ions are chemical species that are present as ions on both sides of a chemical equation and so they are not included in the ionic equation.

2. B. Water molecules are covalent compounds and therefore must be included in ionic equations.

3. A. Ionic salts contain a positively charged cation (Na^+) in an ionic bond with a negatively charged anion (Cl^-). CH_4 is a covalent substance and HCl is an acid.

4. B. The general equation for this reaction is: Acid + Metal hydroxide → Ionic salt + H_2O(l).

5. A. Metal carbonates are compounds of a metal cation and the carbonate ion, CO_3^{2-}. When reacting with acids, they will produce an ionic salt, carbon dioxide gas and liquid water.

6. B. In the reaction between acids and soluble metal hydroxides, only the H^+ and OH^- ions are reacting together.

7. B. Bases neutralise acids when added in excess.

8. A. All antacids can have side-effects.

Deconstructed

9. B

10. D

11. **a.** I^-(aq) and K^+(aq)

b. $2H^+(aq) + CO_3^{2-}(aq) \rightarrow CO_2(g) + H_2O(l)$

Exam-style

12. D. Both Fe and Al are present in different states on the left and right hand sides of the equation.

13. C. ClO_3^- ions would not change state in the reaction, and would hence be omitted from the ionic equation.

14. **a.** $CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$

b. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

c. [CH_3COO^-(aq) and Na^+(aq) are present as ions throughout the whole reaction and so are not included in the ionic equation as they are spectator ions.[1]][The H^+(aq) ion from ethanoic acid and the OH^-(aq) ion from sodium hydroxide combine to form a water molecule, H_2O(l), and so are included in the ionic equation.[2]]

✓ ✗ I have explained my reasoning for the omission of ions in the ionic equation.[1]

✓ ✗ I have explained my reasoning for the inclusion of ions in the ionic equation.[2]

d. Neutralisation reaction

15. **a.** [Word: magnesium hydroxide + hydrochloric acid → magnesium chloride + water.[1]][Symbol: $Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O$.[2]]

✓ ✗ I have written a complete word equation that contains no element symbols.[1]

✓ ✗ I have written and balanced the symbol equation.[2]

b. [Carbon dioxide.[1]][Burping or flatulence.[2]]

✓ ✗ I have identified the gas that is produced.[1]

✓ ✗ I have named a specific side-effect.[2]

c. [Antacids are weak bases.[1]][Calcium and sodium hydroxide are both strong bases (alkalis) and would be very corrosive to the body, making them ineffective as a treatment for indigestion.[2]]

✓ ✗ I have identified the nature of antacids.[1]

✓ ✗ I have given a reason as to why they are dangerous.[2]

Key science skills

16. **a.** [When $MgCO_3$(s) reacts with the nitric acid, CO_2(g) will be produced.[1]][This CO_2(g) will rise out of the test tube with $MgCO_3$(s) and be deposited into the test tube with $Ca(OH)_2$(aq) (limewater), causing the limewater to become a milky/cloudy colour.[2]][The only test tube that has caused this reaction with the limewater is test tube 1, which must, therefore, be the test tube containing $MgCO_3$(s).[3]]

✓ ✗ I have identified the key product from the reaction between $MgCO_3$(s) and nitric acid.[1]

✓ ✗ I have explained how this product will affect the results of the investigation.[2]

✓ ✗ I have linked my answer to the question.[3]

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b. [The data collected is related to the presence/absence of an observed change (fizzing/no fizzing and milky limewater/clear solution), not numerical descriptions of these data (e.g. number of bubbles evolved per second or light absorbance of the limewater test tube).[1]][Therefore, the data collected is qualitative.[2]]

✓ ✗ I have explained the nature of the data collected.[1]

✓ ✗ I have classified the data as qualitative or quantitative.[2]

FROM LESSON 16B

17. [Fatima is not able to make a valid conclusion.[1]][The mass is not known of the active ingredient in each of the tablets or teaspoons of liquid.[2]]

✓ ✗ I have decided whether the conclusion is valid or not.[1]

✓ ✗ I have explained my choice.[2]

FROM LESSON 16D

Questions from multiple lessons

18. a. $H_2SO_4(aq) + ZnCO_3(s) \rightarrow ZnSO_4(aq) + CO_2(g) + H_2O(l)$

b. $2H^+(aq) + ZnCO_3(s) \rightarrow Zn^{2+}(aq) + CO_2(g) + H_2O(l)$

c. [In this reaction, carbon dioxide gas is produced from the zinc carbonate.[1]][This gas will leave the solution and spread into the atmosphere, decreasing the volume (water level) of the solution in the beaker.[2]]

✓ ✗ I have identified the gas that is evolved in this reaction.[1]

✓ ✗ I have explained how this results in the observed change to the water level.[2]

d. The difference in masses will be the mass of $CO_2(g)$ that is produced.

$H_2SO_4(aq) + ZnCO_3(s) \rightarrow ZnSO_4(aq) + CO_2(g) + H_2O(l)$

Once all of the $ZnCO_3(s)$ has reacted, no more $CO_2(g)$ can be produced.

$n(ZnCO_3(s)) = \frac{m}{M}$

$n(ZnCO_3(s)) = \frac{40.3 \text{ g}}{(65.4 + 12.0 + (3 \times 16.0)) \text{ g mol}^{-1}}$

$n(ZnCO_3(s)) = 0.3214 \text{ mol}$ (1 MARK)

$n(CO_2(g)) : n(ZnCO_3(s))$

$1 : 1$

$n(CO_2(g)) = 0.3214 \text{ mol}$

$m(CO_2) = n \times M$

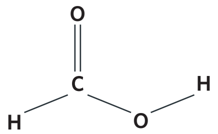
$m(CO_2) = 0.3214 \text{ mol} \times (2 \times 16.0 + 12.0) \text{ g mol}^{-1}$

$m(CO_2) = 14.14 \text{ g}$

Therefore, the change in mass would be a decrease of 14.1 g. (1 MARK)

FROM LESSON 4B

19. a.



b. Carboxyl

c. i. $2HCOOH(aq) + Sn(s) \rightarrow (HCOO)_2Sn(aq) + H_2(g)$

ii. $2H^+(aq) + Sn(s) \rightarrow Sn^{2+}(aq) + H_2(g)$

d. $2HCOOH(aq) + Sn(s) \rightarrow (HCOO)_2Sn(aq) + H_2(g)$

$n(Sn(s)) = \frac{m}{M}$

$n(Sn(s)) = \frac{40.0 \text{ g}}{118.7 \text{ g mol}^{-1}}$

$n(Sn(s)) = 0.3369 \text{ mol}$ (1 MARK)

$n(Sn(s)) : n(HCOOH(aq))$

$1 : 2$

$n(HCOOH(aq)) = 2 \times n(Sn(s))$

$n(HCOOH(aq)) = 0.6739 \text{ mol}$ (1 MARK)

$n(HCOOH(aq)) = n \times M$

$n(HCOOH(aq)) = 0.6739 \text{ mol} \times (2 \times 1.0 + 12.0 + 2 \times 16.0) \text{ g mol}^{-1}$

$n(HCOOH(aq)) = 31.00 \text{ g}$ (1 MARK)

Number of ants $= \frac{31.00 \text{ g}}{(0.600 \times 10^{-6}) \text{ g}}$

Number of ants $= 5.167 \times 10^7$ ants (1 MARK)

e. $n(Sn(s)) = 0.3369 \text{ mol}$

$n(H_2(g)) : n(Sn(s))$

$1 : 1$

$n(H_2(g)) = 0.3369 \text{ mol}$ (1 MARK)

$N(H_2(g)) = 0.3369 \text{ mol} \times 6.02 \times 10^{23} \text{ mol}^{-1}$

$N(H_2(g)) = 2.03 \times 10^{23}$ molecules of $H_2(g)$ (1 MARK)

FROM LESSONS 4B, 6B & 7D

20. a. Monoprotic

b. i. $3HCl(aq) + Al(OH)_3(s) \rightarrow AlCl_3(aq) + 3H_2O(l)$

ii. $3H^+(aq) + Al(OH)_3(s) \rightarrow Al^{3+}(aq) + 3H_2O(l)$

c. $n(Al(OH)_3)$ per tablet $= \frac{m(Al(OH)_3)}{M(Al(OH)_3)}$

$n(Al(OH)_3)$ per tablet $= \frac{(250 \times 10^{-3}) \text{ g}}{(27.0 + 3 \times (16.0 + 1.0)) \text{ g mol}^{-1}}$

$n(Al(OH)_3)$ per tablet $= 0.003205 \text{ mol}$ (1 MARK)

Number of antacid tablets required $= \frac{n(Al(OH)_3)_{\text{required}}}{n(Al(OH)_3)_{\text{per tablet}}}$

Number of antacid tablets required $= \frac{0.05 \text{ mol}}{0.003205 \text{ mol}}$

Number of antacid tablets required $= 15.6$ (1 MARK)

Therefore, 16 antacid tablets would be required. (1 MARK)

FROM LESSONS 3B, 6B & 10A

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10C Calculating pH

Progress questions

1. B. The ionic product of water gives the concentrations of hydronium and hydroxide ions in pure water at low concentrations.

2. A. The individual concentrations of hydronium and hydroxide ions may change due to the addition of an acid or base, but the product of their concentrations is constant at 25 °C.

3. A. Acidic solutions have higher concentrations of hydronium ions and lower concentrations of hydroxide ions than basic solutions.

4. B. Acidic solutions have higher concentrations of hydronium ions than a neutral solution, and therefore have a pH below 7. Basic solutions have lower concentrations of hydronium ions than a neutral solution, and therefore have a pH above 7.

5. B. Since the pH scale is a logarithmic scale, a decrease in the pH by 2 is caused by an increase in the concentration of hydronium ions by a factor of 10^2.

6. A. An acid with a higher concentration will have a higher concentration of hydronium ions, and hence a lower pH value.

Deconstructed

7. D

8. C

9. $1.3 = -\log_{10}[H_3O^+]$

$[H_3O^+] = 10^{-1.3}$ (1 MARK)

$[H_3O^+] = 0.050\ M$ (1 MARK)

Exam-style

10. B. The pH of a solution can be approximated using its concentration, and whether it is an acid or a base.

11. **a.** Basic

b. [Basic solutions contain a higher concentration of hydroxide ions, OH^-, than hydronium ions, H_3O^+.[1]][Therefore, given human blood is basic, it would contain a higher concentration of OH^- ions.[2]]

✓ ✗ I have identified the relative concentrations of hydronium and hydroxide ions in basic solutions.[1]

✓ ✗ I have linked my answer to the question.[2]

c. Highest: $[H_3O^+] = 10^{-pH}$

$[H_3O^+] = 10^{-7.35}$ (1 MARK)

$[H_3O^+] = 4.47 \times 10^{-8}\ M$ (1 MARK)

Lowest: $[H3O+] = 10^{-pH}$

$[H_3O^+] = 10^{-7.45}$ (1 MARK)

$[H_3O^+] = 3.55 \times 10^{-8}\ M$ (1 MARK)

12. **a.** $[H_3O^+] = [HNO_3]$

$[H_3O^+] = 0.50\ M$ (1 MARK)

$pH = -\log_{10}[H_3O^+]$

$pH = -\log_{10}(0.50\ M)$ (1 MARK)

$pH = 0.30$ (1 MARK)

b. [Since the value of K_w has increased, the concentration of hydronium ions would increase also.[1]][As a result, the pH would decrease.[2]]

✓ ✗ I have described the effect of an increased K_w value on $[H_3O^+]$.[1]

✓ ✗ I have linked my answer to the question.[2]

c. $[H_3O^+] \times [OH^-] = 51.3 \times 10^{-14}\ M^2$

Given $[H_3O^+] = [OH^-]$,

$[H_3O^+] \times [H_3O^+] = 51.3 \times 10^{-14}\ M^2$ (1 MARK)

$[H_3O^+] = \sqrt{51.3 \times 10^{-14}\ M^2}$

$[H_3O^+] = 7.162 \times 10^{-7}\ M$ (1 MARK)

$pH = -\log_{10}(7.16 \times 10^{-7})\ M$ (1 MARK)

$pH = 6.14$ (1 MARK)

13. **a.** $11 = -\log_{10}[H_3O^+]$

$[H_3O^+] = 1 \times 10^{-11}\ M$ (1 MARK)

$[OH^-] = \dfrac{1.00 \times 10^{-14}\ M^2}{1 \times 10^{-11}\ M}$ (1 MARK)

$[OH^-] = 0.0010\ M$ (1 MARK)

Therefore, $[NaOH] = 0.0010\ M$ (1 MARK)

b. [Jamie's sodium hydroxide solution is not safe for him to use.[1]][0.1 M is significantly higher than the highest concentration of 0.0010 M that Jamie can safely use.[2]]

✓ ✗ I have identified whether the solution Jamie has is safe for him to use.[1]

✓ ✗ I have explained my answer using molar concentrations.[2]

Key science skills

14. **a.** The beaker of deionised water acts as a control for this experiment.

b. [The independent variable is the identity of the acid.[1]][The dependent variable is the brightness of the lamp in lumens.[2]]

✓ ✗ I have identified the independent variable.[1]

✓ ✗ I have identified the dependent variable.[2]

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c. [The solutions in beakers 1–3 are acidic solutions, and therefore have a pH below 7, whilst the solution in beaker 4, the deionised water, will have the highest pH of approximately 7.[1]][Given electrical conductivity increases with ion concentration, solutions that have higher concentrations of ions will have greater electrical conductivity, and will thus have the brightest lamps.[2]][The solutions in beakers 1–3 are all acidic, and will thus have higher concentrations of H_3O; according to the ionic product of water, solutions with higher $[H_3O^+]$ will have lower pH.[3]][Therefore, since greater brightness means greater electrical conductivity means lower pH, the pH of the solutions in the beakers, from highest to lowest, is: 4, 1, 3, 2.[4]]

✓ ✗ I have compared the pH of the solutions in beakers 1–3 to the pH of the control beaker.[1]

✓ ✗ I have described the relationship between electrical conductivity, ion concentration, and lamp brightness.[2]

✓ ✗ I have explained the relationship between ion concentration and pH.[3]

✓ ✗ I have linked my answer to the question.[4]

d. **i.** If the concentrations of each of the solutions in beakers 1–3 are all increased, then the brightness of each lamp will increase also.

ii. Quantitative

FROM LESSONS 16A & 16B

Questions from multiple lessons

15. **a.** Diprotic acid

b. For strong acids, $[H_3O^+] = [H_2SeO_3]$

$pH = -\log_{10}[H_3O^+]$

$pH = -\log_{10}(0.080\ M)$ (1 MARK)

$pH = 1.10$ (1 MARK)

c. [The deprotonated selenous acid, $HSeO_3^-$, has a negative charge that can be thought of as electrostatically attracting and ‘holding on’ to the positive H^+ ion.[1]][As such, it becomes more difficult for the positive H^+ ion to escape the pull of the remaining compound’s negative charge, resulting in fewer H_2SeO_3 molecules successfully losing both protons.[2]]

✓ ✗ I have explained the relevant theory regarding deprotonated polyprotic acids.[1]

✓ ✗ I have linked my answer to the question.[2]

d. [If selenous acid were a stronger acid, it would donate its protons more readily.[1]][This would lead to a greater amount of H_3O^+ ions in the solution, as more of the H_2SeO_3 would deprotonate either once or twice.[2]][A higher concentration of H_3O^+ ions in the solution would correspond to a lower pH.[3]]

✓ ✗ I have described how the strength of an acid and its willingness to donate protons are related.[1]

✓ ✗ I have explained the effect of acid strength on the $[H_3O^+]$ in the solution.[2]

✓ ✗ I have linked my answer to the question.[3]

FROM LESSON 10A

10D Comparing pH measurements

Progress questions

1. B. Thymolphthalein turns blue in basic solutions.

2. A. Red litmus paper is red in acidic solutions.

3. A. Methyl orange turns red in strongly acidic solutions.

4. B. A pH meter gives a numerical reading for pH.

5. B. Thymol blue turns yellow in a solution with a pH between 2.8 and 8.0.

6. A. Phenolphthalein is colourless in acidic solutions.

7. B. Weak bases have a pH range of approximately 7–10, and the universal indicator turns blue within that range.

8. B. Indicators provide relatively inaccurate and relatively imprecise measurements of pH.

9. A. Calibrated pH meters typically provide both accurate and precise measurements of pH.

Deconstructed

10. C

11. C

12. [Red litmus paper can only indicate whether the soil has a pH of above or below 7.[1]][Given this is only a range of values, using red litmus paper will provide an inaccurate measurement of soil pH.[2]][The farmer can use a pH meter instead, which will give her a more accurate reading of the pH of her soil.[3]]

✓ ✗ I have described the pH indication ability of red litmus paper.[1]

✓ ✗ I have explained the inaccuracy of pH measurements using red litmus paper.[2]

✓ ✗ I have provided an alternative, more accurate technique for measuring pH.[3]

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Exam-style

13. C. Phenolphthalein is colourless at a pH below 8.3. Methyl red is yellow at a pH above 6.2. Thymol blue is yellow at a pH somewhere between 2.8 and 8.0. Therefore, the pH of the solution is between 6.2 and 8.0.

Key science skills

14. a. [Blue litmus paper will only tell Jean-luc whether the pH is above or below 7.^{1}][The universal indicator will produce a coloured solution that he can use to roughly estimate the pH.^{2}][A pH meter will give him a more accurate and precise reading of pH, and is therefore best to determine the pH of the pool water.^{3}]

✓ ✗ I have described the pH indication ability of blue litmus paper.^{1}

✓ ✗ I have described the pH indication ability of universal indicator.^{2}

✓ ✗ I have identified the pH meter as the best instrument to use.^{3}

b. [Given Jean-luc is only taking one reading, there is a higher chance for random error to skew the result.^{1}][As such, the results obtained from taking only one reading would be relatively inaccurate.^{2}]

✓ ✗ I have identified the higher chance for random error to affect results.^{1}

✓ ✗ I have described the accuracy of the results obtained from Jean-luc's technique.^{2}

c. [Jean-luc can take several readings of the pH and average them out.^{1}][Averaging the results reduces the potential effect of random error on the results, thereby making them more accurate.^{2}]

✓ ✗ I have described a method for improving the accuracy of the results.^{1}

✓ ✗ I have explained that the method reduces the effect of random errors.^{2}

FROM LESSONS 16C & 16D

Questions from multiple lessons

15. a. Blue

b. [Bromophenol blue is acting as an acid.^{1}][As per the Brønsted-Lowry theory of acids and bases, bromophenol blue, $C_{19}H_{10}Br_4O_5S(aq)$ acts as an acid by donating a proton to a base and becoming $C_{19}H_9Br_4O_5S^-(aq)$.^{2}]

✓ ✗ I have identified whether bromophenol blue would be an acid or base in the given reaction.^{1}

✓ ✗ I have explained my answer in relation to the Brønsted-Lowry theory of acids and bases.^{2}

FROM LESSON 10A

10E Applications of acid-base reactions in society

Progress questions

1. B. Combustion reactions involve the oxidation of a fuel.

2. A. Acid rain is formed from the naturally occurring reaction between carbon dioxide gas and water in the atmosphere.

3. A. Acid rain is a result of human and natural processes.

4. B Acid rain lowers the pH of water.

5. A The gain of oxygen is oxidation and the loss of oxygen by a substance is reduction.

6. B. Human activities are increasing the amount of carbon dioxide in the atmosphere and, therefore, in ocean water.

7. A. Shells are made of calcium carbonate which is a base.

8. A. This is the process by which calcium ions are released from calcium carbonate.

Deconstructed

9. B

10. C

11. [The graphs show that as the pH becomes more acidic (decreases),^{1}][all three of the coral species^{2}][decrease their percentage mass increase per month (calcification).^{3}]

✓ ✗ I have identified the changing pH from the graphs.^{1}

✓ ✗ I have identified which of the coral species fit a trend.^{2}

✓ ✗ I have described the trend in percentage mass increase.^{3}

Exam-style

12. C. As more carbon dioxide is absorbed by the ocean, the ocean's pH drops, becoming more acidic.

13. C. Burning sulfur in dry air can produce sulfur dioxide.

14. a. [Fossil fuels.^{1}][Combustion or burning.^{2}]

✓ ✗ I have identified the source.^{1}

✓ ✗ I have identified the chemical process.^{2}

b. $2N_2(g) + 5O_2(g) + 2H_2O(l) \rightleftharpoons 4HNO_3(aq)$ (1 MARK)

$2SO_2(g) + O_2(g) + 2H_2O(l) \rightleftharpoons 2H_2SO_4(aq)$ (1 MARK)

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c. [Damage or death to vegetation.[1]][Aluminium ions can be released in the soil and absorbed by plants.[2]][$Al(OH)_3(s) + 3H^+(aq) \rightleftharpoons Al^{3+}(aq) + 3H_2O(l)$.[3]][Damage to buildings.[4]][Limestone is a common building material and is a base.[5]][$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + CO_2(g) + H_2O(l)$ or $CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$.[6]]

✓ ✗ I have given an example of a negative effect.[1]

✓ ✗ I have explained the negative effect.[2]

✓ ✗ I have used an equation and/or chemical terminology.[3]

✓ ✗ I have given a different example of a negative effect.[4]

✓ ✗ I have explained the negative effect.[5]

✓ ✗ I have used an equation and/or chemical terminology.[6]

15. [Excess carbon dioxide gas emitted by human activities [1]][is being absorbed by the oceans [2]][which results in the formation of carbonic acid [3]][which lowers the pH of the oceans [4]][and leads to the death of many marine creatures.[5]]

✓ ✗ I have identified one of the major greenhouse gas.[1]

✓ ✗ I have described how it enters the oceans.[2]

✓ ✗ I have identified the compound.[3]

✓ ✗ I have explained the effect of this compound.[4]

✓ ✗ I have explained the consequences on marine life.[5]

16. a. The reversible nature of some chemical reactions

b. Decalcification

c. $CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$

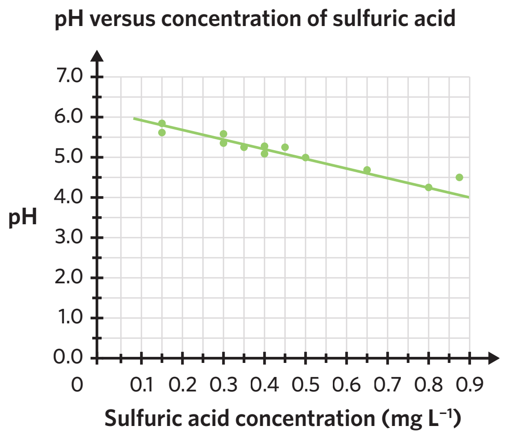
Key science skills

17. a. Quantitative

b. Independent – the rain event
Dependent – sulfuric acid concentration or pH

c. 2 significant figures

d.



✓ ✗ I have plotted all the data points accurately.[1]

✓ ✗ I have drawn an appropriate line.[2]

✓ ✗ I have created a sensible scale on both axes that fills the graph.[3]

e. [As the concentration of sulfuric acid increases in the rainwater,[1]][the pH decreases.[2]]

✓ ✗ I have identified a pattern on the graph.[1]

✓ ✗ I have linked this pattern to a relationship between pH and concentration.[2]

FROM LESSONS 16B, 16C & 16D

Questions from multiple lessons

18. a. [Metallic bonding is the electrostatic attraction between positively charged metal ions and a sea of delocalised electrons.[1]][Pure metal cations exist in ordered layers which are surrounded by delocalised electrons in what is referred to as a lattice structure.[2]]

✓ ✗ I have described metallic bonding with respect to ions and electrons.[1]

✓ ✗ I have identified the structure of metals.[2]

b. [Heat conduction is the ability for a substance to transfer heat.[1]][Metals are good heat conductors due to the delocalised electrons and close proximity of metal ions to each other which enables heat, often in the form of kinetic energy, to be rapidly transferred throughout the metal by the electrons moving about or by vibrating metal ions.[2]]

✓ ✗ I have defined heat conduction.[1]

✓ ✗ I have described the properties of metals that enable heat conduction.[2]

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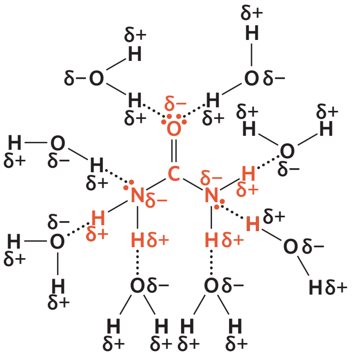
c. [Specific heat capacity is the amount of energy required to raise the temperature of 1 g of a substance by 1 °C.^{1}][Metals have a low specific heat capacity because they conduct heat very effectively which means they require a small amount of heat energy to raise their temperature.^{2}]

✓ ✗ I have defined specific heat capacity.^{1}

✓ ✗ I have explained why metals have low specific heat capacities.^{2}

FROM LESSONS 3A & 9A

19. a.



[Urea is very soluble in water as it can form hydrogen bonds with water molecules as shown in the diagram.^{1}][This means that the solute-solute and solvent-solvent forces can be overcome by the strength of the solute-solvent attractive forces.^{2}][This enables urea to dissolve in water.^{3}]

✓ ✗ I have described the hydrogen bonding interactions between urea and water molecules with a diagram.^{1}

✓ ✗ I have explained the mechanism behind the process of a substance dissolving.^{2}

✓ ✗ I have linked this to the solubility of urea in water.^{3}

b. [Blood and urine consist mainly of water.^{1}][Therefore, since urea is soluble in water, it can freely dissolve into the bloodstream and into urine, allowing it to be transported around the body and then excreted.^{2}]

✓ ✗ I have recognised the significance of water in the human body.^{1}

✓ ✗ I have described the solubility properties of urea which facilitate its transport.^{2}

FROM LESSONS 2A & 2B

Chapter 10 review

Multiple choice

1. C. The higher the pH of a solution, the greater the concentration of $OH^-(aq)$ ions in the solution.

2. D. Sulfur dioxide and nitrogen oxides are released into the atmosphere, where they react with water molecules to form acid rain.

3. B. Brønsted-Lowry acid-base reactions involve the transfer of a proton (H^+) from an acid to a base.

4. C. $HPO_4^{2-}(aq)$ is an amphiprotic species.

5. C. Ionic equations do not include spectator ions and will only include chemical species that undergo a chemical change.

6. A. $H_2PO_4^-(aq)$ is an amphiprotic species as it can both donate and accept protons and is polyprotic as it can donate two protons.

7. B. Pure water has a pH of 7. Ethanoic acid is a weak acid and so the pH will decrease from 7 to a value significantly higher than 1.23, the final pH in options C and D.

8. B. $H_2SeO_3(aq)$ is a Brønsted-Lowry acid that donates a proton to form its conjugate base, $HSeO_3^-(aq)$.

9. B. The reactive metal magnesium will react with iodic acid to form an ionic salt and hydrogen gas.

10. D. Human and turkey vulture stomachs may just have different concentrations of the same acid (1000 times more concentrated), not necessarily different acids. A lower pH corresponds to a higher concentration of $H^+(aq)$ ions.

Short answer

11. [Phenolphthalein was added to solution A.^{1}][Phenol red was added to solution B.^{2}][Bromophenol blue was added to solution C.^{3}]

✓ ✗ I have identified the indicator added to solution A.^{1}

✓ ✗ I have identified the indicator added to solution B.^{2}

✓ ✗ I have identified the indicator added to solution C.^{3}

12. a. $2HBr(aq) + MgCO_3(s) \rightarrow MgBr_2(aq) + CO_2(g) + H_2O(l)$

b. $2H^+(aq) + MgCO_3(s) \rightarrow Mg^{2+}(aq) + CO_2(g) + H_2O(l)$

c. [The conjugate base of hydrobromic acid is $Br^-(aq)$.^{1}] [The stronger an acid, the weaker its conjugate base.^{2}] [As $HBr(aq)$ is a very strong acid, its conjugate base $Br^-(aq)$ will be a base of negligible strength and so will not accept protons from water to form $OH^-(aq)$ ions.^{3}][This means that the concentration of $OH^-(aq)$ ions would not increase by any noticeable extent.^{4}]

✓ ✗ I have identified the conjugate base of hydrobromic acid.^{1}

✓ ✗ I have explained the relationship between acids and their conjugate bases.^{2}

✓ ✗ I have explained how this relationship would affect how bromide acts in solution.^{3}

✓ ✗ I have linked my answer to the question.^{4}

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13. **a.** [AcidAway was least effective.^{1}][The trial with AcidAway resulted in the lowest, or most acidic pH, meaning it had the smallest neutralising effect on the stomach acid.^{2}]

✓ ✗ I have identified the least effective antacid brand.^{1}

✓ ✗ I have explained my answer in reference to the final pH of the trial.^{2}

b. [pHarmaCalm and NeutrapH both resulted in the same final pH of 5, equally higher than the result achieved with AcidAway.^{1}][However, the pHarmaCalm trial reached the final pH 6 minutes faster than the NeutrapH trial.^{2}][Therefore, pHarmaCalm would provide more rapid relief and would hence be most effective for treating indigestion.^{3}]

✓ ✗ I have identified that the two brands had the same final pH.^{1}

✓ ✗ I have identified that pHarmaCalm worked more quickly than NeutrapH.^{2}

✓ ✗ I have linked my answer to the question.^{3}

c. $Mg(OH)_2(aq) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$

d. [Magnesium hydroxide is a base.^{1}][Since magnesium hydroxide is used as an antacid, it must be basic to neutralise the hydrochloric acid in the stomach.^{2}]

✓ ✗ I have identified magnesium hydroxide as a base.^{1}

✓ ✗ I have explained my answer using neutralisation reactions.^{2}

14. **a.** Carbon dioxide

b. [$2SO_2(g) + O_2(g) + 2H_2O(l)$^{1}][$\rightleftharpoons 2H_2SO_4(aq)$.^{2}]

✓ ✗ I have given the equation that produces sulfuric acid.^{1}

✓ ✗ I have include state symbols and the reversible arrow.^{2}

c. Calcification

d. Bicarbonate/Hydrogencarbonate, HCO_3^- or Hydrogen ion, H^+

e. [Metals ions like aluminium that are normally bonded in a compound become aqueous ions that are able to be absorbed by plants and are poisonous.^{1}][$Al(OH)_3(s) + 3H^+(aq) \rightleftharpoons Al^{3+}(aq) + 3H_2O(l)$ ^{2}]

✓ ✗ I have identified and explained the chemical.^{1}

✓ ✗ I have provided a relevant ionic equation.^{2}

15. **a.** $2HNO_3(aq) + Zn(s) \rightarrow Zn(NO_3)_2(aq) + H_2(g)$

b. $2H^+(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + H_2(g)$

c. [Ingot A is gold as it does not react with the nitric acid.^{1}][Ingot B is Prince's metal as the zinc present in the alloy is reacting with the nitric acid to produce hydrogen gas. This is observed by the fizzing/bubbling, as well as the corrosion, of ingot B.^{2}]

✓ ✗ I have identified the metal of ingot A and linked my answer to the experimental observations.^{1}

✓ ✗ I have identified the metal of ingot B and linked my answer to the experimental observations.^{2}

d. [The pH of the solution will increase ^{1}][as $H^+(aq)$ ions are being consumed in the reaction to form $H_2(g)$.^{2}]

✓ ✗ I have identified the change in pH of the solution.^{1}

✓ ✗ I have identified the change in H^+ ions.^{2}

Key science skills

16. **a.**

1. Add 500 mL of 1.0 M $NH_3(aq)$ to a clean, dry 750 mL glass beaker.
2. Place the pH probe in the solution so that it is fully submerged.
3. Record the pH of the solution once the reading stabilises.
4. Remove the pH probe from the solution, rinse it with deionised water and then dry it. (1 MARK)
5. Repeat steps 1–4 with 500 mL of 1.0 M $HONH_2(aq)$.
6. Measure 250 mL of deionised water with the 500 mL measuring cylinder.
7. Add 250 mL of deionised water to the 250 mL of 2.0 M $N_2H_4(aq)$ in a clean, dry 750 mL beaker.
8. Repeat steps 2–4 with this solution. (1 MARK)

b. [The litmus paper will turn blue in all of the basic solutions and so cannot be used to determine their relative strengths.^{1}][The phenolphthalein indicator can only undergo one colour change (between colourless and pink). This means that at least two of the solutions will be the same colour and their relative strength will not be able to be determined.^{2}][The pH probe and recording device allow for a precise quantitative measurement of degree of ionisation of each base in solution and so can be used to determine each base's relative strength.^{3}]

✓ ✗ I have explained why litmus paper should not be used.^{1}

✓ ✗ I have explained why phenolphthalein indicator should not be used.^{2}

✓ ✗ I have explained why the pH probe and recording device should be used.^{3}

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c. [The strongest base will ionise to the greatest extent, resulting in the highest concentration of $OH^-(aq)$ ions and the highest pH as measured by the pH probe.[1]][Weaker bases will ionise to a lower extent, leading to a lower concentration of $OH^-(aq)$ ions and a lower pH.[2]][This information can be used to determine which bases are stronger/weaker depending on the pH of the solutions they form.[3]]

✓ ✗ I have explained the link between a stronger base and pH.[1]

✓ ✗ I have explained the link between a weaker base and pH.[2]

✓ ✗ I have linked my answer to the question.[3]

d. [[OH^-]: Ammonia is highest, hydrazine is middle, and hydroxylamine is lowest.[1]][[H_3O^+]: Hydroxylamine is highest, hydrazine is middle, and ammonia is lowest.[2]]

✓ ✗ I have accurately ordered the hydroxide ion concentration in the 3 solutions.[1]

✓ ✗ I have accurately ordered the hydronium ion concentration in the 3 solutions.[2]

FROM LESSONS 16B & 16E

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11A Redox reactions and equations

Progress questions

1. B. A redox reaction involves both reduction and oxidation.
2. A. Reduction is the gain of electrons.
3. B. Oxidising agents cause another chemical species to undergo oxidation, and are themselves reduced.
4. A. $x + 3 \times (+1) = 0$

   $x = -3$
5. A. An increase in oxidation state is due to the loss of negatively charged electrons, which occurs during oxidation.
6. A. Remember the acronym OILRIG.
7. B. Redox half-equations must be balanced with respect to both charge and the number of each element on each side of the equation.
8. A. Additional electrons are not consumed or created during a redox reaction, and therefore should not be present in the overall redox equation.
9. B. Overall redox equations will not explicitly show the movement of electrons, but will involve the transfer of electrons.

Deconstructed

10. B

11. A

12. Oxidation: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ (1 MARK)

Reduction: $NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(l)$ (1 MARK)

Overall: $3Cu(s) + 2NO_3^-(aq) + 8H^+(aq) \rightarrow 3Cu^{2+}(aq) + 2NO(g) + 4H_2O(l)$ (1 MARK)

Exam-style

13. C 14. B 15. A 16. C

17. a. [Oxidising agents cause another chemical species to be oxidised, and are themselves reduced.[1]][Therefore, since $O_2(g)$ is the oxidising agent, it will be reduced.[2]]

✓ ✗ I have described what occurs to oxidising agents.[1]

✓ ✗ I have identified the chemical species being reduced.[2]

b. Since this is not a redox reaction, electrons are not transferred between chemical species.

18. a. i. [Fe in $FeCl_3(s)$ has an oxidation state of +3 and an oxidation state of +2 in $FeCl_2(s)$.[1]][This means that $FeCl_3(s)$ must be the oxidising agent as it has undergone reduction, shown by the decrease in oxidation state due to the gain of negatively charged electrons.[2]][S in $H_2S(g)$ has an oxidation state of −2 and an oxidation state of 0 in S(s).[3]][This means that $H_2S(g)$ must be the reducing agent as it has undergone oxidation, shown by the increase in oxidation state due to the loss of negatively charged electrons.[4]]

✓ ✗ I have described the change in oxidation state for Fe.[1]

✓ ✗ I have identified the oxidising agent in this reaction, with reference to the change in oxidation state.[2]

✓ ✗ I have described the change in oxidation state for S.[3]

✓ ✗ I have identified the reducing agent in this reaction, with reference to the change in oxidation state.[4]

ii. [N in $N_2(g)$ has an oxidation state of 0 and an oxidation state of −3 in $Mg_3N_2(s)$.[1]][This means that $N_2(g)$ must be the oxidising agent as it has undergone reduction, shown by the decrease in oxidation state due to the gain of negatively charged electrons.[2]][Mg in Mg(s) has an oxidation state of 0 and an oxidation state of +2 in $Mg_3N_2(s)$.[3]][This means that Mg(s) must be the reducing agent as it has undergone oxidation, shown by the increase in oxidation state due to the loss of negatively charged electrons.[4]]

✓ ✗ I have described the change in oxidation state for N.[1]

✓ ✗ I have identified the oxidising agent in this reaction, with reference to the change in oxidation state.[2]

✓ ✗ I have described the change in oxidation state for Mg.[3]

✓ ✗ I have identified the reducing agent in this reaction, with reference to the change in oxidation state.[4]

b. i. Oxidation: $H_2S(g) \rightarrow S(s) + 2H^+(aq) + 2e^-$ (1 MARK)

Reduction: $Fe^{3+}(s) + e^- \rightarrow Fe^{2+}(s)$ (1 MARK)

ii. Oxidation: $Mg(s) \rightarrow Mg^{2+}(s) + 2e^-$ (1 MARK)

Reduction: $N_2(g) + 6e^- \rightarrow 2N^{3-}(s)$ (1 MARK)

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Key science skills

19. a. Colour of the solution

b.

	1	2	3	4	5
Colour	Pink	Green	Blue	Yellow	Orange
Main species present	MnO_4^- (aq)	MnO_4^{2-} (aq)	MnO_4^{3-} (aq)	MnO(s)	MnO_2(s)
Mn oxidation state	+7 (1 MARK)	+6 (1 MARK)	+5 (1 MARK)	+4 (1 MARK)	+4 (1 MARK)

c. $MnO_4^{2-}(aq) + e^- \rightarrow MnO_4^{3-}(aq)$

d. [As the experiment proceeds, a higher amount of MnO_4^-(aq) is reduced to MnO_2(s) by the sugars in the lollipop.[1]] [As a low amount of MnO_2(s) leads to a yellow coloured solution, the orange colour could be a result of this increased amount of MnO_2(s) present in the beaker.[2]]

✓ ✗ I have described the relationship between the experimental procedure and the amount of MnO_2(s) present in the beaker.[1]

✓ ✗ I have suggested a reason for the observed change in the solution's colour.[2]

FROM LESSONS 16A, 16C & 16E

Questions from multiple lessons

20. a. [Chemical compounds which are oxidised lose electrons.[1]] [Therefore, if antioxidants are oxidised in preference to another compound, they must be able to lose electrons easily.[2]]

✓ ✗ I have defined the process of oxidation.[1]

✓ ✗ I have linked the ability to be oxidised to the ability to lose electrons.[2]

b. [The student is incorrect.[1]] [This is because oxidising agents cause another chemical species to be oxidised and are, therefore, reduced themselves as they gain electrons from a chemical species that has lost electrons (i.e. been oxidised).[2]]

✓ ✗ I have commented on the accuracy of the student's statement.[1]

✓ ✗ I have explained why oxidising agents are reduced themselves.[2]

c. The hydroxyl functional group, OH

d. [Antioxidants are chemical species which are easily able to lose electrons and be oxidised themselves.[1]] [The hydroxyl function group, with its 2 lone pairs of electrons on each oxygen atom, means that there are plenty of electrons that can be lost, allowing it to readily undergo oxidation which is the loss of electrons.[2]]

✓ ✗ I have defined the purpose of an antioxidant.[1]

✓ ✗ I have explained how the structure of the hydroxyl functional group relates to its ability to be oxidised.[2]

FROM LESSONS 2A & 7C

21. a. [+1 +5 −2 0 +1 −2 [1]]
[$12H^+(aq) + 2ClO_3^-(aq) + 10e^- \rightarrow Cl_2(g) + 6H_2O(l)$[2]]

✓ ✗ I have deduced the oxidation states.[1]

✓ ✗ I have written a balanced reduction half equation.[2]

b. [+2 −2 +1 −2 +5 −2 +1 [1]]
[$NO(g) + 2H_2O(l) \rightarrow NO_3^-(aq) + 4H^+(aq) + 3e^-$[2]]

✓ ✗ I have deduced the oxidation states.[1]

✓ ✗ I have written a balanced reduction half equation.[2]

c. $10NO(g) + 2H_2O(l) + 6ClO_3^-(aq) \rightarrow 10NO_3^-(aq) + 4H^+(aq) + 3Cl_2(g)$

d. [In the balanced overall redox reaction, there is a net production of H^+(aq) ions.[1]] [This increase in the concentration of H^+(aq) ions will result in a decrease in pH.[2]]

✓ ✗ I have described the effect of the redox reaction on the production of H^+(aq) ions.[1]

✓ ✗ I have explained how this will affect the pH of the solution.[2]

FROM LESSON 10C

11B Reactivity of metals – part 2

Progress questions

1. A. Highly reactive metals readily undergo oxidation, making them strong reducing agents.

2. A. The strongest reducing agents (most reactive metals) are found on the bottom right of the reactivity series.

3. B. Au^+(aq) will gain an electron, undergoing reduction to form Au(s), a weak reducing agent. This makes Au^+(aq) a strong oxidising agent.

4. B. Potassium, K(s), is so highly reactive that it will react vigorously with O_2(g) from the atmosphere.

5. A. Al(s) is lower on the reactivity series than Cu(s), making it a stronger reducing agent and a more reactive metal.

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6. A. Highly reactive metals readily undergo oxidation, donating electrons to the cation in solution to form their own cations and, in doing so, displace the less reactive metal from solution.

7. B. Metal displacement reactions can only occur if the two species have a stronger oxidising (top left) and stronger reducing agent (bottom right) from the reactivity series.

8. B. Zinc is a more reactive metal than lead, and so solid zinc will displace lead ions from solution.

Deconstructed

9. C

10. B

11. Oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ (1 MARK)

Reduction: $Co^{2+}(aq) + 2e^- \rightarrow Co(s)$ (1 MARK)

Overall: $Zn(s) + Co^{2+}(aq) \rightarrow Zn^{2+}(aq) + Co(s)$ (1 MARK)

Exam-style

12. C

13. D

14. B

15. a. $Pb^{2+}(aq) + Co(s) \rightarrow Pb(s) + Co^{2+}(aq)$

b. $2Ag^+(aq) + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}(aq)$

c. $3Sn^{2+}(aq) + 2Al(s) \rightarrow 3Sn(s) + 2Al^{3+}(aq)$

d. No reaction

e. $Mg(s) + 2Au^+(aq) \rightarrow Mg^{2+}(aq) + 2Au(s)$

f. No reaction

16. D

17. a. [No reaction would occur.[1]][$Fe^{2+}(aq)$ will not form from $Fe(s)$ as it is a stronger oxidising agent than $Zn^{2+}(aq)$.[2]]

✓ ✗ I have identified whether a reaction would occur.[1]

✓ ✗ I have described the difference in oxidising strength between zinc ions and iron (II) ions on the reactivity series of metals.[2]

b. Solid lead, $Pb(s)$

c. Oxidation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$ (1 MARK)

Reduction: $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$ (1 MARK)

Overall: $Fe(s) + Ni^{2+}(aq) \rightarrow Fe^{2+}(aq) + Ni(s)$ (1 MARK)

Key science skills

18. a. $4Wa(s) + 5Kt^{4+}(aq) \rightarrow 4Wa^{5+}(aq) + 5Kt(s)$

b. [Kt, Sk, Wa.[1]][In experiment 1, $Wa(s)$ displaces $Kt^{4+}(aq)$ from the solution, shown by the green metallic coating and the solution's change in colour. This means that $Wa(s)$ is a stronger reducing agent than $Kt(s)$, making it a more reactive metal.[2]][In experiment 2, $Sk(s)$ cannot displace $Wa^{5+}(aq)$ from its solution as there is no reaction. This means that $Wa(s)$ is also a stronger reducing agent than $Sk(s)$, making it a more reactive metal.[3]][In experiment 3, $Sk(s)$ displaces $Kt^{4+}(aq)$ from the solution, shown by the green metallic coating and the solution's change in colour. This means that $Sk(s)$ is a stronger reducing agent than $Kt(s)$, making it a more reactive metal.[4]] [Therefore, Sparkonium is more reactive than Kostrethium, but less reactive than Wallacium.[5]]

✓ ✗ I have ranked the metals by their reactivity.[1]

✓ ✗ I have explained the results of experiment 1 with reference to any observations.[2]

✓ ✗ I have explained the results of experiment 2 with reference to any observations.[3]

✓ ✗ I have explained the results of experiment 3 with reference to any observations.[4]

✓ ✗ I have linked these results to each metal's relative reactivity.[5]

FROM LESSONS 16C & 16E

Questions from multiple lessons

19. a. Oxidation: $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$ (1 MARK)

Reduction: $Ag^+(aq) + e^- \rightarrow Ag(s)$ (1 MARK)

b. $Sn(s) + 2Ag^+(aq) \rightarrow Sn^{2+}(aq) + 2Ag(s)$

c. i. [During the metal displacement reaction, $Sn(s)$ is lost from the tin figurine to form $Sn^{2+}(aq)$ ions in the solution,[1]] [but $Ag(s)$ is also gained as a silver deposit from $Ag^+(aq)$ ions forms on the figurine as well. This results in the observed change in mass.[2]]

✓ ✗ I have identified that tin is lost from the figure.[1]

✓ ✗ I have identified that silver is deposited onto the figurine.[2]

ii. $n(Sn(s)\ \text{lost}) = \frac{1}{2} \times n(Ag(s)\ \text{gained})$ (1 MARK)

$\Delta m(\text{figurine}) = m(Ag(s)\ \text{gained}) - m(Sn(s)\ \text{lost})$

$37.4\ g = n(Ag(s)\ \text{gained}) \times M(Ag)) - n(Sn(s)\ \text{lost}) \times M(Sn))$

$37.4\ g = n(Ag(s)\ \text{gained}) \times 107.9\ g\ mol^{-1} - \frac{1}{2} \times n(Ag(s)\ \text{gained}) \times 118.7\ g\ mol^{-1}$

$37.4\ g = 48.55\ g\ mol^{-1} \times n(Ag(s)\ \text{gained})$ (1 MARK)

$n(Ag(s)\ \text{gained}) = 0.770\ mol$ (1 MARK)

$m(Ag(s)\ \text{gained}) = n(Ag(s)\ \text{gained}) \times M(Ag)$

$m(Ag(s)\ \text{gained}) = 0.770\ mol \times 107.9\ g\ mol^{-1}$

$m(Ag(s)\ \text{gained}) = 83.1\ g$ (1 MARK)

FROM LESSONS 6B & 6C

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11C Applications of redox reactions in society

Progress questions

1. A. The cathode is the positively charged electrode in a galvanic cell.

2. A. Negatively charged electrons flow towards the positively charged cathode from the anode in a galvanic cell.

3. B. In a galvanic cell, oxidation always occurs at the anode.

4. A. The salt bridge completes the circuit in a galvanic cell.

5. B. Dry corrosion occurs in the presence of oxygen only.

6. B. Sodium reacts vigorously with water to produce hydrogen gas.

7. A. Oxygen is reduced ($O_2 = 0$ to O^{2-}) in the formation of rust, whilst iron is oxidised ($Fe = 0$ to Fe^{2+}).

8. B. The anode is sacrificially oxidised to protect iron from corrosion.

Deconstructed

9. D

10. C

11. Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ (1 MARK)
Oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ (1 MARK)
Overall: $Cu^{2+}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq)$ (1 MARK)

Exam-style

12. **a.** $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

b. $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$

c. $Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$

d. $4Fe(OH)_2(s) + O_2(g) + 2H_2O(l) \rightarrow 4Fe(OH)_3(s)$

13. A. In the corrosion of iron, Fe(s) is oxidised to $Fe^{2+}(aq)$ at the anode.

14. **a.** [Oxygen (Air).[1]][Water.[2]]

✓ ✗ I have identified that oxygen is required for rusting to occur.[1]

✓ ✗ I have identified that water is required for rusting to occur.[2]

b. [According to the reactivity series of metals, zinc is a stronger reducing agent than iron.[1]][As such, zinc is oxidised more readily than iron.[2]][The preferential oxidation of zinc is shown in the following equation:
$2Zn(s) + O_2(g) + 2H_2O(l) \rightarrow 2Zn^{2+}(aq) + 4OH^-(aq)$.[3]]

✓ ✗ I have identified the stronger reducing agent using the reactivity series of metals.[1]

✓ ✗ I have identified the correlation between reducing agent strength and ease of oxidation.[2]

✓ ✗ I have provided the equation for the preferential oxidation of zinc.[3]

15. B. Zinc is a stronger reducing agent and will therefore be oxidised, producing more $Zn^{2+}(aq)$ ions.

16. **a.** From the manganese electrode to the nickel electrode

b. $Mn(s) \rightarrow Mn^{2+}(aq) + 2e^-$

c. Chemical energy → electrical energy

Key science skills

17. **a.** Quantitative

b. [A greater amount of oxygen present means a greater amount of corrosion of iron.[1]][According to the graph, there is a greater amount of oxygen present at 3000 m below sea level than at 1000 m below sea level.[2]][Therefore, shipwrecks at 3000 m below sea level would experience more corrosion than shipwrecks at 1000 m below sea level.[3]]

✓ ✗ I have explained that more oxygen causes more corrosion.[1]

✓ ✗ I have identified that there is more oxygen present at 3000 m than at 1000 m below sea level.[2]

✓ ✗ I have linked my answer to the question.[3]

c. An anecdote

FROM LESSONS 16A, 16B & 16C

Questions from multiple lessons

18. **a.** [The chloride ions found in the salt bridge form an insoluble precipitate with the Ag^+ ions.[1]][This would hinder the function of the cell as silver ions are necessary for the redox reaction to occur.[2]]

✓ ✗ I have identified the formation of an insoluble precipitate.[1]

✓ ✗ I have linked this to the effect on the function of the galvanic cell.[2]

b. No electricity was generated because, without the salt bridge, the circuit is not completed.

c. Anode: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ (1 MARK)
Cathode: $Ag^+(aq) + e^- \rightarrow Ag(s)$ (1 MARK)
Overall: $2Ag^+(aq) + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}(aq)$ (1 MARK)

FROM LESSONS 4C & 11B

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Chapter 11 review

Multiple choice

1. C. This is an oxidation half-reaction as the oxidation number of nitrogen increases from −3 to +4 due to the loss of negatively charged electrons.

2. B. The oxidising agent undergoes reduction, decreasing in oxidation number. The reducing agent undergoes oxidation, increasing in oxidation number.

3. D. An increase in oxidation number is due to an oxidation process occurring.

4. B. The solid of a more reactive metal (a stronger reducing agent and lower on the reactivity series) will displace a less reactive (a weaker reducing agent) metal's cations from solution.

5. D. In options A and B, sulfur has an oxidation state of +6. In option C, sulfur has an oxidation state of +4. In option D, sulfur has an oxidation state of −2.

6. A. $PbO_2(s)$ is reduced to $Pb^{2+}(aq)$ and this reaction must be balanced with respect to both charge and number of each element. The oxidation number of iodine increases.

7. B. Whether a metal displacement reaction will occur depends on the oxidising and reducing agent strengths of the species involved. K(s) is highly reactive so it will react vigorously even with water.

8. C. Electrons move from the site of oxidation to the site of reduction (from Q to S), and S is immersed in solution R.

9. A. Only solid magnesium is a stronger reducing agent than iron, and would therefore be the only metal more readily oxidised than iron.

10. D. Half-equations are balanced according to the KOHES sequence of steps.

Short answer

11. **a.** Carbon: $H_2C_2O_4(aq)$ (+3), $CO_2(g)$ (+4) (1 MARK)

Manganese: $MnO_4^{-}(aq)$ (+7), $Mn^{2+}(aq)$ (+2) (1 MARK)

b. [$MnO_4(aq)$ is reduced, as the oxidation state of Mn decreases from +7 to +2 due to the gain of negatively charged electrons.[1]] [$H_2C_2O_4(aq)$, however, is oxidised, as the oxidation state of C increases from +3 to +4 due to the loss of negatively charged electrons.[2]]

✓ ✗ I have identified the species undergoing reduction through oxidation states.[1]

✓ ✗ I have identified the species undergoing oxidation through oxidation states.[2]

c. Oxidation: $H_2C_2O_4(aq) \rightarrow 2CO_2(g) + 2H^{+}(aq) + 2e^{-}$ (1 MARK)

Reduction: $MnO_4^{-}(aq) + 8H^{+}(aq) + 5e^{-} \rightarrow Mn^{2+}(aq) + 4H_2O(l)$ (1 MARK)

12. **a.** [A metal displacement reaction is occurring in experiment 2.[1]] [This is a spontaneous redox reaction.[2]]

✓ ✗ I have identified the type of reaction occurring.[1]

✓ ✗ I have classed the reaction as either spontaneous or non-spontaneous.[2]

b. Oxidation: $V(s) \rightarrow V^{2+}(aq) + 2e^{-}$. Reduction: $Fe^{2+}(aq) + 2e^{-} \rightarrow Fe(s)$ (1 MARK)

Overall equation: $V(s) + Fe^{2+}(aq) \rightarrow V^{2+}(aq) + Fe(s)$ (1 MARK)

c. [V(s) displaced $Fe^{2+}(aq)$ cations from solution but Fe(s) was unable to displace $V^{2+}(aq)$ ions from solution, as shown by the vigorous reaction in Experiment 2, but the lack of a reaction in Experiment 1.[1]] [This means that V(s) is a more reactive metal than Fe(s), and therefore a stronger reducing agent.[2]]

✓ ✗ I have used experimental observations to justify my answer.[1]

✓ ✗ I have described how these experimental results explain the relative reactivity of V(s) and Fe(s).[2]

d. [From Experiment 2, we know that vanadium is a more reactive metal than Fe(s) and so will also be more reactive than Cu(s) as Cu(s) is less reactive than Fe(s).[1]] [This means that V(s) is also a stronger reducing agent than Cu(s) and so will displace $Cu^{2+}(aq)$ ions from a solution of $CuSO_4(aq)$. Hence, a reaction will be observed.[2]]

✓ ✗ I have determined the relative reactivity of V(s) and Cu(s).[1]

✓ ✗ I have explained how this will affect whether a reaction will occur between V(s) and $Cu^{2+}(aq)$ ions.[2]

13. **a.** [$Fe_2O_3{\cdot}xH_2O$ [1]] [Rust [2]]

✓ ✗ I have given the accurate chemical formula for hydrated iron (III) oxide.[1]

✓ ✗ I have given the common name for iron (III) oxide.[2]

b. [The corrosion of iron occurs in the presence of oxygen and water.[1]] [A more electrically conductive solution of water will increase the rate of corrosion.[2]] [Since seawater contains more ions and is therefore more electrically conductive, it will corrode iron more quickly than freshwater.[3]]

✓ ✗ I have identified the conditions necessary for the corrosion of iron.[1]

✓ ✗ I have explained the effect of solution conductivity on the rate of corrosion.[2]

✓ ✗ I have linked my answer to the question.[3]

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c. [The zinc coating on the bolt would be oxidised more readily than the iron bolt itself, as zinc is a stronger reducing agent than iron.^{1}] [Therefore, Viktória would observe oxidation of the zinc coating, but minimal corrosion on the iron bolt itself.^{2}]

✓ ✗ I have identified that zinc is more readily oxidised than iron.^{1}

✓ ✗ I have linked my answer to the question.^{2}

d. [Since aluminium forms an overlying layer of aluminium oxide, Al_2O_3, when oxidised, it is well protected from corrosion.^{1}] [Iron, on the other hand, forms a layer of rust when oxidised, which flakes away easily.^{2}] [Therefore, the aluminium parts on their car would be much less corroded than the iron bolt from Viktória's boat.^{3}]

✓ ✗ I have explained the Al_2O_3 protective layer formed from the oxidation of aluminium.^{1}

✓ ✗ I have explained how the layer of rust that forms on iron is not protective.^{2}

✓ ✗ I have linked my answer to the question.^{3}

14. a. $\overset{+5\,-2}{NO_3^-}(aq) + \overset{+1\,+1\,-2}{HBrO}(aq) \rightarrow \overset{+4\,-2}{N_2O_4}(g) + \overset{+5\,-2}{BrO_3^-}(aq)$ (2 MARKS)

b. $4NO_3^-(aq) + 3H^+(aq) + HBrO(aq) \rightarrow 2N_2O_4(g) + 2H_2O(l) + BrO_3^-(aq)$ (2 MARKS)

c. $n(N_2O_4(g)) = 2.78$ mol

$n(e^-) = \frac{4}{2} \times n(N_2O_4)$

$n(e^-) = 2 \times 2.78$ mol

$n(e^-) = 5.56$ mol (1 MARK)

$N(e^-) = n \times N_A$

$N(e^-) = 5.56 \text{ mol} \times (6.02 \times 10^{23} \text{ mol}^{-1})$

$N(e^-) = 3.35 \times 10^{24}$ electrons (1 MARK)

Key science skills

15. a. [The white sheet of paper allows all colour descriptions and absorption readings to be taken with a similar background.^{1}] [This allows some variability in these readings to be controlled.^{2}]

✓ ✗ I have described the purpose of the white sheet of paper for data collection.^{1}

✓ ✗ I have explained how this influences the variability of the data obtained.^{2}

b. $2Ag^+(aq) + Co(s) \rightarrow 2Ag(s) + Co^{2+}(aq)$ (1 MARK)

c. [No, Meredith could not use this observation to estimate the surface area of the cobalt wire in her phone.^{1}] [This is because the 'ripe tomato' red description of the $Co^{2+}(aq)$ solution is purely qualitative and so cannot be used to make inferences about quantitative measurements such as the wire's surface area.^{2}]

✓ ✗ I have stated whether Meredith could use this observation to estimate the wire's surface area.^{1}

✓ ✗ I have explained why this is the case with reference to qualitative and quantitative data.^{2}

FROM LESSON 16B

16. a. [Electrons flow from the anode to the cathode in a galvanic cell.^{1}] [Since zinc is a stronger reducing agent than copper, it would be oxidised at the anode.^{2}] [Therefore, electrode R would be the zinc electrode, and electrode S would be the copper electrode.^{3}]

✓ ✗ I have identified the direction of electron flow in a galvanic cell.^{1}

✓ ✗ I have identified which metal would be oxidised at which electrode.^{2}

✓ ✗ I have linked my answer to the question.^{3}

b. [Goggles should be worn to protect eyes from solutions in beakers.^{1}] [Gloves should be worn to protect skin from solutions in beakers that may be corrosive.^{2}]

✓ ✗ I have identified one relevant safety precaution.^{1}

✓ ✗ I have identified a second relevant safety precaution.^{2}

FROM LESSON 16B

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12A Measuring solutions

Progress questions

1. A. Solutions with the same concentration have the same solute : solution ratio.

2. B. Solutions with the same concentration have the same solute : solution ratio.

3. A. A solution with less solute per unit of volume is comparatively more dilute.

4. B. Molarity is expressed as mol L^{-1}.

5. B. There are 1 000 000 milligrams in a litre, which is equal to ppm.

6. A. There are 1 000 000 milligrams in a litre, which is equal to ppm.

7. B. For the values to be compared, the volumes and weights need to be expressed in the same units.

8. B. Given the mass of solute is given, and the volume of solution is given, %(m/v) would be most suitable to express the concentration of the solution.

9. A. Dilution is the process of adding more solvent to a solution.

10. B. Dilution does not change the amount of solute present in a solution.

Deconstructed

11. A

12. B

13. $n(\text{NaOH}) = c \times V$

$n(\text{NaOH}) = 2.5 \text{ mol L}^{-1} \times \frac{500 \text{ mL}}{1000}$

$n(\text{NaOH}) = 1.25 \text{ mol}$ (1 MARK)

$m(\text{NaOH}) = 1.25 \text{ mol} \times (23.0 + 16.0 + 1.0) \text{ g mol}^{-1}$

$m(\text{NaOH}) = 50 \text{ g}$ (1 MARK)

Exam-style

14. a. i. $n(\text{CaCl}_2) = \frac{m}{M}$

$n(\text{CaCl}_2) = \frac{1.47 \text{ g}}{(40.1 + 2 \times 35.5) \text{ g mol}^{-1}}$

$n(\text{CaCl}_2) = 0.01323 \text{ mol}$ (1 MARK)

$c(\text{CaCl}_2) = \frac{n}{V}$

$c(\text{CaCl}_2) = \frac{0.01323 \text{ mol}}{\frac{200 \text{ mL}}{1000}}$

$c(\text{CaCl}_2) = 0.0662 \text{ mol L}^{-1}$ (1 MARK)

ii. $m(\text{NaOH}) = 2.0 \text{ g}$

$m(\text{NaOH}) = (2.0 \times 1000) \text{ mg}$

$m(\text{NaOH}) = 2000 \text{ mg}$ (1 MARK)

$c(\text{NaOH}) = \frac{2000 \text{ mol}}{\frac{200 \text{ mL}}{1000}}$

$c(\text{NaOH}) = 1.0 \times 10^4 \text{ ppm}$ (1 MARK)

iii. $m(\text{NaOH}) = n \times M$

$m(\text{NaOH}) = 2.0 \text{ mol} \times (23.0 + 16.0 + 1.0) \text{ g mol}^{-1}$

$m(\text{NaOH}) = 80.0 \text{ g}$

$m(\text{NaOH}) = (80.0 \times 1000) \text{ mg}$

$m(\text{NaOH}) = 80\,000 \text{ mg}$ (1 MARK)

$c(\text{NaOH}) = \frac{80\,000 \text{ mg}}{\frac{200 \text{ mL}}{1000}}$

$c(\text{NaOH}) = 4.0 \times 10^5 \text{ ppm}$ (1 MARK)

iv. $m(\text{FeSO}_4) = n \times M$

$m(\text{FeSO}_4) = 0.50 \text{ mol} \times (55.8 + 32.1 + 4 \times 16) \text{ g mol}^{-1}$

$m(\text{FeSO}_4) = 76.0 \text{ g}$ (1 MARK)

$\%(\text{m/v}) = \frac{76.0 \text{ g}}{100 \text{ mL}} \times 100\%$

$\%(\text{m/v}) = 76\%$ (1 MARK)

b. i. $1337 \text{ ppm} = 1337 \text{ mg L}^{-1}$

$\frac{1337 \text{ mg L}^{-1}}{1000} = 1.337 \text{ g L}^{-1}$

ii. $981 \text{ ppm} = \frac{981 \text{ mg}}{1 \text{ L}}$

$m(\text{KCl}) = \frac{981 \text{ mg}}{1000}$

$m(\text{KCl}) = 0.981 \text{ g}$

$n(\text{KCl}) = \frac{m}{M}$

$n(\text{KCl}) = \frac{0.981 \text{ g}}{(39.1 + 35.5) \text{ g mol}^{-1}}$

$n(\text{KCl}) = 0.01315 \text{ mol}$ (1 MARK)

$c(\text{KCl}) = \frac{n}{V}$

$c(\text{KCl}) = \frac{0.01315 \text{ mol}}{1 \text{ L}}$

$c(\text{KCl}) = 0.0132 \text{ mol L}^{-1}$ (1 MARK)

iii. $1.5 \text{ M} = \frac{1.5 \text{ mol}}{1 \text{ L}}$

$m(\text{PbI}_2) = n \times M$

$m(\text{PbI}_2) = 1.5 \text{ mol} \times (207.2 + 2 \times 126.9) \text{ g mol}^{-1}$

$m(\text{PbI}_2) = 691.5 \text{ g}$ (1 MARK)

$\%(\text{m/v}) = \frac{691.5 \text{ g}}{(1 \times 1000) \text{ mL}} \times 100\%$

$\%(\text{m/v}) = 69\%$ (1 MARK)

c. i. $c_1 \times V_1 = c_2 \times V_2$

$c_1 \times \frac{20.0 \text{ mL}}{1000} = 2.0 \text{ mol L}^{-1} \times \frac{50.0 \text{ mL}}{1000}$ (1 MARK)

$c_1 = 5.0 \text{ mol L}^{-1}$ (1 MARK)

ii. $c_1 \times V_1 = c_2 \times V_2$

$1.05 \text{ mol L}^{-1} \times \frac{150 \text{ mL}}{1000} = c_2 \times \frac{(200 + 150) \text{ mL}}{1000}$ (1 MARK)

$c_2 = 0.450 \text{ mol L}^{-1}$ (1 MARK)

15. a. 8.7 g in every 100 mL

$\%(\text{m/v}) = \frac{8.7}{100} \times 100\%$

$\%(\text{m/v}) = 8.7\%$

b. $n(\text{sugar}) = \frac{m}{M}$

$n(\text{sugar}) = \frac{8.7 \text{ g}}{180 \text{ g mol}^{-1}}$

$n(\text{sugar}) = 0.0483 \text{ mol}$ (1 MARK)

$c(\text{sugar}) = \frac{n}{V}$

$c(\text{sugar}) = \frac{0.0483 \text{ mol}}{\frac{100 \text{ mL}}{1000}}$

$c(\text{sugar}) = 0.48 \text{ mol L}^{-1}$ (1 MARK)

16. a. $1 \text{ ppm} = 1 \text{ mg kg}^{-1}$

$1 \text{ ppm} = \frac{0.5 \text{ mg}}{500 \text{ g}}$ (1 MARK)

Therefore, the maximum amount of mercury is 0.5 mg for a fish weighing 500 g. (1 MARK)

b. $1.76 \text{ mol L}^{-1} = 1.76$ mol of mercury per litre of solution.

$m(\text{Hg}) = n \times M$

$m(\text{Hg}) = 1.76 \text{ mol} \times 200.6 \text{ g mol}^{-1}$

$m(\text{Hg}) = 353.1$ g of mercury per litre of solution (1 MARK)

Therefore, in 50 mL,

$m(\text{Hg}) = 353.1 \text{ g} \times \frac{50 \text{ mL}}{1000}$

$m(\text{Hg}) = 17.7 \text{ g}$ (1 MARK)

c. $m(\text{Hg}) = (17.7 \times 1000) \text{ mg}$

$m(\text{Hg}) = 17\,700 \text{ mg}$

Concentration in ppm $= \frac{17\,700 \text{ mg}}{2.67 \text{ kg}}$

Concentration in ppm $= 6.63 \times 10^3$ ppm (1 MARK)

[The fish has a mercury concentration of 6.63×10^3 ppm, a concentration much higher than the Australian requirement of 1 ppm.[2]][Therefore, the mercury concentration in this fish would not meet the Australian standards.[3]]

✓ ✗ I have calculated the concentration of mercury, in ppm, in the fish.[1]

✓ ✗ I have compared the mercury concentration in the fish with the Australian requirement.[2]

✓ ✗ I have linked my answer to the question.[3]

d. $c_1 = 1.76 \text{ mol L}^{-1}$

$V_1 = \frac{50 \text{ mL}}{1000}$

$V_1 = 0.050 \text{ L}$

$c_2 = 1.0 \text{ mol L}^{-1}$

$V_2 = ?$

$c_1 \times V_1 = c_2 \times V_2$

$1.76 \text{ mol L}^{-1} \times 0.050 \text{ L} = 1.0 \text{ mol L}^{-1} \times V_2$

$V_2 = \frac{1.76 \text{ mol L}^{-1} \times 0.050 \text{ L}}{1.0 \text{ mol L}^{-1}}$

$V_2 = 0.088 \text{ L}$ (1 MARK)

Required volume $= V_2 - V_1$

Required volume $= (0.088 - 0.050) \text{ L}$

Required volume $= 0.038 \text{ L}$ (1 MARK)

17. a. 50 %(m/v) = 50 g of sugar per 100 mL of cordial

Therefore, in 1 L,

$m(C_6H_{12}O_6) = 50 \text{ g} \times \frac{1000}{100} \text{ L}$

$m(C_6H_{12}O_6) = 500 \text{ g}$

b. $n(C_6H_{12}O_6) = \frac{m}{M}$

$n(C_6H_{12}O_6) = \frac{500 \text{ g}}{((6 \times 12.0) + (12 \times 1.0) + (6 \times 16.0)) \text{ g mol}^{-1}}$

$n(C_6H_{12}O_6) = 2.78 \text{ mol}$ (1 MARK)

$c(C_6H_{12}O_6) = \frac{n}{V}$

$c(C_6H_{12}O_6) = \frac{2.78 \text{ mol}}{1.0 \text{ L}}$

$c(C_6H_{12}O_6) = 2.78 \text{ mol L}^{-1}$ (1 MARK)

c. $c_1 = 2.78 \text{ mol L}^{-1}$

$V_1 = \frac{50 \text{ mL}}{1000}$

$V_1 = 0.050 \text{ L}$

$c_2 = ?$

$V_2 = \frac{200 \text{ mL}}{1000}$

$V_2 = 0.200 \text{ L}$

$c_1 \times V_1 = c_2 \times V_2$

$2.78 \text{ mol L}^{-1} \times 0.050 \text{ L} = c_2 \times 0.200 \text{ L}$ (1 MARK)

$c_2 = \frac{2.78 \text{ mol L}^{-1} \times 0.050 \text{ L}}{0.200 \text{ L}}$

$c_2 = 0.70 \text{ mol L}^{-1}$ (1 MARK)

18. a. $7.0 \text{ ppm} = 7.0 \text{ mg kg}^{-1}$

$m(\text{aldicarb}) = 7.0 \text{ mg kg}^{-1} \times 13.6 \text{ kg}$

$m(\text{aldicarb}) = 95 \text{ mg}$

b. $n(\text{aldicarb}) = c \times V$

$n(\text{aldicarb}) = 1.82 \times 10^{-5} \text{ mol L}^{-1} \times \frac{200 \text{ mL}}{1000}$

$n(\text{aldicarb}) = 3.64 \times 10^{-6} \text{ mol}$ (1 MARK)

$m(\text{aldicarb}) = n \times M$

$m(\text{aldicarb}) = 3.64 \times 10^{-6} \text{ mol} \times 190.3 \text{ g mol}^{-1}$

$m(\text{aldicarb}) = 6.926 \times 10^{-4} \text{ g}$ (1 MARK)

$m(\text{aldicarb}) = (6.926 \times 10^{-4} \times 1\,000) \text{ mg}$

$m(\text{aldicarb}) = 0.693 \text{ mg}$ (1 MARK)

[By consuming 200 mL of the $1.82 \times 10^{-5} \text{ mol L}^{-1}$ aldicarb water sample, the toddler would have consumed a total of 0.693 mg.[4]][This does not exceed the safe consumption amount of 95 mg indicated by the guidelines.[5]]

✓ ✗ I have included calculations in my answer.

✓ ✗ I have identified the total aldicarb consumed.[4]

✓ ✗ I have linked my answer to the question.[5]

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Key science skills

19. **a.** Random error

b. [Since there was some mass of NaCl lost during the experiment, there would be less NaCl than required.[1]] [Therefore the actual concentration of the solution would be less than the expected value.[2]]

✓ ✗ I have identified the effect of the error on the sample.[1]

✓ ✗ I have identified the effect of the error on the results.[2]

c. $n(\text{NaCl}) = \frac{m}{M}$

$n(\text{NaCl}) = \frac{4.00 \text{ g}}{(23.0 + 35.5) \text{ g mol}^{-1}}$

$n(\text{NaCl}) = 0.06838 \text{ mol}$ (1 MARK)

$c(\text{NaCl}) = \frac{n}{V}$

$c(\text{NaCl}) = \frac{0.06838 \text{ mol}}{\frac{250 \text{ mL}}{1000}}$

$c(\text{NaCl}) = 0.274 \text{ mol L}^{-1}$ (1 MARK)

FROM LESSON 16D

Questions from multiple lessons

20. **a.** Ionic compound

b. $\text{KCl(s)} \rightarrow \text{K}^+\text{(aq)} + \text{Cl}^-\text{(aq)}$

c. $n(\text{KCl}) = c \times V$

$n(\text{KCl}) = 2.5 \text{ mol L}^{-1} \times \frac{100 \text{ mL}}{1\,000}$

$n(\text{KCl}) = 0.25 \text{ mol}$ (1 MARK)

$m(\text{KCl}) = n \times M$

$m(\text{KCl}) = 0.25 \text{ mol} \times (39.1 + 35.5) \text{ g mol}^{-1}$

$m(\text{KCl}) = 18.7 \text{ g}$ (1 MARK)

d. $\text{AgNO}_3\text{(aq)} + \text{KCl(aq)} \rightarrow \text{AgCl(s)} + \text{KNO}_3\text{(aq)}$

FROM LESSONS 4B & 5A

12B Predicting solubility

Progress questions

1. B. Solubility is the extent to which a solute is able to dissolve in a solvent.

2. B. Solubility tables are used to predict the solubility of ionic compounds.

3. A. The solubility of gaseous solutes and some solid solutes decreases as the temperature increases.

4. B. Increasing the temperature tends to decrease the solubility of gases.

5. A. The solubility of KClO_3 increases exponentially with temperature, but the solubility of $\text{Pb(NO}_3)_2$ increases linearly with temperature.

6. B. A saturated solution is unable to dissolve any more solute.

7. A. Crystallisation takes place when a supersaturated solution is disturbed or when a saturated solution is cooled.

8. A. Unsaturated solutions are capable of dissolving more solute.

9. B. The solubility curve represents a saturated solution. Therefore, areas above the curve indicate supersaturated solutions.

10. A. Eutrophication occurs when nutrients accumulated in bodies of water cause excessive growth of algae.

11. A. Fertiliser runoff contains phosphates.

Deconstructed

12. C

13. A

14. [At 90 °C, the solubility of $\text{K}_2\text{Cr}_2\text{O}_7$ is 70 g per 100 g of H_2O. As there are 200 g of H_2O in the solution, 140 g of $\text{K}_2\text{Cr}_2\text{O}_7$ will be found in a saturated solution.[1]] [In a supersaturated solution, the amount of solute dissolved is greater than the amount found in a saturated solution for a given amount of solvent.[2]] [Given there is 180 g of $\text{K}_2\text{Cr}_2\text{O}_7$ dissolved in the 200 g of solvent, and only 140 g of $\text{K}_2\text{Cr}_2\text{O}_7$ can be dissolved in a saturated solution, there will be 40 g of crystals formed when the solution is disturbed.[3]]

✓ ✗ I have identified the amount of $\text{K}_2\text{Cr}_2\text{O}_7$ present in a saturated solution.[1]

✓ ✗ I have defined a supersaturated solution.[2]

✓ ✗ I have identified the mass of crystals formed from this solution.[3]

Exam-style

15. **a.** The most common trend between temperature and solubility is an increase in solubility as temperature increases.

b. Unlike the other compounds on the graph, the solubility of $\text{Ce}_2(\text{SO}_4)_3$ decreases as the temperature increases.

c. **i.** Unsaturated

ii. Saturated

iii. Saturated

iv. Supersaturated

v. Supersaturated

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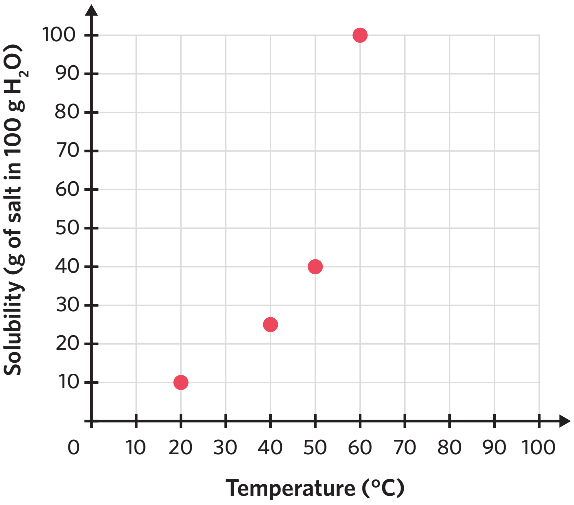
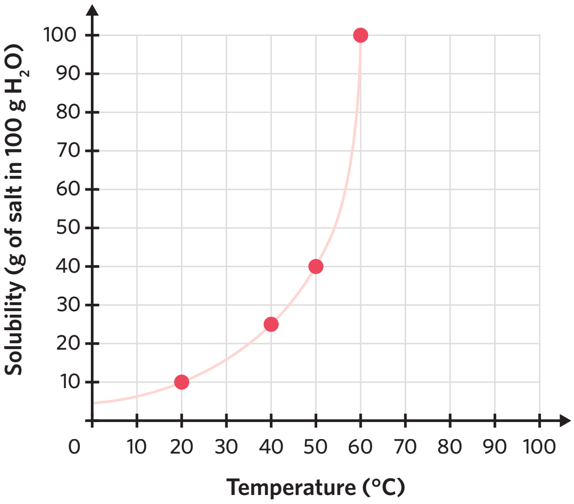
16. a.



✓ ✗ I have labelled the axes appropriately.^{1}

✓ ✗ I have plotted the values on the graph accurately.^{2}

b.



c. [The solubility curve increases exponentially.^{1}][This implies that, as the temperature increases, the solubility of the solute will increase exponentially.^{2}]

✓ ✗ I have commented on the shape of the solubility curve.^{1}

✓ ✗ I have identified the implication of the solubility curve's shape.^{2}

17. a. Eutrophication

b. [Alum can be added to water to produce aluminium hydroxide.^{1}][Aluminium hydroxide then reacts with phosphates in the water, forming hydroxides and an insoluble aluminium phosphate precipitate.^{2}][This precipitate settles in the lake bed and can no longer be used as a nutrient to support the growth of algae.^{3}]

✓ ✗ I have identified the formation of aluminium hydroxide from alum and water.^{1}

✓ ✗ I have explained how an aluminium phosphate precipitate is formed.^{2}

✓ ✗ I have explained that the aluminium phosphate precipitate can not be used to support algal growth.^{3}

Key science skills

18. a. Random error

b. [Variation in the position of the thermometer in the beaker.^{1}]
[Variation in the volume of the water (ratio of solute to solvent).^{2}]

✓ ✗ I have identified one possible source of error.^{1}

✓ ✗ I have identified another possible source of error.^{2}

c. [As the results are all centred around the theoretical result, the results are accurate.^{1}][However, as the results are all different to each other, they are not precise.^{2}]

✓ ✗ I have commented on the accuracy of the results.^{1}

✓ ✗ I have commented on the precision of the results.^{2}

FROM LESSONS 16C & 16D

Questions from multiple lessons

19. a. i. A solute is a substance that is dissolved in a solution.

ii. A solvent is a substance in which a solute is dissolved.

b. [Ethanol will have a higher solubility in water than ethane.^{1}]
[This is because ethanol possesses a polar hydroxyl group, —OH, enabling it to form strong intermolecular hydrogen bonds with water, and therefore dissolve in water.^{2}][In comparison, ethane is non-polar and only forms weaker dispersion forces. Consequently, ethane is unable to overcome the intermolecular forces between water molecules, making it less soluble in water.^{3}]

✓ ✗ I have compared the solubility of ethanol and ethane in water.^{1}

✓ ✗ I have described the intermolecular forces of ethanol and how they relate to dissolution.^{2}

✓ ✗ I have described the intermolecular forces of ethane and how they relate to dissolution.^{3}

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c. [The solubility curve for ethane has an expected relationship between temperature and solubility.[1]] [As ethane has a boiling point of −89 °C, the molecule will be gaseous for all the temperatures shown on the graph.[2]] [As the general trend for the solubility of gaseous molecules is decreasing as the temperature increases, the solubility curve of ethane has a conventional shape.[3]]

✓ ✗ I have identified the trend of the solubility curve of ethane.[1]

✓ ✗ I have described the state of ethane based on the conditions of the solubility curve.[2]

✓ ✗ I have described the common relationship between temperature and solubility for gaseous molecules.[3]

d. **i.** According to the graph, 30 g of $K_2Cr_2O_7$ dissolves in 100 g of H_2O at 50 °C.

Therefore, in 500 g of H_2O:

$30\text{ g} \times \frac{500\text{ g}}{100\text{ g}} = 150$ g would dissolve at 50 °C

ii. $n(K_2Cr_2O_7) = \frac{m}{M}$

$n(K_2Cr_2O_7) = \frac{150\text{ g}}{(2 \times 39.1 + 2 \times 52.0 + 7 \times 16.0)\text{ g mol}^{-1}}$

$n(K_2Cr_2O_7) = 0.510\text{ mol}$

iii. 500 g of water = 500 mL of water

500 mL of water = 0.500 L of water (1 MARK)

$c(K_2Cr_2O_7) = \frac{m}{M}$

$c(K_2Cr_2O_7) = \frac{0.51\text{ mol}}{0.500\text{ L}}$

$c(K_2Cr_2O_7) = 1.02\text{ mol L}^{-1}$ (1 MARK)

FROM LESSONS 5A, 6B, 7A & 12A

Chapter 12 review

Multiple choice

1. A. Stirring a solution does not change the solubility of a solute but instead only increases the rate at which the solute dissolves.

2. D. $c = \frac{m}{V}$

$c = \frac{400\text{ g}}{2.00\text{ L}}$

$c = 200\text{ g L}^{-1}$

$c = 2.00 \times 10^5\text{ mg L}^{-1}$

$c = 2.00 \times 10^5\text{ ppm}$

3. D. Acid rain is not a source of phosphate in large bodies of water.

4. D. 0.00130 %(m/v) = 0.00130 g/100 mL

Dilution factor	30 × 0.00130 g = 0.0390 g/100 mL
Convert to mg	1000 × 0.0390 g = 39 mg/100 mL
Convert to µg	1000 × 39 mg = 39 000 µg/100 mL
Convert to $\mu\text{g mL}^{-1}$	39 000 µg ÷ 100 mL = $390\ \mu\text{g mL}^{-1}$

5. B. From the curve, KNO_3 has a solubility of 30 g per 100 g of H_2O at 20 °C and a solubility of 65 g per 100 g of H_2O at 40 °C.

6. D. Supersaturated solutions have more dissolved solute than in saturated solutions.

7. D. $\frac{250\text{ mg L}^{-1}}{1000} = 0.25\text{ g L}^{-1}$

8. C. $M(Ca(OH)_2) = (40.1 + (16.0 \times 2) + (1.0 \times 2))\text{ g mol}^{-1}$

$M(Ca(OH)_2) = 74.1\text{ g mol}^{-1}$

$n(Ca(OH)_2) = \frac{m}{M}$

$n(Ca(OH)_2) = \frac{0.25\text{ g}}{74.1\text{ g mol}^{-1}}$

$n(Ca(OH)_2) = 3.37 \times 10^{-3}\text{ mol}$

Rearrange $n = c \times V$ to $V = \frac{n}{c}$

$V((CaOH)_2) = \frac{3.37 \times 10^{-3}\text{ mol}}{0.25\text{ M}}$

$V((CaOH)_2) = 1.35 \times 10^{-3}\text{ L}$

$V((CaOH)_2) = 13\text{ mL}$

9. C. Solubility curves are generated by recording the point at which crystals start to form when cooling a solution down.

10. B. $n(H_2SO_4) = c \times V$

$n(H_2SO_4) = 0.25\text{ mol L}^{-1} \times \frac{50}{1000}\text{ L}$

$n(H_2SO_4) = 0.0125\text{ mol}$

$n(H_2SO_4) : n(NaOH)$

1 : 2

$n(NaOH) = 0.025\text{ mol}$

$V = \frac{n}{c}$

$V = \frac{0.025\text{ mol}}{0.50\text{ mol L}^{-1}}$

$V = 0.050\text{ L}$

$V = 50\text{ mL}$

or, $c_1 \times V_1 = c_2 \times V_2$

$0.25\text{ mol L}^{-1} \times 50\text{ mL} = 0.5\text{ mol L}^{-1} \times V_2$

$V_2 = 25\text{ mL}$

1 mole of sulfuric acid reacts with 2 moles of sodium hydroxide.

$V_2 = 50\text{ mL}$

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Short answer

11. **a.** A supersaturated solution is a solution that contains more than the normal amount of solute dissolved at a given temperature.

b. A concentration greater than approximately 37.8 g of NaCl per 100 mL of water is required to create a supersaturated solution at 80 °C.

c. $c(\text{NaCl}) = 37.0 \text{ g NaCl per 100 mL of } H_2O$ (1 MARK)

$n(\text{NaCl}) = \frac{m}{M}$

$n(\text{NaCl}) = \dfrac{37.0 \text{ g}}{(23.0 + 35.5) \text{ g mol}^{-1}}$

$n(\text{NaCl}) = 0.6325 \text{ mol per 100 mL of } H_2O$ (1 MARK)

$c(\text{NaCl}) = \frac{n}{V}$

$c(\text{NaCl}) = \dfrac{0.6325 \text{ mol}}{\frac{100}{1000} \text{L}}$

$c(\text{NaCl}) = 6.33 \text{ mol L}^{-1}$ (1 MARK)

d. $n(\text{NaCl}) = c \times V$

$n(\text{NaCl}) = 6.33 \text{ mol L}^{-1} \times \frac{193}{1000} \text{L}$

$n(\text{NaCl}) = 1.22 \text{ mol}$ (1 MARK)

$m(\text{NaCl}) = n \times M$

$m(\text{NaCl}) = 1.22 \text{ mol} \times (23.0 + 35.5) \text{ g mol}^{-1}$

$m(\text{NaCl}) = 71.4 \text{ g}$ (1 MARK)

12. **a.** $2.0 \text{ ppm} = 2.0 \text{ mg L}^{-1}$

$2.0 \text{ mg L}^{-1} = 2.0 \times 10^{-3} \text{ g L}^{-1}$ (1 MARK)

$n(\text{Cl}^-) \text{ in } 1.0 \text{ L} = \frac{m}{M}$

$n(\text{Cl}^-) \text{ in } 1.0 \text{ L} = \dfrac{2.0 \times 10^{-3} \text{ g}}{35.5 \text{ g mol}^{-1}}$

$n(\text{Cl}^-) \text{ in } 1.0 \text{ L} = 5.63 \times 10^{-5} \text{ mol}$ (1 MARK)

$n(\text{KCl}) \text{ in } 1.0 \text{ L} = n(\text{Cl}^-) \text{ in } 1.0 \text{ L}$

Therefore, $n(\text{KCl}) \text{ in } 1.0 \text{ L} = 5.63 \times 10^{-5} \text{ mol}$ (1 MARK)

$m(\text{KCl}) = n \times M$

$m(\text{KCl}) = 5.63 \times 10^{-5} \text{ mol} \times (35.5 + 39.1) \text{ g mol}^{-1}$

$m(\text{KCl}) = 4.2 \times 10^{-3} \text{ g}$ (1 MARK)

b. $2.0 \text{ ppm} = 2.0 \text{ mg L}^{-1}$

$2.0 \text{ mg L}^{-1} = 2.0 \times 10^{-3} \text{ g L}^{-1}$

$2.0 \times 10^{-3} \text{ g L}^{-1} = 2.0 \times 10^{-6} \text{ kg L}^{-1}$ (1 MARK)

$2.0 \times 10^{-6} \text{ kg L}^{-1} = 2.0 \text{ kg ML}^{-1}$ (1 MARK)

$m(\text{KCl}) = c \times V$

$m(\text{KCl}) = 2.0 \text{ kg ML}^{-1} \times 900 \text{ ML}$

$m(\text{KCl}) = 1800 \text{ kg}$ (1 MARK)

c. A solid

d. 60 °C

e. $c_1 = 2.0 \text{ mol L}^{-1}$

$V_1 = 1.5 \text{ L}$

$c_2 = ?$

$V_2 = 1.5 \text{ L} + \frac{750}{1000} \text{L}$

$V_2 = 2.25 \text{ L}$ (1 MARK)

$c_1 \times V_1 = c_2 \times V_2$

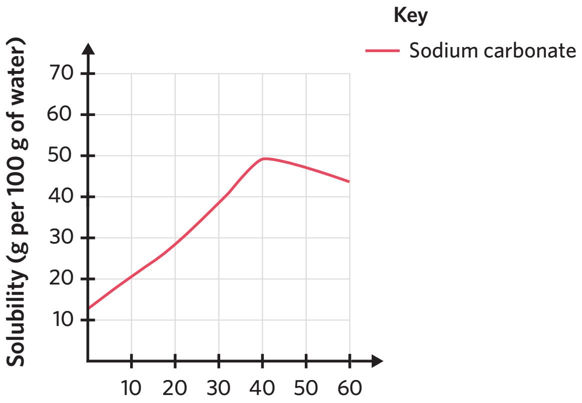
$2.0 \text{ mol L}^{-1} \times 1.5 \text{ L} = c_2 \times 2.25 \text{ L}$ (1 MARK)

$c_2 = \dfrac{2.0 \text{ mol L}^{-1} \times 1.5 \text{ L}}{2.25 \text{ L}}$

$c_2 = 1.3 \text{ mol L}^{-1}$ (1 MARK)

13. **a.** The solubility of sodium carbonate increases as the temperature increases from 0 °C to 40 °C, and then decreases as the temperature increases from 40 °C to 60 °C.

b.



✓ ✗ I have drawn an appropriate curve that has a peak solubility somewhere between 40 - 50 °C.[1]

✓ ✗ I have used an appropriate scale for the graph.[2]

✓ ✗ I have labelled the axes of the graph appropriately.[3]

c. At 40 °C, 49 g of Na_2CO_3 can be dissolved in 100 g of H_2O.

At 20 °C, 29 g of Na_2CO_3 can be dissolved in 100 g of H_2O. (1 MARK)

49 g - 29 g = 20 g of crystals would form if the volume of H_2O was 100 g. (1 MARK)

Volume of H_2O = 300 g

Therefore, $\frac{300 \text{ g}}{100 \text{ g}} \times 20 \text{ g} = 60 \text{ g}$ crystals would form. (1 MARK)

d. From most soluble to least soluble at 50 °C : NaOH, $NaBrO_3$.

e. At 60 °C, 64 g of $NaBrO_3$ can dissolve in 100 g of H_2O. (1 MARK)

Volume of H_2O = 200 g

Therefore, $\frac{200 \text{ g}}{100 \text{ g}} \times 64 \text{ g} = 128 \text{ g}$ of $NaBrO_3$ can dissolve in 200 g of H_2O. (1 MARK)

$m(\text{Undissolved NaBrO}_3) =$
$m(\text{NaBrO}_3 \text{ added}) - m(\text{NaBrO}_3 \text{ dissolved})$

$m(\text{Undissolved NaBrO}_3) = 200 \text{ g} - 128 \text{ g}$

$m(\text{Undissolved NaBrO}_3) = 72 \text{ g}$ (1 MARK)

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Key science skills

14. a. [The independent variable is the temperature of the solution.[1]] [The dependent variable is the solubility of helium gas.[2]]

✓ ✗ I have correctly identified the independent variable.[1]

✓ ✗ I have correctly identified the dependent variable.[2]

b. [Pressure of helium or amount of helium.[1]] [Volume of water or surface area of water.[2]]

✓ ✗ I have identified one variable that should be controlled.[1]

✓ ✗ I have identified a second variable that should be controlled.[2]

c. [According to the SDS, helium gas may explode.[1]] [Therefore, it is appropriate to keep it in a sealed vessel and away from any external sources of ignition or heat.[2]]

✓ ✗ I have interpreted the SDS.[1]

✓ ✗ I have identified some safety considerations.[2]

d. [A solution is supersaturated when the amount of dissolved solute exceeds that of the solubility curve at a given temperature.[1]] [Given 0.0016 g of helium can dissolve in a saturated solution of 1 kg of H_2O at 10 °C, and given the solubility of helium in water decreases with increasing temperature, 0.0016 g of helium in 1 kg of H_2O would constitute a supersaturated solution at temperatures above 10 °C.[2]]

✓ ✗ I have explained what a supersaturated solution is.[1]

✓ ✗ I have identified the correct temperature range.[2]

e. [At higher temperatures, gas molecules have more kinetic energy.[1]] [This means that the gas molecules are able to move about more and escape the solution, decreasing their solubility at higher temperatures.[2]]

✓ ✗ I have identified the effect of temperature on kinetic energy.[1]

✓ ✗ I have explained why gas solubility decreases with temperature.[2]

FROM LESSONS 16A & 16B

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13A Volumetric analysis

Progress questions

1. B. The appropriate unit of concentration is mol L^{-1}

2. A. The equation $V = \frac{n}{c}$ can be used to find volume

3. A. The purpose of a primary standard is to create a solution of known concentration.

4. A. Primary standard solutions are solutions of a precisely known concentration.

5. B. Strong acids and bases like NaOH and HCl are very reactive with the atmosphere and the surrounding environment.

6. B. The meniscus is the curvature of the liquid due to the difference in the forces of attraction between solvent molecules and the walls of the burette.

7. B. Concordant titres are titres that differ by 0.10 mL from the highest value to the lowest value.

8. B. Burettes and conical flasks are designed to suit their functions.

9. B. The end point indicates the approximate point at which the stoichiometric ratios of the reactants are met, noted by the colour change of the indicator.

10. B. Indicators make changes in the pH of a solution visible.

11. B. The equivalence point on a titration curve is where the titre added is just enough to neutralise the unknown solution.

12. B. The choice of indicator is based on having the end point and equivalence point of a titration as close to each other as possible.

13. B. HCl and NaOH react in a 1:1 ratio, so the number of moles of NaOH reacting will be equal to that of HCl. From this, concentration of HCl can be calculated.

14. A. Rinsing the burette with water will reduce the $c(\text{HCl})$ added and so increase the $V(\text{HCl})$ in the titre, leading to a lower calculated $c(\text{HCl})$.

Deconstructed

15. A

16. D

17. Average titre $= \frac{20.00\text{ mL} + 20.05\text{ mL} + 19.95\text{ mL}}{3}$

Average titre = 20.00 mL

$n(\text{HCl}) = c \times V$

$n(\text{HCl}) = 1.0\text{ mol L}^{-1} \times 0.02000\text{ L}$

$n(\text{HCl}) = 0.020\text{ mol}$ (1 MARK)

$n(\text{HCl}) : n(\text{NaOH})$

1 : 1

$n(\text{NaOH}) = n(\text{HCl})$

$n(\text{NaOH}) = 0.020\text{ mol}$ (1 MARK)

$c(\text{NaOH}) = \frac{n}{V}$

$c(\text{NaOH}) = \frac{0.020\text{ mol}}{\left(\frac{20.00}{1000}\right)\text{L}}$

$c(\text{NaOH}) = 1.0\text{ mol L}^{-1}$ (1 MARK)

The concentration of the sodium hydroxide present in SutherClean^{TM} is 1.0 mol L^{-1}.

Exam-style

18. a. Number of wasps: $\frac{20.00\text{ mL}}{0.30\text{ mL}}$ per wasp = 67 wasps

b. The purpose of the indicator is to signify the end point of the titration through a colour change.

c. [These are concordant titres.[1]][This is because the difference between the highest value (22.15 mL) and lowest value (22.10 mL) is less than 0.10 mL.[2]]

✓ ✗ I have stated whether the titres are concordant or not.[1]

✓ ✗ I have justified my answer with reference to the definition of concordant titres.[2]

d. Average titre $= \frac{22.10\text{ mL} + 22.15\text{ mL} + 22.10\text{ mL}}{3}$

Average titre = 22.12 mL

$n(\text{HCl}) = 0.0010\text{ mol L}^{-1} \times \left(\frac{22.12}{1000}\right)\text{L}$

$n(\text{HCl}) = 2.212 \times 10^{-5}\text{ mol}$ (1 MARK)

$Mg(OH)_2(aq) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$

$n(Mg(OH)_2) : n(\text{HCl})$

1 : 2

$n(Mg(OH)_2) = \frac{1}{2} \times (2.122 \times 10^{-5}\text{ mol})$

$n(Mg(OH)_2) = 1.106 \times 10^{-5}\text{ mol}$ (1 MARK)

$c(Mg(OH)_2) = \frac{n}{V}$

$c(Mg(OH)_2) = \frac{1.106 \times 10^{-5}\text{ mol}}{\left(\frac{20.00}{1000}\right)\text{L}}$

$c(Mg(OH)_2) = 5.53 \times 10^{-4}\text{ mol L}^{-1}$ (1 MARK)

e. $n(Mg(OH)_2) = c \times V$

$n(Mg(OH)_2) = 5.53 \times 10^{-4}\text{ mol L}^{-1} \times \left(\frac{0.30}{1000}\right)\text{L}$

$n(Mg(OH)_2) = 1.659 \times 10^{-7}\text{ mol}$

$n(\text{HCl}) : n(Mg(OH)_2)$

2 : 1

$n(\text{HCl}) = 2 \times (1.659 \times 10^{-7}\text{ mol})$

$n(\text{HCl}) = 3.318 \times 10^{-7}\text{ mol}$ (1 MARK)

$V(\text{HCl}) = \frac{3.318 \times 10^{-7}\text{ mol}}{0.0010\text{ mol L}^{-1}}$

$V(\text{HCl}) = 3.318 \times 10^{-4}\text{ L}$

$V(\text{HCl}) = 0.33\text{ mL}$ (1 MARK)

19. a. Traditional methods of pH measurement require aqueous solutions rather than oils.

b. [The end point of the titration occurs when the indicator in the conical flask changes colour.[1]][In this case, it is when bromothymol blue transitions from a yellow to blue colour.[2]][The equivalence point is when there is the exact stoichiometric ratio of acid to base present in the conical flask.[3]][In this case, it is when the vegetable oil solution's acidity has been neutralised by the NaOH.[4]]

- ✓ ✗ I have given the definition of the end point.[1]
- ✓ ✗ I have linked the definition of the end point to the experiment.[2]
- ✓ ✗ I have given the definition of the equivalence point.[3]
- ✓ ✗ I have linked the definition of the equivalence point to the experiment.[4]

c. [Primary standard solutions are solutions of accurately known concentrations. [1]][They are used to standardise solutions of NaOH because NaOH is not a primary standard, and therefore, when it is made up into a solution, an accurate value for the concentration is not able to be calculated. [2]]

- ✓ ✗ I have defined a primary standard solution.[1]
- ✓ ✗ I have explained the importance of primary standard solutions in standardising a solution of NaOH.[2]

20. a. [Adding more indicator would decrease the accuracy of the experiment.[1]][Since indicators are weak acids and bases, they react with the known and unknown solutions in an acid-base titration.[2]][Since all the indicator has to react in order for the colour change to be observed, adding more indicator delays the onset of the end point, even when the equivalence point has been reached.[3]]

- ✓ ✗ I have identified the impact on adding more indicator to the accuracy of the experiment.[1]
- ✓ ✗ I have explained the fact that indicators react with the known and unknown solutions in a reaction.[2]
- ✓ ✗ I have linked the effect of this to the equivalence point and end point.[3]

b. [The best choice of indicator is when the end point, signalled by the change in colour, occurs as close as possible to the equivalence point.[1]][Since the equivalence point is 6.8, bromothymol blue will be the best choice of indicator to use.[2]] [This is because its end point pH range of 6.0–7.6 and colour transition from yellow to blue matches the equivalence point of the titration reaction.[3]]

- ✓ ✗ I have defined the purpose of the indicator.[1]
- ✓ ✗ I have identified the most suitable indicator.[2]
- ✓ ✗ I have justified my answer with reference to the equivalence point and the end point.[3]

c. [The 5 indicators mentioned above have different colours and have different pH ranges for the transition.[1]] [Therefore, a rainbow of colours can be created from the mixtures of different colours and the different pH ranges that they transition in.[2]]

- ✓ ✗ I have identified the different colours of the indicators present and their different pH ranges.[1]
- ✓ ✗ I have explained how the spectrum of different colours could arise.[2]

21. a. [Primary standards are required to have no water present to eliminate errors arising from weighing.[1]][Therefore, the sodium carbonate has to be dried (for example, by heating) before its use to make a primary standard solution.[2]]

- ✓ ✗ I have explained why no water can be present in primary standards.[1]
- ✓ ✗ I have identified a strategy to reduce the water content of the sodium carbonate.[2]

b. Unreactive with the atmosphere, readily available and soluble in solution.

c. A volumetric flask is a highly accurate piece of glassware used to create a solution of accurate volume.

d. n(potassium phthalate) required for a 250.00 mL sample at 1.000 mol L^{-1}

$n(\text{potassium phthalate}) = c \times V$

$n(\text{potassium phthalate}) = 1.0 \text{ mol L}^{-1} \times \left(\frac{250.00}{1000}\right) \text{ L}$

$n(\text{potassium phthalate}) = 0.2500 \text{ mol}$ (1 MARK)

$m(\text{potassium phthalate}) = n(\text{potassium phthalate}) \times M(\text{potassium phthalate})$

$m(\text{potassium phthalate}) = 0.2500 \text{ mol} \times 204.2 \text{ g mol}^{-1}$

$m(\text{potassium phthalate}) = 51.05 \text{ g}$ (1 MARK)

1. Accurately weigh 51.05 g potassium hydrogen phthalate. (1 MARK)
2. Transfer the potassium hydrogen phthalate into a 250.00 mL volumetric flask. (1 MARK)
3. Add deionised water into the volumetric flask to approximately halfway. (1 MARK)
4. Swirl the volumetric flask to dissolve the solute completely. (1 MARK)
5. Add enough deionised water to make the solution to 250.00 mL. (1 MARK)

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Key science skills

22. a. [Washing the burette with water results in some residual water being left in the burette.[1]][This means that the volume of HCl in the burette delivered will be less than what the burette says due to the water diluting the solution.[2]][Consequently, a greater volume of this diluted HCl from the burette will be required, leading to the calculated concentration of NaOH being greater than the true concentration.[3]]

✓ ✗ I have identified the impact of washing the burette with water.[1]

✓ ✗ I have described the effect of washing with water to the concentration of HCl delivered by the burette.[2]

✓ ✗ I have described the effect of washing with water on the accuracy of the calculated concentration of NaOH.[3]

b. [The conical flask can have residual water because it is not the concentration of the unknown solution that is important in the flask, but the amount of solute.[1]][Therefore, the conical flask cannot be washed with the unknown solution because residual NaOH could be left, leading to the calculated concentration of NaOH being greater than the actual concentration.[2]]

✓ ✗ I have explained why water can be used to wash the conical flask.[1]

✓ ✗ I have explained why NaOH cannot be used to wash the conical flask.[2]

c. [PPE should be worn, such as gloves and eye protection.[1]] [The lab should have easily accessible dilute bases that can be used to neutralise the acid.[2]][The broken glass is safely cleaned up.[3]]

✓ ✗ I have identified one safety precaution that could be taken.[1]

✓ ✗ I have identified a second safety precaution that could be taken.[2]

✓ ✗ I have identified a third safety precaution that could be taken.[3]

FROM LESSONS 16B & 16D

Questions from multiple lessons

23. a. $CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$

b. Carboxyl group

c. Titration numbers 1, 3 and 5 produced the concordant titres.

$$\text{Average titre} = \frac{21.55\text{ mL} + 21.60\text{ mL} + 21.65\text{ mL}}{3}$$

$$\text{Average titre} = 21.60\text{ mL}$$

d. $n(NaOH) = 1.0\text{ mol L}^{-1} \times 0.02160\text{ L}$

$n(CH_3COOH) = n(NaOH)$

$n(CH_3COOH) = 0.0216\text{ mol}$ (1 MARK)

$$c(CH_3COOH) = \frac{0.0216\text{ mol}}{\left(\frac{25.00}{1000}\right)\text{ L}}$$

$c(CH_3COOH) = 0.864\text{ mol L}^{-1}$ (1 MARK)

e. $c(CH_3COOH) = 0.864\text{ mol L}^{-1}$

$c(CH_3COOH) = 0.864\text{ mol L}^{-1} \times ((2 \times 12.0\text{ g mol}^{-1})$

$+ (4 \times 1.0\text{ g mol}^{-1}) + (2 \times 16.0\text{ g mol}^{-1}))$

$c(CH_3COOH) = 51.84\text{ g L}^{-1}$

This means there are 51.84 grams of ethanoic acid per litre of vinegar in the sample (1 MARK)

$$\frac{51.84\text{ g}}{1.05\text{ g mL}^{-1}} = 49.37\text{ mL}$$

$$\%\ (v/v) = \frac{49.37\text{ mL}}{1000\text{ mL}} \times 100\%$$

$\%\ (v/v) = 4.9\%$ (1 MARK)

Therefore, the vinegar will meet the regulatory standards. (1 MARK)

FROM LESSONS 7B & 12A

24. a. A polyprotic acid is an acid that can donate multiple protons.

b. [There are two equivalence points relating to both deprotonation events.[1]][Since a diprotic acid can donate a proton twice, there will be the first equivalence point when all the diprotic acid has donated the first proton, and a second equivalence point when all the deprotonated diprotic acid has donated its second proton.[2]]

✓ ✗ I have referenced the ability for diprotic acids to donate multiple protons.[1]

✓ ✗ I have explained this observation with reference to equivalence points.[2]

c. [For the first equivalence point, the indicator methyl red could be used because the first equivalence point lies within methyl red's pH range (4.4–6.2).[1]][For the second equivalence point, the indicator phenolphthalein (pH range of 8.3–10.0) could be used because it transitions from colourless to pink in the pH range of the second equivalence point.[2]]

✓ ✗ I have identified a suitable indicator for the first equivalence point.[1]

✓ ✗ I have identified a suitable indicator for the second equivalence point.[2]

d. $n(\text{base}) = 0.01600\text{ L} \times 1.0\text{ mol L}^{-1}$

$n(\text{base}) = 0.01600\text{ mol}$

$n(\text{diluted acid}) = n(\text{base})$

$n(\text{diluted acid}) = 0.01600\text{ mol}$ (1 MARK)

$$c(\text{diluted acid}) = \frac{n}{V}$$

$$c(\text{diluted acid}) = \frac{0.01600\text{ mol}}{0.02000\text{ L}}$$

$c(\text{diluted acid}) = 0.80\text{ mol L}^{-1}$ (1 MARK)

$c(\text{undiluted acid}) = \text{dilution factor} \times c(\text{diluted acid})$

$c(\text{undiluted acid}) = 10 \times 0.80\text{ mol L}^{-1}$

$c(\text{undiluted acid}) = 8.0\text{ mol L}^{-1}$ (1 MARK)

FROM LESSON 10A

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Chapter 13 review

Multiple choice

1. B. The burette dispenses solution incrementally into the conical flask.

2. D. A high molecular weight makes the calculation of the number of moles more accurate.

3. A. Thymol blue changes colour at a slightly basic pH, which is appropriate for the reaction between a weak acid and a strong base.

4. A. Curve begins at very basic pH, and ends at slightly acidic pH. Equivalence point is greater than 7.

5. B. The solution of unknown concentration can be in the conical flask, or be dispensed from the burette.

6. D. All of the above are true – there are specific methods by which a burette should be read.

7. C. The only relevant factor in the volume of NaOH required for neutralisation is the availability of H^+ ions from each acid. As the two acids are present in equal volumes and concentration, the NaOH titres will be equal.

8. A. The presence of water in the conical flask would not affect the number of moles of NaOH in the flask, and therefore would not affect the titre volume.

9. D. According to the 2:1 mole ratio between HCl and $Ca(OH)_2$, 160 mL of HCl is required.

10. B. A higher titre volume is consistent with a higher acid concentration.

Short answer

11. a. [The point at which the reactants have reacted completely in accordance with the stoichiometric ratio of the balanced equation.[1]][The point at which the indicator undergoes a permanent colour change.[2]][Concordant titres.[3]]

✓ ✗ I have defined equivalence point.[1]

✓ ✗ I have defined end point.[2]

✓ ✗ I have identified the key term used.[3]

b. $n(\text{NaOH}) = c \times V$

$n(\text{NaOH}) = 1.0\text{ M} \times \frac{25}{1000}\text{ L}$

$n(\text{NaOH}) = 0.025\text{ mol}$

$n(\text{NaOH}) : n(C_2H_2O_4)$

$2 : 1$

$n(C_2H_2O_4) = \frac{1}{2} \times n(\text{NaOH})$

$n(C_2H_2O_4) = \frac{1}{2} \times 0.025\text{ mol}$

$n(C_2H_2O_4) = 0.0125\text{ mol}$ (1 MARK)

$c(C_2H_2O_4) = \frac{n}{V}$

$c(C_2H_2O_4) = \frac{0.0125\text{ mol}}{0.02740\text{ L}}$

$c(C_2H_2O_4) = 0.46\text{ M}$ (1 MARK)

12. a. Statements 3, 4, 7, 9 and 10 are correct.

✓ ✗ I have identified one correct statement.[1]

✓ ✗ I have identified a second correct statement.[2]

✓ ✗ I have identified a third correct statement.[3]

✓ ✗ I have identified a fourth correct statement.[4]

b. [Statement 1 is incorrect/Statement 2 is incorrect/Statement 5 is incorrect./Statement 6 is incorrect/Statement 8 is incorrect.[1]] [Statement 1– as the equivalence point and end point represents different reaction states, and so will sometimes occur at slightly different times.[2]][Statement 2 – as methyl orange undergoes a colour change at a low pH.[3]][Statement 5 – as 2 moles of KOH would release 2 moles of potassium ions and 2 moles of hydroxide ions.[4]][Statement 6 – as NaOH is highly reactive with the atmosphere, and so the solution would decrease in concentration over time.[5]][Statement 8 – as titre volumes within ± 0.1 mL are considered concordant, not ± 0.2 mL.[6]]

✓ ✗ I have identified one incorrect statement.[1]

✓ ✗ I have explained why this statement is incorrect.[2]

✓ ✗ I have identified a second incorrect statement.[3]

✓ ✗ I have explained why this statement is incorrect.[4]

✓ ✗ I have identified a third incorrect statement.[5]

✓ ✗ I have explained why this statement is incorrect.[6]

13. a. $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$ or $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

b. $n(\text{NaOH})\text{ in flask} = c \times V$

$n(\text{NaOH})\text{ in flask} = 0.0222\text{ M} \times \frac{900.0}{1000}\text{ L}$

$n(\text{NaOH})\text{ in flask} = 0.0200\text{ mol}$

$n(\text{HCl})\text{ reacted} = 0.0200\text{ mol}$

$n(\text{HCl})\text{ added to flask} = c \times V$

$n(\text{HCl})\text{ added to flask} = 1.00\text{ mol L}^{-1} \times \frac{100.0}{1000}\text{ L}$

$n(\text{HCl})\text{ added to flask} = 0.100\text{ mol}$

$n(\text{HCl})\text{ remaining} = 0.100\text{ mol} - 0.0200\text{ mol}$

$n(\text{HCl})\text{ remaining} = 0.0800\text{ mol}$ (1 MARK)

$c(\text{HCl}) = \frac{n}{V}$

$c(\text{HCl}) = \frac{0.0800\text{ mol}}{1.00\text{ L}}$

$c(\text{HCl}) = 0.0800\text{ mol L}^{-1}$ (1 MARK)

c. [Greater.[1]][A larger volume of the HCl solution will be required.[2]]

✓ ✗ I have indicated the change.[1]

✓ ✗ I have justified the change.[2]

14. a.

Glassware used	Rinse with water	Rinse with diluted vinegar solution	Rinse with NaOH solution
Volumetric flask	✓		
Burette			✓
20.00 mL pipette		✓	
Conical flask	✓		

b. So as not to use too large a volume of the NaOH solution (or to use and obtain a reasonably sized titre).

c. To minimise the impact of random errors on the accuracy of the results (or to ensure reliability).

d. $n(\text{NaOH}) = c \times V$

$n(\text{NaOH}) = 0.11 \text{ mol L}^{-1} \times 0.01535 \text{ L}$

$n(\text{NaOH}) = 0.0017 \text{ mol}$ (1 MARK)

$n(\text{acetic acid})$ in 20.00 mL of vinegar $= 0.0017$ mol (1 MARK)

$c(\text{diluted vinegar}) = \frac{n}{V}$

$c(\text{diluted vinegar}) = \frac{0.0017 \text{ mol}}{0.020 \text{ L}}$

$c(\text{diluted vinegar}) = 0.0844 \text{ mol L}^{-1}$ (1 MARK)

$c(\text{undiluted vinegar}) = c(\text{diluted vinegar}) \times \text{dilution factor}$

$c(\text{undiluted vinegar}) = 0.0844 \text{ mol L}^{-1} \times 10$

$c(\text{undiluted vinegar}) = 0.84 \text{ mol L}^{-1}$ (1 MARK)

Key science skills

15. a. [Calculation of the average titre was appropriate.[1]][It involved three concordant titres that differed by less than 0.1 mL / were within a range of 0.1 mL.[2]]

✓ ✗ I have identified that the calculation used was appropriate.[1]

✓ ✗ I have justified my answer with reference to concordant titres.[2]

b. [The impact of inappropriate rinsing of glassware, for example: – if the 500 mL volumetric flask was not rinsed with deionised water, impurities may have an impact on the $c(C_4H_6O_6)$, decreasing the accuracy and validity of the results – if the burette was rinsed with water instead of 0.1 M NaOH(aq), the c(NaOH) will be decreased, as will the accuracy and validity of results.[1]][Improper preparation of tartaric acid solution, for example: if the $C_4H_6O_6$ was not fully dissolved before the 500 mL flask was topped up to the calibration mark, some may remain undissolved, thus decreasing the accuracy and validity of the results or inadequate mixing of the 500 mL volumetric flask contents, leading to variation in the $c(C_4H_6O_6)$ in different aliquots and so decreasing the accuracy and validity of results.[2]][NaOH issues, for example: if the NaOH(aq) was not standardised prior to the titrations its concentration would be unreliable due to reaction with atmospheric CO_2, thus the decreasing accuracy of results or the c(NaOH) was only stated to one significant figure, the fewest significant figures, and hence the least accurate of all the data in the exercise. This reduces the accuracy and validity of results.[3]]

[Titration issues, for example: 3 concordant titres were obtained; this increases the accuracy of results or 4 drops of phenolphthalein were used in each titration; this increases the likelihood of getting a consistent endpoint colour and so increases the accuracy and validity of the results.4]

✓ ✗ I have explained how rinsing can affect the results.[1]

✓ ✗ I have explained how the preparation of the solution can affect the results.[2]

✓ ✗ I have explained how the nature of NaOH can affect the results.[3]

✓ ✗ I have explained how the indicator can affect the results.[4]

c. [The conclusion is based on the results of one experimental group, so even though concordant titres were obtained there may have been an error in either $c(C_6H_4O_6)$ or c(NaOH), which limits validity which could be checked by either collating results from a number of groups or repeating the procedure with fresh solutions or the percentage purity calculated depends on, and its validity is limited to some degree by, the accuracy of identification of the colour change at the endpoint. This could be remedied by using a digital pH meter to record endpoint pH.[1]][Percentage purity 92.5% is quoted to three significant figures, yet the c(NaOH) used in its calculation only has one significant figure, which could be addressed by standardising the NaOH to three significant figures or the sample tested may not have been truly representative of the manufacturer's product. This could be addressed by testing a sample from a different container or one student's results cannot be conclusive; there may have been a consistent error in the technique. This could be addressed by using a different analytical technique such as HPLC.[2]]

✓ ✗ I have identified a limitation of the student's conclusion.[1]

✓ ✗ I have explained how this limitation could be addressed.[2]

d. i. Preparing the $C_4H_6O_6$ solution: fume hood or breathing apparatus or equivalent/safety goggles/gloves

ii. Conducting the titration: safety goggles/gloves

FROM LESSONS 16B, 16D & 16E

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14A Gases

Progress questions

1. A Water molecules are able to absorb infrared radiation.

2. B. Infrared radiation is radiated heat energy.

3. B. Most of the infrared radiation from the sun escapes back into space.

4. B. To perform calculations, assumptions of gas behaviour have to be made, so ideal gases do not behave identically to real gases.

5. B. Pressure = force divided by area.

6. A. One newton per metre squared (Nm^{-2}) is equal to one Pascal (Pa).

7. B. SLC are fixed at 298 K and 100 kPa.

8. A. 0.987 atm is equal to 100 kPa.

9. A. Under the same conditions, the same molar amount of any two gases will occupy the same volumes.

10. A. Under the same conditions, the same molar amount of any two gases will occupy the same volumes.

11. B. The equation $n = \frac{m}{M}$ can be substituted into the universal gas equation $pV = nRT$.

Deconstructed

12. D

13. C

14. Temperature = 25 °C

Temperature = (25 + 273) K

Temperature (in K) = 298 K (1 MARK)

$$p = \frac{nRT}{V}$$

$$p = \frac{0.0455 \text{ mol} \times 8.31 \times 298 \text{ K}}{50 \text{ L}}$$

$p = 2.3$ kPa (1 MARK)

Exam-style

15. [Farming livestock.[1]][Burning fossil fuels.[2]]

✓ ✗ I have identified an activity for methane production.[1]

✓ ✗ I have identified an activity for carbon dioxide production.[2]

16. [The carbon dioxide molecule absorbs infrared radiation from the sun.[1]][The carbon dioxide molecule re-emits the infrared radiation as heat, some of which is re-emitted back towards Earth.[2]][Increased levels of CO_2 means more heat that would otherwise escape into space is trapped on Earth, warming the Earth.[3]]

✓ ✗ I have described how infrared radiation arrives in our atmosphere.[1]

✓ ✗ I have explained what happens to this absorbed infrared radiation.[2]

✓ ✗ I have explained how increased levels of carbon dioxide influence the Earth's temperature.[3]

17. A. One mole of any gas occupies the same volume at SLC.

18. $n(Cl_2) = \frac{m}{M}$

$$n(Cl_2) = \frac{10.0 \text{ g}}{71.0 \text{ g}}$$

$n(Cl_2) = 0.1408$ mol (1 MARK)

$$M = \frac{m}{n}$$

$$M(\text{unknown gas}) = \frac{2.85 \text{ g}}{0.1408 \text{ mol}}$$

$M(\text{unknown gas}) = 20.2 \text{ g mol}^{-1}$ (1 MARK)

Neon (1 MARK)

19. a. Volume: 250 mL = 0.250 L (1 MARK)

Temperature: 65.0 °C, (65.0 + 273) = 338 K (1 MARK)

Number of moles: $\frac{4.60 \text{ g}}{28.0 \text{ g mol}^{-1}} = 0.164$ mol (1 MARK)

Convert $PV = nRT$ to $p = \frac{nRT}{V}$

$$p = \frac{0.164 \text{ mol} \times 8.31 \times 338 \text{ K}}{0.250 \text{ L}}$$

$p = 1.845 \times 10^3$ kPa (1 MARK)

b. $(1.845 \times 10^3 \text{ kPa}) \times \frac{0.987 \text{ atm}}{100 \text{ kPa}} = 18.2$ atm

20. a. $n(CO_2) = \frac{m}{M}$

$$n(CO_2) = \frac{5.00 \text{ g}}{44.0 \text{ g mol}^{-1}}$$

$n(CO_2) = 0.1136$ mol

$V = n \times V_m$

$V = 0.1136 \text{ mol} \times 24.8 \text{ L mol}^{-1}$

$V = 2.82$ L (1 MARK)

$$n(SO_2) = \frac{5.00 \text{ g}}{64.0 \text{ g mol}^{-1}}$$

$n(SO_2) = 0.07813$ mol

$V = n \times V_m$

$V = 0.0781 \text{ mol} \times 24.8 \text{ L mol}^{-1}$

$V = 1.94$ L (1 MARK)

$$n(NO_2) = \frac{5.00 \text{ g}}{28.0 \text{ g mol}^{-1}}$$

$n(NO_2) = 0.1786$ mol

$V = n \times V_m$

$0.1786 \times 24.8 \text{ L mol}^{-1} = 4.43$ L (1 MARK)

Therefore, the gases all occupy different volumes. (1 MARK)

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b. [The universal gas law is $pV = nRT$. Rearranging for pressure, $p = \frac{nRT}{V}$[1]][We can see that as temperature increases, so too does the pressure, indicating that a temperature increase from 25 °C to 40 °C would increase the pressure.[2]]

✓ ✗ I have given the relevant form of the universal gas law.[1]

✓ ✗ I have explained how this demonstrates the relationship between temperature and pressure.[2]

21. Convert units: $\frac{55}{1000} = 0.055\ \text{L}$ (1 MARK)

Use correct formula: $M = m \times \frac{RT}{pV}$

$M = 0.135 \times \frac{8.31 \times 363\ \text{K}}{100\ \text{kPa} \times 0.055\ \text{L}}$ (1 MARK)

$M = 74.0\ \text{g mol}^{-1}$ (1 MARK)

22. Using $n = \frac{m}{M}$, calculate the moles of chlorine gas

$n(Cl_2) = \frac{2.00\ \text{g}}{71.0\ \text{g mol}^{-1}}$

$n(Cl_2) = 0.02816\ \text{mol}$ (1 MARK)

Rearrange $pV = nRT$ to $T = \frac{pV}{nR}$

$T = \frac{100\ \text{kPa} \times 1.00\ \text{L}}{0.02816\ \text{mol} \times 8.31}$ (1 MARK)

$T = 427\ \text{K}$ (1 MARK)

23. $n(CH_4) = \frac{0.150\ \text{g}}{16.0\ \text{g mol}^{-1}}$

$n(CH_4) = 0.009375\ \text{mol}$ (or 9.375×10^{-3} mol) (1 MARK)

$n(CH_4) : n(O_2)$

1 : 2

$n(O_2) = 0.009375\ \text{mol} \times \frac{2}{1}$

$n(O_2) = 0.01875\ \text{mol}$ (1 MARK)

Rearrange $n = \frac{V}{V_m}$ to $V = n \times V_m$

$V(O_2) = 0.01875\ \text{mol} \times 24.8\ \text{L mol}^{-1}$ (1 MARK)

$V(O_2) = 0.465\ \text{L}$ (1 MARK)

Key science skills

24. a. The temperature of the laboratory (in kelvin), the volume of the room, and the amount of gas in the room (in moles).

b. $M = m \times \frac{RT}{pV}$

c. [Hypodermic needles are very sharp.[1]][They must be disposed of in the proper container and handled with care.[2]][The syringe is potentially going to be very hot.[3]][Gloves should be worn.[4]]

✓ ✗ I have identified a hazard.[1]

✓ ✗ I have described a precaution for this hazard.[2]

✓ ✗ I have identified a second hazard.[3]

✓ ✗ I have described a precaution for this hazard.[4]

d. [The resolution of the scale on the syringe is very large.[1]][The lack of significant figures will result in a calculation that has low accuracy.[2]]

✓ ✗ I have explicitly addressed resolution.[1]

✓ ✗ I have explicitly addressed accuracy.[2]

FROM LESSONS 16B & 16D

Questions from multiple lessons

25. a. [S has an oxidation state of −2 in ZnS(s) and an oxidation state of +4 in $SO_2(g)$.[1]][This increase in oxidation state is due to the loss of negatively charged electrons, meaning that ZnS(s) must have undergone oxidation.[2]]

✓ ✗ I have identified the oxidation state of sulfur in ZnS(s) and $SO_2(g)$.[1]

✓ ✗ I have linked the change in oxidation state to whether ZnS(s) has undergone reduction or oxidation.[2]

b. $S^{2-}(s) + 2H_2O(l) \rightarrow SO_2(g) + 4H^+(aq) + 6e^-$

c. $O_2(g) + 4e^- \rightarrow 2O^{2-}(s)$

d. One mole of ZnO(s) will be produced for every mole of ZnS(s) consumed.

FROM LESSON 11A

Chapter 14 review

Multiple choice

1. A. Oxygen bonds are too simple for infrared radiation to be absorbed by the double bond.

2. C. Ozone is a potent greenhouse gas as it has a much more complex bonding arrangement than oxygen.

3. C. Burning fossil fuels, which all contain carbon, produces the largest amount of carbon dioxide.

4. B. Infrared radiation is emitted from the Sun.

5. D. Gases will occupy the total volume of any container.

6. B. Units must all be converted to kPa, L, mol and K.

7. B. Divide both sides by RT, then switch the left and right sides of the equation.

8. D. Temperature = (17 + 273) K

Temperature = 290 K

Volume = 4.2 L

Amount = 0.33 mol

$R = 8.31$

$p = n \times \frac{RT}{V}$

$p = 0.33\ \text{mol} \times \frac{8.31 \times 290\ \text{K}}{4.2\ \text{L}}$

$p = 1.9 \times 10^2\ \text{kPa}$

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9. C. According to the equation, the explosion of 2 mol of $C_7H_5N_3O_6$ produces 20 mol of gaseous products (12 mol CO + 5 mol H_2 + 3 mol N_2). Hence 1 mol $C_7H_5N_3O_6 \rightarrow$ 10 mol gases.

$$V(\text{gas}) = n(\text{gas}) \times \frac{RT}{p}$$

$$V(\text{gas}) = 10\text{ mol} \times \frac{8.31 \times (27 + 273\text{ K})}{1.00 \times 10^2\text{ kPa}}$$

$$V(\text{gas}) = 249\text{ L}$$

10. C. $V = \frac{nRT}{p}$

$$n(\text{He}) = \frac{pV}{RT}$$

$$n(\text{He}) = \frac{\left(\frac{10.33\text{ atm}}{0.987\text{ atm}} \times 100\text{ kPa}\right) \times 5.65\text{ L}}{8.31 \times (25.0 + 273)\text{ K}}$$

$$n(\text{He}) = 2.39\text{ mol}$$

Short answer

11. $n = 5.00$ mol

$T = (26 + 273)$ K

$T = 299$ K (1 MARK)

$p = 520$ kPa

$$V = \frac{nRT}{p}$$

$$V = \frac{5.00\text{ mol} \times 8.31 \times 299\text{ K}}{520\text{ kPa}}\quad\text{(1 MARK)}$$

$V = 23.9$ L (1 MARK)

12. **a.** $n(Al_2S_3)_{reacted} : n(H_2S)_{produced}$

1 : 3 (1 MARK)

$n(H_2S) = 3 \times 0.200$ mol

$n(H_2S) = 0.600$ mol (1 MARK)

b. $V = n \times V_m$

$V(H_2S) = 0.600\text{ mol} \times 24.8\text{ L mol}^{-1}$

$V(H_2S) = 14.9$ L

13. **a.** That all the water is condensed back into the container and none was lost out of the can.

b. Loss of carbon dioxide into the environment

c. $m(CO_2\ lost) = initial\ mass - final\ mass$

$m(CO_2) = 1375\text{ g} - 1350\text{ g}$

$m(CO_2) = 25$ g (1 MARK)

$$n(CO_2) = \frac{25\text{ g}}{44.0\text{ g mol}^{-1}}$$

$n(CO_2) = 0.568$ mol (1 MARK)

$V = m \times V_m$

$V = 0.568\text{ mol} \times 24.8\text{ L mol}^{-1}$ (1 MARK)

$V(CO_2) = 14$ L (1 MARK)

d. $pV = nRT$

$400\text{ kPa} \times 14.1\text{ L} = n \times 8.31 \times (30 + 273)\text{ K}$ (1 MARK)

$$n(CO_2) = \frac{400\text{ kPa} \times 14.1\text{ L}}{8.31 \times 303\text{ K}}\quad\text{(1 MARK)}$$

$n(CO_2) = 2.2$ mol (1 MARK)

14. **a.** $n(O_2)_{required} = \frac{pV}{RT}$

$$n(O_2)_{required} = \frac{76.0\text{ kPa} \times 1.60\text{ L}}{8.31 \times 292\text{ K}}\quad\text{(1 MARK)}$$

$n(O_2)_{required} = 0.0501$ mol (1 MARK)

b. $n(NaClO_3)_{required} : n(O_2)_{produced}$ (1 MARK)

2 : 3

$$n(NaClO_3)_{required} = \frac{2}{3} \times n(O_2)_{produced}$$

$$n(NaClO_3)_{required} = \frac{2}{3} \times 0.0501\text{ mol}$$

$n(NaClO_3)_{required} = 0.0334$ mol (1 MARK)

c. $m = n(NaClO_3) \times M(NaClO_3)$

$m(NaClO_3)_{required} = 0.0334\text{ mol} \times 106.5\text{ g mol}^{-1} = 3.56$ g per minute

$m(NaClO_3)_{required} = 3.56$ g per minute (1 MARK)

35.6 g per 10 minutes (1 MARK)

15. **a.** [The $V(N_2)$ present is at 100 kPa,[1]][and 36.6 °C or 309.6 K.[2]]

✓ ✗ I have stated the pressure value.[1]

✓ ✗ I have stated the temperature value.[2]

b. The temperature must be at SLC, which is at 298 K (25 °C), not 36.6 °C, to use this formula.

c. $n(N_2) = \frac{p(N_2) \times V(N_2)}{RT}$

$$n(N_2) = \frac{100\text{ kPa} \times 62.0\text{ L}}{8.31 \times (36.6 + 273)\text{ K}}\quad\text{(1 MARK)}$$

$n(N_2) = 2.41$ mol (1 MARK)

d. $\frac{n(NaN_3)}{n(N_2)} = \frac{10\text{ mol}}{16\text{ mol}}$ simplified to $\frac{5}{8}$ (1 MARK)

$$n(NaN_3) = \frac{5}{8} \times n(N_2)$$

$$n(NaN_3) = \frac{5}{8} \times 2.410\text{ mol}$$

$n(NaN_3) = 1.51$ mol (1 MARK)

e. $n = \frac{m}{M}$; rearrange to:

$m = n \times M$

$m(NaN_3) = n(NaN_3) \times M(NaN_3)$

$m(NaN_3) = 1.51\text{ mol} \times 65.0\text{ g mol}^{-1}$ (1 MARK)

$m(NaN_3) = 97.9$ g (1 MARK)

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Key science skills

16. a. i. $n = \frac{V}{V_m}$

$n(\text{butane}) = \frac{88.0 \times 10^{-2}\ \text{L}}{24.8\ \text{L mol}^{-1}}$ (1 MARK)

$n(\text{butane}) = 3.548 \times 10^{-3}\ \text{mol}$ (1 MARK)

$M = \frac{m}{n}$

$M(\text{butane}) = \frac{0.200\ \text{g}}{3.548 \times 10^{-3}\ \text{mol}}$

$M(\text{butane}) = 56.4\ \text{g mol}^{-1}$ (1 MARK)

ii. $M(\text{butane}) = m \times \frac{RT}{pV}$

$M(\text{butane}) = 0.200\ \text{g} \times \frac{8.31 \times 290.1\ \text{K}}{95.0\ \text{kPa} \times 0.088\ \text{L}}$ (1 MARK)

$M(\text{butane}) = 57.7\ \text{g mol}^{-1}$ (1 MARK)

b. $\%\ similarity = \frac{56.4\ \text{g mol}^{-1}}{57.7\ \text{g mol}^{-1}} \times 100\%$

$\%\ similarity = 97.7\%$ (1 MARK)

$\%\ difference = 100\% - 97.7\%$

$\%\ difference = 2.3\%$ (1 MARK)

c. The value of 24.8 L mol^{-1} for molar volume is only valid when the conditions are at 298 K and 100 kPa.

d. This would make the experiment invalid, as the volume of gas would be an unknown mixture of air and butane.

e. Butane gas is both dangerous and very flammable, and should be released into a fume cupboard.

FROM LESSONS 16A, 16C, 16D & 16E

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15A Analysing ions using gravimetric analysis

Progress questions

1. A. Ions conduct electricity, so electrical conductivity testing provides information about the concentration of ions in water.

2. A. Salt ions conduct electricity better than pure water.

3. B. The light intensity from a lightbulb indicates the amount of electricity passing through the circuit.

4. A. This word translates to 'without water'.

5. A. The water of crystallisation molecules are bonded to the salt and these can be broken by heat energy.

6. A. Mole ratios indicate the ratio of salt to water of crystallisation molecules.

7. B. Gravimetric analysis involves precipitating ions out of solution and weighing the precipitate.

8. A. Water molecules remaining from improper drying will increase the recorded mass of a precipitate.

9. A. The product of the reaction between the analyte and the reactant is insoluble in water.

10. B. Mass-mass stoichiometry in gravimetric analysis is used to measure the mass of an analyte.

11. B. There are four NH_3 molecules reacting for every three O_2 molecules.

12. A. Mole ratio of $H_2 : O_2$ is 2 : 1, and $5 \div 3 = 1.67 < 2$, so H_2 is the limiting reactant.

13. B. Molar ratios indicate the relative numbers of moles of substances, and do not reflect their masses.

Deconstructed

14. C

15. A

16. [Advantages of using gravimetric analysis include that it requires fewer materials for chemical analysis to be carried out. Specifically, gravimetric analysis only requires an analyte and a means of drying and weighing the precipitate.[1]][However, some disadvantages include the possibility of incorrectly weighing the sample due to the presence of water contaminants. Additionally, the reaction between the solute and the analyte may be slow.[2]] [Advantages of using electrical conductivity testing as a form of chemical analysis include that it can be carried out quickly and easily, and the equipment used is portable.[3]][However, a disadvantage of using electrical conductivity testing is that it cannot be used to determine the concentration of a particular element, instead providing an indication of the overall ionic content of a sample.[4]]

✓ ✗ I have identified an advantage of gravimetric analysis.[1]

✓ ✗ I have identified a disadvantage of gravimetric analysis.[2]

✓ ✗ I have identified an advantage of electrical conductivity testing.[3]

✓ ✗ I have identified a disadvantage of electrical conductivity testing.[4]

Exam-style

17. The mass of hydrated salt = 10.00 g

The mass of anhydrous salt after heating = 3.70 g

Mass of water in the salt = 10.00 g – 3.70 g

Mass of water in the salt = 6.30 g (1 MARK)

The mass ratio of $Na_2CO_3 : H_2O$

$3.70 \quad : 6.30$

The mole ratio of $Na_2CO_3 : H_2O$ using

$n = \frac{m}{M}$

The mole ratio of $\quad Na_2CO_3 \quad : \quad H_2O$

$$\frac{3.70 \text{ g}}{106.0 \text{ g mol}^{-1}} : \frac{6.30 \text{ g}}{18.0 \text{ g mol}^{-1}}$$

The mole ratio of $\quad Na_2CO_3 \quad : \quad H_2O$

3.49×10^{-2} mol : 0.350 mol (1 MARK)

Mole ratio of $\quad H_2O : Na_2CO_3 = \frac{0.35 \text{ mol}}{3.49 \times 10^{-2} \text{ mol}}$

$H_2O : Na_2CO_3 = 10.0$

$x = 10$ (1 MARK)

$Na_2CO_3 \bullet 10H_2O(s)$ (1 MARK)

18. **a.** **i.** $n(MgSO_4) : n(Ca(NO_3)_2)$

$1 \quad : \quad 1$

$n(MgSO_4) = 0.332$ mol, $n(Ca(NO_3)_2) = 1.98$ mol (1 MARK)

0.332 mol < 1.98 mol

$MgSO_4$ is the limiting reactant. (1 MARK)

ii. $n(AgNO_3) : n(CaCl_2)$

$2 \quad : \quad 1$

$n(AgNO_3) = 3.71$ mol, $n(CaCl_2) = 1.94$ mol

To use all of the $AgNO_3$ available, the $CaCl_2$ required is:

$n(CaCl_2) = \frac{1}{2} \times 3.71$ mol

$n(CaCl_2) = 1.86$ mol (1 MARK)

$AgNO_3$ is the limiting reactant. (1 MARK)

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iii. $n(\text{Pb(NO}_3)_2) : n(\text{NaCl})$

$1 \quad : \quad 2$

$n(\text{Pb(NO}_3)_2) = \frac{m}{M}$

$n(\text{Pb(NO}_3)_2) = \frac{0.866\text{ g}}{(207.2 + (2 \times 14.0) + (6 \times 16.0))\text{ g mol}^{-1}}$

$n(\text{Pb(NO}_3)_2) = 0.00261\text{ mol}$

$n(\text{NaCl}) = \frac{m}{M}$

$n(\text{NaCl}) = \frac{2.34\text{ g}}{(23.0 + 35.5)\text{ g mol}^{-1}}$

$n(\text{NaCl}) = 0.0400\text{ mol}$

To use all the $\text{Pb(NO}_3)_2$ available, the NaCl required is:

$n(\text{NaCl}) = \frac{2}{1} \times 0.00261\text{ mol}$

$n(\text{NaCl}) = 0.00523\text{ mol}$ (1 MARK)

$\text{Pb(NO}_3)_2$ is the limiting reactant. (1 MARK)

iv. $n(\text{Al}) : n(\text{O}_2)$

$4 \quad : \quad 3$

$n(\text{Al}) = \frac{m}{M}$

$n(\text{Al}) = \frac{1.45\text{ g}}{27.0\text{ g mol}^{-1}}$

$n(\text{Al}) = 0.0537\text{ mol}$

$n(\text{O}_2) = \frac{m}{M}$

$n(\text{O}_2) = \frac{3.11\text{ g}}{32.0\text{ g mol}^{-1}}$

$n(\text{O}_2) = 0.0972\text{ mol}$

To use all of the Al available, the O_2 required is:

$n(\text{O}_2) = \frac{3}{4} \times 0.0537\text{ mol}$

$n(\text{O}_2) = 0.0403\text{ mol}$ (1 MARK)

Al is the limiting reactant. (1 MARK)

b. i. $n(\text{CaSO}_4) : n(\text{MgSO}_4)$

$1 \quad : \quad 1$

$n(\text{CaSO}_4) = \frac{1}{1} \times n(\text{MgSO}_4)$

$n(\text{CaSO}_4) = 0.332\text{ mol}$ (1 MARK)

$m(\text{CaSO}_4) = n \times M$

$m(\text{CaSO}_4) = 0.332\text{ mol} \times (40.1 + 32.1 + (4 \times 16.0))\text{ g mol}^{-1}$

$m(\text{CaSO}_4) = 45.2\text{ g}$ (1 MARK)

ii. $n(\text{AgNO}_3) : n(\text{AgCl})$

$2 \quad : \quad 2$

$n(\text{AgCl}) = \frac{2}{2} \times 3.71\text{ mol}$

$n(\text{AgCl}) = 3.71\text{ mol}$ (1 MARK)

$m(\text{AgCl}) = n \times M$

$m(\text{AgCl}) = 3.71\text{ mol} \times (107.9 + 35.5)\text{ g mol}^{-1}$

$m(\text{AgCl}) = 532\text{ g}$ (1 MARK)

iii. $n(\text{PbCl}_2) : n(\text{Pb(NO}_3)_2)$

$1 \quad : \quad 1$

$n(\text{PbCl}_2) = 0.00261\text{ mol}$ (1 MARK)

$m(\text{PbCl}_2) = n \times M$

$m(\text{PbCl}_2) = 0.00261\text{ mol} \times (207.2 + (2 \times 35.5))\text{ g mol}^{-1}$

$m(\text{PbCl}_2) = 0.727\text{ g}$ (1 MARK)

iv. $n(\text{Al}_2\text{O}_3) : n(\text{Al})$

$2 \quad : \quad 4$

$n(\text{Al}_2\text{O}_3) = \frac{2}{4} \times 0.0537\text{ mol}$

$n(\text{Al}_2\text{O}_3) = 0.0269\text{ mol}$ (1 MARK)

$m(\text{Al}_2\text{O}_3) = n \times M$

$m(\text{Al}_2\text{O}_3) = 0.0269\text{ mol} \times ((2 \times 27.01) + (3 \times 16.0))\text{ g mol}^{-1}$

$m(\text{Al}_2\text{O}_3) = 2.74\text{ g}$ (1 MARK)

Key science skills

19. a. [Accuracy refers to the closeness of measurements to the true value.[1]][Precision refers to the closeness of measurements to each other.[2]]

✓ ✗ I have defined the term accuracy.[1]

✓ ✗ I have defined the term precision.[2]

b. i. The results are precise, but inaccurate.

ii. [The student has made a systematic error.[1]][This is because the results deviate from the true value by a consistent amount in the one direction.[2]]

✓ ✗ I have identified the type of error.[1]

✓ ✗ I have explained why the error is systematic.[2]

FROM LESSON 16D

20. a. Analyte

b. [It will not be possible to use octane as the analyte in this experiment.[1]][Gravimetric analysis requires the analyte to be dissolved in a solvent so that it can be precipitated out through the addition of a reactant to the solution.[2]] [Since octane is a non-polar molecule, it will not be able to dissolve in the water. This is because octane is unable to form sufficiently strong intermolecular forces with the water molecules in order to overcome the intermolecular forces between the water molecules.[3]]

✓ ✗ I have identified that octane cannot be used as the analyte in this experiment.[1]

✓ ✗ I have described the importance of the type of analyte used in gravimetric analysis.[2]

✓ ✗ I have explained the inability of octane to dissolve in water.[3]

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c. [A potential change that could be made would be using a non-polar solvent instead of water.[1]][This would enable octane to dissolve in the solvent as it would not need to overcome the strong intermolecular forces found in water.[2]]

✓ ✗ I have identified the nature of the solvent required.[1]

✓ ✗ I have explained the benefit of this change.[2]

d. [It is not necessary to carry out this experiment.[1]][The purpose of gravimetric analysis is to determine the mass of an analyte present in either a mixed sample or in a solution through precipitating out that particular analyte with a reactant and measuring the mass.[2]][Since Vinh has proposed to use a sample of pure octane for gravimetric analysis this process is unnecessary, as by weighing the pure sample, the mass of the octane sample can be determined.[3]]

✓ ✗ I have identified the necessity of performing the experiment.[1]

✓ ✗ I have described the purposes of gravimetric analysis.[2]

✓ ✗ I have explained that the pure sample of octane can be weighed.[3]

e. i. [Presence of water particles on the precipitate as a proper technique of washing, filtering, heating and cooling has not been followed.[1]][Some precipitate may be lost while drying with a paper towel or a sufficient amount of time may not have passed to form all of the precipitate.[2]]

✓ ✗ I have identified one possible source of error.[1]

✓ ✗ I have identified a second possible source of error.[2]

ii. $MgCl_2(aq) + 2AgNO_3(aq) \rightarrow 2AgCl(s) + Mg(NO_3)_2(aq)$

iii. $n(MgCl_2) = \frac{m}{M}$

$n(MgCl_2) = \frac{33.2\text{ g}}{(24.3 + (2 \times 35.5))\text{ g mol}^{-1}}$

$n(MgCl_2) = 0.348\text{ mol}$ (1 MARK)

$n(AgNO_3) = \frac{m}{M}$

$n(AgNO_3) = \frac{108.0\text{ g}}{(107.9 + 14.0 + (16.0 \times 3))\text{ g mol}^{-1}}$

$n(AgNO_3) = 0.636\text{ mol}$

$n(MgCl_2) : n(AgNO_3)$

1 : 2

To use all of the $MgCl_2$, the $AgNO_3$ required is:

$n(AgNO_3) = \frac{2}{1} \times 0.348\text{ mol}$

$n(AgNO_3) = 0.697\text{ mol}$ (1 MARK)

Limiting reactant: $AgNO_3$ (1 MARK)

Excess reactant: $MgCl_2$

iv. $n(AgCl) : n(AgNO_3)$

2 : 2

$n(AgCl) = \frac{2}{2} \times 0.636\text{ mol}$

$n(AgCl) = 0.636\text{ mol}$ (1 MARK)

$m(AgCl) = n \times M$

$m(AgCl) = 0.636\text{ mol} \times (107.9 + 35.5)\text{ g mol}^{-1}$

$m(AgCl) = 91.2\text{ g}$ (1 MARK)

FROM LESSON 16D

Questions from multiple lessons

21. a. Aluminium phosphate

b. $M(AlPO_4) = 27.0\text{ g mol}^{-1} + 31.0\text{ g mol}^{-1} + (4 \times 16.0)\text{ g mol}^{-1}$

$M(AlPO_4) = 122.0\text{ g mol}^{-1}$ (1 MARK)

$n(AlPO_4) = \frac{m}{M}$

$n(AlPO_4) = \frac{7.78\text{ g}}{122.0\text{ g mol}^{-1}}$

$n(AlPO_4) = 6.38 \times 10^{-2}\text{ mol}$ (1 MARK)

c. The molar ratio is 1:1

d. $M(Al(NO_3)_3) = (27.0 + (3 \times (14.0 + (3 \times 16.0 \times 9))))\text{ g mol}^{-1}$

$M(Al(NO_3)_3) = 213.0\text{ g mol}^{-1}$ (1 MARK)

$m(Al(NO_3)_3) = n \times M$

$m(Al(NO_3)_3) = (6.38 \times 10^{-2})\text{ mol} \times 213.0\text{ g mol}^{-1}$

$m(Al(NO_3)_3) = 13.6\text{ g}$ (1 MARK)

FROM LESSONS 4B, 4C & 6A

22. a. [Silver nitrate is required in excess.[1]][This is to ensure that all chloride ions react to form precipitate.[2]]

✓ ✗ I have identified the substance required in excess.[1]

✓ ✗ I have explained my answer based on the requirements of the reaction.[2]

b. [The precipitate must have water particles removed from it before it can be weighed.[1]][This is because the precipitate may have water molecules on it when it is taken out of the solution.[2]][If still present when the precipitate is weighed, an inaccurate measurement of the mass of the sample would be obtained.[3]][The mass of the precipitate recorded would be greater than the true value, which would result in the ion concentration calculated to be higher than expected.[4]]

✓ ✗ I have identified the necessary step prior to weighing.[1]

✓ ✗ I have explained the effect of this step on the precipitate.[2]

✓ ✗ I have explained how the presence of water molecules may lead to weighing inaccuracies.[3]

✓ ✗ I have described the effect of water molecules on the results of the experiment.[4]

c. i. $M(\text{AgCl}) = 107.9 \text{ g mol}^{-1} + 35.5 \text{ g mol}^{-1}$

$M(\text{AgCl}) = 143.4 \text{ g mol}^{-1}$ (1 MARK)

$n(\text{AgCl}) = \frac{167.4 \text{ g}}{143.4 \text{ g mol}^{-1}}$

$n(\text{AgCl}) = 1.167 \text{ mol}$

$n(\text{AgCl}) : n(\text{Cl}^-)$

1 : 1 (1 MARK)

$m(\text{Cl}^-) = n \times M$

$m(\text{Cl}^-) = 1.167 \text{ mol} \times 35.5 \text{ g mol}^{-1}$

$m(\text{Cl}^-) = 41.4 \text{ g}$ (1 MARK)

ii. $c(\text{Cl}^-) = \frac{n}{V}$

$c(\text{Cl}^-) = \frac{1.167 \text{ mol}}{2.0 \text{ L}}$

$c(\text{Cl}^-) = 0.58 \text{ mol L}^{-1}$

FROM LESSONS 4C & 13A

23. a. [Stearate anions (from the soap)are less effective in hard water as they react with cations present in the water.[1]] [Stearate anions are required to be in their ionic form in order to carry out their dirt-removal function.[2]] [However, when they react with cations they form an insoluble and unreactive precipitate, making them less effective.[3]]

✓ ✗ I have identified why stearate anions (from the soap) are less effective in hard water.[1]

✓ ✗ I have explained the role of stearate ions in soaps.[2]

✓ ✗ I have explained the effect of the inability of stearate ions to function.[3]

b. i. [If the identity of the ions in the precipitate was unknown, then this would be an ineffective method.[1]] [This is because gravimetric analysis provides quantitative data, and requires the molar mass of the precipitate which can only be determined if the identity of the constituent ions are known.[2]]

✓ ✗ I have identified the effectiveness of the method.[1]

✓ ✗ I have explained the requirements of this analytical method.[2]

ii. [The efficacy of this method in determining the concentration of soap ions depends on which ions are in excess.[1]] [If salt cations from the water are in excess, then the method will be effective as all of the soap ions will be present in the precipitate and measurable.[2]] [However, if the soap ions are in excess, then not all of the soap ions will react and therefore this would not be an effective method of measuring the concentration of soap ions.[3]]

✓ ✗ I have identified a factor influencing the efficacy of the method.[1]

✓ ✗ I have explained the effectiveness of the method if salt ions are in excess.[2]

✓ ✗ I have explained the effectiveness of the method if soap ions are in excess.[3]

iii. If the experiment is conducted correctly, the analyte will be the limiting reactant,and the added reactant will be the excess reactant.

c. i. $m(\text{Mg}(\text{C}_{18}\text{H}_{35}0_2)_2) = 61.0 \text{ g}$ (1 MARK)

$n(\text{Mg}(\text{C}_{18}\text{H}_{35}0_2)_2) = \frac{m}{M}$

$n(\text{Mg}(\text{C}_{18}\text{H}_{35}0_2)_2) =$

$$\frac{61.0 \text{ g}}{(24.3 + 2 \times (18 \times 12.0 + 35 \times 1.0 + 2 \times 16.0)) \text{ g mol}^{-1}}$$

$n(\text{Mg}(\text{C}_{18}\text{H}_{35}0_2)_2) = 0.103 \text{ mol}$

$n(\text{Mg}(\text{C}_{18}\text{H}_{35}0_2)_2) : n(\text{C}_{18}\text{H}_{35}0_2{}^-)$

1 : 2

$n(\text{C}_{18}\text{H}_{35}0_2{}^-) = \frac{2}{1} \times 0.1033 \text{ mol}$

$n(\text{C}_{18}\text{H}_{35}0_2{}^-) = 0.207 \text{ mol}$ (1 MARK)

$c(\text{C}_{18}\text{H}_{35}0_2{}^-) = \frac{n}{V}$

$c(\text{C}_{18}\text{H}_{35}0_2{}^-) = \frac{0.207 \text{ mol}}{0.500 \text{ L}}$

$c(\text{C}_{18}\text{H}_{35}0_2{}^-) = 0.413 \text{ mol L}^{-1}$ (1 MARK)

ii. $m(\text{C}_{18}\text{H}_{35}0_2{}^-) = n \times M$

$m(\text{C}_{18}\text{H}_{35}0_2{}^-) = 0.207 \text{ mol} \times ((18 \times 12.0) + (35 \times 1.0) + (2 \times 16.0)) \text{ g mol}^{-1}$

$m(\text{C}_{18}\text{H}_{35}0_2{}^-) = 58.5 \text{ g}$ (1 MARK)

Percentage of soap made of stearate by mass $= \frac{58.5 \text{ g}}{300.0 \text{ g}} \times 100\%$

Percentage of soap made of stearate by mass $= 19.5\%$ (1 MARK)

FROM LESSON 13A

24. a. [Calcium is the limiting reactant.[1]] [This is because if calcium is the limiting reactant, the reaction will stop when all calcium has reacted, enabling all of the calcium present to be weighed and the amount measured.[2]]

✓ ✗ I have identified the limiting reactant.[1]

✓ ✗ I have explained my answer with respect to the precipitate that must form.[2]

b. i. Ensure that all of the precipitate has been collected or ensure that all water has been removed from the precipitate.

ii. $M(\text{Ca}_3(\text{PO}_4)_2) = ((3 \times 40.1) + 2 \times (31.0 + (4 \times 16.0)))) \text{ g mol}^{-1}$

$M(\text{Ca}_3(\text{PO}_4)_2) = 310.3 \text{ g mol}^{-1}$ (1 MARK)

$n(\text{Ca}_3(\text{PO}_4)_2) = \frac{m}{M}$

$n(\text{Ca}_3(\text{PO}_4)_2) = \frac{87.4 \text{ g}}{310.3 \text{ g mol}^{-1}}$

$n(\text{Ca}_3(\text{PO}_4)_2) = 0.2816 \text{ mol}$ (1 MARK)

$n(\text{Ca}^{2+}) : \text{n}(\text{Ca}_3(\text{PO}_4)_2)$

3 : 1

$n(\text{Ca}^{2+}) = \frac{3}{1} \times 0.2816 \text{ mol}$

$n(\text{Ca}^{2+}) = 0.845 \text{ mol}$ (1 MARK)

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iii. $m(\text{Ca}^{2+}) = n \times M$

$m(\text{Ca}^{2+}) = 0.845\ \text{mol} \times 40.1\ \text{g mol}^{-1}$

$m(\text{Ca}^{2+}) = 33.9\ \text{g}$ (1 MARK)

$\%\ \text{calcium in gypsum} = \dfrac{33.9\ \text{g}}{150.0\ \text{g}} \times 100\%$

$\%\ \text{calcium in gypsum} = 22.6\%$ (1 MARK)

FROM LESSON 6B

✓ ✗ I have identified that colorimetry and UV-visible spectroscopy provide both qualitative and quantitative data.[1]

✓ ✗ I have explained that ions strongly absorb particular wavelengths of light.[2]

✓ ✗ I have explained that absorbance is directly proportional to ion concentration.[3]

15B Analysing ions using light

Progress questions

1. B. Visible light is a very small proportion of the electromagnetic spectrum.

2. B. Complementary colours will have opposite interactions with a substance.

3. A. Calibration curves are constructed on the basis of the direct relationship between ion concentration and absorbance to determine the concentration of ions in a solution with known absorbance.

4. B. The unit on the x-axis is the concentration which can be determined from the absorbance value.

5. B. The colour of an object is determined by the wavelength reflected or transmitted through the object to the greatest extent.

6. B. Spectrophotometers can scan a number of different electromagnetic wavelengths to choose the one with highest absorbance.

7. B. A monochromator is used to specifically select the wavelength of light to be used in UV-visible spectroscopy.

8. A. The absorbance is calculated from a series of solutions with a specific concentration of solute.

9. B. Colorimetry and UV-visible spectroscopy enable assessment of the contents in blood, such as haemoglobin, bilirubin, cholesterol and sugars.

10. A. Metal complexes enable the absorption of electromagnetic radiation by a sample.

Deconstructed

11. A

12. B

13. [In both colorimetry and UV-visible spectroscopy, the wavelength of maximum absorbance is used to determine the substances present; the absorbance at a particular wavelength of light is used to measure the concentration of ions in a solution.[1]] [This information is able to be obtained since specific ions strongly absorb particular wavelengths of electromagnetic radiation.[2]] [The amount of absorbance increases linearly as the concentration of ions increases. Hence, the concentration of the solution can be ascertained from absorbance data obtained from the colorimeter or spectrophotometer.[3]]

Exam-style

14. **a.** A calibration curve is a graph that visualises the relationship between the concentration of several standard solutions and their corresponding absorbance values.

b. [A calibration curve is produced by preparing a series of standard solutions containing known concentrations of the solute being analysed.[1]] [By measuring the absorbance at a certain wavelength of light using these standard solutions, the relationship between concentration (on the x-axis) and absorbance (on the y-axis) can be plotted to construct a calibration curve.[2]]

✓ ✗ I have identified the use of standard solutions in constructing a calibration curve.[1]

✓ ✗ I have described how a calibration curve is developed.[2]

c. [Since a calibration curve represents the relationship between concentration and absorbance of a substance, it can be used to determine the concentration of a solution with a known absorbance value.[1]] [The known value of absorbance can be matched up with the corresponding concentration value using the calibration curve.[2]]

✓ ✗ I have identified the use of a calibration curve.[1]

✓ ✗ I have explained how a calibration curve can be used to determine a solution's concentration.[2]

d. 0.17 to 0.19 M

e. [A solution that appears blue absorbs the complimentary colour to blue, which is orange.[1]] [The copper in the solution will absorb orange as a result of the chemical properties of the copper.[2]] [This means that orange wavelengths of light will need to be used when performing colorimetry as orange is the complementary colour to blue.[3]]

✓ ✗ I have identified that orange light will be used.[1]

✓ ✗ I have explained why the solution absorbs orange light.[2]

✓ ✗ I have explained why the orange light will be useful for colorimetry.[3]

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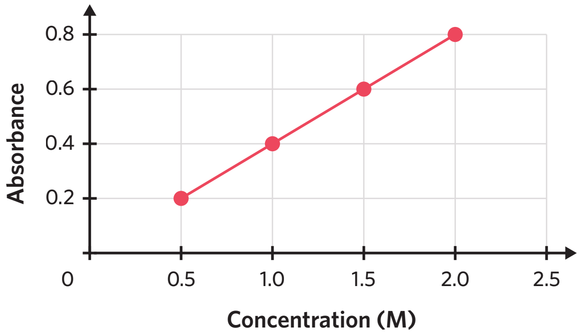
15. a. Absorbance in colorimetry does not have a unit.

b.



- ✓ ✗ I have correctly labelled the axes.[1]
- ✓ ✗ I have used the correct units.[2]
- ✓ ✗ I have drawn a correctly shaped curve.[3]

c. [Ryan’s claim is incorrect.[1]][This is because colorimetry measures the amount of absorbance of a particular wavelength of light, and therefore, the wavelength most readily absorbed should be used.[2]][As cobalt chloride appears violet in colour, this is the colour most strongly reflected, and therefore, the complementary colour of violet, which is yellow,is the ideal colour for colorimetry, as it will be the most strongly absorbed.[3]]

- ✓ ✗ I have identified that Ryan is incorrect.[1]
- ✓ ✗ I have described the relationship between the colour used for analysis and absorbance in colorimetry.[2]
- ✓ ✗ I have described the absorption of the complementary colour of violet.[3]

d. i. 1.75 M

ii. [This calibration curve cannot be used for the chosen solution as the absorbance is greater than the absorbance of any of the standard solutions used to produce the calibration curve.[1]][We cannot extrapolate the shape of the graph beyond the maximum measured value as the graph will not accurately provide information of solutions with higher concentrations than 2.0 M.[2]]

- ✓ ✗ I have compared the absorbance relative to the calibration curve.[1]
- ✓ ✗ I have explained the limitations of the graph.[2]

16. a. [A monochromator is used to select a specific wavelength of light for UV-visible spectroscopy.[1]][In colorimetry, a light filter is used instead of a monochromator.[2]]

- ✓ ✗ I have identified the function of a monochromator in UV-visible spectroscopy.[1]
- ✓ ✗ I have identified the alternative to a monochromator used in colorimetry.[2]

b. i. [As a UV-visible spectrophotometer scans different wavelengths of light, it can be used to select the wavelength with the highest absorbance for a particular solution.[1]][Choosing the wavelength which is absorbed by the specific solution to the greatest extent enhances the accuracy of the absorbance data from the spectroscopic analysis.[2]]

- ✓ ✗ I have identified the function of scanning different light wavelengths in UV-visible spectroscopy.[1]
- ✓ ✗ I have described the advantage of the ability to choose the wavelength with the greatest absorbance.[2]

ii. Green light would be best suited for analysis of this solution.

c. [In a solution where the metal ions do not produce a colour, a metal complex must be formed to perform UV-visible spectroscopy.[1]][Producing a metal complex will render the colourless solution coloured such that sufficient absorbance of light for analysis can occur.[2]]

- ✓ ✗ I have identified that a metal complex would need to be produced.[1]
- ✓ ✗ I have explained why a metal complex is necessary.[2]

d. Assessment of contents of blood, including bilirubin, cholesterol, iron, haemoglobin and sugars or assessment of contents of urine, such as lead or assessment of milk quality or air quality assessment for presence of chromium and lead or assessment of phosphate concentration in waterways.

17. [An advantage of colorimetry is that it is cheap and requires little equipment.[1]][A disadvantage is that the less advanced equipment (in comparison to UV-visible spectroscopy) can lead to inaccuracies.[2]][An advantage of UV-visible spectroscopy is that it is highly accurate due to the use of a spectrophotometer including the monochromator.[3]][A disadvantage is that it is expensive due to the use of advanced equipment.[4]]

- ✓ ✗ I have identified an advantage of colorimetry.[1]
- ✓ ✗ I have identified a disadvantage of colorimetry.[2]
- ✓ ✗ I have identified an advantage of UV-visible spectroscopy.[3]
- ✓ ✗ I have identified a disadvantage of UV-visible spectroscopy.[4]

18. [The correct option is III, as the amount of light absorbed by a solution is used to determine the concentration of the solution.[1]][Option I is incorrect as UV-visible spectroscopy measures the amount of light absorbed by a solution, not the amount of light emitted from it.[2]][Option II is incorrect as there are peaks not only at 210 nm, but also at approximately 400 nm.[3]]

- ✓ ✗ I have explained why option III is correct.[1]
- ✓ ✗ I have explained why option I is incorrect.[2]
- ✓ ✗ I have explained why option II is incorrect.[3]

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19. [As the solution is transparent, the complementary colour to the colour most strongly absorbed by the solution will be transmitted through the solution.[1]][This gives the solution the appearance of the transmitted colour.[2]]

✓ ✗ I have identified the colour that will be transmitted.[1]

✓ ✗ I have explained the appearance of the solution.[2]

Key science skills

20. a. [The non linear shape of the calibration curve is likely due to the non-standardised test tubes.[1]][As some of the test tubes are transparent and some are translucent, this means that different amounts of light will be absorbed depending on which type of test tube is used.[2]]

✓ ✗ I have identified the effect of the test tubes used.[1]

✓ ✗ I have explained the different interactions between light and the test tubes.[2]

b. [The use of both transparent and translucent test tubes for analysis is a systematic error.[1]][This is because the absorbance values from the standard solutions in the translucent test tubes would be consistently higher than those from the transparent tubes. In this way, there would be a constant bias in results.[2]]

✓ ✗ I have identified the error as systematic.[1]

✓ ✗ I have explained why this is a systematic error.[2]

c. The student should ensure that all of the test tubes are transparent.

FROM LESSON 16D

Questions from multiple lessons

21. a. [Silver is considered insoluble when reacting with iodide, and therefore would form a precipitate.[1]][As a result, this decreases the amount of silver ions available in the solution to absorb light.[2]][Therefore, the concentration of silver determined by the detector would be inaccurate, and less than the true value.[3]]

✓ ✗ I have explained that silver and iodide will form a precipitate.[1]

✓ ✗ I have explained that the formation of a precipitate decreases the silver ion concentration.[2]

✓ ✗ I have connected the decreased silver concentration with the detector's readings.[3]

b. $Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$

$n(Ag^+) : n(I^-)$

$1 \quad : \quad 1$

$n(Ag^+) = c \times V$

$n(Ag^+) = 1.0 \text{ mol L}^{-1} \times \frac{20}{1000} \text{ L}$

$n(Ag^+) = 0.020 \text{ mol}$ (1 MARK)

$n(I^-) = 1.2 \text{ mol L}^{-1} \times \frac{15}{1000} \text{ L}$

$n(I^-) = 0.018 \text{ mol}$ (1 MARK)

$n(Ag^+) > n(I^-)$, therefore I^- is the limiting reactant (1 MARK)

c. $n(Ag^+$ available to absorb light$) = n(Ag^+$ available$) - n(Ag^+$ consumed in reaction$)$

$n(Ag^+$ available to absorb light$) = 0.020 \text{ mol} - 0.018 \text{ mol}$

$n(Ag^+$ available to absorb light$) = 0.002 \text{ mol}$

FROM LESSONS 12A & 15A

22. a. [The wavelength of light used should be approximately 600 nm.[1]][This is because only solution B absorbs 600 nm strongly, while another wavelength, such as 325 nm, would likely be absorbed by both solutions.[2]][Therefore, the absorbance of solution B could not be accurately determined.[3]]

✓ ✗ I have identified the appropriate wavelength for solution B.[1]

✓ ✗ I have described the absorption of solution B at different wavelengths.[2]

✓ ✗ I have identified the potential error associated with selecting an alternative wavelength.[3]

b. At an absorbance of 0.20, $c(Co(NO_3)_2) = 4.1 \text{ mM}$ (1 MARK)

$c(Co(NO_3)_2) = \frac{4.1 \text{ mM}}{1000}$

$c(Co(NO_3)_2) = 4.1 \times 10^{-3} \text{ M}$ (1 MARK)

c. i. $Co(NO_3)_2(aq) + CuSO_4(aq) \rightarrow CoSO_4(aq) + Cu(NO_3)_2(aq)$

ii. $n(Co(NO_3)_2) : n(CuSO_4)$

$1 \quad : \quad 1$

$n(Co(NO_3)_2)$ in 10.00 mL:

$n(Co(NO_3)_2) = c \times V$

$n(Co(NO_3)_2) = 4.1 \times 10^{-3} \text{ M} \times \frac{10}{1000} \text{ L}$

$n(Co(NO_3)_2) = 4.1 \times 10^{-5} \text{ mol}$ (1 MARK)

$c(CuSO_4) = \frac{n}{V}$

$c(CuSO_4) = \frac{4.1 \times 10^{-5} \text{ mol}}{\frac{20.0}{1000} \text{ L}}$

$c(CuSO_4) = 2.1 \times 10^{-3} \text{ M}$ (1 MARK)

FROM LESSON 12A

Chapter 15 review

Multiple choice

1. C. In order to determine the salt concentration, the soil must first be dissolved in water and then filtered to remove insoluble components of soil.

2. A. Although some salts are used as fertilisers, sodium chloride can be toxic to most plants and does not promote plant growth.

3. B. Electrical conductivity testing is effective in testing salinity since conductivity is correlated with the quantity of ions in solution.

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4. C. Water of crystallisation molecules can be removed from the hydrated salt by heating it and removing the water through evaporation.

5. B. The salts are named normally – $Ba(OH)_2$ reads as barium hydroxide. Then, the number of water molecules bonded to it are then quantified as hydrates – because there are 8 water molecules, the substance is an octahydrate.

6. A. The highest intensity absorbance is at a wavelength of around 420 nm. On the visible light spectrum, a wavelength of about 420 nm corresponds to a purple colour. Hence, the filter should be a purple colour for optimum analysis.

7. D. Since the ions are not removed from the solution, the colour of the solution remains the same.

8. D. $M(Na_2SO_4{\bullet}10H_2O) = ((23.1 \times 2) + 32.1 + (16.0 \times 4) + (18.0 \times 10))\text{ g mol}^{-1}$

$M(Na_2SO_4{\bullet}10H_2O) = 322.3\text{ g mol}^{-1}$

9. D. 6 moles of H_2 will react in a ratio of 2:1 with oxygen which means 3 moles of oxygen will react leaving 9 – 3 = 6 moles (excess oxygen).

10. C. $n(CaCl_2) = \frac{m}{M}$

$n(CaCl_2) = \frac{0.42\text{ g}}{(40.1 + (35.5 \times 2))\text{ g mol}^{-1}}$

$n(CaCl_2) = 3.78 \times 10^{-3}\text{ mol}$

$m(Ca(OH)_2) = n \times M$

$m(Ca(OH)_2) = (3.78 \times 10^{-3}\text{ mol}) \times (40.1 + (16.0 \times 2) + (1.0 \times 2))\text{ g mol}^{-1}$

$m(Ca(OH)_2) = 0.28\text{ g}$

Short answer

11. a. [If the sample contained compounds other than magnesium sulfate and water, all calculations would be incorrect.[1]][This is because a specific molar mass is necessary to determine the molar ratio of each compound.[2]]

✓ ✗ I have identified that an unknown compound would present.[1]

✓ ✗ I have explained why an exact molar mass must be known.[2]

b. Mass of water lost = 3.383 g −1.656 g

Mass of water lost = 1.727 g

c. $M(MgSO_4) = (24.3 + 32.1 + (16.0 \times 4))\text{ g mol}^{-1}$

$M(MgSO_4) = 120.4\text{ g mol}^{-1}$ (1 MARK)

$n(MgSO_4) = \frac{m}{M}$

$n(MgSO_4) = \frac{1.656\text{ g}}{120.4\text{ g mol}^{-1}}$

$n(MgSO_4) = 1.375 \times 10^{-2}\text{ mol}$ (1 MARK)

d. $M(MgSO_4{\bullet}x\,H_2O) = \frac{m}{n}$

$M(MgSO_4{\bullet}x\,H_2O) = \frac{3.383\text{ g}}{1.375 \times 10^{-2}\text{ mol}}$

$M(MgSO_4{\bullet}x\,H_2O) = 246.0\text{ g mol}^{-1}$

e. $M(H_2O) = M(MgSO_4{\bullet}xH_2O) - M(MgSO_4)$

$M(H_2O) = 246.0\text{ g mol}^{-1} - 120.4\text{ g mol}^{-1}$

$m(H_2O) = 125.6\text{ g mol}^{-1} \times 1\text{ mol}$

$m(H_2O) = 125.6\text{ g}$

f. $n(H_2O) = \frac{m}{M}$

$n(H_2O) = \frac{126.0\text{ g}}{18.0\text{ g mol}^{-1}}$

$n(H_2O) = 6.975\text{ mol}$

$n(H_2O) = 7\text{ mol}$

g. [$MgSO_4{\bullet}7H_2O$.[1]][Magnesium sulfate heptahydrate.[2]]

✓ ✗ I have used the correct conventions for the formula.[1]

✓ ✗ I have added the number of water molecules to the salt's name.[2]

12. a. i. [The ions under analysis should be the limiting reactant.[1]] [This is because it is essential that all the ions react with the precipitating agent so its quantity can be accurately measured.[2]]

✓ ✗ I have identified that the ions in the water sample are the limiting reactant.[1]

✓ ✗ I have explained that all of the ions in the water sample must be measured.[2]

ii. $n(Pb(NO_3)_2) : n(CaCl_2)$

1 : 1

$n(Pb(NO_3)_2) = \frac{m}{M}$

$n(Pb(NO_3)_2) = \frac{6.34\text{ g}}{(207.2 + (2 \times 14.0) + (6 \times 16.0))\text{ g mol}^{-1}}$

$n(Pb(NO_3)_2) = 0.01914\text{ mol}$ (1 MARK)

$n(CaCl_2) = \frac{m}{M}$

$n(CaCl_2) = \frac{2.51\text{ g}}{(40.1 + (2 \times 35.5))\text{ g mol}^{-1}}$

$n(CaCl_2) = 0.02259\text{ mol}$ (1 MARK)

$n(Pb(NO_3)_2) < n(CaCl_2)$

$Pb(NO_3)_2$ is the limiting reactant. (1 MARK)

iii. $n(PbCl_2) = 0.01914\text{ mol}$

$m(PbCl_2) = n \times M$

$m(PbCl_2) = 0.01914\text{ mol} \times (207.2 + (2 \times 35.5))\text{ g mol}^{-1}$

$m(PbCl_2)$ formed = 5.33 g

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b. [The electrical conductivity in the river water sample is greater since it has a higher concentration of ions than pure water.[1]][This is because water does not conduct electricity well, whereas ions are effective conductors of electricity.[2]][The electrical conductivity of a solution increases as the concentration of ions increases.[3]]

- ✓ ✗ I have identified the main reason for the difference in electrical conductivity.[1]
- ✓ ✗ I have explained the difference in conductivity between water and ions.[2]
- ✓ ✗ I have explained the effect of increasing ion concentration on electrical conductivity.[3]

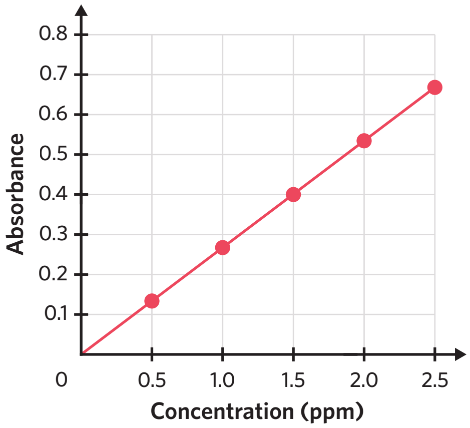
13. a. [Industrial run-off/Agricultural run-off.[1]]
[Mining operations/Sewage.[2]]

- ✓ ✗ I have given one example.[1]
- ✓ ✗ I have given a second example.[2]

b. [Colorimetry involves passing light of a particular wavelength through a coloured sample and using a detector to determine the amount of light absorbed. The absorbance data is used to determine the concentration of the substance.[1]][On the other hand, UV-Vis spectroscopy involves passing of UV and visible light into compounds that may not have a visible colour.[2]][Therefore, both colorimetry and UV-Vis utilise the absorbance of light to obtain qualitative and quantitative data about a sample.[3]]

- ✓ ✗ I have outlined the method of colorimetry.[1]
- ✓ ✗ I have outlined the method of UV-Vis.[2]
- ✓ ✗ I have identified the similarities and difference between colorimetry and UV-Vis.[3]

c. i.



- ✓ ✗ I have the correct labels on the axes.[1]
- ✓ ✗ I have the units of concentration in ppm.[2]
- ✓ ✗ I have an appropriate scale.[3]
- ✓ ✗ I have plotted the data accurately and drawn a line of best fit.[4]

ii. [For the Yenisei River, the absorbance value of 0.37 corresponds to a fluoride concentration of 1.4 ppm. Since this is less than the 1.5 ppm limit, the water in this river is safe to consume.[1]][For the Yangtze River, the absorbance value of 0.64 corresponds to a fluoride concentration of 2.4 ppm. Since this is greater than the 1.5 ppm limit, it is unsafe to drink the water.[2]]

- ✓ ✗ I have explained that the Yenisei River water is safe to consume.[1]
- ✓ ✗ I have explained that the Yangtze River water is unsafe to consume.[2]

Key science skills

14. a. $CaCO_3(s) + (COOH)_2(aq) \rightarrow CaC_2O_4(s) + H_2O(l) + CO_2(g)$ (2 MARKS)

b. Neutralisation

c. $CaC_2O_4(s) + \frac{1}{2}O_2(g) \rightarrow CaO(s) + 2CO_2(g)$ (2 MARKS)

d. $n(\text{CaO}) \text{ produced} = \frac{m}{M}$

$n(\text{CaO}) \text{ produced} = \frac{1.24 \text{ g}}{56.1 \text{ g mol}^{-1}}$

$n(\text{CaO}) \text{ produced} = 0.02210 \text{ mol}$ (1 MARK)

$\text{m}(CaCO_3) \text{ in dolomite} = n \times M$

$\text{m}(CaCO_3) \text{ in dolomite} = 0.02210 \text{ mol} \times 100.1 \text{ g mol}^{-1}$

$\text{m}(CaCO_3) \text{ in dolomite} = 2.212 \text{ g}$ (1 MARK)

$\% \ CaCO_3 \text{ in dolomite} = \frac{m(CaCO_3)}{m(\text{dolomite})} \times 100\%$

$\% \ CaCO_3 \text{ in dolomite} = \frac{2.212 \text{ g}}{3.72 \text{ g}} \times 100\%$

$\% \ CaCO_3 \text{ in dolomite} = 59.5\%$ (1 MARK)

e. [If the CaO precipitate is not rinsed with water, acid soluble components of the ore may be trapped in and collected with the precipitate, hence increasing the total mass of the precipitate.[1]][If the precipitate is not heated to constant mass, some water will remain and the weighed mass of the precipitate will be higher than the mass of CaO.[2]]

- ✓ ✗ I have identified two sources of error.[1]
- ✓ ✗ I have justified that these two errors will result in an overestimation.[2]

Adapted from VCAA 2005 Exam 1 Multiple choice Q,8,9

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16A Scientific investigation

Progress questions

1. B. Scientific ideas are rigorously tested through the scientific method.
2. B. Scientific ideas need to be able to be replicated by others.
3. B. Scientific ideas are based on objective evidence.
4. B. The basic scientific method begins with a brief introduction, followed by method, results, discussion, and conclusion.
5. A. Scientific knowledge is built over time on top of previous discoveries.
6. B. Not all conclusions are accurate and therefore need to be reviewed by others.
7. B. Textbooks present information that has been interpreted from original research articles.
8. B. Results are not always in line with theory.
9. B. A hypothesis highlights the potential effect of the independent variable on the dependent variable.
10. B. The independent variable is the only variable that is deliberately changed in an experiment. As a result, this variable causes the changes observed in an experiment.

Deconstructed

11. B
12. D
13. The degree of breakdown of the egg shell.

Exam-style

14. [The idea that Zodiac signs can influence current or future events has not been established using the scientific method.[1]] [Therefore, the ideas are currently considered to be non-scientific.[2]]

    ✓ ✗ I have identified the need for the scientific method to classify an idea as scientific or not.[1]

    ✓ ✗ I have linked my answer to the question.[2]

15. **a.** The wavelengths/colours used in the experiment.

    b. [If the attraction of bugs to light depends on the wavelength of light used,[1]] [then the colour that attracts the most bugs will reflect the wavelength that bugs are most attracted to.[2]]

    ✓ ✗ I have identified the effect of the independent variable on the dependent variable.[1]

    ✓ ✗ I have identified a possible outcome based on the relationship between the independent and dependent variable.[2]

16B Conducting an experiment

Progress questions

1. A. Qualitative analysis gives observational data.
2. A. Litmus paper tells whether something is acid, neutral or alkali.
3. A. Primary data is collected directly by the experimenter.
4. B. Quantitative data refers to numerical data.
5. A. Primary data is collected directly by the experimenter.
6. B. Secondary data refers to data collected by others, not the type of data that can be collected.
7. B. Findings from research studies can have negative impacts on different areas in society.
8. B. Scientists have to consider all of the possible uses of their research, however are not responsible for how it is used by others.
9. B. Previous work that contributes to research should always be referenced.
10. B. Research findings are generally tested multiple times and shared within the scientific community first so that they can be checked by other experts before it is shared to the wider community.
11. B. Risks associated with substances are already documented in the SDS.
12. B. Researchers need to take into account potential safety risks prior to conducting research.

Deconstructed

13. D
14. B
15. [Although CO_2 increases can be seen in the increase in the rate of crop growth, the idea that an increased concentration of CO_2 can result in a decrease in the amount of nutrients present in the crop can have a negative economic impact.[1]] [To get the same amount of nutrients, animals and humans would need to eat more crops.[2]] [This would require more land so that it can be used to grow more crops and this would affect the livelihood of the farmers.[3]]

    ✓ ✗ I have identified the key idea from the passage.[1]

    ✓ ✗ I have identified the effect of the idea on nutrient consumption.[2]

    ✓ ✗ I have described the socio-economic impact of this idea.[3]

Exam-style

16. **a.** Secondary data

    b. Plagiarism

17. **a.** Qualitative

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b. [In his experiment, Luis is collecting data based on the different colour changes.[1]][This is classified as qualitative data.[2]] [In order to calculate the amount of a substance present, Luis would need to collect quantitative data that gives numerical amounts.[3]][Unless Luis uses an analysis technique that allows the colours and their intensities to be translated into numerical values, Luis won't be able to calculate the amount of substance present by only using the data he collected.[4]]

✓ ✗ I have identified the data collected from the experiment.[1]

✓ ✗ I have identified the type of data collected from the experiment.[2]

✓ ✗ I have described the type of data that needs to be collected to calculate the amount of substance present.[3]

✓ ✗ I have identified what is required for Luis to calculate the amount of substance present.[4]

18. a. Ethical

b. [As vaccines are generally produced for human use, testing on humans would allow for the development of an effective vaccine in a shorter period of time.[1]][As a result, the vaccine would be available to affected individuals in a shorter period of time.[2]]

✓ ✗ I have identified the purpose of vaccines.[1]

✓ ✗ I have described the impact of human testing.[2]

c. [Although human experimentation allows us to develop vaccines quickly using less resources, there is a chance that the individuals that are involved in the testing can become sick themselves.[1]][The alternative testing method that does not involve humans takes a significant amount of time, which could not only have high economic costs, but also means that those affected by disease may not be able to receive the appropriate treatment in time.[2]][When choosing a method, it is important that we consider and understand all of the immediate and secondary implications of the method.[3]]

✓ ✗ I have identified a disadvantage of human testing.[1]

✓ ✗ I have described the disadvantages of alternative methods.[2]

✓ ✗ I have identified the considerations when choosing an appropriate method.[3]

19. a. [Lab coat to avoid contact on the skin.[1]][Safety goggles to avoid contact with eyes.[2]]

✓ ✗ I have identified one item of PPE.[1]

✓ ✗ I have identified a second item of PPE.[2]

b. [Different experiments use different reagents, which carry their own risks.[1]][As a result, experimenters are required to develop a risk assessment specific to their experiment.[2]] [Therefore, it is the experimenter and not the lab technician that should be writing the risk assessment.[3]]

✓ ✗ I have identified the nature of experimental risk.[1]

✓ ✗ I have identified the role of the experimenter in the risk management of an experiment.[2]

✓ ✗ I have linked my answer to the question.[3]

20. a. The brands of bottled water

b. [The volume or temperature of the water used.[1]][The type of technique used for analysis.[2]][The types of reagents used or the amounts of reagent used.[3]]

✓ ✗ I have identified a variable.[1]

✓ ✗ I have identified a second variable.[2]

✓ ✗ I have identified a third variable.[3]

21. a. [If the level of sugar in drinks increases then the tax on manufacturers will increase.[1]][To avoid costs, there will be a decrease in the proportion of drinks with over 5 g sugar per 100 mL.[2]]

✓ ✗ I have referred to the independent variable in the hypothesis.[1]

✓ ✗ I have described the effect of the independent variable on the dependent variable.[2]

b. External ethical concern

c. [As mentioned in the passage, sugar can have negative health implications such as obesity and type 2 diabetes.[1]][The SDIL can create an incentive for companies to reduce the amount of sugar in their drinks.[2]][This means that there would be more drink options with less sugar available for people to consume.[3]] [Therefore, the number of people with sugar-related disease would decrease.[4]]

✓ ✗ I have identified an impact.[1]

✓ ✗ I have described the effect of SDIL.[2]

✓ ✗ I have explained the effect of the SDIL on the availability of drinks with less sugar.[3]

✓ ✗ I have identified the overall impact on population health.[4]

16C Interpreting data

Progress questions

1. B. Experimental data is represented in a way that showcases the data best.

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2. A. Qualitative and quantitative data are better represented in different ways.

3. A. Without values, pie charts show relative comparisons.

4. B. The point of scientific research is to identify the effect of the independent variable on the dependent variable.

5. A. For the graph, the x-axis represents the independent variable.

6. B. The dependent variable is affected by the independent variable.

7. B. Control groups are generally not subjected to the independent variable.

8. B. Not all relationships seen in experiments are causal.

9. B. Other variables can impact the experiment, and therefore the conclusion drawn.

10. B. The supporting or refuting of a hypothesis depends on the results obtained in the experiment.

Deconstructed

11. C

12. A

13. [At the beginning of the experiment, the pH of the solution is close to 0.[1]][Therefore the original solution can be considered acidic, as acids have a pH lower than 7.[2]]

✓ ✗ I have used the graph to determine the pH of the original solution.[1]

✓ ✗ I have explained why the solution would be considered as acidic.[2]

Exam-style

14. **a.** As the reaction proceeds, the mass of the beaker and its contents remain the same.

b. [According to the law of conservation of mass, no mass is lost or gained in a chemical reaction.[1]][Therefore, for the reaction in the experiment which is sealed, there would be no loss or gain in mass recorded as the reaction proceeds.[2]][This can be shown by the data presented in the graph, where there was no change in mass over time.[3]]

✓ ✗ I have identified the underlying chemical concept.[1]

✓ ✗ I have described the effect of the chemical concept.[2]

✓ ✗ I have used data to justify my response.[3]

15. **a.** According to the graph, the world derives most of its electrical energy from fossil fuels, followed by hydro, nuclear, wind, other renewable resources then solar.

b. [The use of fossil fuels both disrupts the earth (due to drilling) and produces greenhouse gases such as carbon dioxide, as well as dangerous compounds such as sulfur dioxide and nitric oxides.[1]][Due to the growing population, it is expected that more fuels will need to be consumed to increase the supply of energy.[2]][This could result in negative environmental impacts, which may cause a shift towards more environmentally conscious fuel alternatives.[3]]

✓ ✗ I have identified the main factors involved in the use of fossil fuels.[1]

✓ ✗ I have described the impact of a growing population on energy consumption.[2]

✓ ✗ I have identified a major impact of the use of energy resources over time.[3]

16. **a.** This allows the data to be plotted to a larger, more visible scale. If there were no breaks in the axes the plotted line would have to be very small.

b. The density of water decreases with increasing temperature.

17. **a.** Quantitative data

b. Different energy drinks contain different levels of caffeine.

c. [There are major health concerns related to the consumption of energy drinks with high caffeine content.[1]][Hence it is inappropriate to have these drinks readily available for and tested on teenagers.[2]]

✓ ✗ I have identified the negative effects related to energy drinks.[1]

✓ ✗ I have identified an ethical consideration.[2]

d. The control group would involve the consumption of non-caffeinated energy drinks.

18. **a.** Number of colours (in the highlighter)

b. [The amount of highlighter pen used or the amount of time the experiment was conducted for.[1]][The brand of highlighter pen used or the type of paper used in the experiment.[2]]

✓ ✗ I have identified one controlled variable.[1]

✓ ✗ I have identified a second controlled variable.[2]

16D Experimental factors affecting data

Progress questions

1. A. Spilling a solution is a random error.

2. A. Forgetting a reagent is a random error.

3. B. Random errors do not affect results in consistent ways.

4. A. Precise data are within close proximity of each other.

5. A. Values that are accurate are close to the theoretical value.

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6. A. The zeros after the first significant figure are considered significant.

7. A. Addition requires answers to be expressed to the lowest number of decimal points in the calculation.

8. B. Multiplication requires answers to be expressed to the least number of significant figures in the calculation.

9. A. Reproducibility describes results collected by different experimenters.

10. A. Results can be repeated under the same conditions.

11. A. In an experiment, only the independent variable should be changed.

Deconstructed

12. D

13. D

14. [For the experiment to be valid, all variables outside the independent variable should be controlled.[1]][For this experiment, the independent variable is the conditions in which the banana was kept.[2]][Despite this, many other variables were also changed in the experiment. This included the size of the bananas and the time intervals between measurements.[3]][Therefore the conclusion isn't very valid as other variables may have affected the results obtained.[4]]

✓ ✗ I have identified the parameters of validity.[1]

✓ ✗ I have identified the independent variable.[2]

✓ ✗ I have identified the changes in other variables.[3]

✓ ✗ I have linked my answer to the question.[4]

Exam-style

15. [Systematic.[1]][Random.[2]][Parallax.[3]][Mistake.[4]]

✓ ✗ I have connected an error to the uncalibrated thermometer.[1]

✓ ✗ I have connected an error to the loss of reagent.[2]

✓ ✗ I have connected an error to the meniscus.[3]

✓ ✗ I have connected an error to the reagent.[4]

16. **a.** Tests 1, 2 and 4.

b. [It was expected that the results would be around 15.00 mL of HCl.[1]][However, most of the results collected were much higher than this value.[2]][Although there was one result that was accurate, the average of all the results, 17.8 mL, is still higher than the true value.[3]][Therefore, the results are not very accurate.[4]]

✓ ✗ I have identified the expected results.[1]

✓ ✗ I have identified the main difference between the expected and actual results.[2]

✓ ✗ I have described the the data collected.[3]

✓ ✗ I have linked my answer to the question.[4]

c. Two

17. **a.** [Accessibility of reagents.[1]][The quantity of reagents available for testing.[2]]

✓ ✗ I have identified one limitation associated with the experiment.[1]

✓ ✗ I have identified a second limitation associated with the experiment.[2]

b. [Reproducibility is a measure of how close results are of the same experiment being conducted in a different environment by different experimenters.[1]][Due to the fact the results obtained in the experiments were different,[2]][the experiment is not considered to be very reproducible.[3]]

✓ ✗ I have defined reproducibility.[1]

✓ ✗ I have identified the key difference between the experiments.[2]

✓ ✗ I have linked my answer to the question.[3]

c. The reproducibility can be improved through the use of more precise equipment.

d. [Reproducibility describes the ability of results to be replicated using the same method but conducted in different conditions by different experimenters.[1]][Repeatability, however, measures the ability of results to be replicated when conducted under the same conditions by the same experimenter.[2]]

✓ ✗ I have defined reproducibility.[1]

✓ ✗ I have compared reproducibility to repeatability.[2]

18. B. A systematic error would affect all the results by the same amount so the line would not be kinked.

19. A. Precise results may be biased because even though they are close together they may have been obtained from a preferred sample and not reflect the true characteristics of the substance under analysis. Sensitive instruments should give more accurate measurements, but may be subject to random error associated with environmental fluctuations. The method must also be appropriate for the analysis to be valid. Repeating a procedure using the same equipment will not remove the uncertainty associated with systematic errors. Systematic errors are consistent repeatable errors associated with faulty equipment or flaws in experimental method or design.

20. **a.** % of fat in the chocolate.

b. Chocolate sample with no fat.

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c. As the % of fat in the chocolate increases, the melting temperature decreases.

d. Three

21. a. The change in blood glucose levels.

b. The higher the sugar content in food, the higher the blood glucose level.

c. [Thomas's glucose metre was consistently measuring the glucose levels 0.20 mmol L^{-1} less than the true value, likely underestimating the true sugar content of each snack.[1]] [This would decrease the accuracy of his results and is a systematic error.[2]]

✓ ✗ I have identified the key factor affecting the results.[1]

✓ ✗ I have linked my answer to the question.[2]

16E Writing scientific material

Progress questions

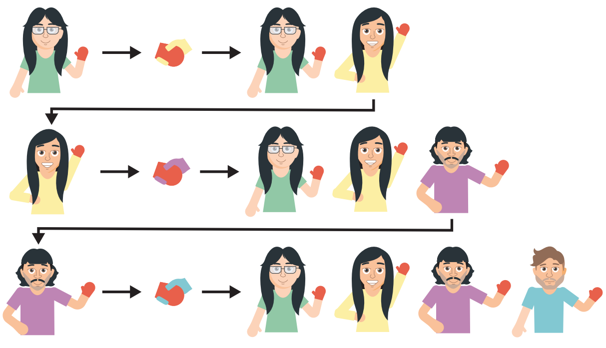
1. B. Clear writing allows for the findings to be easily understood.
2. A. It is better to be concise in scientific writing.
3. B. A methods section is required in all scientific papers.
4. B. Chemical terminology can represent a large concept that contains many sub-ideas.
5. B. Units depend on the information being expressed.
6. B. Measuring distance requires a unit that can indicate a distance.
7. A. Chemical equations represent a process involved in chemical reactions.
8. B. Visual models represent a concept/idea.
9. B. Models cannot show all of the details of a concept.
10. B. The details regarding how each heat transfer method operates aren't shown in the model.

Deconstructed

11. B

12. C

13.



Exam-style

14. a. Different colours of light have different wavelengths/patterns.

b. nm

c. We are unable to see why each colour has different waves.

15. The size of the atoms or molecules or size of bond length or exact shape or amount of space between atoms

16. C

L is Litres which is used when measuring the volume of a liquid.

M is Molarity of solution in mol L^{-1}.

N is a Newton which is used to measure forces in physics, not to be confused with N_A which is Avogadro's constant.

K is Kelvin which is used to measure temperature.

17. a. Days

b. Test A

c. The length of carrot

18. a. Size of rock used

b. To test the effect of the use of different sized rocks on the size of crystals that are able to grow.

c. Title, introduction, materials and methods, results, discussion, conclusion and references.

Chapter 16 review

Multiple choice

1. A. Qualitative data is based on descriptive information.
2. B. In this experiment, the relationship between the independent and dependent variables is correlated (as long as the experiment is appropriately controlled).
3. C. Controls show that any changes in the dependent variable were due to the independent variable and cannot improve the accuracy of the results.
4. B. Errors in reading measuring equipment affect all readings in the same way, however results can still be measured as being close to each other.
5. B. There are factors that can limit the ability of an experiment to be conducted. Errors are not considered as limitations.
6. D. SDS provides information regarding the hazards of substances and techniques used.
7. C. Random errors are unpredictable and occur irregularly.
8. C. All variables except the independent and dependent variable need to be controlled.
9. D. Data represented in a graph shows a pattern between the independent and dependent variable.
10. B. Precise data are within close proximity of each other.

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Short answer

11. a. Amount of corn flour

b. Qualitative data

c. [As shown in the table, every sample contains a different amount of corn flour and water.[1]][The experimenter has changed multiple variables (corn flour and water), and therefore the results are not valid. [2]]

- ✓ ✗ I have identified the change in multiple variables.[1]
- ✓ ✗ I have identified the relationship between changing variables and validity.[2]

12. a. It is expected that the extra tax placed on unhealthy foods would encourage people to consume healthier food(s).

b. [The taxes placed on unhealthy foods would mean that the foods would cost more to purchase.[1]][As a result, people would be less likely to buy these foods, leading to a decrease in the consumption of unhealthy foods.[2]]

- ✓ ✗ I have identified the impact of taxes on food prices.[1]
- ✓ ✗ I have described the effect of taxes on the consumption of unhealthy foods.[2]

c. [By increasing food prices, this could make it more difficult for people to afford and therefore access foods.[1]][This could have the opposite effect, where people could become less healthy due to not being able to eat enough food.[2]][Also, introducing this tax is effectively trying to influence human behaviour, whereby the government is making decisions on what people can and cannot access. This could be considered highly unethical.[3]]

- ✓ ✗ I have identified the impact of taxes on food accessibility.[1]
- ✓ ✗ I have identified an adverse impact on individual health.[2]
- ✓ ✗ I have described an ethical concern related to the government influencing human behaviour.[3]

13. a. A bubble solution that does not contain any additional substances.

b. i. Bar graph

ii. Corn syrup

iii. [Reproducibility measures the ability of an experiment to collect similar results when conducted by different experimenters.[1]][In this experiment, there is a lot of variability in the way that the data is collected. For example, when taking the photo that is used to measure bubble size, the positioning of the experimenter relative to the bubbles could impact the size of the bubbles measured.[2]][Considered the results recorded for both glycerin and corn syrup are relatively close to each other, this could affect the overall conclusion of the experiment.[3]][Also, the concentration of the solutions and the size of the wand were not included, therefore the experiment cannot be replicated properly.[4]][As a result, the results are not very reproducible.[5]]

- ✓ ✗ I have defined reproducibility.[1]
- ✓ ✗ I have described an area of variability in the experiment.[2]
- ✓ ✗ I have identified an impact of the variability on the conclusion drawn.[3]
- ✓ ✗ I have described the impact of the exclusion of key information related to the materials used.[4]
- ✓ ✗ I have linked my answer to the question.[5]

iv. The use of corn syrup produces the largest bubbles, followed by glycerin and dish washing detergent.

14. a. [The law of conservation of energy states that energy is transferred or transformed, which does not mean that the energy needs to be used for a specific purpose.[1]][The energy available in the coal was still transformed into thermal energy, even if some escaped to heat up the surrounding air rather than the food.[2]][If the total energy before and after was measured, even in its different forms, it would be the same.[3]] [Therefore, the student's statement is inaccurate.[4]]

- ✓ ✗ I have identified the main error in the statement relative to the law of conservation of energy.[1]
- ✓ ✗ I have described the energy change(s) during this scenario.[2]
- ✓ ✗ I have identified a strategy that can be used to confirm the law of conservation of energy.[3]
- ✓ ✗ I have linked my answer to the question.[4]

b. i. Being able to measure the energy lost to the environment or being able to measure the energy transferred into the food or being able to measure the chemical energy in the coal or the availability of coal/food for the experiment.

ii. [The type of coal used or the type of food used.[1]] [The barbecue used or the distance between the food and the coal.[2]][The length of time of cooking/burning coal.[3]]

- ✓ ✗ I have listed one variable that needs to be controlled.[1]
- ✓ ✗ I have listed a second variable that needs to be controlled.[2]
- ✓ ✗ I have listed a third variable that needs to be controlled.[3]

15. a. The concentration of vinegar used.

b. i. [This test is being conducted on humans,[1]][where the effect of the experiment may cause the subjects some discomfort.[2]]

- ✓ ✗ I have identified a key ethical consideration.[1]
- ✓ ✗ I have identified the effect of conducting research in this way.[2]

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ii. [Since the results were derived from many different individuals, the self rating used is subjective and would be based on the previous experiences of each subject.[1]] [As a result, the interpretation of the results would vary between subjects.[2]] [Facial expressions from each subject in response to eating lemon would not be an accurate indicator of sourness as this would also vary between subjects and their experiences.[3]] [Therefore, the results from this experiment are are not valid.[4]]

✓ ✗ I have identified the subjectivity of the self rating system.[1]

✓ ✗ I have identified the impact of the self rating system.[2]

✓ ✗ I have described the subjective nature of the use of facial expressions.[3]

✓ ✗ I have linked my answer to the question.[4]

iii. There's no direct relationship between the concentration of vinegar consumed and a person's reaction when subsequently consuming lemons.

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GLOSSARY

A

accuracy how close measured values are to their true value p. 545

acid-base reaction chemical reaction where a proton, H^+, is transferred between two chemical species p. 314

acid species that donates a proton, H^+, during an acid-base reaction p. 314

acidic solution solution containing a greater concentration of hydronium ions than hydroxide ions p. 340

addition polymerisation chemical reaction between monomers resulting in the formation of a polymer p. 263

addition reaction one molecule combining with another molecule to form a larger molecule p. 208

adhesivity property of being sticky p. 252

adsorption adhering of a component onto the stationary phase p. 144

aim purpose of an experiment p. 518

alcohol organic compound that has a hydroxyl (–OH) group bonded to a carbon atom in the parent chain p. 231

aliquot portion of a sample solution used for analysis during titration p. 444

alkali a base that is soluble in water and neutralises acids p. 316

alkane hydrocarbon that contains only single carbon to carbon bonds p. 206

alkene hydrocarbon with at least one double carbon to carbon bond p. 206, 263

alkyl group a side-chain attached to the parent carbon chain, containing only carbon and hydrogen atoms p. 216

alkyne hydrocarbon with at least one triple carbon to carbon bond p. 263

allotrope one of the different physical forms an element can take due to different structural arrangements of atoms p. 67

alum inorganic chemical compound composed of a hydrated aluminium salt p. 430

amphiprotic species that can either donate or accept a proton, H^+ p. 317

analyte substance which is quantified or identified p. 486

anecdote the communication of an abstract idea about a person, place, or thing through the use of a story or narrative p. 536

anhydrous without water p. 484

anions negative ion formed when an atom gains one or more valence electrons p. 108, 394

anode negatively charged electrode in a galvanic cell where oxidation occurs p. 394

antacid a weak base that is used to neutralise stomach acid p. 332

aqueous solution solution in which the solvent is water p. 141

aqueous dissolved in water p. 130

atom economy mass of desired products divided by mass of all reactants given as a percentage conversion p. 97

atom smallest unit of matter used in VCE chemistry p. 4

atomic number number of protons in the nucleus of an atom p. 4

Aufbau principle rule that states subshells are filled by electrons from the lowest to the highest energy level p. 13

autoionisation process by which an atom or molecule spontaneously forms ions p. 339

Avogadro's constant (N_A) the number of atoms in 12 g of ^{12}C, $6.02 \times 10^{23}\ mol^{-1}$ p. 178

B

base species that accepts a proton, H^+, during an acid-base reaction and are a group of substances that neutralise acids p. 314

basic solution solution containing a greater concentration of hydroxide ions than hydronium ions p. 340

bias difference between the average of a large set of measurements and the true value p. 544

bioaccumulation phenomenon of increasing contaminant concentration in an organism over time p. 250

biodegradable able to decompose in the environment p. 272

biomagnification phenomenon of higher contaminant concentration in organisms higher in the food chain p. 250

biomass any organic material made of plant or animal matter p. 243

bioplastics plastics produced from biomass p. 273

blocks assortment of elements into discrete categories based on which subshell their valence electrons are found in p. 18

boiling point temperature at which a liquid transitions into a gas p. 61, 296

bond pair pair of valence electrons that is shared between atoms in a covalent bond p. 50

brittle tendency of a material to fracture (break) when subjected to pressure p. 110

Brønsted-Lowry theory definition of an acid as a chemical species which donates a proton, H^+, to a base p. 314

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C

calcification a chemical process that forms calcium carbonate p. 358

calibration curve graph depicting the relationship between concentration and light absorbance for a particular substance p. 500

carbon chain linear set of carbon atoms in a hydrocarbon p. 205

carbon negative describes a process that absorbs more carbon dioxide than it produces p. 274

carbonic acid formed by the reaction of carbon dioxide and water p. 354

carboxyl group functional group that has a carbon double bonded to an oxygen and singly bonded to a hydroxyl group p. 234

carboxylic acid organic compound that contains a carboxyl functional group (–COOH) p. 234

carcinogenic having the potential to cause cancer p. 249

cathode positively charged electrode in a galvanic cell where reduction occurs p. 394

cation positive ions formed when an atom loses its valence electron(s) p. 79, 394

chemical symbol abbreviation used to represent a chemical element p. 5

chemolysis use of solely chemical substances to decompose organic substances into simpler ones p. 275

chromatogram visual output of chromatography p. 144

circular economy a continuous cycle that focuses on the optimal use and re-use of resources from the extraction of raw materials through to production of new materials, followed by the consumption and re-purposing of unused and waste materials p. 96, 277

colorimeter instrument used to undertake colorimetric analysis p. 498

colorimetry method of determining the concentration of a known substance which exploits the light absorbance properties of substances p. 497

combustion when a fuel reacts with oxygen (also called burning) p. 355

commercial indicator an artificially prepared indicator p. 347

complementary colours colours opposite to each other on the colour wheel p. 498

compostable describes a polymer that can be at least 90% decomposed after 180 days in a composting environment p. 276

compound two or more atoms of different elements bonded together p. 5

concentrated solution solution with a relatively high amount of solute in comparison to the volume of solution p. 410

concentration amount of chemical substance in a unit of volume p. 410, 440

concordant titres three or more titres that fall within approximately 0.1 mL of each other p. 445

condensation polymerisation chemical reaction between monomers resulting in the formation of a polymer and water p. 263, 281

condensation phase change from gas to liquid p. 295

control group (in chemistry) sample that is subjected to the same conditions as all other samples without the independent variable p. 536

controlled variable(s) variable(s) held constant throughout the experiment p. 518

coordination number the number of atoms, molecules or ions bonded to a chemical species p. 109

core charge attractive force between the positive nucleus and valence electrons, measured as the difference in the number of protons and inner electrons p. 25

corrosion occurs when some metals react with gases in the atmosphere (mainly oxygen) p. 87, 393

corrosive a highly reactive substance that causes obvious damage to tissue p. 315

covalent bonding bonding resulting from the sharing of electrons between non-metal atoms p. 47

covalent layer lattice atoms covalently bonded together in a two-dimensional arrangement to form layers, held together by weak dispersion forces p. 68

covalent network lattice regular three-dimensional arrangement of atoms covalently bonded together extending outwards p. 68

critical elements elements that are vital and considered in short supply unless they are recycled as sources could become depleted in 50–100 years p. 34

cross-link covalent bond between polymer chains p. 268

cross-linked (thermosetting) polymer polymer containing cross-links that degrades or chars when heated p. 268

crude oil mixture of differently sized hydrocarbons found in underground deposits and used as a resource p. 242

crystal lattice (ionic) symmetrical, three-dimensional arrangement of ions p. 108

crystal lattice (metallic) atoms of one type of metal element that are metallically bonded and organised in a pattern p. 78

crystallisation process by which solute molecules leave solution to form a highly organised structure known as a crystal p. 428

D

decalcification a chemical process that breaks down calcium carbonate p. 358

dehydrate remove H_2O from a molecule or ion p. 274

dehydrogenate remove a hydrogen atom or atoms from a molecule or ion p. 273

dependent variable variable that is measured by the experimenter p. 518

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deprotonation loss of a proton, H^+, from an acid p. 317

desorption release of a component from the stationary phase by dissolving p. 144

diatomic molecule a covalently bonded pair of atoms of the same or a different non-metal element p. 47

dilute solution solution with a relatively low amount of solute in comparison to the volume of solution p. 411

dilution process of decreasing concentration of a solution by adding more solvent p. 417

diprotic species that can donate up to two protons, H^+, during ionisation p. 318

dissociation process by which a molecule separates into smaller particles p. 320

dissolved incorporated into a liquid to form a solution p. 127, 141

dry corrosion oxidation of metals due to only oxygen in the air p. 397

ductility ability to be hammered or stretched into a thin shape without breaking p. 80

E

economic considerations taking into account the associated costs of time and money p. 526

ecosystem community of life forms interacting with their environment p. 243

elastomer polymer that forms occasional cross-links and is elastic p. 268

electrical conductivity ability to allow an electric current to flow through p. 80

electrode electrically conductive medium p. 394

electromagnetic radiation waves of the electromagnetic field including visible and ultraviolet light p. 498

electron configuration arrangement of electrons in shells and/or subshells p. 11

electron negatively charged particle that exists outside the nucleus of an atom p. 4

electronegativity how strongly an atom attracts electrons towards itself p. 23

electrostatic attraction between opposite charges p. 47, 79

element pure substance containing only a single type of atom p. 4

eluent solvent fluid that moves through a chromatography system p. 146

empirical formula chemical formula depicting the lowest whole number ratio of atoms of different elements in a compound p. 187

end point point at which a colour change occurs during titration p. 446

endangered elements that have a limited supply left on Earth p. 34

energy shells or energy levels orbits containing different levels of energy, around the nucleus of an atom where electrons are found according to the Bohr and Schrodinger models p. 13

entities atoms, molecules, compounds, ions, electrons etc. p. 178

equivalence point point at which the reaction has occurred as per the stoichiometric ratio of the balanced equation p. 446

error difference between the measured value and the true value p. 543

ethical considerations taking into account the effect on other living organisms p. 526

eutrophication excessive algal growth due to increased nutrient concentration in water p. 430

evaporation alternative term for vaporisation p. 295

exoskeleton the outer shell of a marine creature that is made of calcium carbonate p. 358

F

feedstock raw material used for producing another product p. 275

first ionisation energy energy required to remove the first valence electron from an atom p. 25

food chain a series of organisms each dependent on the next as a source of food p. 248

fossil fuel hydrocarbon-containing material formed in the Earth's crust from plant and animal remains p. 242

full equation chemical equation with all chemical species present p. 130

functional group specific groups of atoms within a compound that affect the properties of the compound p. 228, 263

G

galvanic cell electrochemical cell in which chemical energy from spontaneous redox reactions is converted into electrical energy p. 393

galvanised coated with zinc p. 397

global warming a long-term increase in the average global temperature of Earth's climate system since the industrial revolution due to human activities p. 465

gravimetric analysis chemical analysis involving the weighing of a precipitate formed from an analyte p. 482

green chemistry principles aimed at reducing the chemical-related impact on both humans and the environment through dedicated sustainability management programs p. 96

greenhouse effect the natural warming of the earth that results when gases in the atmosphere trap heat from the sun that would otherwise escape into space. p. 464

greenhouse gases molecules that absorb infrared radiation from the sun p. 274, 464

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group 16 hydride group 16 element covalently bonded to hydrogen p. 296

groups columns in the periodic table p. 17

H

half-equation reduction or oxidation equation of a complete redox reaction p. 374

haloalkane organic compound that has one or more halogen atom bonded to a carbon atom in the carbon chain p. 228

halogen element in group 17 of the periodic table (fluorine, chlorine, bromine, iodine) p. 228

hardness resistance to deformation when subjected to pressure p. 109

heartburn a burning sensation in the chest, caused by acid regurgitation into the oesophagus. p. 322

heat capacity relationship between the heat absorbed by a substance and its temperature change p. 298

heat conductivity ability to allow heat to pass through p. 80

high density polyethene (HDPE) tightly packed polymer produced at low pressures p. 266

homogeneous a mixture with uniform composition and properties p. 141

homologous series series of hydrocarbons that have a similar chemical structure and chemical properties p. 205

HPLC chromatography technique in which the mobile phase is pumped through a tightly packed stationary phase under pressure p. 146

humus dark, nutrient-rich, organic material produced from composting p. 277

hydrated salt contains water of crystallisation molecules p. 484

hydrated surrounded by water molecules p. 142

hydrocarbon compound consisting of hydrogen and carbon p. 204

hydrogen bonding is a special type of permanent dipole-dipole attraction between molecules (not a covalent bond) that results from the attractive force between a hydrogen atom covalently bonded to a very electronegative atom such as an N, O, or F atom p. 60

hydrogen pop test test used to indicate a reaction between a metal and an acid p. 89

hydrolysis reaction involving the cleavage of bonds using water, for example to break up polymers into monomers p. 283

hydronium ion positively charged chemical species with molecular formula H_3O^+ p. 315

hydroxide ion negatively charged chemical species with molecular formula OH^- p. 315

hydroxyl group functional group consisting of an oxygen atom covalently bonded to a hydrogen atom (–OH) p. 231

hygroscopic able to absorb or adsorb water from its surroundings p. 484

hypothesis testable statement which predicts the outcome of an experiment p. 518

I

ideal gas substance assumed to be composed of molecules of negligible volume with no intermolecular interactions p. 466

independent variable variable that is deliberately manipulated by the experimenter p. 518

indicator a substance that changes colour to indicate a pH range p. 347, 444

indigestion pain or discomfort in the stomach p. 322

inertness property of being unreactive p. 252

infrared radiation heat energy emitted from the sun p. 464

inhalant abusable chemical whose vapours can be concentrated and breathed in to produce intoxication p. 249

innovation introduction of new methods, procedures or products to traditional industries p. 281

insoluble unable to be dissolved in a given solvent p. 127

instantaneous dipole partial positive and negative charge that suddenly appears in an atom/molecule due to the random movement of electrons p. 58

intermolecular forces forces that attract molecules to each other p. 57

intramolecular bonds bonding between atoms within a molecule p. 47

ion-dipole attraction electrostatic attraction between a permanent dipole and an ion p. 143

ion atom that has lost or gained electrons to become a charged particle. p. 5, 25, 108

ionic bond electrostatic force of attraction between cations and anions p. 120

ionic compound compound made up of cations and anions held together by electrostatic forces of attraction p. 115

ionic equation chemical equation in which aqueous species are expressed as dissociated ions p. 131

ionic product of water K_w, expression of the concentration of ions present in water at 25 °C p. 338

ionisation process by which a chemical species gains or loses an electric charge p. 315, 142

isotope atoms of the same element that have a different mass due to a different number of neutrons p. 7, 164

K

kinetic energy the level of movement of a particle, molecule or molecular substance p. 61

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L

lanthanoids 15 naturally occurring metallic chemical elements, whose atomic numbers fall between 57 (Lanthanum) and 71 (Lutetium) p. 35

latent heat of vaporisation amount of energy required to convert 1 mol of a substance from its liquid state to its gaseous state at the boiling point of the substance p. 301

lattice a series of atoms arranged in a regular pattern p. 67

law of conservation of mass law that states that in a closed system, the total mass of reactants equals the total mass of products p. 488

legal and/or political factors a combination of factors such as the current political party in power, the efficiency of the current government and their policies, current legal framework, and public attitude towards the economy. p. 526

Lewis structure representation of a molecule where valence electrons are shown as dots p. 47

life cycle the management of the production and processing, storage, transportation, use, and disposal of a chemical product p. 96

linear (thermoplastic) polymer polymer with no cross-links that can be remoulded when heated p. 268

linear economy operates on a 'take-make-dispose' model, making use of resources to produce products that will be discarded after use p. 96, 277

litmus paper paper used to indicate the acidity or basicity of a substance p. 347

lone pair electron pairs of valence electrons not shared with another atom in a covalent bond p. 50

low density polyethene (LDPE) branched polymer produced at high pressures p. 266

lustre shiny and glossy appearance p. 80

M

macromolecule a molecule containing a very large number of atoms such as a giant atomic covalent lattice of diamond or graphite p. 67

malleability ability to deform under force without breaking p. 80

marine invertebrates ocean and sea creatures that do not have a backbone p. 358

mass number mass of an atom, equal to the sum of protons and neutrons p. 6

mass spectrometry analytical technique used to measure the mass of ions relative to their charge p. 170

mass-mass stoichiometry calculation of reactants and products in a chemical reaction using the law of conservation of mass p. 488

mass-to-charge ratio (m/z) the mass of an ion divided by its charge p. 170

melting point temperature at which a substance changes its state from solid to liquid p. 61, 110

meniscus curved upper surface of liquid p. 444

metal carbonate compound that neutralises acids to form a salt, water and carbon dioxide p. 328

metal displacement reaction redox reaction in which a more reactive metal displaces a less reactive metal's cation from solution p. 386

metal hydrogencarbonate compound that neutralises acids to form a salt, water and carbon dioxide p. 329

metal hydroxide compound that neutralises acids to form a salt and water p. 328

metal oxide formed as a result of a reaction between a metallic element and oxygen p. 91

metallic bonding the electrostatic force of attraction between delocalised electrons and cations in a metallic lattice structure p. 79

metallic character chemical properties associated with metals p. 23

metalloids elements which possess properties that are in between those of metals and non-metals p. 28

methodology different ways in which the planning and conducting of scientific investigations can be carried out p. 544

microbeads manufactured microplastics less than 1 mm in length used in exfoliating products and toothpastes p. 251

microplastics polluting fragments of plastic less than 5 mm in length p. 252

mistake sometimes called personal errors, mistakes should not be included in reporting and analysis p. 544

mobile phase fluid flowing through a chromatography system that carries the sample over the stationary phase p. 144

molar mass (M) the mass, in grams, per mol of substance p. 180

molarity (M) also called molar concentration, is a measure of the concentration of a chemical species in terms of amount of substance per unit volume of solution, measured in moles per litre (mol L^{-1}) p. 412

mole (n) amount of substance, in mol p. 178

molecular formula actual number of atoms in a molecule p. 49, 168, 194

molecular substance substance with two or more atoms joined by covalent bonds p. 58

molecule when two or more non-metal atoms bond by sharing electrons p. 5, 46

molten a pure substance that has been liquefied by heat p. 110

monatomic ions ions made up of only one atom p. 115

monochromator device used to select a particular wavelength of light p. 502

monomer molecule that can react with other molecules to form larger molecules (polymers) p. 263

monoprotic species that can only donate one proton, H^+, during ionisation p. 317

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N

natural indicator an indicator made from naturally sourced substances p. 347

neutral solution solution containing an equal concentration of hydronium and hydroxide ions p. 340

neutralisation reaction a reaction between an acid and an alkali or base that forms a salt and water (and carbon dioxide when a metal carbonate is a reactant) p. 329

neutralise the reaction between an acid and a base to produce water p. 329

neutron neutral particle inside the nucleus of an atom p. 4

nitric acid can be formed from a series of reactions in the atmosphere between nitrogen gas, oxygen gas and water p. 355

noble gases unreactive gases found in group 18 of the periodic table p. 30

nomenclature system developed in order to identify organic compounds p. 212

non-conductivity the inability of a molecule to conduct electricity p. 61

non-polar covalent bond covalent bond between two atoms of similar electronegativity p. 52

non-renewable unable to be naturally replenished p. 243

non-scientific ideas ideas that are not developed by following the scientific method p. 515

nucleus region at the centre of an atom that contains protons and neutrons p. 4

O

octet rule general principle which states that atoms like to have 8 electrons in their valence shell for stability p. 24

opinion a view or judgement formed about something, not necessarily based on fact or knowledge p. 536

orbitals regions with the highest probability of finding electrons p. 13

ore deposit in Earth's outermost layer containing metals and other minerals p. 78

organic chemical any compound that contains carbon-carbon and/or carbon-hydrogen bonds p. 241

organic compounds compounds consisting of carbon atoms most commonly covalently bonded to hydrogen, nitrogen, phosphorus or oxygen atoms p. 205

origin line at which samples are placed in paper and thin-layer chromatography p. 144

outlier data points or observations that differ significantly from other data points or observations p. 544

oxidation number number assigned to an atom that can be used to determine the movement of electrons in redox reactions p. 372

oxidation chemical reaction where a chemical species loses electrons p. 87, 370

oxidise cause a chemical species to lose electrons p. 371

oxidising agent chemical species that oxidises another substance by accepting electrons p. 371

ozone layer layer within stratosphere with high concentration of ozone (O_3) that absorbs UV radiation p. 249

P

paper chromatography chromatography technique using absorbent paper p. 144

parent chain longest continuous chain of carbon-carbon bonds p. 213

parent name component of systematic name that refers to the length of the parent chain p. 213

parts per billion (ppb) one in one billion parts of a solution p. 414

parts per million (ppm) one in one million parts of a solution p. 414

Pauli exclusion principle rule that states an orbital can't hold more than 2 electrons p. 13

peak area the area enclosed between the peak and the baseline on a chromatogram p. 147

percentage composition percentage by mass of an element in a compound p. 187

periodic table table of chemical elements in which elements are arranged in order of increasing atomic number p. 5

periodicity characteristics of elements in a period p. 16

periods rows in the periodic table p. 16

permanent dipole-dipole attraction electrostatic force of attraction between the permanent dipoles of adjacent polar molecules p. 59

permanent dipole a permanent area of partial positive or partial negative charge on a molecule p. 52

petrochemical derived from crude oil p. 245

pH meter an instrument used to measure the pH of a substance p. 347

pH scale pH values arranged on a logarithmic scale p. 338

pH measure of the concentration of hydronium ions in a solution p. 338

phase change transition of a substance from one state to another p. 295

phase distinct form in which matter can exist p. 130

plastic dissolution dissolving plastic to extract polymers and separate them from their additives p. 275

plastic molecule produced synthetically from monomers bonded together p. 264

polar covalent bond covalent bond between atoms of differing electronegativities p. 52

polar molecule an asymmetrical molecule with a partially positively charged end and a partially negatively charged end p. 53

polarity property of having a partial positive and negative charge p. 46

polyatomic ion ions made up of a group of atoms covalently bonded together p. 116, 372

polymer molecule produced naturally or synthetically from monomers bonded together p. 263

polyprotic species that donates multiple protons, H^+, during ionisation p. 318

potable alternative term for drinkable p. 295

precipitate solid compound formed in a precipitation reaction p. 127

precipitation reaction chemical reaction resulting in the formation of a solid from an aqueous solution p. 127, 423

precision how close measured values are to each other p. 544

pressure force a gas exerts on the walls of its container p. 464

primary data original data collected firsthand by researchers p. 523

primary source source that provides the original raw data p. 517

primary standard solution solution of precisely known concentration p. 443

primary standard substance used to make a primary standard solution p. 443

proton positively charged particle inside the nucleus of an atom p. 4

protonated water another name for the hydronium ion, H_3O^+ p. 315

pyrolysis decomposition brought about by high temperatures p. 275

Q

qualitative analysis technique that determines a non-numerical result p. 523

qualitative data non-numerical (descriptive) data collected based on observations taken during an experiment p. 523

qualitatively describes non-numerical (descriptive) data collected based on observations taken during an experiment p. 87

quantitative analysis technique that identifies the amount of substance present p. 523

quantitative data numerical data collected during experiments p. 524

R

random error error in measurement that differs in amount each time the experiment is conducted and is usually a 'one-off' error p. 543

ratio numerical relationship between the amounts of two or more elements p. 187

reactivity series of metals organised scale of metals and their cations, ranked according to their strength as reducing and oxidising agents p. 383

reactivity tendency of atoms to undergo a chemical reaction p. 23

recycle convert a waste product into a reusable product p. 95

recycled converting a waste product into a reusable product p. 34

redox reaction chemical reaction involving the transfer of electrons between chemical species p. 370

reduce cause a chemical species to gain electrons p. 371

reducing agent chemical species that reduces another substance by giving away electrons p. 371

reduction chemical reaction in which a chemical species gains electrons p. 370

relative abundance percentage of a particular isotope found in a naturally occurring sample of an element p. 164

relative atomic mass (A_r) weighted average of the masses of an element's isotopes on a scale on which a carbon-12 atom is assigned a value of 12 exactly p. 165

relative formula mass (M_r) weighted average of a substance's masses of the formula units on a scale on which the mass of a carbon-12 atom is assigned a value of 12 exactly p. 169

relative isotopic mass the mass of the isotope on the scale in which the relative isotopic mass of the carbon-12 isotope is assigned a value of 12 exactly p. 165

relative molecular mass (M_r) weighted average of a molecule's masses of the formula units on a scale on which the mass of a carbon-12 atom is assigned a value of 12 exactly p. 168

renewable resources comes from sources that naturally renew themselves at a rate that allows them not to become depleted p. 97

renewable comes from sources that naturally renew themselves at a rate that allows them not to become depleted p. 243

repeatability closeness of the agreement between the results of successive measurements of the same quantity being measured, carried out under the same conditions of measurement p. 545

reproducibility closeness of the agreement between the results of measurements of the same quantity being measured, carried out under changed conditions of measurement p. 545

resolution the smallest change in quantity that is measurable p. 550

retention factor the ratio of the distance moved by a component to the distance moved by the solvent front p. 144

retention time time taken for a component to pass through a HPLC column p. 147

rust a flaky, reddish precipitate formed from the corrosion of iron with water and oxygen p. 397

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S

sacrificial anode a readily oxidised anode used to coat other metals for protection against corrosion p. 397

sacrificial protection the use of a protective layer that is readily oxidised p. 397

safety data sheet (SDS) document that outlines the health and safety information associated with different materials and chemicals p. 527

salt bridge device used to separate the solutions in each half-cell whilst allowing the flow of ions to complete the circuit p. 394

saturated hydrocarbon a hydrocarbon possessing only single bonds between carbon atoms p. 207

saturated solution solution containing the maximum amount of dissolved solute p. 425

scientific ideas ideas that are developed by following the scientific method p. 514

scientific method procedure used to investigate scientific ideas p. 515

sea of delocalised electrons electrons that freely move between metal cations in the metallic bonding model p. 79

secondary data data that has been previously collected that is now accessible to different researchers p. 523

secondary source source that has interpreted primary sources p. 517

significant figures number of digits required to express a number to a certain level of accuracy p. 547

social considerations taking into account the effect on others p. 526

solar energy energy produced by the sun p. 302

solubility curve graph plotting the solubility of a substance (vertical axis) against temperature (horizontal axis) p. 423

solubility table table listing the solubilities of compounds formed from different ions when dissolved in water p. 423

solubility ability to be dissolved in a given solvent p. 127, 423

solute substance that dissolves in a solvent p. 141

solution homogeneous mixture in which a substance is dissolved in another p. 141

solvent front the wet moving edge of the solvent in chromatography p. 145

solvent substance that dissolves another substance p. 141, 249

specific heat capacity energy (J) required to raise the temperature of 1 gram of a given substance by 1 °C p. 298

spectator ions cations or anions that remain unchanged on both sides of a chemical equation p. 130

spectroscopy analysis of the interaction between matter and electromagnetic radiation p. 498

spontaneous redox reaction feasible redox reaction that occurs without the addition of any extra heat or energy p. 386

standard laboratory conditions (SLC) set of conditions used as a standard for experiments (298 K and 100 kPa) p. 469

standard solution solution with a precisely known concentration p. 500

stationary phase solid onto which the components of a sample adsorb p. 144

steam cracking breaking down of larger saturated hydrocarbons into smaller, often unsaturated ones p. 273

stoichiometry the relationship between relative quantities of reactants and products in a chemical reaction p. 488

stratosphere layer of earth's atmosphere between 20–50 km above sea level p. 249

strength how readily an acid or base will donate or accept a proton, H^+ p. 328

structural formula shows all the atoms and bonds in a molecule or compound p. 49

structural isomers that have the same molecular formula but the atoms are arranged in a different spatial arrangement p. 216

subatomic particle particle that exists inside the atom p. 4

substitution reaction one atom or group of atoms replacing a part of another molecule p. 208

sulfuric acid can be formed from a series of reactions in the atmosphere between sulfur, oxygen gas and water p. 355

supersaturated solution solution containing more than the maximum amount of solute that can be dissolved p. 425

sustainable can be produced at a rate that is greater than consumption without compromising future generations p. 37, 95, 241

synthesise produce a chemical through reaction with other compounds p. 242

synthetic produced artificially by humans using chemical reactions p. 241

systematic error error in measurement by the same amount in the same direction every time which also includes errors that are inherent in the experiment p. 543

systematic name standardised name used to identify organic compounds p. 212

T

temporary dipole-dipole attraction (dispersion force) electrostatic force of attraction between two instantaneous dipoles of adjacent molecules p. 58

tetravalent describes an atom with four electrons in its valence shell p. 67

thin-layer chromatography chromatography technique using a thin stationary phase supported by an inert backing p. 144

titration curve graphic representation of an acid-base titration, with titre volume on the x-axis and pH on the y-axis p. 449

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titration quantitative technique used to find the concentration and/or amount of substance in a solution p. 440

titre volume of the solution delivered from the burette to reach the end point of a titration p. 445

transition metals elements that occupy groups 3 to 12 on the periodic table p. 35

triprotic species that can donate up to three protons, H^+, during ionisation p. 318

U

ulcer a break in the lining of an organ like the stomach p. 322

uncertainty level of doubt regarding the data obtained (for VCE purposes, we are only looking at this from a qualitative perspective) p. 543

unpaired electron electron in an orbital by itself p. 48

unsaturated hydrocarbon a hydrocarbon possessing at least one double bond between carbon atoms p. 207

unsaturated solution solution containing less than the maximum amount of dissolved solute p. 425

UV-visible spectrophotometer instrument used in UV-visible spectroscopy p. 502

UV-visible spectroscopy spectroscopy method which uses the ultraviolet light region of the electromagnetic spectrum p. 497

V

valence electrons electrons in the outermost shell of an atom p. 16

valence shell electron pair repulsion theory (VSEPR) model used to predict the different shapes of molecules based on the repulsion between electron pairs p. 50

valence shell outermost energy shell p. 16

validity whether or not the experiment and its components including the results address the aim and hypothesis of the research p. 545

vaporisation phase change from liquid to gas p. 302

volumetric analysis quantitative analytical technique used to determine unknown concentrations of solutions p. 440

W

water of crystallisation water that is bonded into a crystal structure p. 482

wet corrosion oxidation of metals due to oxygen and water in the air p. 397

Symbols

%(m/m) mass percent concentration of solute in solution p. 416

%(m/v) mass percent concentration of solute per unit of volume of solution p. 416

%(v/v) volume percent concentration of solute in solution p. 416

%(w/w) weight percent concentration of solute in solution p. 416

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RESEARCH QUESTIONS FOR UNIT 1 AOS 3

The information below has been taken from your study design and outlines the possible options for your research investigation in Unit 1 AOS 3.

How can chemical principles be applied to create a more sustainable future?

Investigation topic 1: Endangered elements in the periodic table

Today's chemists are involved in many branches of chemistry, covering all 118 elements in the periodic table. Some of these elements are now considered to be critical and endangered, particularly due to the prevalence of modern technologies that rely on many different scarce minerals. It has been estimated that 44 elements will soon be, or are already, facing supply limitations, making a future of continuing technological advancement uncertain.

Questions that may be explored in this investigation include:

- Which chemicals are used in the manufacture of fireworks, what is the environmental impact of the combustion of these chemicals to produce the colourful effects seen in fireworks displays, and what alternatives are available?
- Based on their usefulness for society, how would you compare the value of lanthanoids and actinoids with the value of other metal groups in the periodic table?
- Why is helium classified as a critical and endangered element, and how can it be saved given that its atmospheric recovery is almost impossible?
- How is indium mined and used in the manufacture of products such as LCD screen televisions and computer monitors, mobile phones or photovoltaic panels, and what alternatives are available if indium becomes scarce?
- How do the properties of the metalloids (such as germanium, antimony, tellurium) differ so much to their neighbours on the periodic table, and how have these properties made them highly important for society and consequentially scarce in supply?
- How are precious metals from electronic waste (e-waste) recycled and what are the environmental and economic benefits of these recovery processes?

Investigation topic 2: Producing and using 'greener' polymers

Both natural and synthetic polymers play an important role in everyday life. The cells in animals and plants are built of, and metabolise, natural polymers. Proteins and carbohydrates in our food are both polymers. Synthetic polymers are used for a myriad of purposes in everyday life but may present challenges in terms of the by-products resulting from their manufacture or breakdown, and their persistence in the environment. The sustainability of polymers can be considered in terms of whether these plastics can be avoided by using different products or activities, reduced through design, or replaced by different materials.

Questions that may be explored in this investigation include:

- What are plant-based biopolymers and what are the impacts of their production on the environment?
- How do biodegradable and degradable polymers, compostable polymers and recyclable polymers differ in structure, production and environmental impacts?
- What is the difference between micropolymers and nanopolymers, and how are used plastic materials and litter managed and repurposed?
- Is the recycling of packaging products containing aluminium more sustainable than LDPE polymer-based packaging products?
- Why is the sale of plastic water bottles and single-use plastics banned in many countries?
- How do animal proteins compare with non-animal proteins for different applications, such as meat substitutes and non-animal leather?
- How do the chemical structures of elastomers differ from the structures of thermosetting and thermoplastic polymers, and what are the implications of the production of elastomers for society?
- What impact does the vulcanisation of rubber have on the environment and the communities where rubber is sourced and produced?
- What are the risks and benefits to the environment of the manufacturing, production and application of synthetic fibres for the textile industry (for example, synthetic grass, active wear, shoes and single-use plastics such as takeaway cups, containers, and electrical and electronic products such as mobile phone cords and USB flash drives)?

Investigation topic 3: The chemistry of Aboriginal and Torres Strait Islander peoples' practices

Throughout history, people all over the world have hypothesised, experimented, made empirical observations, gathered evidence, recognised patterns, verified through repetition, and made inferences and predictions to help them to make sense of the world around them and their place within it. Recent research and discussion have confirmed many Aboriginal and Torres Strait Islander groups use the environment and its resources to solve the challenges they face in the different Australian climates in ways that are more sustainable than similar materials produced in Western society. Their solutions can be explained by a variety of organic and non-organic chemical processes.

Questions that may be explored in this investigation include:

- Which plants are important to Aboriginal and Torres Strait Islander peoples for their medicinal properties, how are the plants processed before they are used, and what are the active ingredients (for example, the terpineols, cineoles and pinenes as the active constituents of tea trees and eucalyptus resin)?
- What are the chemical processes that occur when Aboriginal and Torres Strait Islander peoples detoxify poisonous food items: for example, the preparation of nardoo as a food source by heating, and the detoxification of cycad seeds through the removal of cycasins?
- How do Aboriginal and Torres Islander peoples utilise animal fats, calcination and plant pigments to vary the properties of the paints they make, and how does this compare to Western paint production processes and materials?
- How do binders and fixatives work to allow Aboriginal and Torres Islander peoples' paintings to be preserved for thousands of years?

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- How do Aboriginal and Torres Islander peoples' glue formulations parallel the use of modern epoxy resins, and how sustainable are the chemical processes involved in producing these materials?
- How are plant-based toxins such as saponins used in Aboriginal and Torres Strait Islander peoples' fishing practices, and how is this similar to other First Nation Peoples' fishing practices around the world?
- Kakadu plums have long been a component of Aboriginal and Torres Islander Peoples diets. What active ingredients do they contain that may make them a 'super food'?

Investigation topic 4: The sustainability of a commercial product or material

In Australia, new materials that are useful for society tend to be produced through a linear economy in which products are purchased, used and then thrown away. Increasingly, manufacturing companies are moving towards a circular economy, which seeks to reduce the environmental impacts of production and consumption while enabling economic growth through more productive use of natural resources and creation of less waste.

Research questions that may be explored in this investigation include:

- What is 'green steel' and what are the implications of its production for human health and the environment?
- Research a metal mined in Australia: for example, gold, copper or lithium. How is the metal processed and what are its useful properties? To what extent has the metal production and use moved towards a circular economy over the last decade? What innovations have led to the production of the metal being more sustainable over time?
- Select a commercial product that is available in different formulations: for example, vinegar (fermented, synthetic); salt (river salt, sea salt, iodised salt, Himalayan salt); cleaning products (soaps and detergents); oil (fish oil, coconut oil, olive oil); or milk (whole milk, skim milk, low-fat milk, A2 milk, plant milks such as almond, soy and coconut). What ingredients are in the product? How do the ingredients compare in the different product formulations? How is the product made? To what extent does the production of the product involve a linear economy or a circular economy? How does the production and use of the product impact human health and the environment?
- Select a product whose composition has changed over time: for example, hair comb (tortoiseshell to polymer); dental fillings (from silver amalgam and gold to porcelain and composite resin fillings); contact lenses (glass to polymers); paints (lead-based to oil-based and water-based); and tennis racquet strings (from cat gut to nylon and polyester). How have the properties and efficacies of the products changed over time? To what extent have the manufacturing processes become 'greener'?
- Examine the life cycle of a new product or material: for example, unbreakable glass inspired by seashells; new nanomaterials for the treatment of skin infections; and ultra-thin self-healing polymers to make water-resistant coatings. What is the relationship between the properties, structure and the nature and strength of the chemical bonding in the product or material? What are the raw materials used to make the product or material? How is the product or material manufactured? How are the by-products of production treated and managed? Is the product recyclable? Can any wastes during production or at the end of the product's use be repurposed into a useful product or material?

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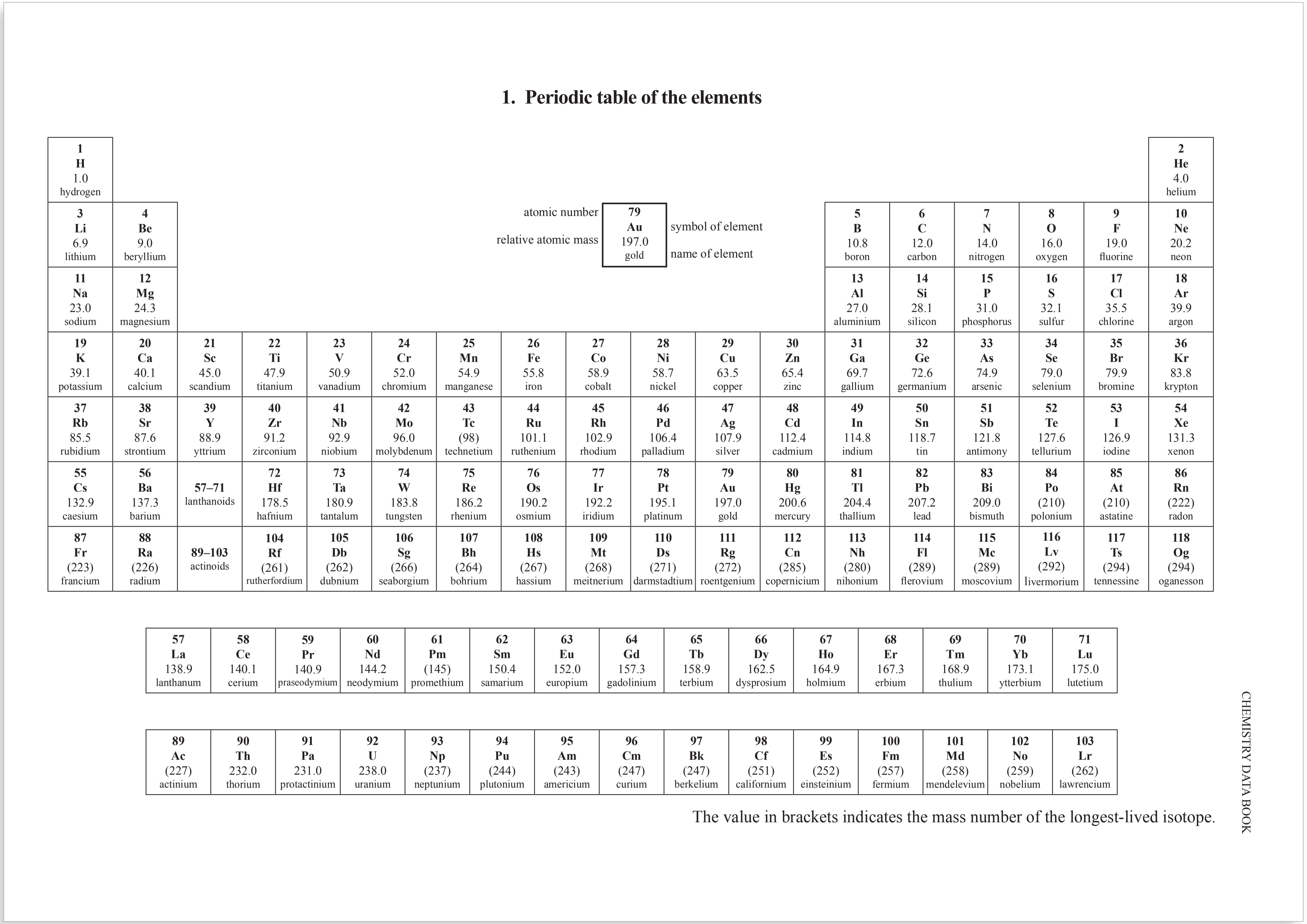
DATA BOOK

For your Year 12 chemistry exam, you will be given a data book that contains lots of useful information that you can use to help ace the exam. We thought it would be a good idea to get you familiar with the information that's in the data book, so we have put together a revised version that only includes things that are relevant to Year 11 chemistry. Have fun!



1. Periodic table of the elements

atomic number relative atomic mass	**79** **Au** 197.0 gold	symbol of element name of element

1 **H** 1.0 hydrogen																	**2** **He** 4.0 helium
3 **Li** 6.9 lithium	**4** **Be** 9.0 beryllium											**5** **B** 10.8 boron	**6** **C** 12.0 carbon	**7** **N** 14.0 nitrogen	**8** **O** 16.0 oxygen	**9** **F** 19.0 fluorine	**10** **Ne** 20.2 neon
11 **Na** 23.0 sodium	**12** **Mg** 24.3 magnesium											**13** **Al** 27.0 aluminium	**14** **Si** 28.1 silicon	**15** **P** 31.0 phosphorus	**16** **S** 32.1 sulfur	**17** **Cl** 35.5 chlorine	**18** **Ar** 39.9 argon
19 **K** 39.1 potassium	**20** **Ca** 40.1 calcium	**21** **Sc** 45.0 scandium	**22** **Ti** 47.9 titanium	**23** **V** 50.9 vanadium	**24** **Cr** 52.0 chromium	**25** **Mn** 54.9 manganese	**26** **Fe** 55.8 iron	**27** **Co** 58.9 cobalt	**28** **Ni** 58.7 nickel	**29** **Cu** 63.5 copper	**30** **Zn** 65.4 zinc	**31** **Ga** 69.7 gallium	**32** **Ge** 72.6 germanium	**33** **As** 74.9 arsenic	**34** **Se** 79.0 selenium	**35** **Br** 79.9 bromine	**36** **Kr** 83.8 krypton
37 **Rb** 85.5 rubidium	**38** **Sr** 87.6 strontium	**39** **Y** 88.9 yttrium	**40** **Zr** 91.2 zirconium	**41** **Nb** 92.9 niobium	**42** **Mo** 96.0 molybdenum	**43** **Tc** (98) technetium	**44** **Ru** 101.1 ruthenium	**45** **Rh** 102.9 rhodium	**46** **Pd** 106.4 palladium	**47** **Ag** 107.9 silver	**48** **Cd** 112.4 cadmium	**49** **In** 114.8 indium	**50** **Sn** 118.7 tin	**51** **Sb** 121.8 antimony	**52** **Te** 127.6 tellurium	**53** **I** 126.9 iodine	**54** **Xe** 131.3 xenon
55 **Cs** 132.9 caesium	**56** **Ba** 137.3 barium	**57–71** lanthanoids	**72** **Hf** 178.5 hafnium	**73** **Ta** 180.9 tantalum	**74** **W** 183.8 tungsten	**75** **Re** 186.2 rhenium	**76** **Os** 190.2 osmium	**77** **Ir** 192.2 iridium	**78** **Pt** 195.1 platinum	**79** **Au** 197.0 gold	**80** **Hg** 200.6 mercury	**81** **Tl** 204.4 thallium	**82** **Pb** 207.2 lead	**83** **Bi** 209.0 bismuth	**84** **Po** (210) polonium	**85** **At** (210) astatine	**86** **Rn** (222) radon
87 **Fr** (223) francium	**88** **Ra** (226) radium	**89–103** actinoids	**104** **Rf** (261) rutherfordium	**105** **Db** (262) dubnium	**106** **Sg** (266) seaborgium	**107** **Bh** (264) bohrium	**108** **Hs** (267) hassium	**109** **Mt** (268) meitnerium	**110** **Ds** (271) darmstadtium	**111** **Rg** (272) roentgenium	**112** **Cn** (285) copernicium	**113** **Nh** (280) nihonium	**114** **Fl** (289) flerovium	**115** **Mc** (289) moscovium	**116** **Lv** (292) livermorium	**117** **Ts** (294) tennessine	**118** **Og** (294) oganesson

57 **La** 138.9 lanthanum	**58** **Ce** 140.1 cerium	**59** **Pr** 140.9 praseodymium	**60** **Nd** 144.2 neodymium	**61** **Pm** (145) promethium	**62** **Sm** 150.4 samarium	**63** **Eu** 152.0 europium	**64** **Gd** 157.3 gadolinium	**65** **Tb** 158.9 terbium	**66** **Dy** 162.5 dysprosium	**67** **Ho** 164.9 holmium	**68** **Er** 167.3 erbium	**69** **Tm** 168.9 thulium	**70** **Yb** 173.1 ytterbium	**71** **Lu** 175.0 lutetium

89 **Ac** (227) actinium	**90** **Th** 232.0 thorium	**91** **Pa** 231.0 protactinium	**92** **U** 238.0 uranium	**93** **Np** (237) neptunium	**94** **Pu** (244) plutonium	**95** **Am** (243) americium	**96** **Cm** (247) curium	**97** **Bk** (247) berkelium	**98** **Cf** (251) californium	**99** **Es** (252) einsteinium	**100** **Fm** (257) fermium	**101** **Md** (258) mendelevium	**102** **No** (259) nobelium	**103** **Lr** (262) lawrencium

The value in brackets indicates the mass number of the longest-lived isotope.

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2. Electrochemical series

Reaction	Standard electrode potential (E^0) in volts at 25 °C
$F_2(g) + 2e^- \rightleftharpoons 2F^-(aq)$	+2.87
$H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l)$	+1.77
$Au^+(aq) + e^- \rightleftharpoons Au(s)$	+1.68
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(1)$	+1.23
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.09
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(aq)$	+0.68
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(g)$	+0.14
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	–0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	–0.14
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	–0.25
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	–0.28
$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$	–0.40
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	–0.44
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	–0.76
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	–0.83
$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s)$	–1.18
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	–1.66
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$	–2.37
$Na^+(aq) + e^- \rightleftharpoons Na(s)$	–2.71
$Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$	–2.87
$K^+(aq) + e^- \rightleftharpoons K(s)$	–2.93
$Li^+(aq) + e^- \rightleftharpoons Li(s)$	–3.04

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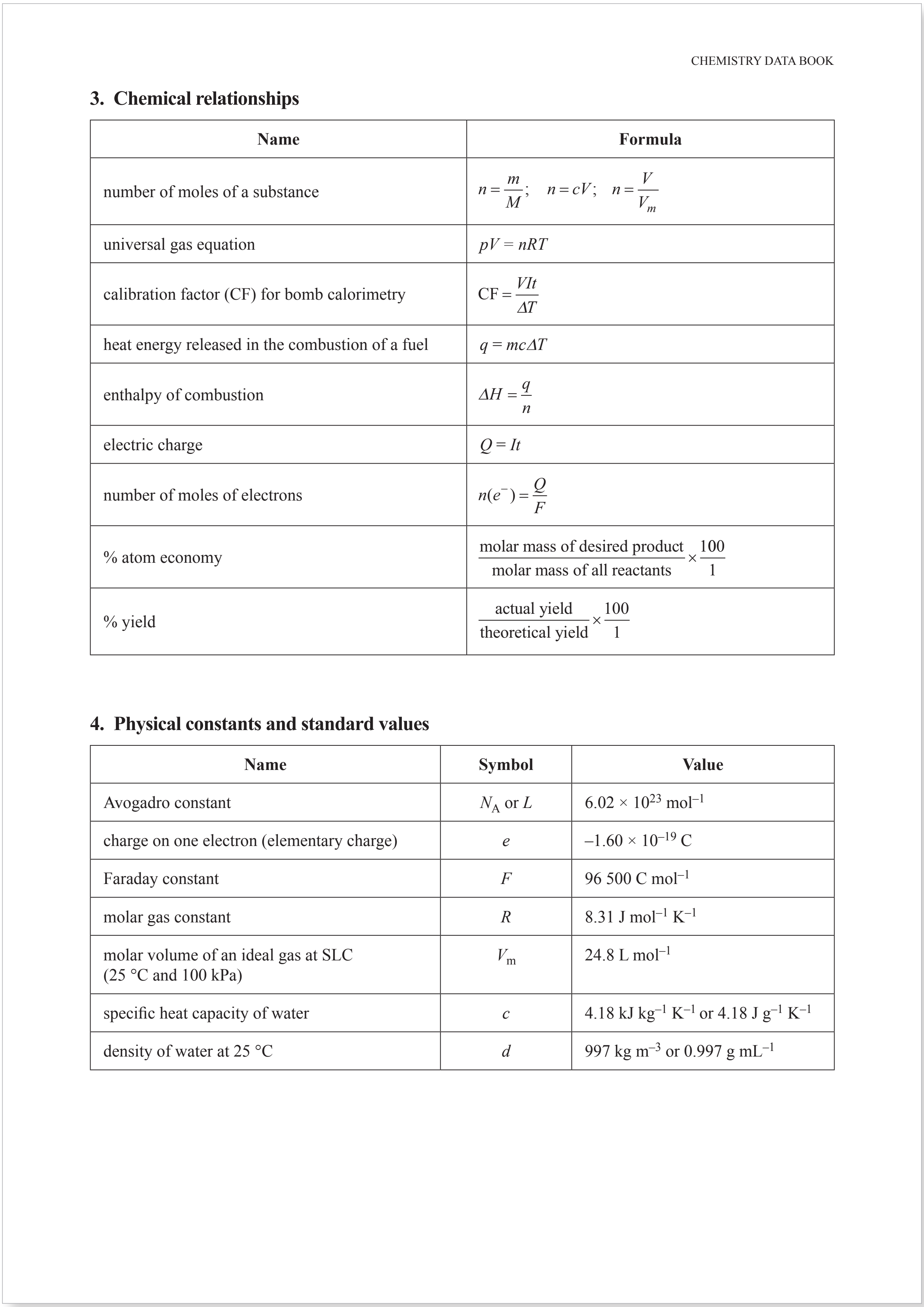

3. Chemical relationships

Name	Formula
number of moles of a substance	$n = \frac{m}{M}; \quad n = cV; \quad n = \frac{V}{V_m}$
universal gas equation	$pV = nRT$
calibration factor (CF) for bomb calorimetry	$\text{CF} = \frac{VIt}{\Delta T}$
heat energy released in the combustion of a fuel	$q = mc\Delta T$
enthalpy of combustion	$\Delta H = \frac{q}{n}$
electric charge	$Q = It$
number of moles of electrons	$n(e^-) = \frac{Q}{F}$
% atom economy	$\frac{\text{molar mass of desired product}}{\text{molar mass of all reactants}} \times \frac{100}{1}$
% yield	$\frac{\text{actual yield}}{\text{theoretical yield}} \times \frac{100}{1}$

4. Physical constants and standard values

Name	Symbol	Value
Avogadro constant	N_A or L	6.02×10^{23} mol^{-1}
charge on one electron (elementary charge)	e	-1.60×10^{-19} C
Faraday constant	F	96 500 C mol^{-1}
molar gas constant	R	8.31 J mol^{-1} K^{-1}
molar volume of an ideal gas at SLC (25 °C and 100 kPa)	V_m	24.8 L mol^{-1}
specific heat capacity of water	c	4.18 kJ kg^{-1} K^{-1} or 4.18 J g^{-1} K^{-1}
density of water at 25 °C	d	997 kg m^{-3} or 0.997 g mL^{-1}

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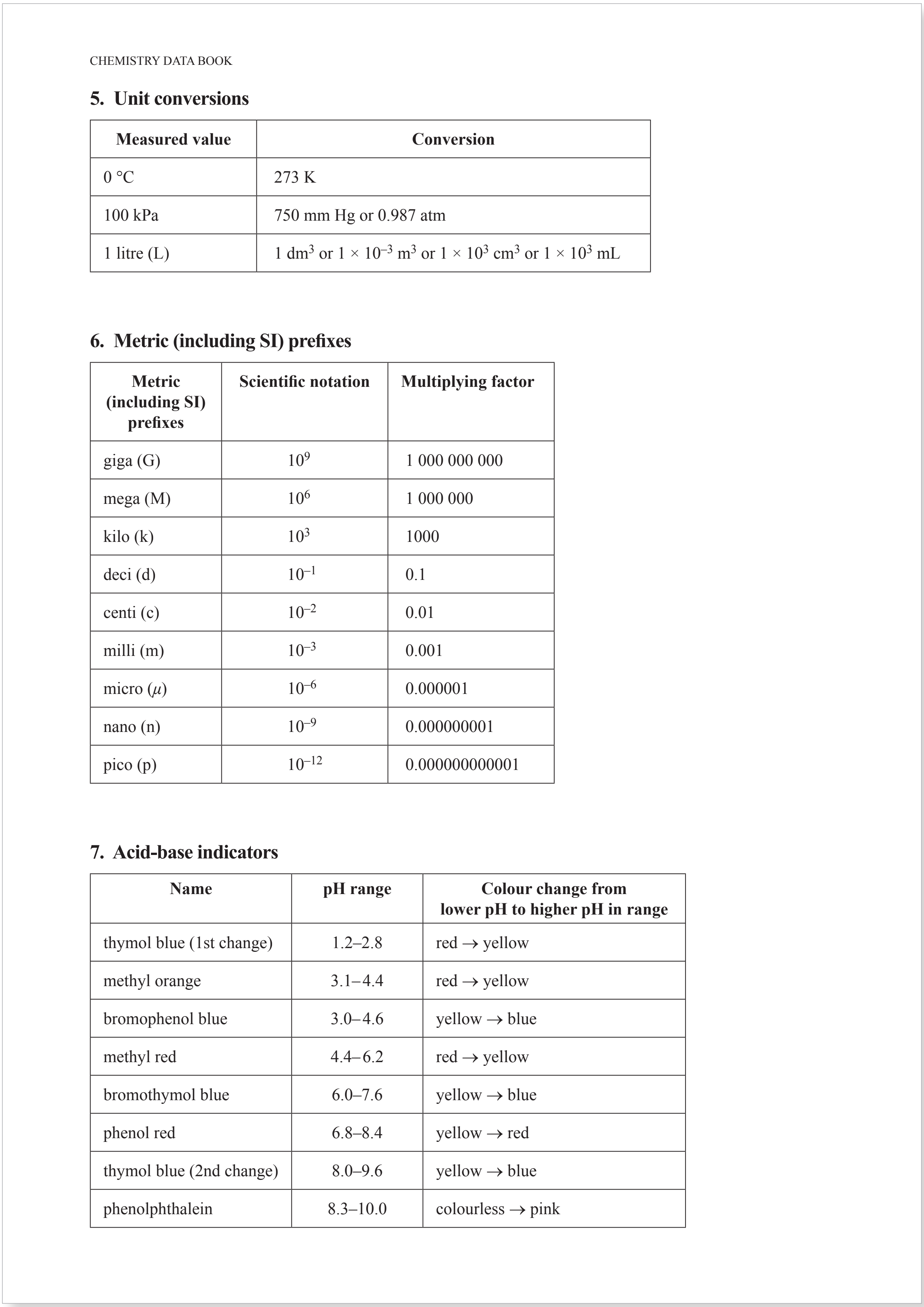


5. Unit conversions

Measured value	Conversion
0 °C	273 K
100 kPa	750 mm Hg or 0.987 atm
1 litre (L)	1 dm^3 or 1×10^{-3} m^3 or 1×10^3 cm^3 or 1×10^3 mL

6. Metric (including SI) prefixes

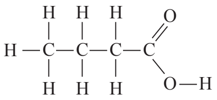
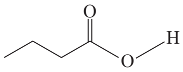
Metric (including SI) prefixes	Scientific notation	Multiplying factor
giga (G)	10^9	1 000 000 000
mega (M)	10^6	1 000 000
kilo (k)	10^3	1000
deci (d)	10^{-1}	0.1
centi (c)	10^{-2}	0.01
milli (m)	10^{-3}	0.001
micro (μ)	10^{-6}	0.000001
nano (n)	10^{-9}	0.000000001
pico (p)	10^{-12}	0.000000000001

7. Acid-base indicators

Name	pH range	Colour change from lower pH to higher pH in range
thymol blue (1st change)	1.2–2.8	red → yellow
methyl orange	3.1–4.4	red → yellow
bromophenol blue	3.0–4.6	yellow → blue
methyl red	4.4–6.2	red → yellow
bromothymol blue	6.0–7.6	yellow → blue
phenol red	6.8–8.4	yellow → red
thymol blue (2nd change)	8.0–9.6	yellow → blue
phenolphthalein	8.3–10.0	colourless → pink

8. Representations of organic molecules

The following table shows different representations of organic molecules, using butanoic acid as an example.

Formula	Representation
molecular formula	$C_4H_8O_2$
structural formula	
semi-structural (condensed) formula	$CH_3CH_2CH_2COOH$ or $CH_3(CH_2)_2COOH$
skeletal structure	

Adapted from VCAA Chemistry Data book

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ACKNOWLEDGEMENTS

Images

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