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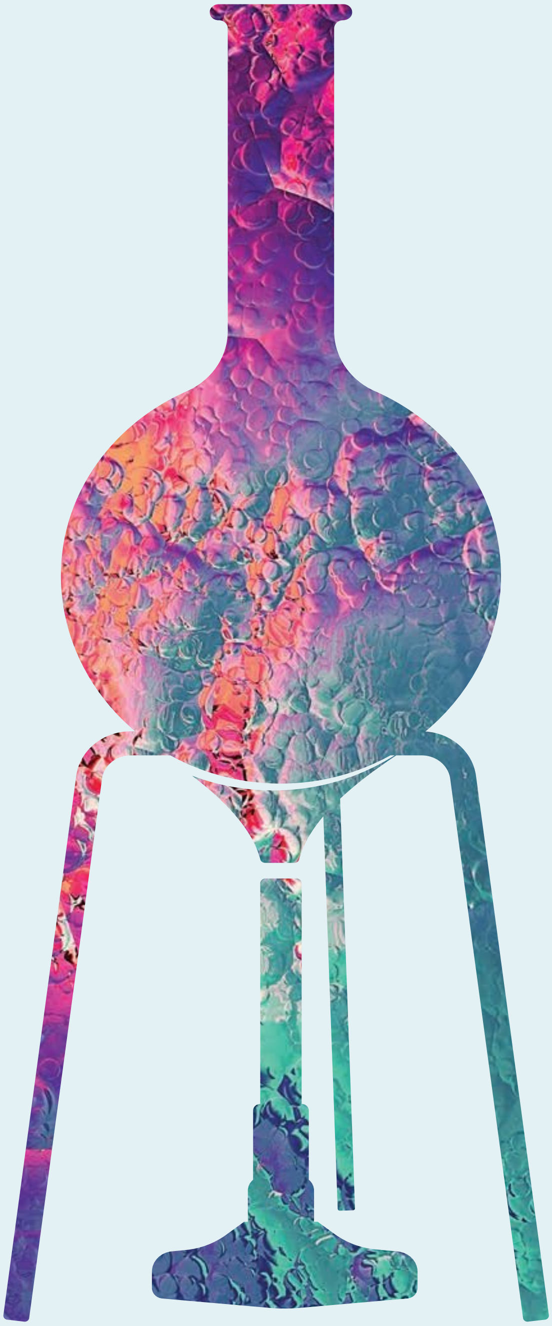
VCE UNITS
1&2



INTERACTIVE
TEXTBOOK
INCLUDED

Cambridge Senior Science

Chemistry

Brett **Drummond**
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Lauren **Angley**
Scott **Fraser**
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About the authors

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Scott Fraser has a wide-ranging set of experiences starting with a biomedical science undergraduate degree, followed by a Grad. Dip. Chemistry and PhD in chemistry. He has taught chemistry and biochemistry subjects across all levels of study and also has his own private teaching company that helps guide and support students towards medical careers.

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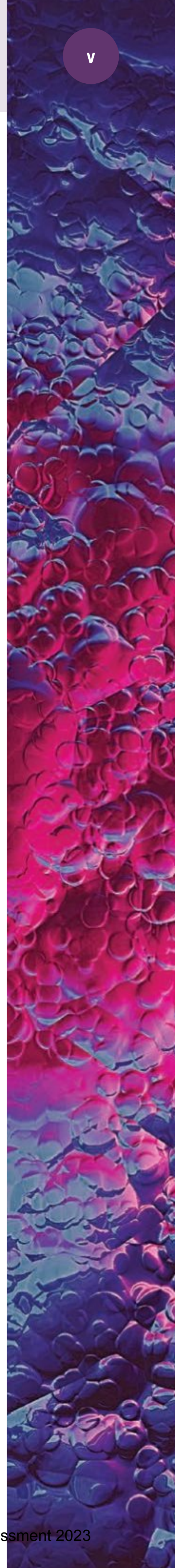
Zhen (James) Wang is a dedicated educator who is passionate about teaching Chemistry to high school students. He holds a PhD in Organic Chemistry and has previously worked as a post-doctoral fellow in multiple universities. He has extensive classroom experience in delivering VCE Chemistry and enjoys the challenge of communicating complex scientific concepts in an elegant manner.

Contents

UNIT 1 How can the diversity of materials be explained?

UNIT 2

How do chemical reactions shape the natural world?



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Answers are available in the Interactive Textbook and the Teacher Resources.

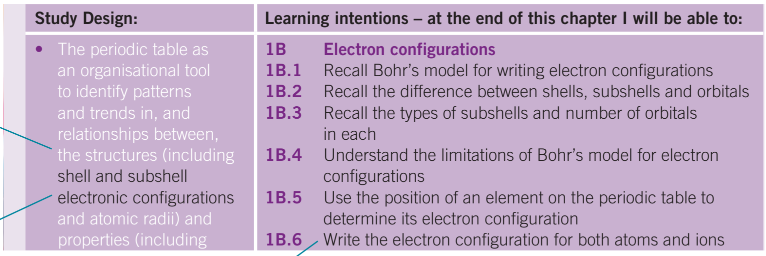
Overview: How to use this resource

This overview guides you through all the components of the **print and PDF textbooks**, the **Interactive Textbook (ITB)** and the teacher resources in the **Online Teaching Suite (OTS)**. Users of the award-winning *Cambridge Science 7–10 for the Victorian Curriculum* will recognise some similarities with this senior science resource, including the hosting of the digital material on the Edjin platform, which was developed by *Cambridge HOTmaths* and is already being used successfully by thousands of teachers and students across Victoria.

Print book features

Learning intentions

In the Curriculum table at the start of each chapter, the Study Design dot points are translated into Learning Intentions, describing what students should be able to do by the end of the chapter.



Study Design:	Learning intentions – at the end of this chapter I will be able to:	
- 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	**1B**	**Electron configurations**
	1B.1	Recall Bohr's model for writing electron configurations
	1B.2	Recall the difference between shells, subshells and orbitals
	1B.3	Recall the types of subshells and number of orbitals in each
	1B.4	Understand the limitations of Bohr's model for electron configurations
	1B.5	Use the position of an element on the periodic table to determine its electron configuration
	1B.6	Write the electron configuration for both atoms and ions

White text indicates the portion of the dot point covered by other sections

Black text indicates the portion of the dot point covered by the section shown in the second column

Learning intentions are turned into Success Criteria (achievement standards) at the end of the chapter and are assessed in the Chapter review and tracked in the Checklists

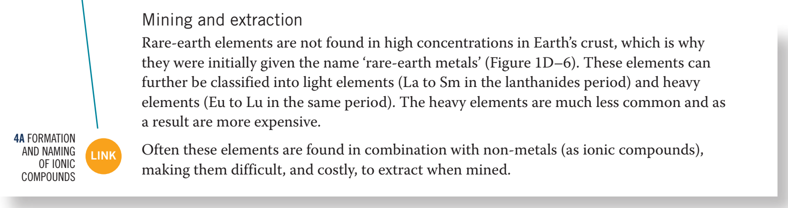
Relevant Study Design dot points are repeated at the start of each section in the chapter, and an overall curriculum grid is provided in the teacher resources.

Concept maps

Concept maps display each chapter's structure with annotations emphasising interconnectedness, providing a great memory aid. The versions in the ITB are hyperlinked and offer an alternative way of navigating through the course. An overall look at all concept maps of Units 1 & 2 is also provided after this overview.

Links

The interconnectedness of topics in VCE Chemistry is demonstrated through links between sections, displayed in the margins. In the ITB, these are hyperlinks that provide an alternative way of navigating through the course.

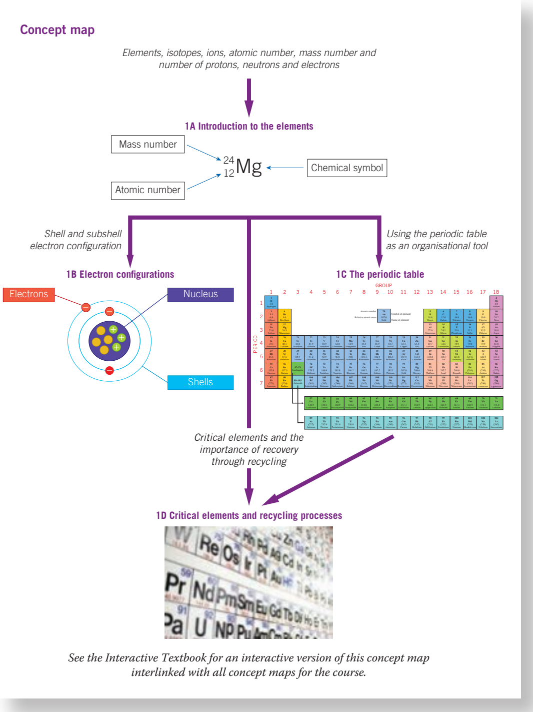


Mining and extraction

Rare-earth elements are not found in high concentrations in Earth's crust, which is why they were initially given the name 'rare-earth metals' (Figure 1D–6). These elements can further be classified into light elements (La to Sm in the lanthanides period) and heavy elements (Eu to Lu in the same period). The heavy elements are much less common and as a result are more expensive.

4A FORMATION AND NAMING OF IONIC COMPOUNDS — LINK

Often these elements are found in combination with non-metals (as ionic compounds), making them difficult, and costly, to extract when mined.



Concept map

See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

Chapter sections

Chapters are divided into numbered sections, each with a consistent set of features.

Engage

At the start of each section, these boxes provide points of interest for the topic, emphasising its place in Chemistry. This material, though not assessable, can be used as examples of applications.

Explain

This icon marks the start of essential content that is assessed.

Glossary

Scientific terms are highlighted in the text; definitions are given in the margin of the print and PDF textbooks, or on mouseover in the ITB; and the terms are listed at the start of each chapter and section.

Check-in questions

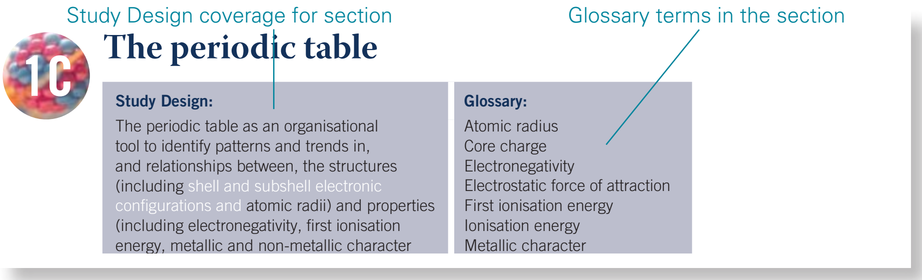
Each section in the chapter has one or more sets of check-in questions, for formative assessment. Full answers are provided in the digital resources.

Skills

Skills boxes in every section provide advice and guidance on how to answer and prepare for questions, especially in examinations. The ITB has guided video versions of these, which provide extra comments and an alternative medium of delivery.

Worked examples

Worked examples are provided for questions requiring computation – for example, in stoichiometry.



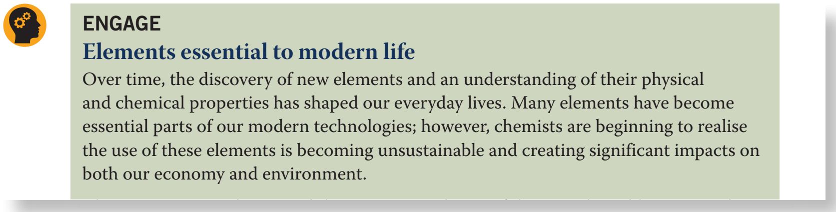
Study Design coverage for section

Glossary terms in the section

1C The periodic table

Study Design:
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

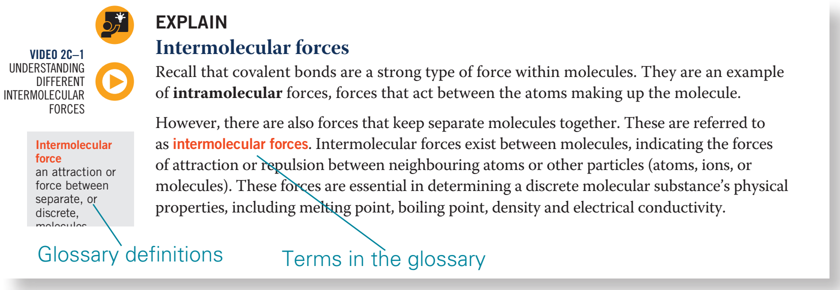
Glossary:
Atomic radius
Core charge
Electronegativity
Electrostatic force of attraction
First ionisation energy
Ionisation energy
Metallic character



ENGAGE

Elements essential to modern life

Over time, the discovery of new elements and an understanding of their physical and chemical properties has shaped our everyday lives. Many elements have become essential parts of our modern technologies; however, chemists are beginning to realise the use of these elements is becoming unsustainable and creating significant impacts on both our economy and environment.



EXPLAIN

Intermolecular forces

VIDEO 2C–1
UNDERSTANDING DIFFERENT INTERMOLECULAR FORCES

Recall that covalent bonds are a strong type of force within molecules. They are an example of **intramolecular** forces, forces that act between the atoms making up the molecule.

However, there are also forces that keep separate molecules together. These are referred to as **intermolecular forces**. Intermolecular forces exist between molecules, indicating the forces of attraction or repulsion between neighbouring atoms or other particles (atoms, ions, or molecules). These forces are essential in determining a discrete molecular substance's physical properties, including melting point, boiling point, density and electrical conductivity.

Intermolecular force an attraction or force between separate, or discrete,

Glossary definitions

Terms in the glossary

Check-in questions – Set 1

1 What is the main difference between Lewis structures and the valence shell electron pair repulsion (VSEPR) model?
2 Apart from linear, name three other molecular geometries according to VSEPR theory.
3 What are two factors that determine molecular shape?

2D SKILLS

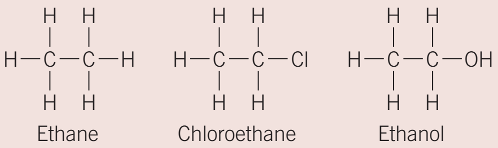
Using Venn diagrams in chemistry

You may well have come across Venn diagrams in your mathematics studies, but did you know they are very useful in comparing structures or properties of different structures in chemistry as well?

If you are not familiar with a Venn diagram, it is an illustration that uses slightly overlapping circles to show the relationships among items. Things that are common between items are included in the overlapping region of the circles, while those that are distinct are in the stand-alone regions of each circle.

Worked example 2C–1: Strength of intermolecular forces

a What type of intermolecular force(s) would occur between the molecules of the following pure substances?

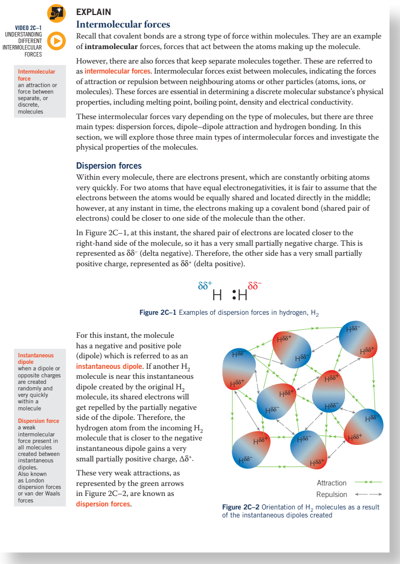


b List the molecules in order of increasing strength of the intermolecular forces acting between them in the pure substances, from weakest to strongest, and explain your response.

ISBN 978-1-009-22964-7 Drummond et al

Charts, diagrams and tables

Detailed charts integrating text and diagrams, and illustrated tables, feature throughout the print book.



VIDEO 2C–1 UNDERSTANDING DIFFERENT INTERMOLECULAR FORCES

EXPLAIN

Intermolecular forces

Recall that covalent bonds are a strong type of force within molecules. They are an example of **intramolecular** forces, forces that act between the atoms making up the molecule.

However, there are also forces that keep separate molecules together. These are referred to as **intermolecular forces**. Intermolecular forces exist between molecules, indicating the forces of attraction or repulsion between neighbouring atoms or other particles (atoms, ions, or molecules). These forces are essential in determining a discrete molecular substance's physical properties, including melting point, boiling point, density and electrical conductivity.

Intermolecular force an attraction or force between separate, or discrete, molecules

These intermolecular forces vary depending on the type of molecules, but there are three main types: dispersion forces, dipole–dipole attraction and hydrogen bonding. In this section, we will explore those three main types of intermolecular forces and investigate the physical properties of the molecules.

Dispersion forces

Within every molecule, there are electrons present, which are constantly orbiting atoms very quickly. For two atoms that have equal electronegativities, it is fair to assume that the electrons between the atoms would be equally shared and located directly in the middle; however, at any instant in time, the electrons making up a covalent bond (shared pair of electrons) could be closer to one side of the molecule than the other.

In Figure 2C–1, at this instant, the shared pair of electrons are located closer to the right-hand side of the molecule, so it has a very small partially negative charge. This is represented as δδ⁻ (delta negative). Therefore, the other side has a very small partially positive charge, represented as δδ⁺ (delta positive).

$$^{\delta\delta^+}H\ {:}H^{\delta\delta^-}$$

Figure 2C–1 Examples of dispersion forces in hydrogen, H_2

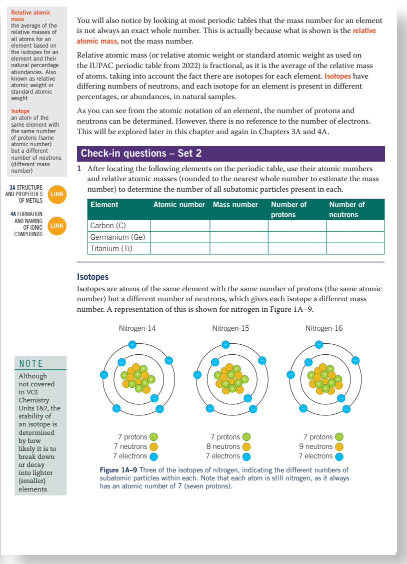
For this instant, the molecule has a negative and positive pole (dipole) which is referred to as an **instantaneous dipole**. If another H_2 molecule is near this instantaneous dipole created by the original H_2 molecule, its shared electrons will get repelled by the partially negative side of the dipole. Therefore, the hydrogen atom from the incoming H_2 molecule that is closer to the negative instantaneous dipole gains a very small partially positive charge, Δδ⁺.

These very weak attractions, as represented by the green arrows in Figure 2C–2, are known as **dispersion forces**.

Instantaneous dipole when a dipole or opposite charges are created randomly and very quickly within a molecule

Dispersion force a weak intermolecular force present in all molecules created between instantaneous dipoles. Also known as London dispersion forces or van der Waals forces

Figure 2C–2 Orientation of H_2 molecules as a result of the instantaneous dipoles created



You will also notice by looking at most periodic tables that the mass number for an element is not always an exact whole number. This is actually because what is shown is the **relative atomic mass**, not the mass number.

Relative atomic mass (or relative atomic weight or standard atomic weight as used on the IUPAC periodic table from 2022) is fractional, as it is the average of the relative mass of atoms, taking into account the fact there are isotopes for each element. **Isotopes** have differing numbers of neutrons, and each isotope for an element is present in different percentages, or abundances, in natural samples.

As you can see from the atomic notation of an element, the number of protons and neutrons can be determined. However, there is no reference to the number of electrons. This will be explored later in this chapter and again in Chapters 3A and 4A.

Relative atomic mass the average of the relative masses of all atoms for an element based on the isotopes for an element and their natural percentage abundances. Also known as relative atomic weight or standard atomic weight

Isotope an atom of the same element with the same number of protons (same atomic number) but a different number of neutrons (different mass number)

3A STRUCTURE AND PROPERTIES OF METALS LINK

4A FORMATION AND NAMING OF IONIC COMPOUNDS LINK

Check-in questions – Set 2

1 After locating the following elements on the periodic table, use their atomic numbers and relative atomic masses (rounded to the nearest whole number to estimate the mass number) to determine the number of all subatomic particles present in each.

Element	Atomic number	Mass number	Number of protons	Number of neutrons
Carbon (C)				
Germanium (Ge)				
Titanium (Ti)				

Isotopes

Isotopes are atoms of the same element with the same number of protons (the same atomic number) but a different number of neutrons, which gives each isotope a different mass number. A representation of this is shown for nitrogen in Figure 1A–9.

Figure 1A–9 Three of the isotopes of nitrogen, indicating the different numbers of subatomic particles within each. Note that each atom is still nitrogen, as it always has an atomic number of 7 (seven protons).

NOTE

Although not covered in VCE Chemistry Units 1&2, the stability of an isotope is determined by how likely it is to break down or decay into lighter (smaller) elements.

Section questions

Summative assessment is provided at the end of each section, again with full answers provided in the digital resources.

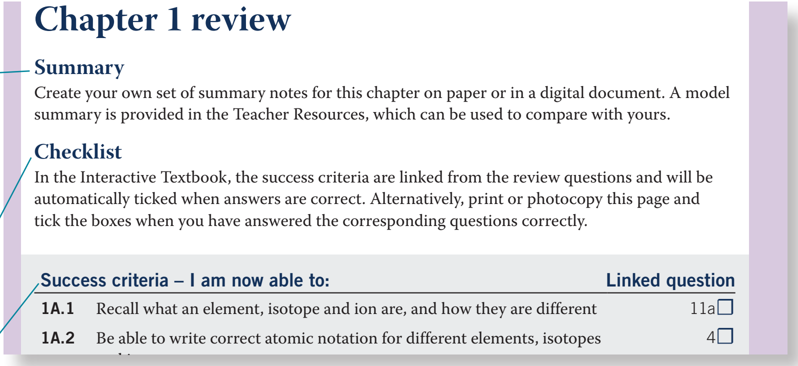
Section 1D questions

1 Define 'critical element'.

2 List three elements in each of the following classifications.

Chapter reviews

Summaries: Students are encouraged to make their own set of summary notes to help them assimilate the material. Model summaries are provided in the Teacher Resources and are to be given to students who need help. Creating summaries can also be turned into an assessment task, with the models serving as the answer. **Checklists** and **Success criteria:** The learning intentions from the front of the chapter are listed again in the form of success criteria linked to the **multiple-choice** and **short-answer questions** that follow. The checklists can be printed from the ITB, and students can tick off their achievement manually. If they do the questions in the ITB, they are ticked automatically when the questions are marked.



Chapter 1 review

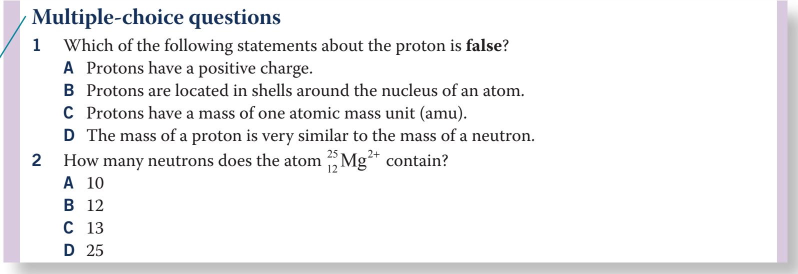
Summary

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

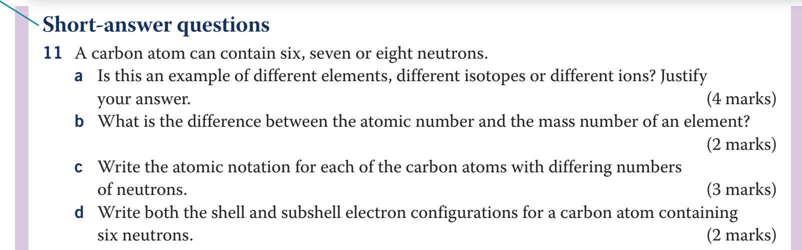
Success criteria – I am now able to:		Linked question
1A.1	Recall what an element, isotope and ion are, and how they are different	11a☐
1A.2	Be able to write correct atomic notation for different elements, isotopes	4☐



Multiple-choice questions

1 Which of the following statements about the proton is **false**?
- **A** Protons have a positive charge.
- **B** Protons are located in shells around the nucleus of an atom.
- **C** Protons have a mass of one atomic mass unit (amu).
- **D** The mass of a proton is very similar to the mass of a neutron.

2 How many neutrons does the atom $^{25}_{12}Mg^{2+}$ contain?
- **A** 10
- **B** 12
- **C** 13
- **D** 25



Short-answer questions

11 A carbon atom can contain six, seven or eight neutrons.
- **a** Is this an example of different elements, different isotopes or different ions? Justify your answer. (4 marks)
- **b** What is the difference between the atomic number and the mass number of an element? (2 marks)
- **c** Write the atomic notation for each of the carbon atoms with differing numbers of neutrons. (3 marks)
- **d** Write both the shell and subshell electron configurations for a carbon atom containing six neutrons. (2 marks)

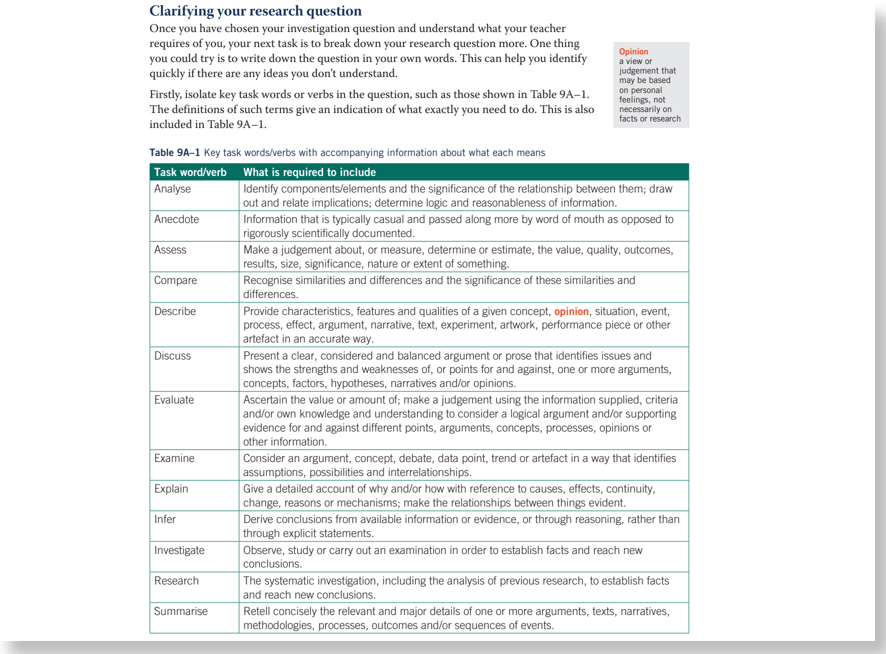
Unit revision exercises

Each Unit has a revision exercise in the print book, with both multiple-choice and short-answer questions.

Special content

Two aspects are highlighted here:

- **Chapter 9 Research investigations** includes information on how to plan for and conduct a research investigation; judge the validity of various sources; link research to knowledge, skills and broader implications; and finally present a research investigation.



Clarifying your research question

Once you have chosen your investigation question and understand what your teacher requires of you, your next task is to break down your research question more. One thing you could try is to write down the question in your own words. This can help you identify quickly if there are any ideas you don't understand.

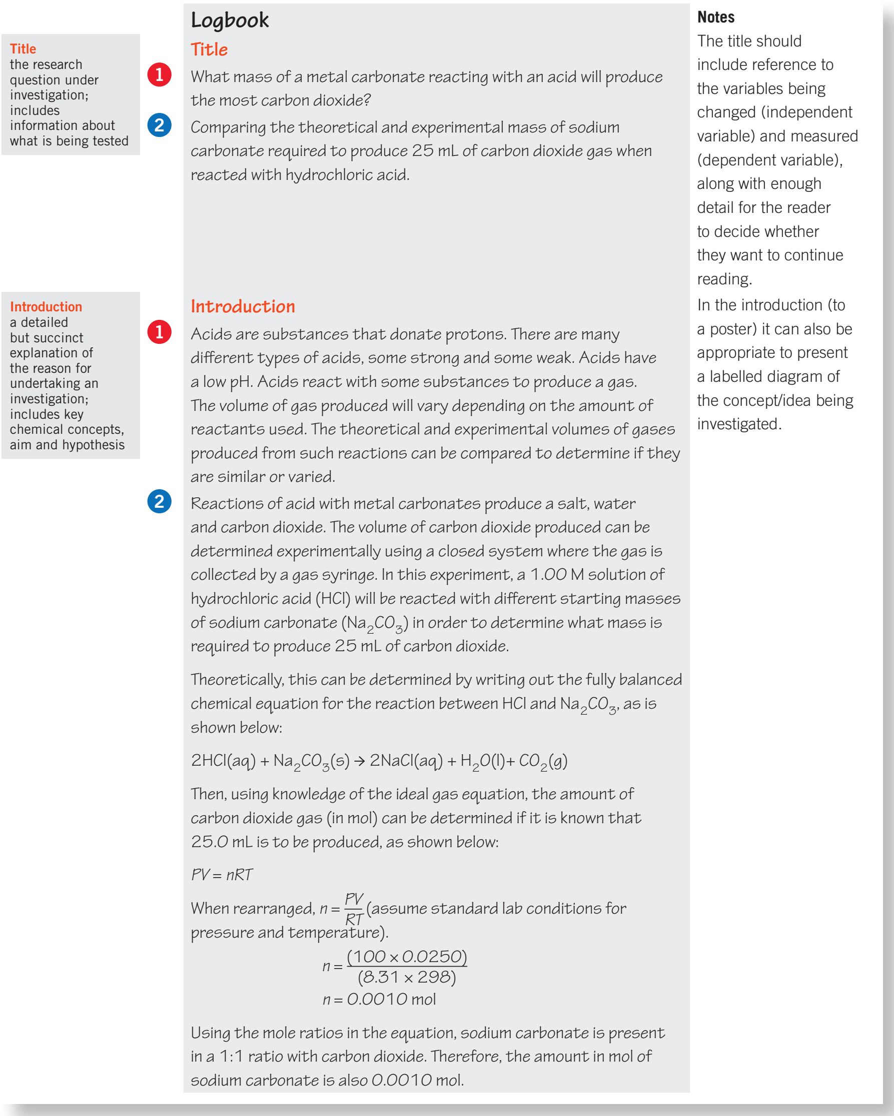
Firstly, isolate key task words or verbs in the question, such as those shown in Table 9A–1. The definitions of such terms give an indication of what exactly you need to do. This is also included in Table 9A–1.

Opinion
a view or judgement that may be based on personal feelings, not necessarily on facts or research

Table 9A–1 Key task words/verbs with accompanying information about what each means

Task word/verb	What is required to include
Analyse	Identify components/elements and the significance of the relationship between them; draw out and relate implications; determine logic and reasonableness of information.
Anecdote	Information that is typically casual and passed along more by word of mouth as opposed to rigorously scientifically documented.
Assess	Make a judgement about, or measure, determine or estimate, the value, quality, outcomes, results, size, significance, nature or extent of something.
Compare	Recognise similarities and differences and the significance of these similarities and differences.
Describe	Provide characteristics, features and qualities of a given concept, **opinion**, situation, event, process, effect, argument, narrative, text, experiment, artwork, performance piece or other artefact in an accurate way.
Discuss	Present a clear, considered and balanced argument or prose that identifies issues and shows the strengths and weaknesses of, or points for and against, one or more arguments, concepts, factors, hypotheses, narratives and/or opinions.
Evaluate	Ascertain the value or amount of; make a judgement using the information supplied, criteria and/or own knowledge and understanding to consider a logical argument and/or supporting evidence for and against different points, arguments, concepts, processes, opinions or other information.
Examine	Consider an argument, concept, debate, data point, trend or artefact in a way that identifies assumptions, possibilities and interrelationships.
Explain	Give a detailed account of why and/or how with reference to causes, effects, continuity, change, reasons or mechanisms; make the relationships between things evident.
Infer	Derive conclusions from available information or evidence, or through reasoning, rather than through explicit statements.
Investigate	Observe, study or carry out an examination in order to establish facts and reach new conclusions.
Research	The systematic investigation, including the analysis of previous research, to establish facts and reach new conclusions.
Summarise	Retell concisely the relevant and major details of one or more arguments, texts, narratives, methodologies, processes, outcomes and/or sequences of events.

- **Chapter 16 Practical investigations** includes the different ways scientific investigations may be undertaken by students. It features the modelling of logbook development for students' own practical investigations, with detailed examples. The digital textbook includes information on other investigation methodologies.



Title
the research question under investigation; includes information about what is being tested

Logbook

Title

① What mass of a metal carbonate reacting with an acid will produce the most carbon dioxide?

② Comparing the theoretical and experimental mass of sodium carbonate required to produce 25 mL of carbon dioxide gas when reacted with hydrochloric acid.

Notes
The title should include reference to the variables being changed (independent variable) and measured (dependent variable), along with enough detail for the reader to decide whether they want to continue reading.

In the introduction (to a poster) it can also be appropriate to present a labelled diagram of the concept/idea being investigated.

Introduction
a detailed but succinct explanation of the reason for undertaking an investigation; includes key chemical concepts, aim and hypothesis

Introduction

① Acids are substances that donate protons. There are many different types of acids, some strong and some weak. Acids have a low pH. Acids react with some substances to produce a gas. The volume of gas produced will vary depending on the amount of reactants used. The theoretical and experimental volumes of gases produced from such reactions can be compared to determine if they are similar or varied.

② Reactions of acid with metal carbonates produce a salt, water and carbon dioxide. The volume of carbon dioxide produced can be determined experimentally using a closed system where the gas is collected by a gas syringe. In this experiment, a 1.00 M solution of hydrochloric acid (HCl) will be reacted with different starting masses of sodium carbonate (Na_2CO_3) in order to determine what mass is required to produce 25 mL of carbon dioxide.

Theoretically, this can be determined by writing out the fully balanced chemical equation for the reaction between HCl and Na_2CO_3, as is shown below:

$$2HCl(aq) + Na_2CO_3(s) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$$

Then, using knowledge of the ideal gas equation, the amount of carbon dioxide gas (in mol) can be determined if it is known that 25.0 mL is to be produced, as shown below:

$$PV = nRT$$

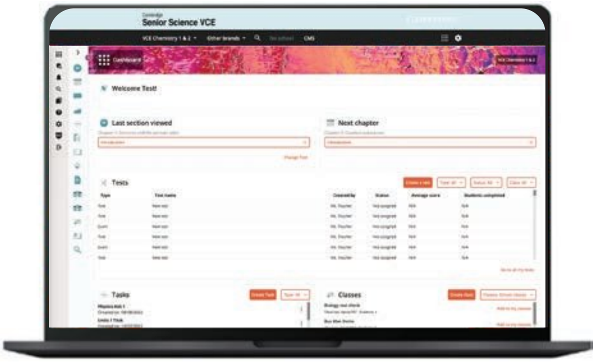
When rearranged, $n = \frac{PV}{RT}$ (assume standard lab conditions for pressure and temperature).

$$n = \frac{(100 \times 0.0250)}{(8.31 \times 298)}$$

$$n = 0.0010 \text{ mol}$$

Using the mole ratios in the equation, sodium carbonate is present in a 1:1 ratio with carbon dioxide. Therefore, the amount in mol of sodium carbonate is also 0.0010 mol.

Interactive Textbook features



The digital version of the textbook is hosted on the Edjin platform, offering easy navigation, excellent on-screen display and multimedia assets, as well as auto-marking of multiple-choice questions, and workspaces for other questions with self-assessment and confidence-rating tools. The different kinds of digital assets are listed below:

- Printable **Worksheets** with extra questions and activities (and content in some cases) are provided for most chapters, marked by an icon in the margin, as shown on the right.
- **Videos** are provided for all chapters and are of two kinds: **concept videos** demonstrate or illustrate important theory, while **skills and example videos** work through the textbook's skills and example boxes, providing extra explanation and guidance. Some videos are provided in the print pages as QR codes for immediate access and review.
- **Answers** (suggested responses) to questions are provided as printable documents in the teacher resources and, if enabled by the teacher, below the question there are workspaces for short-answer questions in the ITB.
- **Prior knowledge** can be tested with an auto-marked quiz with questions from the Year 9 and 10 titles in the *Cambridge Science for the Victorian Curriculum* series.

WORKSHEET 1A–1
ATOMS AND SUBATOMIC PARTICLES

VIDEO 3B–2
SKILLS: WRITING CHEMICAL EQUATIONS

Online Teaching Suite features (teacher resources)

The OTS provides Edjin's learning management system, which allows teachers to set tasks, track progress and scores, prepare reports on individuals and the class, and give students feedback. The assets include:

- **Curriculum Grid** and **teaching programs**
- Editable and printable **Chapter tests** with answers
- **Checklists** with linkage to the success criteria for the chapter question sets and tests
- A **question bank** and test generator, with answers
- **Practice exams** and **assessment tasks**, with answers
- Editable versions of **Worksheets** in the Interactive Textbook, and answers to them
- Downloadable, editable and printable **practicals**
- Editable and printable chapter **summaries** (model answers for the chapter summary activity)
- **Teacher notes** on selected content with additional theory explanation and suggestions for further activities and resources
- **Curated links** to internet resources, such as videos and interactives.

Overview: Aboriginal and Torres Strait Islander knowledge, cultures and history

The VCE Chemistry Study Design includes aspects of Aboriginal and Torres Strait Islander knowledge, cultures and history. This overview is a guide to coverage in this resource.

Aboriginal and Torres Strait Islander peoples' world views are highly integrated: each aspect of culture, history and society connects with all other aspects. Each community has their own personalised system of thinking, doing and knowing based on sharing culture and adapting to the environment around them.

In order to gain an understanding of any system, Indigenous or not, time and effort is needed to appreciate it. That time is limited in this course; and it is wrong to try and generalise the Indigenous culture of Australia, or even of Victoria. Instead, the coverage in the resource should be taken as a collection of examples, and students should read up on or engage with their local Indigenous community to understand their cultural aspects.

This textbook includes examples of Aboriginal and Torres Strait Islander knowledge, cultures and history and in the interactive textbook there is guidance for tackling Unit 1 Area of Study 3 Investigation topic 3: *The chemistry of Aboriginal and Torres Strait Islander peoples' practices*.

In addition, for students, the Interactive Textbook includes an introductory guide prepared by First Nations consultants advising on approaches to studying Aboriginal and Torres Strait Islander knowledge, cultures and history, with links to further reading.

For teachers, the teacher resources include a guide to approaches to teaching Aboriginal and Torres Strait Islander knowledge, cultures and history in the VCE Chemistry course, with links to internet resources.

Guide to terms used in this resource

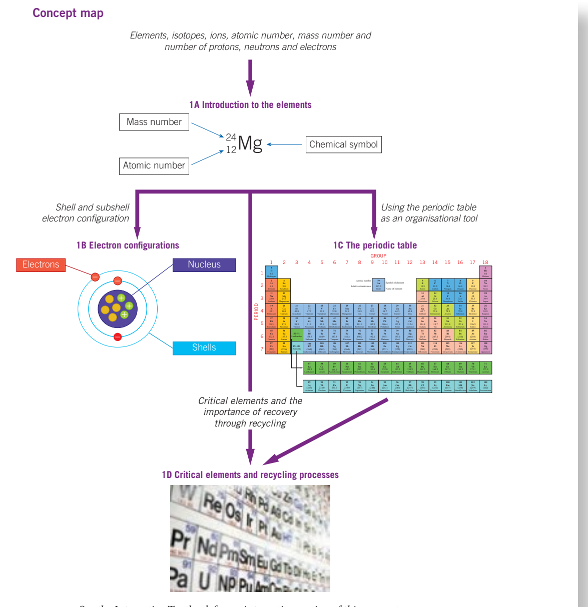
Language is very important in discussing Indigenous issues, especially given the past history of deliberately offensive usage in Australia, where language was used to oppress and control.

Indigenous First Australians and First Peoples of any country	Respectful usage requires a capital 'I'.
First Australians, First Nations or First Peoples Indigenous people of Australia	These terms have become more common in recent years, with 'Indigenous' as the adjective.
Aboriginal an Aboriginal person is someone who is of Aboriginal descent, identifies as being Aboriginal and is accepted as such by the Aboriginal community with which they originally identified	One of the reasons that 'First Nations' and allied forms have become more common is that the term 'Aboriginal' was sometimes used disrespectfully, and still is in some circles.
Aboriginal and Torres Strait Islander peoples the Australian Indigenous population includes Aboriginal People, Torres Strait Islander People, and people who have both Aboriginal and Torres Strait Islander heritage. The term 'Aboriginal and Torres Strait Islander' encompasses all three	While this is still used in official circles and is in the name or title of many organisations and documents, it is tending to be replaced by 'First Australians' and similar terms, especially in everyday use. This is partly because the abbreviation 'ATSI' is considered disrespectful by Indigenous people, who regard it as lazy not to use a full title. The abbreviation should not be used to refer to people.

Concept maps for Units 1 & 2

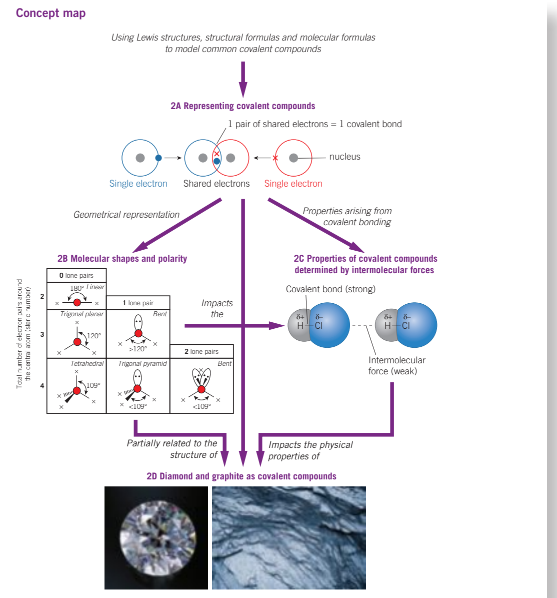
Pages xiii–xvi display the concept maps for topics in Chapters 1–16. Access the digital version of this concept map in the ITB to zoom in on the details and click on hyperlinks to explore the interconnections of the topics.

Chapter 1 Elements and the periodic table



See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

Chapter 2 Covalent substances



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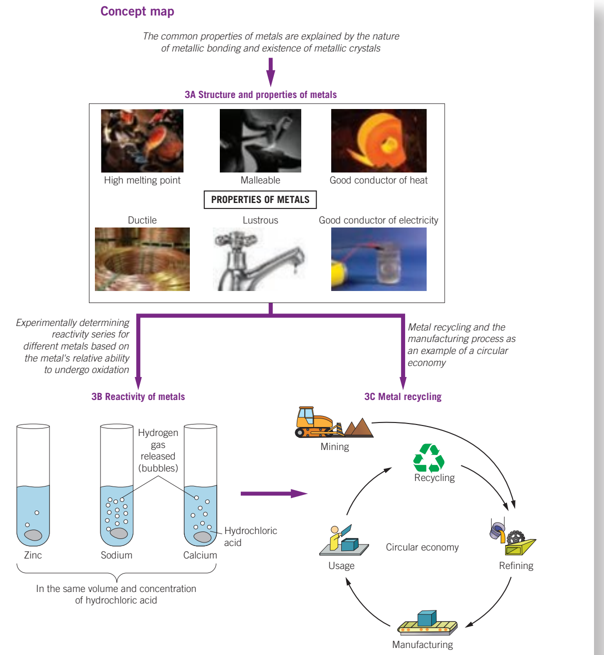
Chapter 3 Metals



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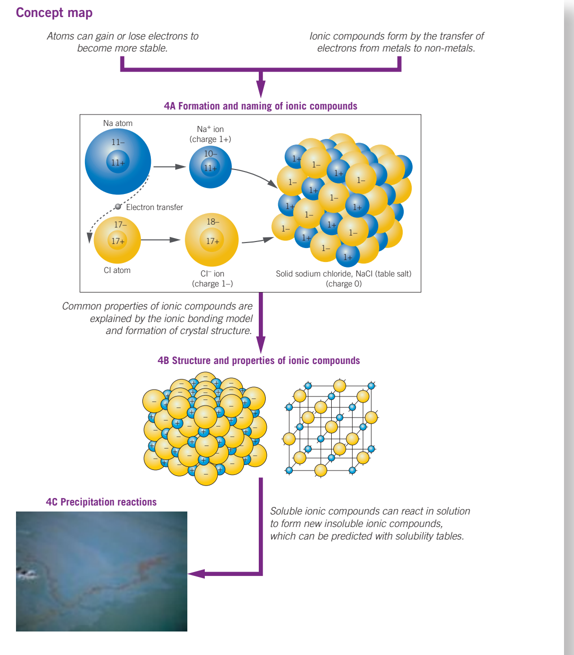
Chapter 4 Ionic compounds



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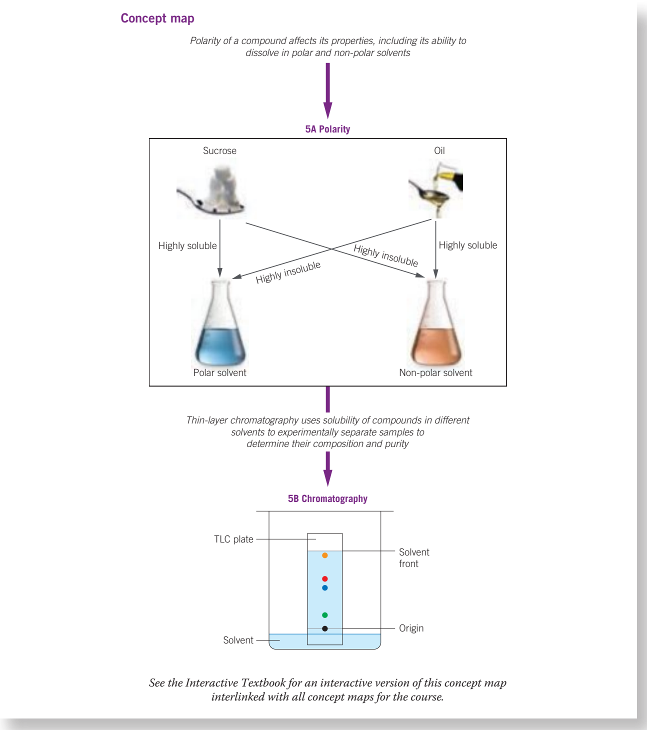
Chapter 5 Separating and identifying compounds in mixtures



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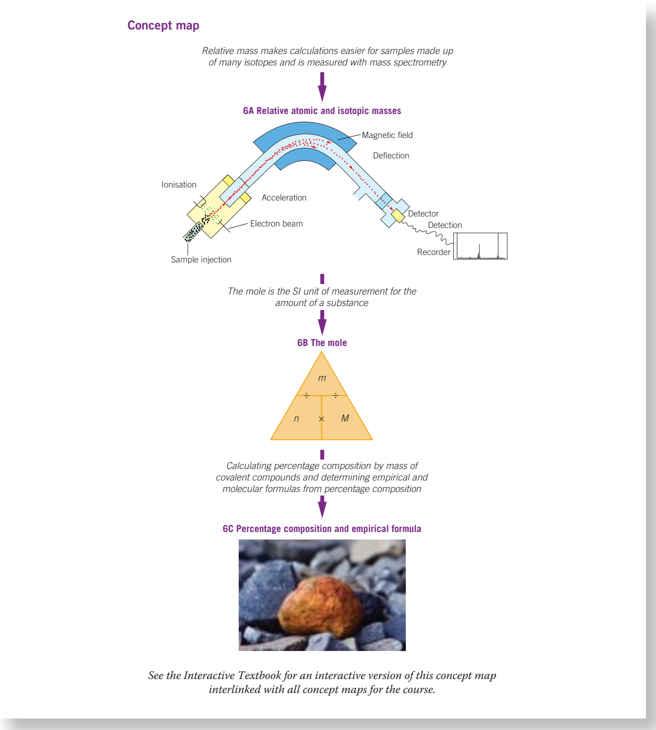
Chapter 6 Quantifying atoms and compounds



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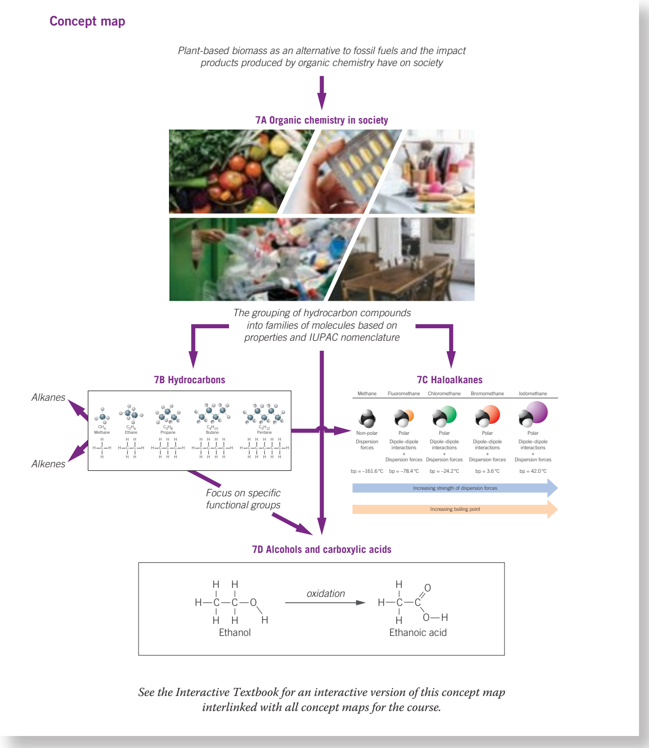
Chapter 7 Organic compounds



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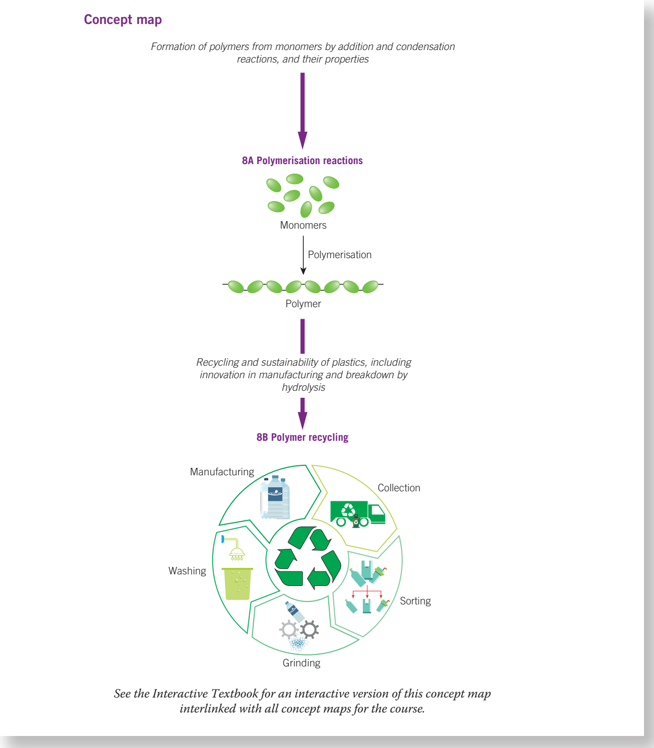
Chapter 8 Polymers



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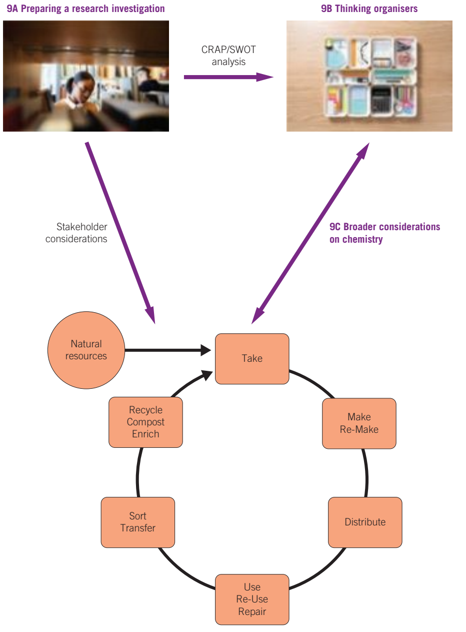
Chapter 9 Research investigations

Concept map



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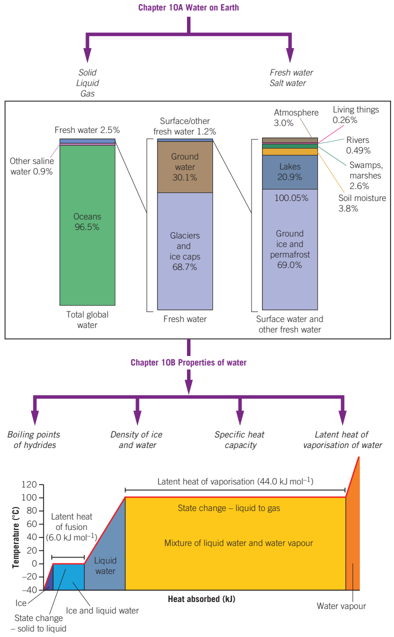
Chapter 10 Water and properties of water

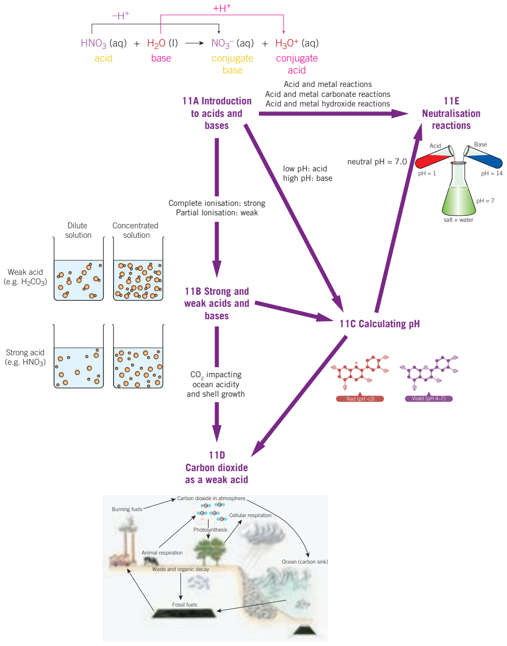
Concept map



See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

Chapter 11 Acids and bases

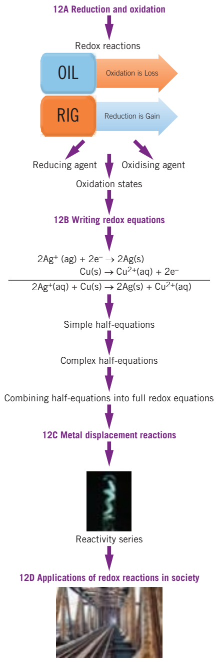
Concept map



See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

Chapter 12 Redox reactions

Concept map



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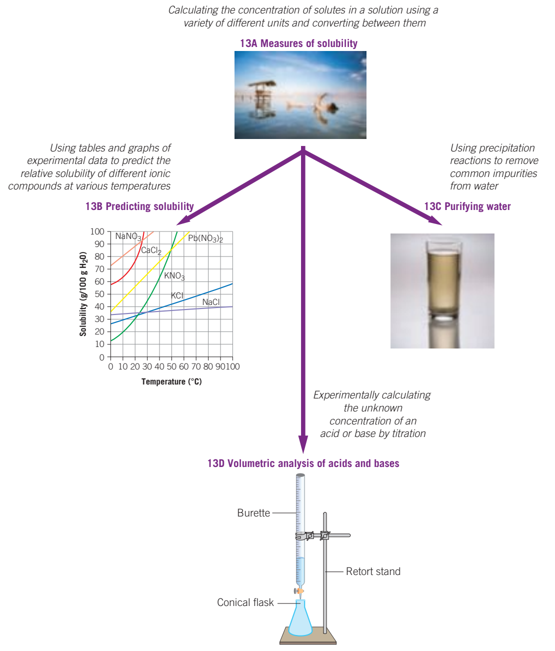
Chapter 13 Measuring solubility and concentration

Concept map



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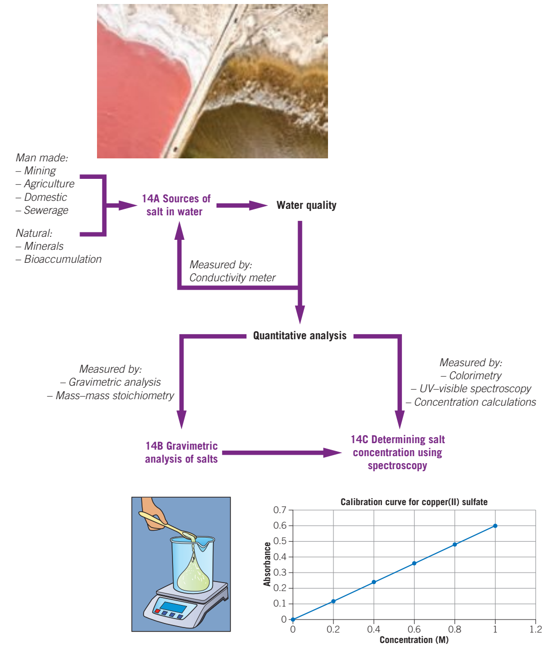
Chapter 14 Analysis of salts in water

Concept map



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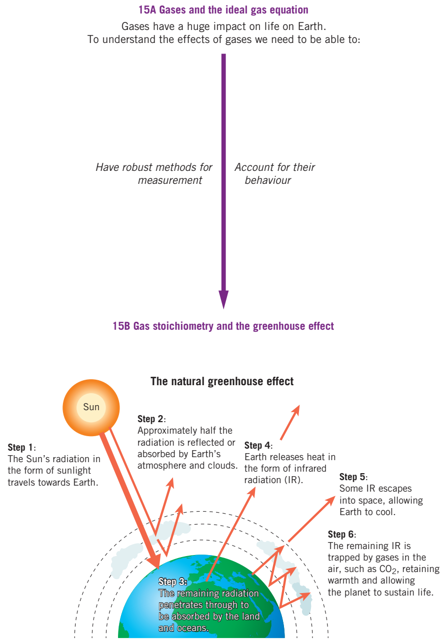
Chapter 15 Gases

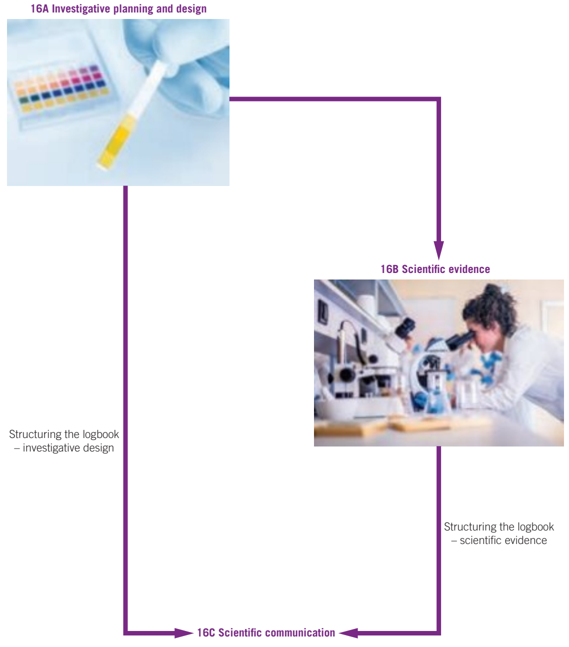
Concept map



See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

Chapter 16 Practical investigations

Concept map



See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

Acknowledgements

The publisher and authors would like to thank Dr Sarah Mann, Janice Stubbings, Dr Brodie Reid and Dr Sul Hwa Yu for their reviews and contributions to this resource.

UNIT 1 HOW CAN THE DIVERSITY OF MATERIALS BE EXPLAINED?

CHAPTER 1 ELEMENTS AND THE PERIODIC TABLE

Introduction

You will already know from previous years of science that atoms are the building blocks of everything in our universe. As you will revisit in this chapter, atoms are composed of smaller particles called protons, neutrons and electrons, and you will recall the placement of these within atoms. This knowledge will then be extended to discuss how different numbers of these smaller particles in atoms may result in the formation of either isotopes or ions.

INTRODUCTION VIDEO
ELEMENTS AND THE PERIODIC TABLE

For elements, you will learn how to represent the numbers of these subatomic particles in atomic notation as well as their electron configurations in both shell and subshell notation. These electron configurations will then be used to look at how the periodic table is constructed and how scientists use this as a 'cheat sheet' for determining the properties of different elements and trends across different areas of the periodic table. The relationship between the properties of some critical elements and how they are recycled will also be investigated.

Curriculum

Area of Study 1 Outcome 1
Elements and the periodic table

Study Design:	Learning intentions – at the end of this chapter I will be able to:	
• The definitions of elements, isotopes and ions, including appropriate notation: atomic number; mass number; and number of protons, neutrons and electrons	**1A**	**Introduction to the elements**
	1A.1	Recall what an element, isotope and ion are and how they are different
	1A.2	Be able to write correct atomic notation for different elements, isotopes and ions
	1A.3	Recall the difference between the atomic number and mass number of an element
	1A.4	Determine the atomic number and/or mass number by looking at the atomic notation of an element
	1A.5	Recall what an electron, proton and neutron are and the differences between them
	1A.6	Determine the number of protons, neutrons and/or electrons for an element, isotope or ion from atomic notation

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Study Design:	Learning intentions – at the end of this chapter I will be able to:
- 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	**1B Electron configurations** **1B.1** Recall Bohr's model for writing electron configurations **1B.2** Recall the difference between shells, subshells and orbitals **1B.3** Recall the types of subshells and number of orbitals in each **1B.4** Understand the limitations of Bohr's model for electron configurations **1B.5** Use the position of an element on the periodic table to determine its electron configuration **1B.6** Write the electron configuration for both atoms and ions
- 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	**1C The periodic table** **1C.1** Recall the different areas and overall structure of the periodic table (blocks, groups and periods) **1C.2** Understand and be able to calculate the core charge of an atom or ion **1C.3** Explain how core charge and the number of electron shells relate to the attraction of the valence electrons to the nucleus **1C.4** Explain the difference in atomic radius when comparing elements down a group and across the same period **1C.5** Recall the definitions of electronegativity, first ionisation energy, reactivity, and metallic and non-metallic character **1C.6** Use the core charge and atomic radius of an element to explain trends in electronegativity, first ionisation energy, reactivity, and metallic and non-metallic character both down a group and across a period
- Critical elements (for example, helium, phosphorus, rare-earth elements and post-transition metals and metalloids) and the importance of recycling processes for element recovery	**1D Critical elements and recycling processes** **1D.1** Explain examples of critical elements, including helium, phosphorus, rare-earth elements, post-transition metals and metalloids, and why they are classified as such **1D.2** Recall examples of the mining, extraction and processing of critical elements **1D.3** Explain the different uses of critical elements **1D.4** Outline examples of current recycling processes for recovering critical elements **1D.5** Explain the importance of element recovery with reference to environmental, economic and/or social implications

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Glossary

- Atom
- Atomic number
- Atomic radius
- Aufbau's principle
- Bohr model
- Boiling point
- Chemical property
- Circular economy
- Compound
- Core charge
- Critical element
- Cryogenics
- Diatomic
- Electron
- Electronegativity
- Electrostatic force of attraction
- Element
- Emission spectrum
- Eutrophication
- Excited state
- First ionisation energy
- Ground state
- Hund's rule
- Ion
- Ionisation energy
- Isotope
- Mass number
- Melting point
- Metallic character
- Monoatomic
- Neutron
- Orbital
- Ore
- Physical property
- Proton
- Relative atomic mass
- Slag
- Smelting
- Subshell
- Subatomic particle
- Valence electron
- Valence shell

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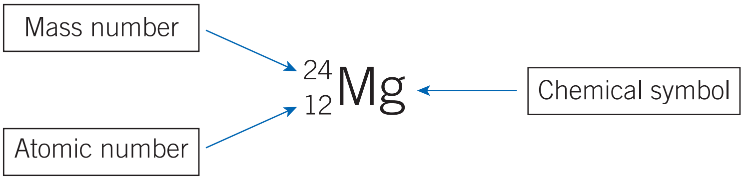
Concept map

Elements, isotopes, ions, atomic number, mass number and number of protons, neutrons and electrons

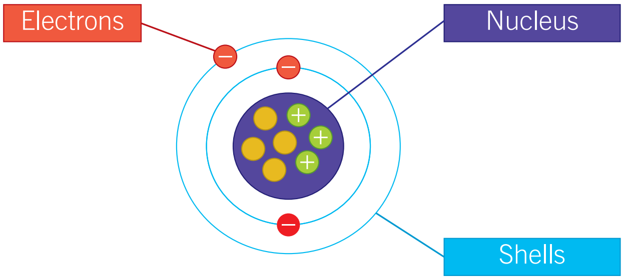
1A Introduction to the elements



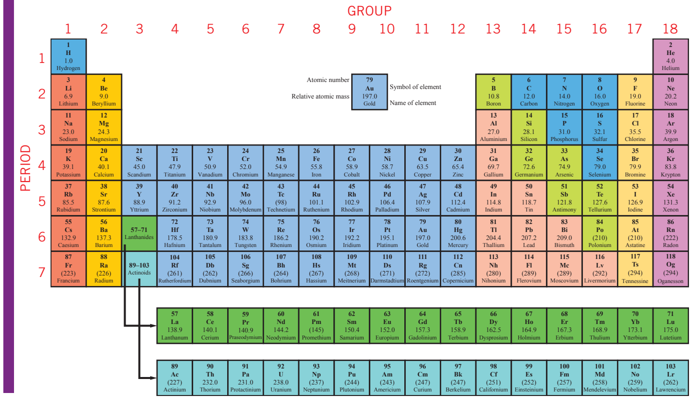
Shell and subshell electron configuration

Using the periodic table as an organisational tool

1B Electron configurations

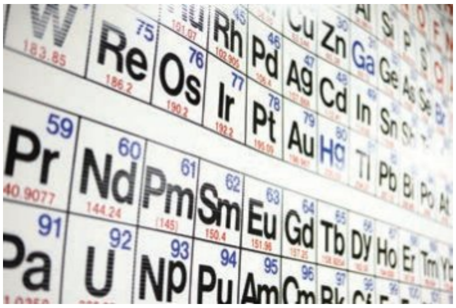


1C The periodic table



Critical elements and the importance of recovery through recycling

1D Critical elements and recycling processes



See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

1A Introduction to the elements

Study Design:

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

Glossary:

Atom
Atomic number
Chemical property
Compound
Diatomic
Electron
Element
Ion
Isotope
Mass number
Monatomic
Neutron
Physical property
Proton
Relative atomic mass
Subatomic particle

ENGAGE

The discovery of the atom

It was 1802 when John Dalton, an English chemist and physicist, presented his first ideas of the atomic theory of matter. In this theory, he proposed the following:

- All matter is made up of tiny spherical particles, which are indivisible.
- Atoms cannot be created or destroyed.
- Elements contain only one type of atom, whereas compounds contain different types of atoms in fixed ratios.

Dalton also developed symbols to represent elements and compounds known at the time.

Figure 1A–1 John Dalton (1766–1844) concluded that matter is composed of atoms.

ELEMENTS

	Wt		Wt
Hydrogen.	1	Strontian	46
Azote	5	Barytes	68
Carbon	5.4	Iron	50
Oxygen	7	Zinc	56
Phosphorus	9	Copper	56
Sulphur	13	Lead	90
Magnesia	20	Silver	190
Lime	24	Gold	190
Soda	28	Platina	190
Potash	42	Mercury	167

Figure 1A–2 Symbols of elements and compounds known to exist at the time

Since Dalton's work in the early nineteenth century, many other scientists have contributed their own ideas through discovery, which have continued to shape the structure of the atom and understanding of elements that we know today.

In 2016, four new elements were named and given a permanent place in the periodic table:

- nihonium (113)
- moscovium (115)
- tennessine (117)
- oganesson (118).

Since 2017, teams of scientists in Japan and in Russia have been trying to create element 119.

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EXPLAIN

Key chemical terms

Before beginning to explore key chemical ideas in greater depth, there are some key terms that you need to be familiar with.

VIDEO 1A–1
ATOMS, IONS AND ISOTOPES

Atoms

Atoms are the building blocks of everything (Figure 1A–3). Your own DNA, cells, the food you eat, fragrances you smell, materials you use every day and the fuel used for transportation are all made of atoms.

Subatomic particles

Subatomic particles determine the characteristics of each type of atom. Figure 1A–3 and Table 1A–1 indicate the different subatomic particles within atoms, the charge they carry and their location.

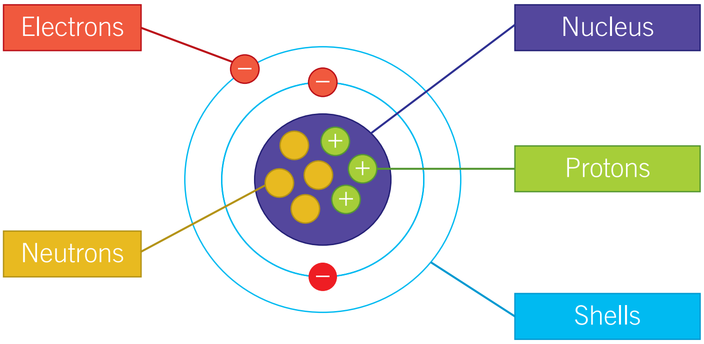


Figure 1A–3 A generalised representation of the structure of an atom, containing the location of the different subatomic particles

Table 1A–1 Subatomic particles and their location within an atom

Subatomic particle	Charge	Location	Relative mass (amu)
Proton	Positive	Nucleus	1
Neutron	Neutral		1
Electron	Negative	Shells	0.0005

You will notice from the table that protons and neutrons have similar relative masses. The actual values of their mass are:

$$\text{Proton} = 1.673 \times 10^{-24} \text{ grams}$$

$$\text{Neutron} = 1.675 \times 10^{-24} \text{ grams}$$

As these are such small values, we use the relative mass – in atomic mass units (amu) – of 1 to represent these. You will also notice that an electron is significantly smaller than each of these, in fact $\frac{1}{1840}$ th the size.

Elements

Elements are made of only one type of atom and therefore are referred to as pure substances. Each element is made of its own unique atoms determined by the number of protons in its nucleus. Depending on the material that an element's atoms form, their **chemical** and **physical properties** are very different.

Atom
the smallest piece of an element that retains the properties of that element

Subatomic particle
a particle that is present within an atom; includes protons, neutrons and electrons

Proton
a positively charged subatomic particle present within the nucleus of an atom

Neutron
an uncharged subatomic particle present within the nucleus of an atom

Electron
a negatively charged subatomic particle that moves around the nucleus of an atom

Element
a pure substance made of only one type of atom

Chemical property
the behaviour of an element or substance when it reacts with another element or substance

Physical property
the features of an element or substance that can be measured without altering the chemical composition of that substance

1D CRITICAL ELEMENTS AND RECYCLING PROCESSES LINK

2A REPRESENTING COVALENT COMPOUNDS LINK

3A STRUCTURE AND PROPERTIES OF METALS LINK

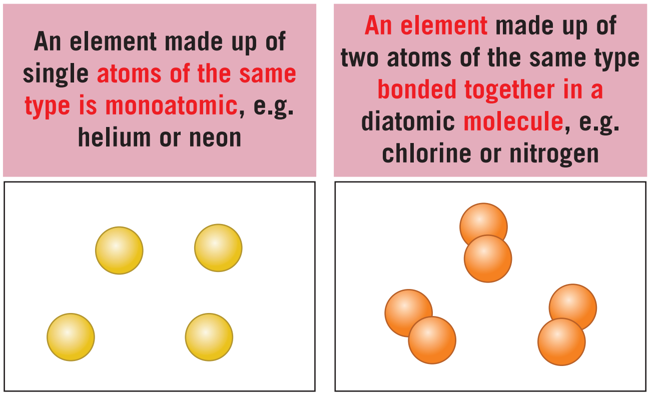
Monoatomic an element that consists of only one atom; the prefix *mono* means 'one'

Diatomic a element that consists of two atoms; the prefix *di* means 'two'

Compound a substance formed from two or more different types of atoms in a fixed ratio

For example:

- Non-metallic elements like helium and neon are **monoatomic**, whereas other non-metallic elements, like chlorine and nitrogen, form molecules that are **diatomic** (Figure 1A–4).
- Other non-metallic elements, like carbon, can form giant networks or layers of atoms, creating substances such as diamond or graphite.
- Metallic elements, like sodium, form strong organised lattices made up of many cations, which form as a result of atoms losing electrons.



An element made up of single atoms of the same type is monoatomic, e.g. helium or neon

An element made up of two atoms of the same type bonded together in a diatomic molecule, e.g. chlorine or nitrogen

Figure 1A–4 Monoatomic (left) and diatomic (right) elements are made of only a single type of atom.

You will explore more of these properties across different chapters in Unit 1.

Compounds

Compounds are formed when two or more different types of atoms combine in fixed ratios to form a new substance (Figure 1A–5).

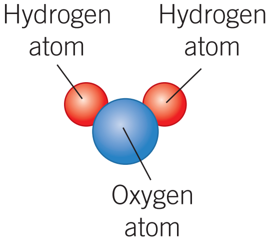


Figure 1A–5 A water molecule is a compound that is formed from two different elements in a fixed ratio. Two hydrogen atoms are covalently bonded to one oxygen atom.

Check-in questions – Set 1

1 Define the following terms: monatomic, diatomic, compound.

2 Using a fully labelled diagram, identify the charge and location of the different subatomic particles.

Atomic notation

There are currently 118 discovered elements, each with a unique name and chemical symbol. The first letter of each chemical symbol is always capitalised, and any subsequent letters, if present, are always lowercase (Table 1A–2).

Table 1A–2 Examples of different elements and their chemical symbols

Element	Chemical symbol
Carbon	C
Hydrogen	H
Sodium	Na
Mercury	Hg

One or two letters where the first letter is capitalised and the subsequent letter is always lowercase

Scientists have a uniform set of rules they follow for communicating their understanding of the representation of different elements. This includes being able to determine the:

- chemical symbol
- atomic number
- mass number.

An example of this is shown in Figure 1A–6.

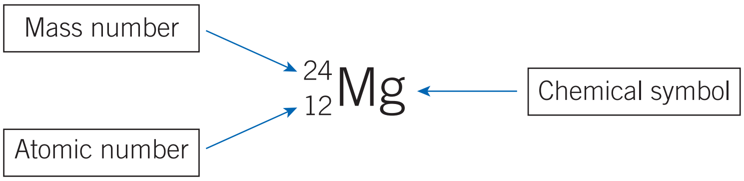


Figure 1A–6 The atomic notation representing the element, magnesium

Atomic number

The **atomic number** represents the number of protons present in the nucleus of an atom. Each element is unique and based on its number of protons (Figure 1A–7).

Atomic number the number of protons present in the nucleus of an atom

The atomic number is usually written as a *subscript* before the chemical symbol in atomic notation. As shown in Figure 1A–6 above, the atomic number of magnesium is 12. Therefore, it has 12 protons.

Similarly, the atomic number of lithium is 3, and its atomic notation is most commonly represented as $^{7}_{3}$Li. Therefore, we know that it has three protons, as represented in Figure 1A–7.

NOTE

On a *periodic table*, the atomic number is usually *above* each chemical symbol and is the smaller of the two numbers, as it does not include any neutrons.

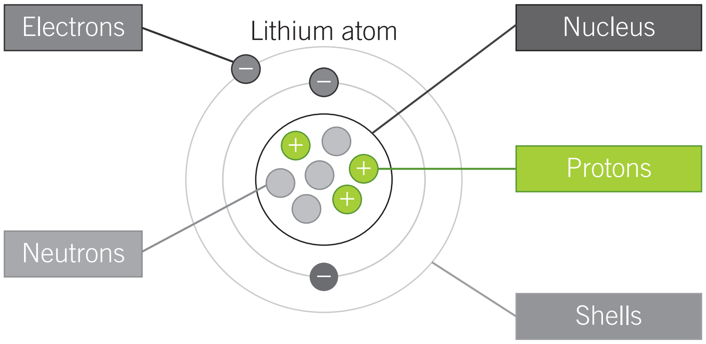


Figure 1A–7 Model of a lithium atom, highlighting the location of its three protons, where the number of these determines the *atomic number* and hence the type of element

Mass number

The **mass number** represents the total number of protons and neutrons present in the nucleus of an atom (Figure 1A–8). The mass number (or atomic mass) is usually written as a *superscript* before the chemical symbol.

Mass number the total number of protons and neutrons present in the nucleus of an atom

Mass number = number of protons + number of neutrons

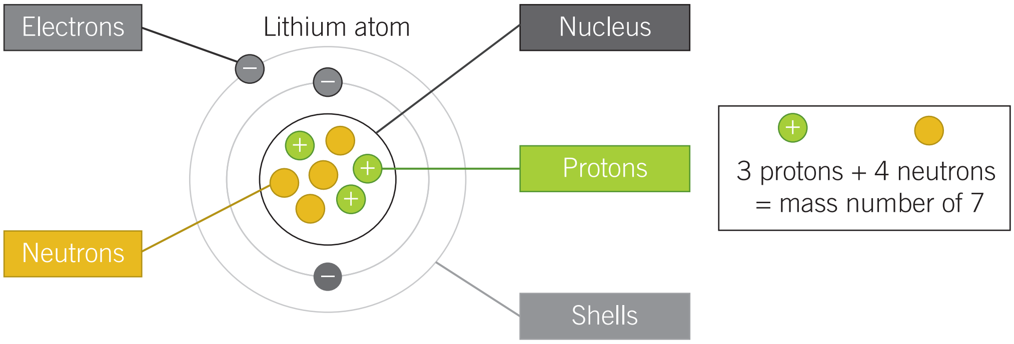


Figure 1A–8 A model of a lithium atom, highlighting the location and number of its three protons (green) and four neutrons (yellow), which add together to give the *mass number* of 7

Relative atomic mass
the average of the relative masses of all atoms for an element based on the isotopes for an element and their natural percentage abundances. Also known as relative atomic weight or standard atomic weight

Isotope
an atom of the same element with the same number of protons (same atomic number) but a different number of neutrons (different mass number)

3A STRUCTURE AND PROPERTIES OF METALS LINK

4A FORMATION AND NAMING OF IONIC COMPOUNDS LINK

You will also notice by looking at most periodic tables that the mass number for an element is not always an exact whole number. This is actually because what is shown is the **relative atomic mass**, not the mass number.

Relative atomic mass (or relative atomic weight or standard atomic weight as used on the IUPAC periodic table from 2022) is fractional, as it is the average of the relative mass of atoms, taking into account the fact there are isotopes for each element. **Isotopes** have differing numbers of neutrons, and each isotope for an element is present in different percentages, or abundances, in natural samples.

As you can see from the atomic notation of an element, the number of protons and neutrons can be determined. However, there is no reference to the number of electrons. This will be explored later in this chapter and again in Chapters 3A and 4A.

Check-in questions – Set 2

1 After locating the following elements on the periodic table, use their atomic numbers and relative atomic masses (rounded to the nearest whole number to estimate the mass number) to determine the number of all subatomic particles present in each.

Element	Atomic number	Mass number	Number of protons	Number of neutrons
Carbon (C)				
Germanium (Ge)				
Titanium (Ti)				

Isotopes

Isotopes are atoms of the same element with the same number of protons (the same atomic number) but a different number of neutrons, which gives each isotope a different mass number. A representation of this is shown for nitrogen in Figure 1A–9.

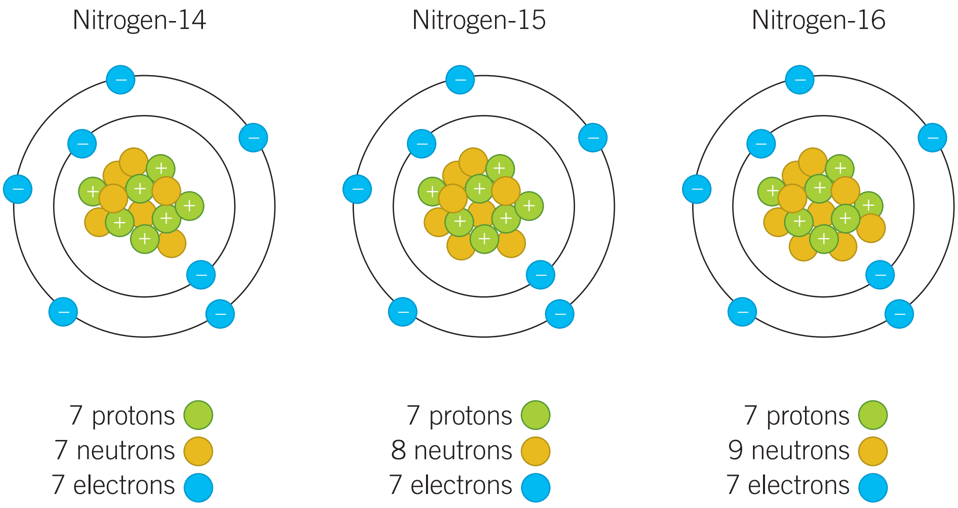


Figure 1A–9 Three of the isotopes of nitrogen, indicating the different numbers of subatomic particles within each. Note that each atom is still nitrogen, as it always has an atomic number of 7 (seven protons).

NOTE

Although not covered in VCE Chemistry Units 1&2, the stability of an isotope is determined by how likely it is to break down or decay into lighter (smaller) elements.

Number of electrons and the formation of ions

Recall that the atomic number is the number of protons in an element's atoms. The atomic number also provides information about the number of electrons in an atom of an element. In a neutral atom, the number of positive protons and negative electrons will always be equal.

Therefore, changing the number of electrons will not change the type of element. However, it will change the chemical properties of that element, often to become more stable. As you will see in Chapters 3A and 4A, changing the number of electrons results in the formation of **ions**.

LINK **3A** STRUCTURE AND PROPERTIES OF METALS

LINK **4A** FORMATION AND NAMING OF IONIC COMPOUNDS

Ion
a positively or negatively charged atom that has either lost or gained electron(s)

Ions are versions of atoms that contain their normal number of protons but have more, or fewer, electrons. This gives them an overall electric charge. The atomic notation for an ion is represented in the same way as for any other atom; however, it includes either a superscript positive or negative charge and the size of this charge following the chemical symbol. This can be seen in Figures 1A–10 and 1A–11.

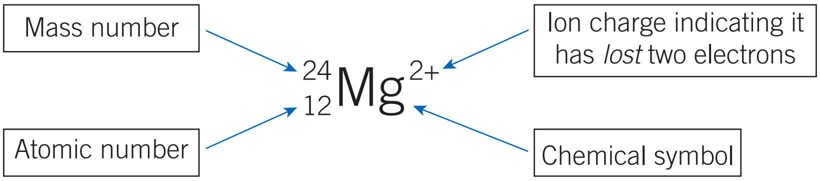


Figure 1A–10 The atomic notation representing a magnesium ion. The '2+' charge indicates that magnesium, a metal, has *lost* two electrons. Overall, it has 12 protons (atomic number) but only has 10 electrons.

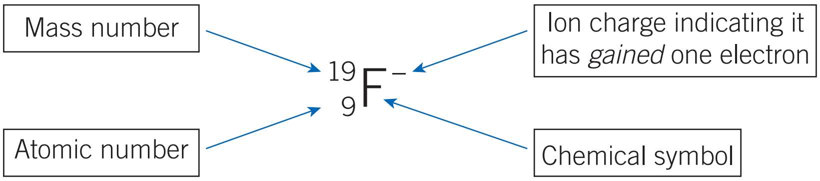


Figure 1A–11 The atomic notation representing a fluoride ion. The '–' charge indicates that fluorine, a non-metal, has *gained* one electron. Overall, it has nine protons (atomic number) but has 10 electrons.

Atoms will either gain or lose electrons based on how many electrons are present in their outer shell. This will be explored further in Chapter 1B.

LINK **1B** ELECTRON CONFIGURATIONS

Check-in questions – Set 3

1 How many protons, neutrons and electrons does each of the atoms or ions below contain?
   - a $^{19}_{9}\text{F}$
   - b $^{65}_{30}\text{Zn}^{2+}$
   - c $^{35}_{17}\text{Cl}^{-}$

2 Define the following key terms.
   - a mass number
   - b atomic number

3 Do isotopes of an element have a different atomic number or mass number?

WORKSHEET 1A–1 ATOMS AND SUBATOMIC PARTICLES

VIDEO 1A–2 SKILLS: DETERMINING THE NUMBER OF NEUTRONS AND ELECTRONS

1A SKILLS

Determining the number of neutrons

A typical exam question will ask you to determine the number of neutrons by looking at an element on the periodic table or by inspecting the atomic notation provided in the question.

The given information tells you the atomic number and the mass number, both of which you can use to determine the number of neutrons.

Number of neutrons = mass number – atomic number

Or broken down slightly further to:

Number of neutrons = (number of protons + neutrons) – (number of protons)

If you recall the example of magnesium used in Figure 1A–10, magnesium has an atomic number of 12 and a mass number of 24. Therefore, using the equation above, you can determine the number of neutrons:

Number of neutrons = mass number – atomic number.
Number of neutrons = 24 – 12
Number of neutrons = 12

Determining the number of electrons in ions

Similar to determining the number of neutrons in an atom, you may also be asked to look at the atomic notation of an element or ion and determine the number of electrons present, or how many electrons an element has lost or gained.

Key points to remember:

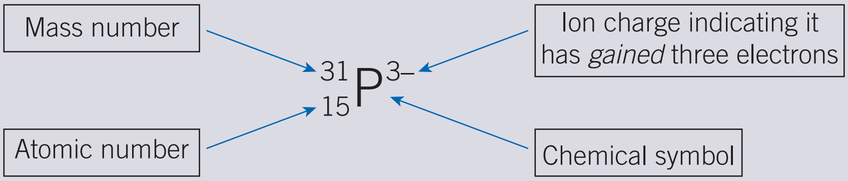
1B ELECTRON CONFIGURATIONS LINK

The following points will help you to identify the key features.

- Determine the number of electrons in a neutral atom – this is equivalent to the atomic number of an element (found on the periodic table)
- Note the charge of the ion – it is written as a superscript following the chemical symbol.
- Determine the number of electrons in the ion:
  - M^{x+}: number of electrons = atomic number – x
  - X^{y-} : number of electrons = atomic number + y

Question:

Look at the atomic notation for an ion of phosphorus (phosphide ion) below and determine the number of electrons present.



Answer:

In a neutral atom, phosphorus has 15 protons and 15 electrons. However, the phosphide ion has a charge of '3–'. This means there is a total of three more negative charges compared to positive charges. It cannot change the number of protons because this will change the element. Therefore, it must have changed its number of electrons. As there is a 3– charge, it gained three electrons, so:

Number of electrons = 15 + 3 = 18

Note that ions will have the number of electrons equivalent to a noble gas (full outer shell of electrons). You will explore this further in Chapter 1B and later in Chapter 4.

3A STRUCTURE AND PROPERTIES OF METALS LINK

4A FORMATION AND NAMING OF IONIC COMPOUNDS LINK

Section 1A questions

1 Using your knowledge from Chapter 1A, determine whether the following statements are true or false.
  a All subatomic particles have the same size and mass.
  b Atoms of a particular element are identical.
  c Different elements would be composed of different atoms.
  d Atoms of different elements combine in simple ratios to form compounds.

2 For each of the atoms below, determine the number of protons, neutrons and electrons.
  a ${}^{27}_{13}\text{Al}$
  b ${}^{40}_{18}\text{Ar}$
  c ${}^{32}_{16}\text{S}^{2-}$
  d ${}^{235}_{92}\text{U}^{3+}$

3 Discuss the difference between the atomic number and the mass number of an atom.

4 Using clear examples, explain the difference between an ion and an isotope.

5 Using a fully labelled diagram, indicate the name and location of the three subatomic particles in an atom.

6 The atomic notation for eight atoms with chemical symbols denoted A through to H are given below. Note that these are not the actual chemical symbols of the elements.
  A ${}^{38}_{88}\text{A}$
  B ${}^{12}_{6}\text{B}$
  C ${}^{56}_{26}\text{C}$
  D ${}^{35}_{17}\text{D}$
  E ${}^{14}_{6}\text{e}$
  F ${}^{4}_{2}\text{F}$
  G ${}^{37}_{17}\text{G}$
  H ${}^{39}_{19}\text{H}$
  a Which of the elements A–H are isotopes?
  b Identify which two atomic notations are written incorrectly.
  c From your answer to part **b**, explain what change is required, and why, to correct each of these.

1B Electron configurations

Study Design:

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

Glossary:

Aufbau's principle
Bohr model
Emission spectrum
Excited state
Ground state
Hund's rule
Orbital
Subshell
Valence electron
Valence shell

1A INTRODUCTION TO THE ELEMENTS LINK

ENGAGE

The discovery of the electron

Over the course of the last 3000 years, beginning with Democritus in the 5th century BCE, many scientists have contributed their own theories on the model of the atom. You will recall that there are three types of subatomic particles – the proton, neutron and electron – located within an atom, and that the electron contributes almost no mass to the overall atom. Surprisingly, then, it was the electron that was the first subatomic particle to be discovered in 1897 by the English physicist Joseph John Thomson (1856–1940) according to the following timeline:

1897 – Thomson discovered 'corpuscles' (later identified as electrons).

1904 – Thomson proposed a model of the atom containing electrons embedded in a positive sphere.

1906 – Thomson awarded Nobel Prize in Physics: 'in recognition of the great merits of his theoretical and experimental investigations on the conduction of electricity by gases.'

By studying 'rays' within a cathode ray tube, shown in Figure 1B–1, Thomson determined that the rays had a mass 1000 times smaller than a hydrogen atom, which was the lightest matter known to scientists at the time. He then claimed that these 'rays' were actually very light particles and were attracted to a positively charged metal plate, so they had to have a negative charge. At the time, Thomson called them 'corpuscles', but they were later named 'electrons'.

From this discovery, his model of the atom consisted of electrons scattered throughout a ball of positive charge, a bit like raisins in a plum pudding. Hence, the name given to the model was the 'plum pudding model'.

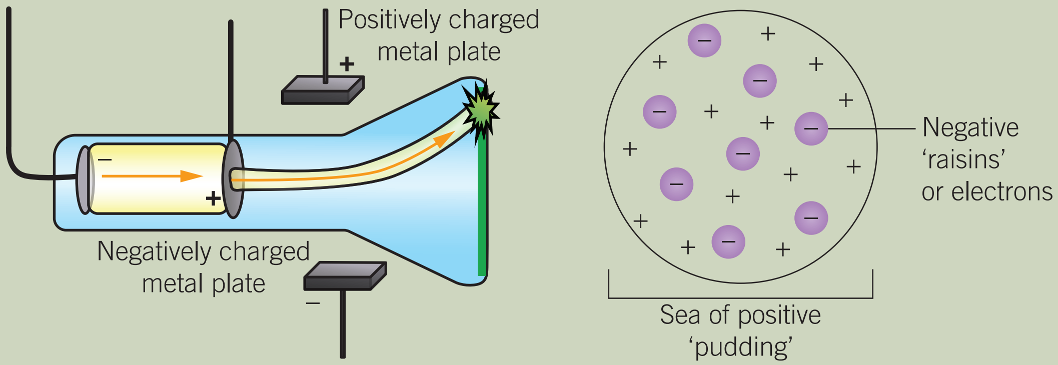


Figure 1B–1 J.J. Thomson's cathode ray tube setup (left) and his plum pudding model of the atom (right)

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EXPLAIN

Electron configurations

Since Thomson's discovery of the electron, much work by later scientists has resulted in our understanding of the position and movement of electrons orbiting the nucleus of an atom. Two of these models, and how we use them to write electron configurations, are outlined below and on the following pages.

VIDEO 1B–1
ELECTRON CONFIGURATIONS

Model 1: The Bohr model (or shell model)

Named after the Danish physicist Niels Bohr, this model of the atom consisted of electrons located in specific levels or 'shells' orbiting the nucleus (Figure 1B–2). This is called the **Bohr model** (also known as the shell model).

Bohr model also called the shell model. Shows the position of electrons in shells around the nucleus

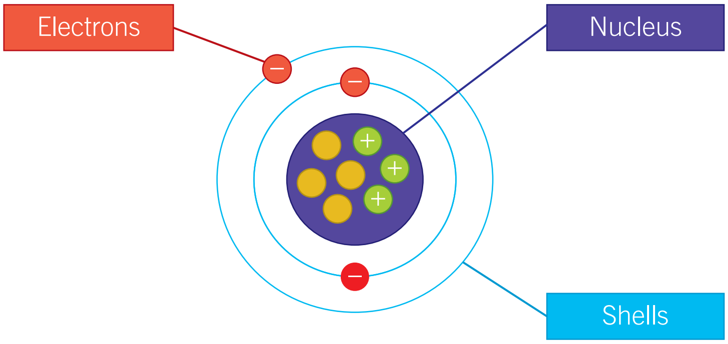

Figure 1B–2 The Bohr model of the atom, indicating the placement of electrons in shells around the nucleus

Bohr developed this model from conclusions made from the **emission spectra** of elements, as shown in Figure 1B–3.

Emission spectrum (plural: spectra) the electromagnetic radiation emitted by a source, such as atoms

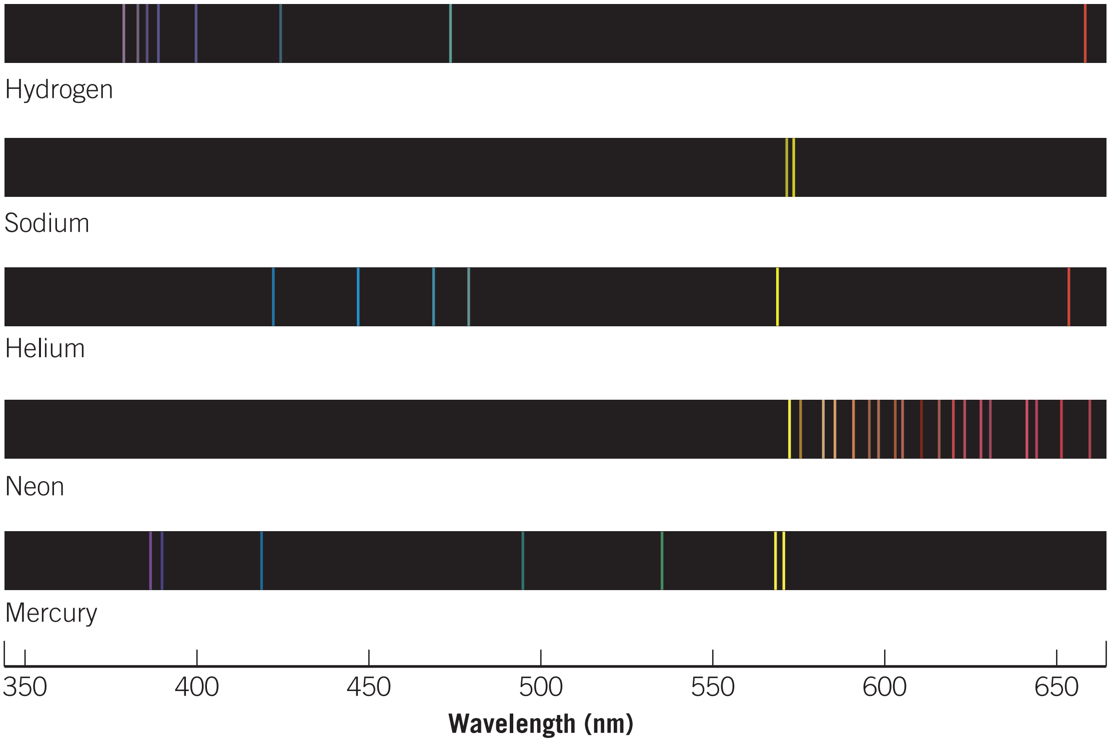

Figure 1B–3 The emission spectra of various elements. Note the unique appearance of lines at different wavelengths (colours) for each element.

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Spectra are produced when electrons from atoms become excited (**excited state**) due to the addition of energy. The electrons eventually lose this energy in the form of light and the return to their lowest possible energy level, or **ground state**. That light is then passed through a prism to separate it into its specific wavelengths (Figure 1B–4).

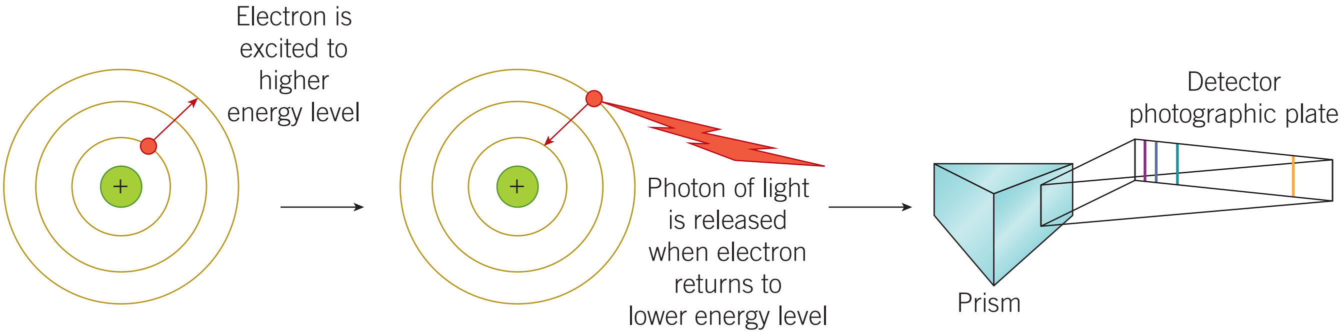


Figure 1B–4 Energy is emitted when electrons in an excited state return from their temporary higher energy level to a lower energy level, or ground state. This is then passed through a prism to produce the emission spectra.

Excited state when electron(s) in an atom have gained energy and are no longer in their lowest possible energy state closest to the nucleus

Ground state when all electrons in an atom are in their lowest possible energy state closest to the nucleus

The emission spectrum is unique for each element. The wavelength(s) of light seen on emission spectra is dependent on the location of an electron in a shell during its excited state and the position of the same electron in its shell in the ground state (Figure 1B–5).

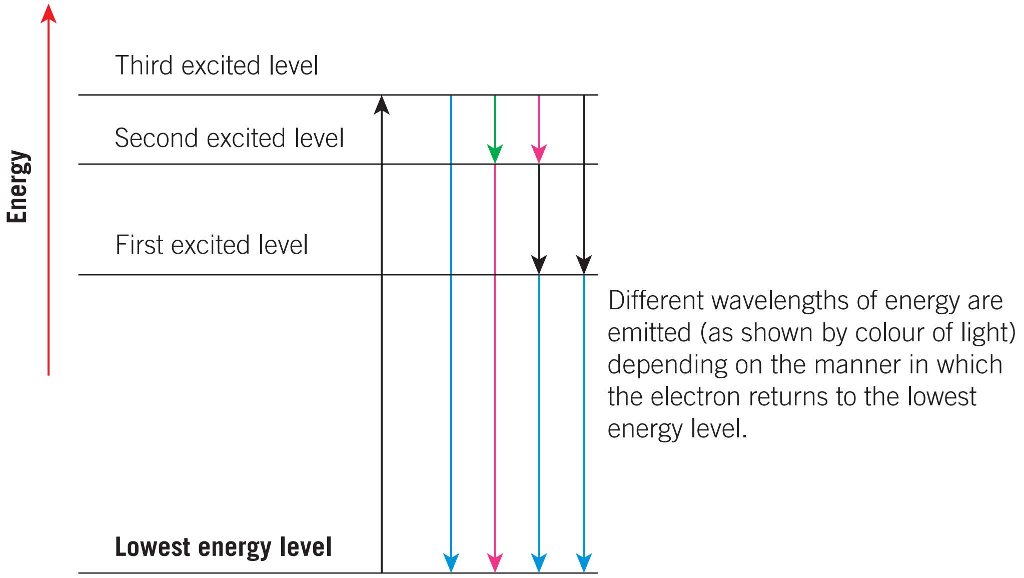


Figure 1B–5 The energy emitted when electrons in an excited state return from their temporary higher energy level to a lower energy level, or ground state. The wavelength(s), or colour(s) of light, seen on an emission spectrum is dependent on the difference in energy of the shells.

1C THE PERIODIC TABLE **LINK**

Scientists studying the ionisation energies (explained in Section 1C) of other elements continued to extend Bohr's model. It was noted that electrons in the same shell are the same distance from the nucleus and have approximately the same energy. Electrons try to get as close to the nucleus as possible, and in doing so, there are particular rules for how the shells are filled and how many electrons each can hold.

Rule 1: Each shell has a maximum number of electrons. And this follows the rule:

$2n^2$ (where n = the shell number)

(e.g. Shell 1 has $2(1)^2 = 2$

Shell 2 has $2(2)^2 = 8$)

Rule 2: Shells closer to the nucleus are occupied first as they have lower energy.

Rule 3: Shells fill in a specific order.

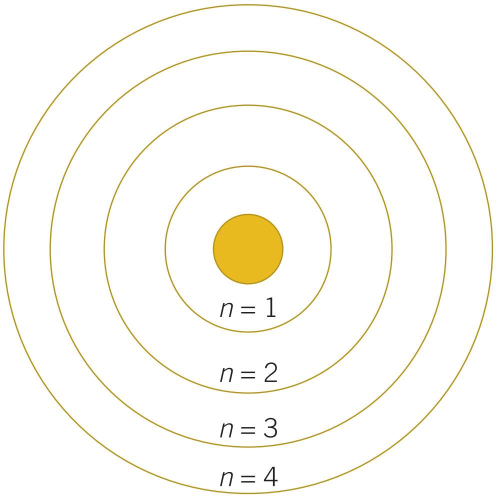


Figure 1B–6 The numbering of electron shells

The outer shell can never contain more than eight electrons, regardless of the maximum possible number for that shell. For example, according to Rule 1 above, the third shell can hold 18 electrons, but will hold no more than eight until there are at least two electrons placed in the fourth shell.

The outermost shell, furthest from the nucleus, is called the **valence shell**. The electrons in this shell are called the **valence electrons**. These are particularly important when we explore the different types of bonding in Chapters 3A and 4A. If you know the number of valence electrons, you can also predict an element's chemical properties, as you will see in Chapter 1C.

Valence shell the outermost shell of an atom

Valence electron an electron found in the outermost shell of an atom

LINK **1C** THE PERIODIC TABLE

LINK **3A** STRUCTURE AND PROPERTIES OF METALS

LINK **4A** FORMATION AND NAMING OF IONIC COMPOUNDS

Writing electron configurations

When we write electron configurations according to the Bohr model, we simply write it in the following way:

Number of electrons in first shell, number of electrons in second shell, and so on.

Worked example 1B–1 shows a more detailed step-by-step process of how to do this.

Worked example 1B–1: Writing electron configurations

Using the Bohr model, write the electron configuration for oxygen.

Solution

Step 1	Look for the element on the periodic table and note down its atomic number.	Oxygen = 8
Step 2	Look for the period (horizontal row) that oxygen appears in on the periodic table.	Period (row) = 2
Step 3	Apply the rules of shell filling to work out the maximum number of electrons that will be located in the first shell and then the second shell.	First shell = $2(1)^2 = 2$ electrons Second shell = $2(2)^2 = 8$ electrons
Step 4	Determine how many electrons are required in the second shell by subtracting the electrons in previously completed shells from the maximum number of electrons that shell can hold.	8 – 2 = 6
Step 5	Write the electron configuration in the following form: number in first shell, *comma*, number in second shell.	2,6

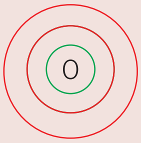
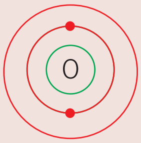
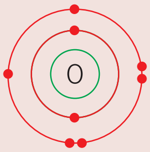
Drawing electron configurations

When drawing electron configurations, the same rules apply as for writing electron configurations. However, a small circle is drawn in the centre, representing the nucleus, with the chemical symbol of the element written in the centre. Sometimes, you will find that the number of protons and neutrons are also written inside this circle. This may be instructed by your teacher. Surrounding this are then circles, representing different electron shells, with the first shell being the closest to the nucleus.

Worked example 1B–2 shows a detailed step-by-step process of how to do this.

Worked example 1B–2: Drawing electron configurations

Using the Bohr model, draw the electron configuration for oxygen.

Step 1	Look for the element on the periodic table and note down its atomic number.	Oxygen = 8
Step 2	Look for the period (horizontal row) that oxygen appears in on the periodic table.	Period (row) = 2
Step 3	Apply the rules of shell filling to work out the maximum number of electrons that will be located in the first shell and then the second shell.	First shell = $2\ (1)^2 = 2$ electrons Second shell = $2\ (2)^2 = 8$ electrons
Step 4	Determine how many electrons are required in the second shell by subtracting the electrons in previously completed shells from the maximum number of electrons that shell can hold.	$8 - 2 = 6$
Step 5	Draw a circle in the middle representing the nucleus and include the uppercase letter ‘O’ (chemical symbol) to indicate it is ‘oxygen’.	
Step 6	Draw two circles surrounding the centre circle in step 5.	
Step 7	Draw two electrons as coloured-in dots or crosses on the first shell.	
Step 8	Draw six electrons as coloured-in dots or crosses on the second shell.	

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Writing the electron configurations for ions

When writing electron configurations for ions, they should have the same electron configuration as that of noble gases.

Worked example 1B–3: Writing electron configurations of ions

Determine the electron configuration for an oxide ion (O^{2-}).

Solution

Step 1	Look for the element on the periodic table and note down its atomic number.	Oxygen = 8
Step 2	Look for the period (horizontal row) that oxygen appears in on the periodic table.	Period (row) = 2
Step 3	Apply the rules of shell filling to work out the maximum number of electrons that will be located in the first shell and then the second shell.	First shell = $2(1)^2 = 2$ electrons Second shell = $2(2)^2 = 8$ electrons
Step 4	Determine how many electrons are required in the second shell by subtracting the electrons in previously completed shells from the maximum number of electrons which that shell can hold.	$8 - 2 = 6$
Step 5	Write the electron configuration in the following form: number in first shell, *comma*, number in second shell.	2,6
Step 6	Since the charge on the oxide ion is 2–, an atom of oxygen has gained two negative charges; that is, it has gained two electrons.	Gain electrons 2 needed
Step 7	Write the electron configuration for the oxide ion.	2,8

Check-in questions – Set 1

1 Use the shell model of the atom to write the electron configuration for the following elements.
- **a** helium
- **b** nitrogen
- **c** calcium
- **d** zinc

2 Draw the electron configurations for the following elements according to the Bohr model.
- **a** lithium
- **b** chlorine

3 How many valence electrons are present in each of the following atoms or ions?
- **a** an atom of aluminium
- **b** an atom of sulfur
- **c** a lithium ion (Li^+)
- **d** a bromide ion (Br^-)

WORKSHEET 1B–1
WRITING AND DRAWING BOHR MODEL ELECTRON CONFIGURATIONS

Model 2: The Schrödinger model (or quantum mechanics model)

At this point you may be asking yourself, ‘Why is there another model of the atom?’ The short answer is that the Bohr model was unable to explain complex emission spectra.

Named after the Austrian–Irish physicist Erwin Schrödinger, this mathematical description treated electrons as a wave with wave-like properties, rather than as particles, which was observed in the Bohr model. He noted that electrons were found in regions of space called orbitals (Figure 1B–7).

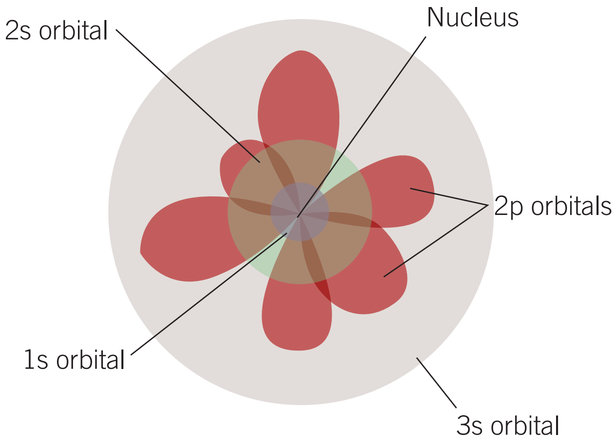


Figure 1B–7 This diagram represents the subshell model (or quantum mechanical model) of the atom, showing electrons located in regions of space called orbitals around the nucleus.

This model has a number of key points.

- Each ‘shell’ is composed of **subshells**.
- Each ‘subshell’ is composed of **orbitals**.
- Each ‘orbital’ can hold a maximum of two electrons.
- Like the Bohr model, electrons occupy the lowest possible energy levels (or subshells) first.

Subshell
a region within a shell composed of certain numbers of orbitals. Subshells are labelled as s, p, d and f

Orbital
a three-dimensional description of the most likely location for an electron within a region of an atom. Each can hold zero, one or two electrons

Determining electron configurations and using the subshell notation

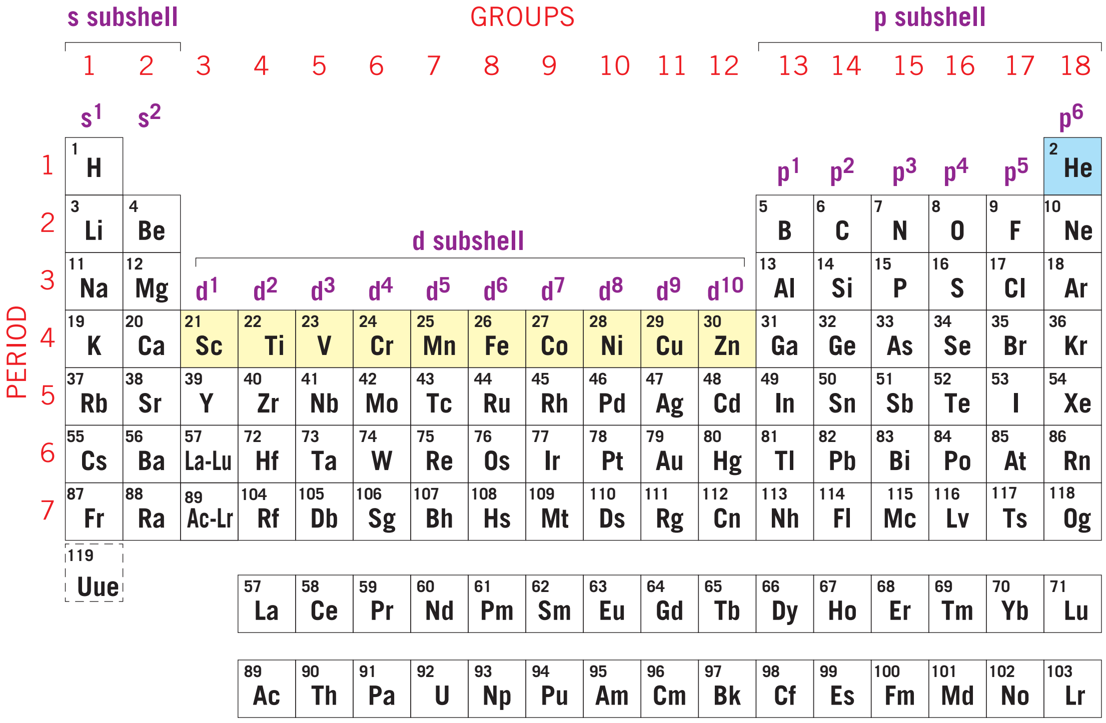
To determine the electron configurations using the subshell model, you need to first know how to read the periodic table and the location of different subshells. The advantage is that the periodic table is like a good recipe; it’s always there for you to refer to, but also when you know it well you can remember the steps of the recipe from recalled knowledge.

Figure 1B–8 shows an annotated version of the periodic table, indicating the location of groups and periods, different subshells and *some* highlighted exceptions.

Determining electron configurations is easiest when working from left to right across a period (row), beginning with the first period and then moving down to the next period.

You also need to recall the rules for this model.

- Each shell is composed of subshells.
- Each subshell is composed of orbitals.
- Each orbital can hold a maximum of two electrons.
- Like the Bohr model, electrons occupy the lowest possible energy levels (or subshells) first.



🟦 Exception: He ends in $1s^2$ not p^6 as it only has two electrons and has a full valence shell.

🟨 This is the 3d subshell not 4d because $4s^2$ fills before 3d.

Figure 1B–8 Periodic table showing the location of different subshells and *some* highlighted exceptions

The types of subshells present in each shell, the number of orbitals present in each subshell and the number of electrons found in each subshell as a result follows a pattern as represented in Table 1B–1.

Table 1B–1 Relationship between shells, subshells, orbitals and the total number of electrons

Shell	Type of subshell(s) present, including order in which they are filled	Number of orbitals in subshell	Maximum number of electrons per subshell	Maximum number of electrons per shell
1	s	1	2	2
2	s p	1 3	2 6	8
3	s p d	1 3 5	2 6 10	18
4	s p d f	1 3 5 7	2 6 10 14	32

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Therefore, using the example of carbon, you know from the Bohr model that it has the electron configuration 2,4. It has a total of six electrons with two in the first shell and four in the valence shell. In the subshell model, the electron configuration is written as:

$$1s^2 2s^2 2p^2$$

Figure 1B–9 shows what each aspect of part of this electron configuration represents.

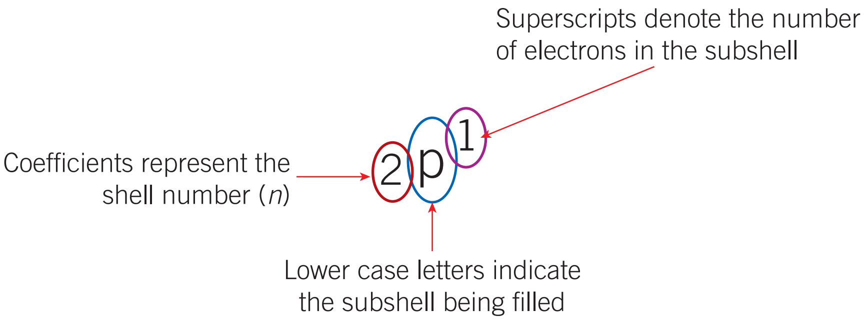
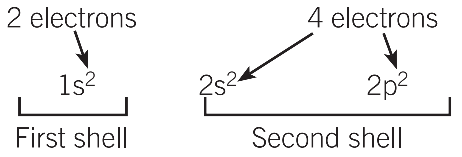


Figure 1B–9 An example of the subshell notation used for writing the electron configuration for Schrödinger's quantum mechanical model

Table 1B–2 helps to explain this example of carbon in more detail.

Table 1B–2 How to write the subshell electron configuration of carbon

Step	Part of electron configuration	Explanation
1	$1s^2$	1 = the shell number. s = the only type of subshell present in the first shell ^{2} = the number of electrons present in the first shell
2	$2s^2$	2 = the shell number. s = the first type of subshell present in the second shell ^{2} = the number of electrons present in that s subshell
3	$2p^2$	2 = the shell number. p = the second type of subshell present in the second shell ^{2} = the number of electrons present in that p subshell
Summary	$1s^2 2s^2 2p^2$	Overall, you can see the following: • Carbon has electrons located in two shells. • The first shell contains only an s subshell. • The second shell contains both an s and p subshell. • There are three subshells occupied by electrons in carbon. • Carbon has a total of six electrons, with two in the first shell and four in the second shell. 

WORKSHEET
1B–2 WRITING SUBSHELL ELECTRON CONFIGURATIONS

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Check-in questions – Set 2

1 Identify the differences between shells, subshells and orbitals.

2 State how many orbitals are present in each of the following subshells.

a s **b** p **c** d

3 Using the rules learned and the periodic table provided in Figure 1B–8 as a guide, write the subshell electron configurations for the following elements.

a He **b** B **c** Mg **d** Ca

Differences in energy levels

Using the periodic table as the guide for determining the subshell electron configurations, you will notice that the first period of transition metals, shown in period 4 in Figure 1B–8, contains electrons that belong in the third shell, specifically the 3d subshell. Recall the Bohr model of electron configurations, where the rule $2n^2$ applied, where n equals the shell number. Using this rule, the third shell can hold up to 18 electrons.

However, you should also recall the rule whereby when the third shell contains eight electrons, two electrons must first be placed in the fourth shell, before the remaining 10 electrons can be placed in the third shell.

Therefore, the last subshell filled for these transition metals is the 3d subshell. This can also be seen in Figure 1B–10, which demonstrates the order of filling subshells. This is also known as **Aufbau's principle**.

Aufbau's principle describes the order of filling subshells with electrons from their lowest possible energy levels first

Hund's rule describes the situation where every orbital in a subshell contains only one electron before any orbital is fully occupied with two electrons

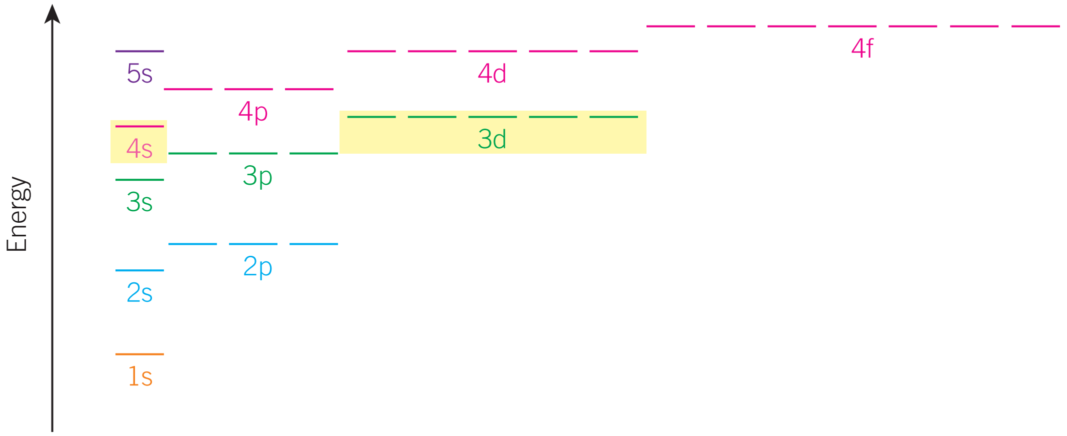

Figure 1B–10 Aufbau's principle – the order of filling subshells from lowest energy first. The dashed lines above each subshell indicate the number of orbitals each subshell contains.

Looking at Figure 1B–10, you will notice that the 3d and 4s subshells are very similar in energy level and that the 4s subshell is actually lower, so it will be filled before the 3d subshell. When you look at the periodic table in Figure 1B–8 and work left to right across each period, as per the steps for determining electron configurations, this is also clear.

Of course, there are exceptions to this in chemistry. The two main exceptions you will encounter include chromium (Cr) and copper (Cu). The electron configurations for these two elements are:

Cr: $1s^22s^22p^63s^23p^64s^13d^5$ Cu: $1s^22s^22p^63s^23p^64s^13d^{10}$

NOTE

These subshell electron configurations should still be written where all subshells within the same shell are grouped together.
Therefore, for copper, its configuration should be written as:
$1s^22s^22p^63s^23p^63d^{10}4s^1$

NOTE

The 4s subshell here is only partially occupied with one electron, rather than fully occupied with two electrons before the 3d subshell has begun to be occupied. Due to the filling of orbitals according to **Hund's rule**, this is more energetically favourable. This is depicted in Figure 1B–11 for the chromium atom.

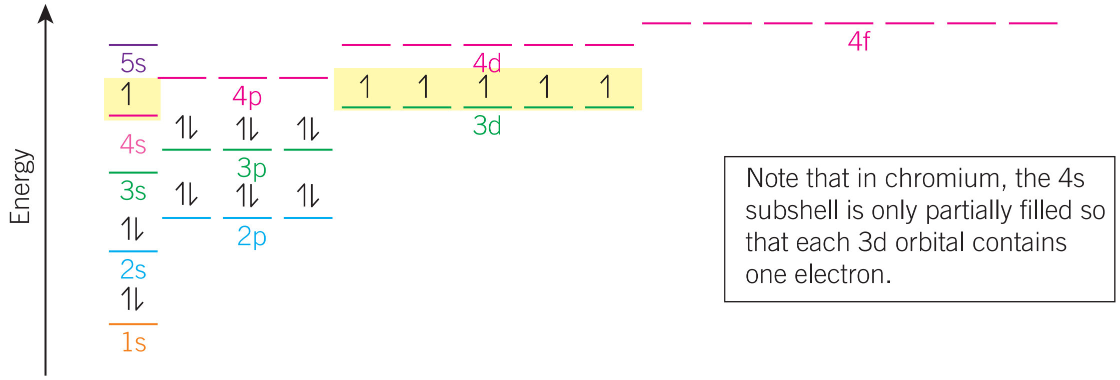

Figure 1B–11 The orbitals within the 4s and 3d subshells in a chromium atom are occupied so that the atom is in a more energy favourable electron configuration.

1B SKILLS

Double-checking subshell electron configurations

When writing electron configurations, there are a few ways you can check your answers to ensure that you have included everything and written the configuration correctly.

The first is to begin by writing the final part of the electron configuration first. To do this, you can use the position of the element on the periodic table. Using the example of carbon provided in this section, carbon is located in period 2 (second row) and therefore contains two shells of occupied electrons. It is located in group 14, which, using the periodic table in Figure 1B–8, corresponds to the subshell ‘p’ being occupied by two electrons. Therefore, the end of carbon’s electron configuration will be $2p^2$. From there, work backwards, filling in each lower energy subshell so that they are completely filled with electrons.

The second – and this can follow on from the first check above – is to add together the superscript numbers in your electron configuration. These should equal the number of electrons present in the atom or ion. Again, using carbon’s electron configuration of $1s^2 2s^2 2p^2$, these superscript numbers add to six, the number of electrons present in carbon.

Shorthand notation for subshell electron configurations

As you may have experienced already, writing out subshell electron configurations in full can be time consuming. Chemists felt the same way and as a result developed a short-hand notation for representing these.

This involves finding the element on the periodic table and then working backwards to find the most recent noble gas in group 18. The chemical symbol of this noble gas is written in square brackets, and then the remainder of the electron configuration is written as per usual.

Using the example of an aluminium atom, this would therefore be written as follows:

$1s^2 2s^2 2p^6 3s^2 3p^1$ … which now becomes … [Ne] $3s^2 3p^1$

Writing the subshell electron configuration for ions

The electron configuration of ions resembles that of a full outer shell. This results from an atom either losing or gaining electrons, and in doing so, forming a positively or negatively charged ion, respectively. For the subshell electron configurations, this means that the configuration will usually end in ‘p^6’.

Aluminium (Al)

$1s^22s^22p^6\underbrace{3s^23p^1}_{\text{Valence electrons lost to form ion}}$

Aluminium ion (Al^{3+})

$1s^2\underbrace{2s^22p^6}_{\text{Full outer shell of electrons}}$

Of course, there are exceptions to this. Lithium, beryllium and boron atoms, which contain one, two and three electrons in their valence shell, respectively, will all lose electron(s), leaving only one shell containing two electrons. Therefore, their electron configurations will be similar to the configuration of helium, $1s^2$. The transition metals are also an exception.

An example of this can be seen with the aluminium ion. Looking at the periodic table, aluminium has an electron configuration of $1s^22s^22p^63s^23p^1$. Further, it contains three valence electrons (in the third shell), which will then be lost to form an ion with a charge of 3+. Removing these three valence electrons results in the electron configuration of $1s^22s^22p^6$, which resembles that of a neon atom.

Section 1B questions

1 Write the electron configurations for the following atoms using both the shell and subshell models.
  - a boron
  - b phosphorus
  - c scandium
  - d krypton

2 Draw the electron configuration for the following atoms or ions according to the Bohr model.
  - a neon
  - b magnesium
  - c oxide (oxygen) ion

3 In Question 2, which of the atoms or ions have the same electron configuration? Write the subshell electron configuration to represent these atoms or ions.

4 Is the electron configuration for the isotopes carbon-12 and carbon-14 the same or different? Explain your answer.

5 Write the electron configuration in subshell notation for the following atoms or ions.
  - a F^-
  - b S
  - c V

6 Explain why the electron configuration for the sodium ion is the same as the electron configuration for neon.

7 Determine if the following atoms would be in the ground state or an excited state. For each, assuming they are not ions, determine the element that the configuration represents.
  - a $[Ar]4s^2$
  - b $1s^22s^23p^1$
  - c $[He]2s^22p^63s^23p^63d^84s^14p^1$
  - d $1s^22s^12p^3$

8 For the atoms determined as being in an excited state from Question 7, explain how you know they are excited by looking at each of their respective electron configurations. In doing so, write out the subshell electron configurations for these elements if they were in the ground state.

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1C The periodic table

Study Design:

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

Glossary:

Atomic radius
Core charge
Electronegativity
Electrostatic force of attraction
First ionisation energy
Ionisation energy
Metallic character

ENGAGE

Mendeleev’s periodic table

Have you heard the story of how the initial version of the modern periodic table was first developed?

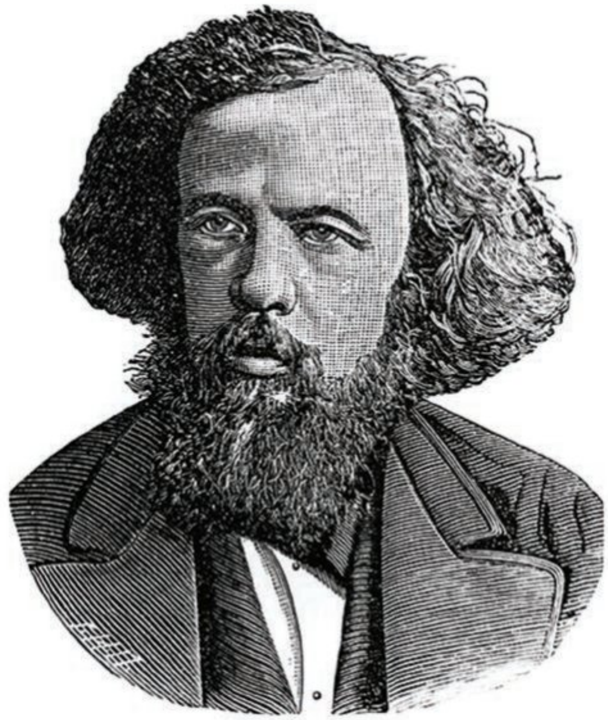
In 1869, a Russian chemist, Dmitri Mendeleev (1834–1907), continued on from the earlier work of other scientists such as Antoine Lavoisier, Johann Döbereiner and John Newlands. Mendeleev jotted down the chemical symbols, atomic weights, and chemical and physical properties of all the elements currently known at the time on separate cards. It was said that he sat for three full days and nights, with no sleep, poring over the cards, trying to arrange them into an order that made sense.

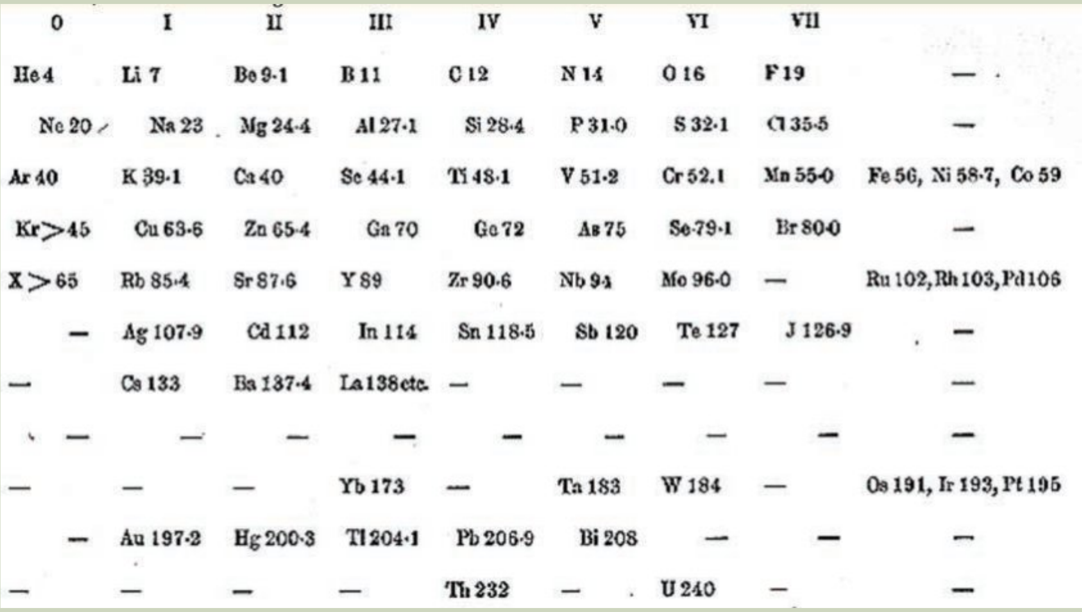
Previously, many chemists had attempted to order the elements based solely on atomic mass or by their chemical and physical properties.

Mendeleev was struggling to find the correct order. Eventually, tired, he fell asleep. It is said that while asleep, all the elements appeared to Mendeleev in a dream and magically arranged themselves into organised positions that formed the foundation of the earliest version of the periodic table. When he woke, he organised each element’s card according to a periodicity he noticed in terms of the element’s chemical and physical properties. This was a huge breakthrough. He noted the repetition of properties every eight elements.

What was more incredible was that Mendeleev realised that many elements had not been discovered yet, so he left spaces in the periodic table for the undiscovered elements. He even predicted the properties and atomic masses of these. Examples of these included gallium, scandium and germanium, all of which were discovered less than 20 years later, and whose atomic masses and properties confirmed his predictions. He also proposed later that some elements, like argon and potassium, which were initially organised based on their increasing atomic mass, should swap positions.

It wasn’t until 1913, six years after Mendeleev’s death, that Henry Moseley (a physicist working in Manchester) demonstrated that an element should be positioned on the periodic table according to its atomic number and not its atomic mass.





0	I	II	III	IV	V	VI	VII	
He 4	Li 7	Be 9·1	B 11	C 12	N 14	O 16	F 19	—
Ne 20	Na 23	Mg 24·4	Al 27·1	Si 28·4	P 31·0	S 32·1	Cl 35·5	—
Ar 40	K 39·1	Ca 40	Sc 44·1	Ti 48·1	V 51·2	Cr 52.1	Mn 55·0	Fe 56, Ni 58·7, Co 59
Kr > 45	Cu 63·6	Zn 65·4	Ga 70	Ge 72	As 75	Se 79·1	Br 80·0	—
X > 65	Rb 85·4	Sr 87·6	Y 89	Zr 90·6	Nb 94	Mo 96·0	—	Ru 102, Rh 103, Pd 106
—	Ag 107·9	Cd 112	In 114	Sn 118·5	Sb 120	Te 127	J 126·9	—
—	Cs 133	Ba 137·4	La 138 etc.	—	—	—	—	—
—	—	—	—	—	—	—	—	—
—	—	—	Yb 173	—	Ta 183	W 184	—	Os 191, Ir 193, Pt 195
—	Au 197·2	Hg 200·3	Tl 204·1	Pb 206·9	Bi 208	—	—	—
—	—	—	—	Th 232	—	U 240	—	—

Figure 1C–1 Dmitri Mendeleev and the original version of his periodic table

EXPLAIN

Organisation of the modern periodic table

The periodic table is the most useful resource available to chemists and chemistry students. If you know how to read it and the power of what it can show you, it minimises the need to memorise the elements as students were instructed to do in days gone by. It provides a framework for organising our understanding through certain trends and properties, which will be explored in this section.

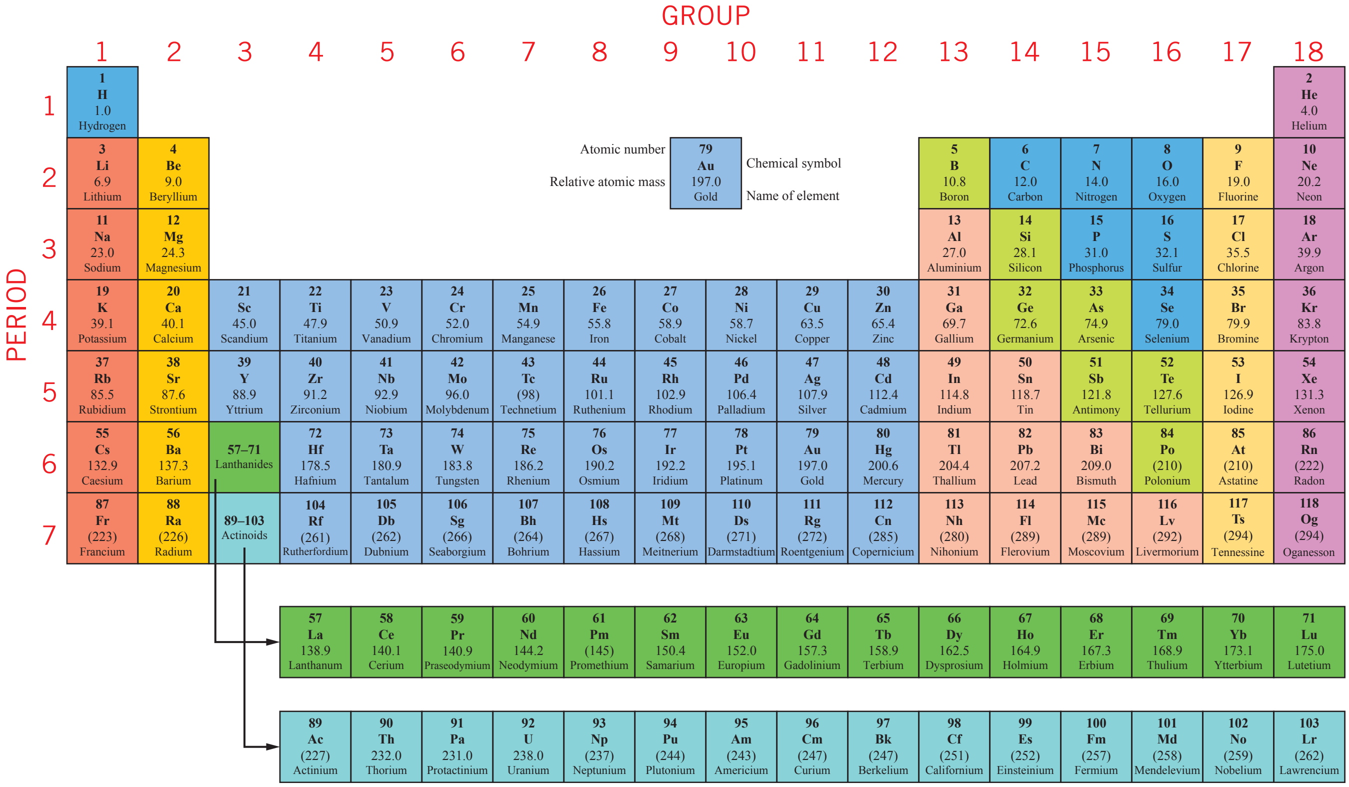
VIDEO 1C–1
UNDERSTANDING THE PERIODIC TABLE

ISBN 978-1-009-22964-7

1A INTRODUCTION TO THE ELEMENTS **LINK**

Compared to Mendeleev's early version of the periodic table, the modern periodic table (Figure 1C–2) is organised by:

- increasing atomic number
- variations in chemical properties (associated with repeating patterns of electron configurations).



GROUP

PERIOD

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** 1.0 Hydrogen																	2 **He** 4.0 Helium
2	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
3	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
4	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
5	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
6	55 **Cs** 132.9 Caesium	56 **Ba** 137.3 Barium	**57–71** Lanthanides	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
7	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

Figure 1C–2 This is the most common form of the modern periodic table.

Structure of the periodic table

Groups

1B ELECTRON CONFIGURATIONS **LINK**

Elements in the periodic table are arranged into vertical columns called groups. As you can see from Figure 1C–2, the periodic table has 18 groups numbered 1–18. Elements in the same group have similar electron configurations, as you learned in Chapter 1B. For example, if you look at the group 17 elements, their electron configurations are as shown in Table 1C–1.

Table 1C–1 Group 17 elements and their subshell electron configurations

Group 17 element	Electron configuration
Fluorine	$1s^2 2s^2 2p^5$
Chlorine	$1s^2 2s^2 2p^6 3s^2 3p^5$
Bromine	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$
Iodine	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^5$
Astatine	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 4f^{14} 5s^2 5p^6 5d^{10} 6p^5$

All elements in a group tend to have very similar chemical properties. Some of these common groups and their properties are listed in Table 1C–2.

Table 1C–2 Common groups, their names, and chemical and physical properties

Group number	Name	Properties
1	Alkali metals	Soft metals – can be cut with a butter knife Very reactive with water
2	Alkaline earth metals	Good conductors of electricity Relatively low density, melting and boiling points (compared to other metals) Soft (but not as soft as alkali metals)
17	Halogens	Coloured gases (with the exceptions of bromine, which is a liquid at room temperature and iodine and astatine, which are solids) Odorous
18	Noble gases	Unreactive

Periods

The horizontal rows in the periodic table are called periods. As you can see from Figure 1C–2, the periodic table has seven periods numbered 1–7. The period an element is in tells you the number of occupied electron shells in the element's atoms. For example, the valence shell of magnesium and chlorine in period 3 is the third shell, so their electron configurations are:

$$\text{Mg } 1s^22s^22p^6\underbrace{3s^2}_{\text{Valence shell}}$$

$$\text{Cl } 1s^22s^22p^6\underbrace{3s^23p^5}_{\text{Valence shell}}$$

Blocks

The periodic table, as you saw in Chapter 1B, is arranged into blocks relating to the last subshell filled in the element's electron configuration (Figure 1C–3).

LINK **1B** ELECTRON CONFIGURATIONS

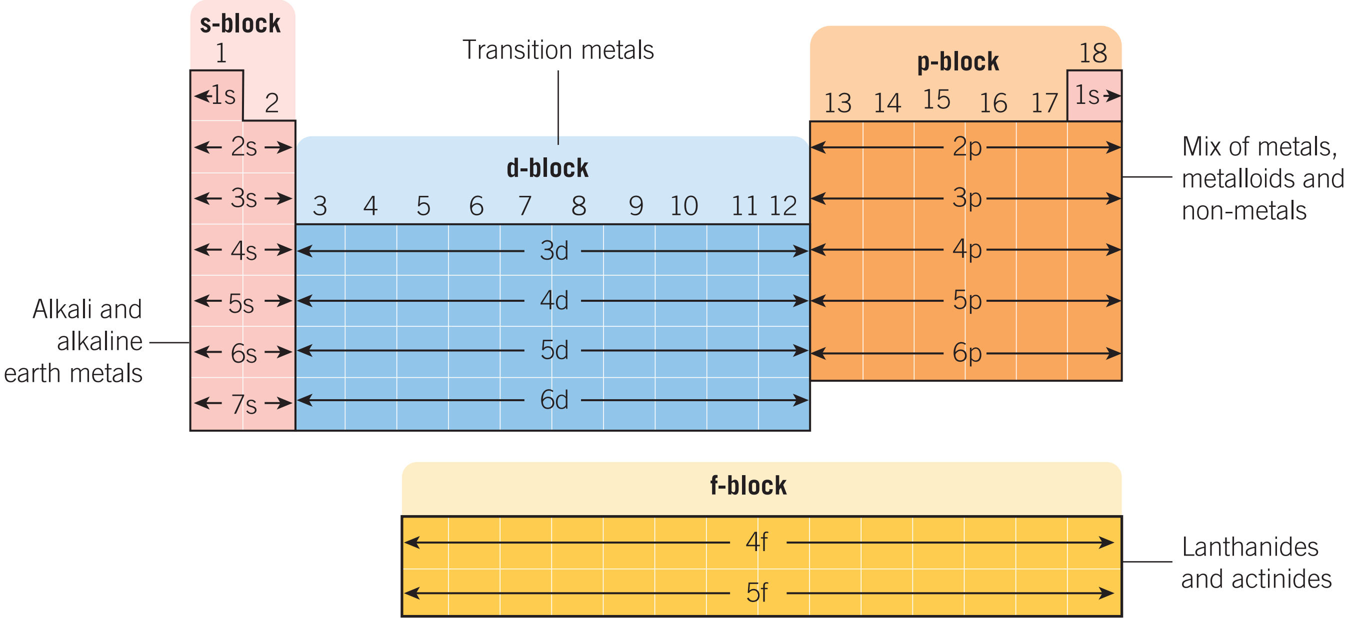


Figure 1C–3 This periodic table highlights the location of the different blocks, based on shells and subshells.

Check-in questions – Set 1

1 What is the difference between groups and periods in the periodic table?

2 What do all elements in the same group have in common? How is this different from all elements in the same period?

Trends in the periodic table

This section explores six key properties of the periodic table and how these can be used as a predictive tool to determine the trend in the properties of elements across a period (left to right) or down a group (top to bottom). We'll explore the comparison between elements using specific examples. What will be important to note here is that the properties are dependent on the **electrostatic force of attraction** between the valence electrons and the positively charged protons in the nucleus. This attractive force is affected by:

- core charge
- atomic radius, determined by the number of electron shells.

Electrostatic force of attraction the attraction between positive and negatively charged particles that are not directly in contact

Core charge the electrostatic force of attraction between valence electrons and the nucleus. It can be calculated by subtracting the number of non-valence electrons from the number of protons

Core charge

Core charge is the electrostatic force of attraction of outer-shell (valence) electrons to the nucleus.

A core charge value can be calculated by the total number of protons of an element minus the sum of the number of inner-shell electrons.

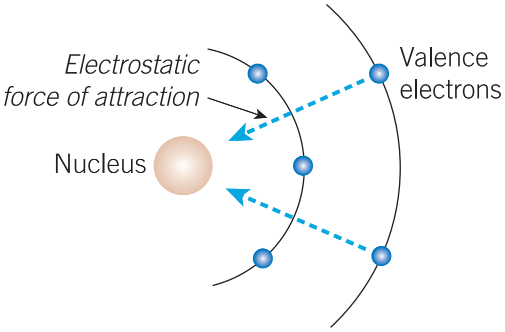


Figure 1C–4 There is an electrostatic force of attraction between valence electrons and protons in the nucleus.

Core charge = total number of protons – total number of inner-shell electrons

Worked examples 1C–1 and 1C–2 for magnesium and sulfur demonstrate this further.

Worked example 1C–1: Finding the core charge of magnesium

Determine the core charge of a magnesium atom.

Solution

Step 1	Determine the atomic number (number of protons) of magnesium.	12
Step 2	Write the Bohr electron configuration for magnesium.	2,8,2
Step 3	Determine the number of inner-shell electrons (i.e. all the electrons not in the valence shell).	2 + 8 = 10
Step 4	Determine core charge using the equation. (answer to Step 1 – answer to Step 3)	Core charge = total number of protons – total number of inner-shell electrons Core charge = 12 – 10 Core charge = +2

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Worked example 1C–2: Finding the core charge of sulfur

Determine the core charge of a sulfur atom.

Solution

Step 1	Determine the atomic number (number of protons) of sulfur.	16
Step 2	Write the Bohr electron configuration for sulfur.	2,8,6
Step 3	Determine the number of inner-shell electrons (i.e. all the electrons not in the valence shell).	2 + 8 = 10
Step 4	Determine core charge using the equation. (answer to Step 1 – answer to Step 3)	Core charge = total number of protons – total number of inner-shell electrons Core charge = 16 – 10 Core charge = +6

NOTE

Core charge is always positive, so the answers above will always be represented as a +2 or +6.

Comparing magnesium and sulfur, which are both elements in period 3, you can clearly see that core charge increases moving from left to right across the period. As the number of protons increases in an atom (and the number of occupied electron shells remains unchanged), there is a greater electrostatic force of attraction between the valence electrons and the protons. Hence, core charge increases.

Moving down a group, the number of occupied electron shells increases; however, the core charge remains the same. This can be seen in Worked example 1C–3, where the calculated core charge of calcium is the same as that of magnesium (both from the same group).

Worked example 1C–3: Finding the core charge of calcium

Determine the core charge of a calcium atom.

Solution

Step 1	Determine the atomic number (number of protons) of calcium.	20
Step 2	Write the Bohr electron configuration for calcium.	2,8,8,2
Step 3	Determine the number of inner-shell electrons (i.e. all electrons not in the valence shell).	2 + 8 + 8 = 18
Step 4	Determine core charge using the equation. (answer to Step 1 – answer to Step 3)	Core charge = total number of protons – total number of inner-shell electrons Core charge = 20 – 18 Core charge = +2

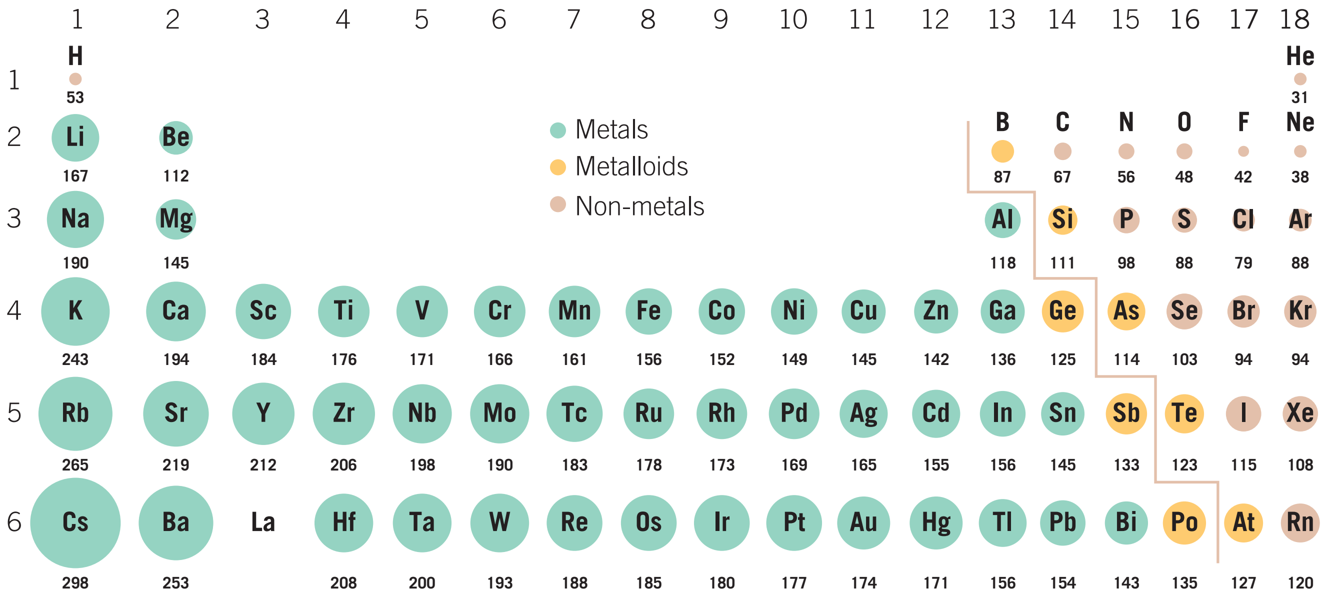
Atomic radius

Atomic radius (plural: radii) a measure of the size of an atom from the centre of its nucleus to its valence shell

The **atomic radius** is a measure of the size of an atom. It may be defined as the distance from the centre of its nucleus to the valence electron shell. An atom's atomic radius is affected by the electrostatic force of attraction (core charge) between the nucleus and valence electrons and also by how many occupied electron shells an atom has.

Core charge does not change as you move down a group. However, the number of occupied shells of electrons increases. Think about what effect this would have on the attraction between the valence electrons and the protons in the nucleus. If you said they would be less strongly attracted, you would be correct.

However, for elements in the same period, where electrons occupy the same number of shells, the increase in core charge moving from left to right means that there is a greater force of attraction between the electrons and the nucleus, and so the atomic radius becomes slightly smaller. This can be seen in Figure 1C–5.



	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 53																	He 31
2	Li 167	Be 112											B 87	C 67	N 56	O 48	F 42	Ne 38
3	Na 190	Mg 145											Al 118	Si 111	P 98	S 88	Cl 79	Ar 88
4	K 243	Ca 194	Sc 184	Ti 176	V 171	Cr 166	Mn 161	Fe 156	Co 152	Ni 149	Cu 145	Zn 142	Ga 136	Ge 125	As 114	Se 103	Br 94	Kr 94
5	Rb 265	Sr 219	Y 212	Zr 206	Nb 198	Mo 190	Tc 183	Ru 178	Rh 173	Pd 169	Ag 165	Cd 155	In 156	Sn 145	Sb 133	Te 123	I 115	Xe 108
6	Cs 298	Ba 253	La	Hf 208	Ta 200	W 193	Re 188	Os 185	Ir 180	Pt 177	Au 174	Hg 171	Tl 156	Pb 154	Bi 143	Po 135	At 127	Rn 120

Figure 1C–5 Pictorial representation of the relative size of atoms. Numerical values represent theoretical calculations of radius in picometres (1 pm = 1 × 10^{-12} m). Notice the trend in size as you move across a period versus down a group.

When answering questions that compare trends, such as atomic radii for two elements, the following steps should be followed:

1. Include a definition of the property.
2. Include a statement about the key difference.
3. Relate your answer back to the original question.

This is also outlined in Worked examples 1C–4 and 1C–5.

Worked example 1C–4: Comparing atomic radius of Li and K

Which element's atom has the larger atomic radius – lithium or potassium? Explain.

Solution

Step 1	Define atomic radius.	Atomic radius is a measure of the size of an atom from the centre of its nucleus to the valence shell.
Step 2	Make a statement about the key difference.	Lithium and potassium are both in group 1 so have a core charge of +1. However, lithium is in period 2 and potassium is in period 4.
Step 3	Relate back to the question.	As potassium has more shells (4) compared to lithium (2), it has a larger atomic radius.

Worked example 1C–5: Comparing atomic radius of C and O

Which element's atom has a larger atomic radius – carbon or oxygen? Explain.

Solution

Step 1	Define atomic radius.	Atomic radius is a measure of the size of an atom from the centre of its nucleus to the valence shell.
Step 2	Make a statement about the key difference.	Carbon and oxygen are both in period 2. However, carbon has a core charge of +4, whereas oxygen has a core charge of +6.
Step 3	Relate back to the question.	As carbon has a smaller core charge, the attraction of its valence electrons to the protons in the nucleus is not as strong, so it has a larger atomic radius.

NOTE

In Worked examples 1C–4 and 1C–5, the following are referred to:
- period and/or group location of the two elements
- core charge.

Check-in questions – Set 2

1. Define 'core charge' and 'atomic radius'.
2. An element further to the right of the periodic table but in the same period will have a greater core charge. True or false?
3. An element further towards the top of a group will have a larger atomic radius. True or false?

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Electronegativity

Electronegativity the ability of an atom to attract electrons towards itself. This is affected by an atom's core charge and the distance electrons are from the nucleus (atomic radius)

Electronegativity is defined as the ability of an atom to attract shared electrons towards itself. Therefore, the higher the attraction of the valence electrons to the nucleus, the higher the electronegativity will be.

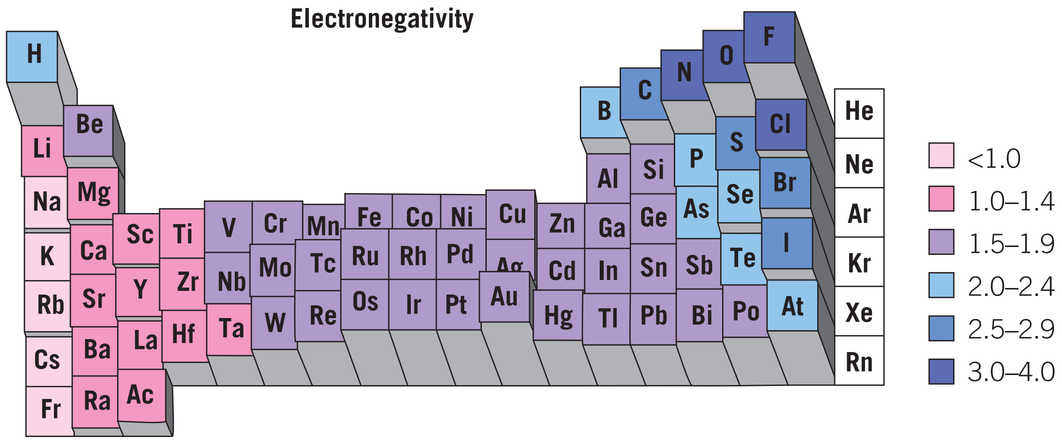
Relating this back to the trends in core charge and atomic radius, for elements in the same period, the greater the core charge (as you move left to right), the greater the attraction of valence electrons to the nucleus. This results in a decrease in atomic radius as electrons are drawn closer to the nucleus, meaning the electrostatic force of attraction between electrons and the nucleus is greater. The same can be applied looking at the trend for electronegativity down a group. Here, as the atomic radius increases due to an increased number of occupied electron shells, the valence electrons are further from the nucleus. Therefore, there is a weaker electrostatic force of attraction, so electronegativity decreases.

1A INTRODUCTION TO THE ELEMENTS LINK

2B MOLECULAR SHAPES AND POLARITY LINK

4A FORMATION AND NAMING IONIC COMPOUNDS LINK

Elements with greater electronegativity are therefore located in the top right of the periodic table. Remember that these elements are the non-metals. This makes sense when you recall that atoms of non-metallic elements want to gain electrons to fill their outer shells, forming negatively charged anions, as opposed to metallic elements, which have few electrons in their valence shell and want to lose electrons, forming positively charged cations. Figure 1C–6 shows the comparison in electronegativity between different elements. The values shown for electronegativity become useful when determining whether a molecule is polar or non-polar.



Electronegativity increases across a period.

Electronegativity decreases down a group.

1	2	13	14	15	16	17
Li 1.0	Be 1.6	B 2.0	C 2.6	N 3.0	O 3.4	F 4.0
Na 0.9	Mg 1.3	Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2
K 0.8	Ca 1.0	Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 3.0
Rb 0.8	Sr 1.0	In 1.8	Sn 2.0	Sb 2.1	Te 2.1	I 2.7
Cs 0.8	Ba 0.9	Tl 2.0	Pb 2.3	Bi 2.0	Po 2.0	At 2.2
Fr 0.7	Ra 0.9					

Figure 1C–6 This periodic table shows the elements as different-sized bars to represent their respective electronegativity values (top). The relative electronegativity values are shown in the table and will be discussed more in Chapter 2 with covalent compounds (bottom).

When answering questions explaining the difference in electronegativity between two elements, it is important that you refer to the other trends discussed so far. Worked example 1C–6 walks you through these steps.

Worked example 1C–6: Comparing electronegativity

Explain whether oxygen or nitrogen has a higher electronegativity.

Solution

Step 1	Define the property/trend.	Electronegativity is the ability of an atom to attract shared electrons to itself.
Step 2	Describe any similarities between the two elements (this includes referring to other trends).	Both oxygen and nitrogen are in period 2, so they have the same number of occupied shells of electrons.
Step 3	Make a statement about the key difference (this includes referring to other trends).	Oxygen is in group 16 and has a core charge of +6, whereas nitrogen is in group 15 and has a core charge of +5. Oxygen will therefore have a smaller atomic radius.
Step 4	Relate back to the question.	This means there is a greater electrostatic force of attraction between the valence electrons for oxygen, so it has a higher electronegativity compared to nitrogen.

Check-in questions – Set 3

1 Using Worked example 1C–6 as a guide, compare the following pairs of elements and explain which element has the greater electronegativity.
   a magnesium and strontium
   b lithium and fluorine

First ionisation energy

First ionisation energy is defined as the amount of energy required to remove the outermost electron from a gaseous atom.

First **ionisation energy** follows a similar trend to electronegativity. The stronger the electrostatic force of attraction between valence electrons and the nucleus, the greater the amount of energy required to remove an electron from the gaseous atom. This can be seen in Figure 1C–7 on the following page.

First ionisation energy the energy required to remove the first, or outermost, electron from a gaseous atom

Ionisation energy the energy required to remove an electron from a gaseous atom or ion

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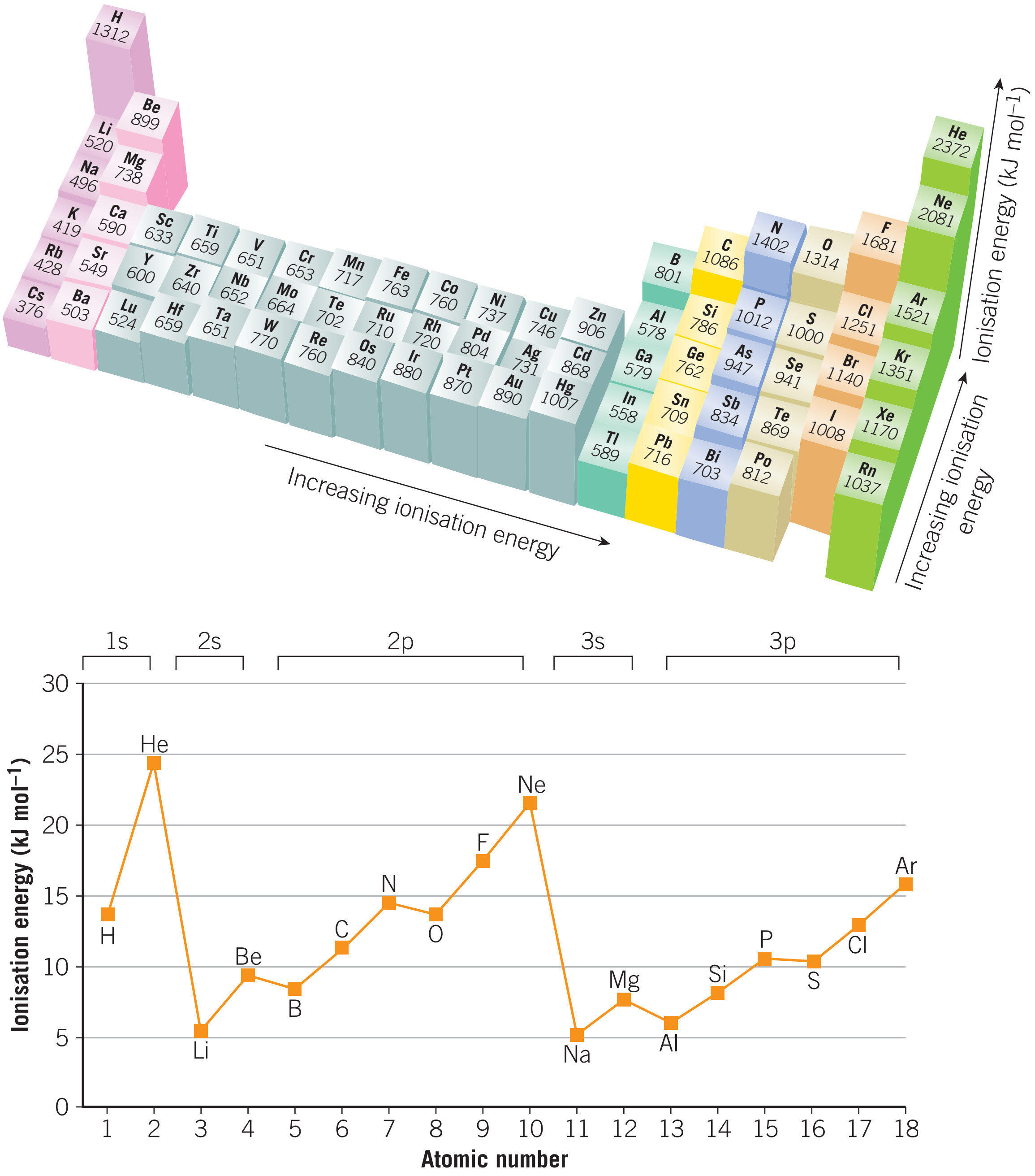


Figure 1C–7 This periodic table represents the elements as different-sized bars to represent their respective first ionisation energy values (top). The graph shows the energy required to remove the first electron of each element up to element 18, argon (bottom).

Figure 1C–7 shows the subshell location and energy required to remove the first electron of the first 18 elements. You should note these points:

- The first ionisation energies for noble gases are all higher than any other element. This is because they have full valence shells of electrons.
- Generally, the trend for ionisation energy increases as you go across the period. There are slight dips when beginning the p subshell in the second and third electron shell and dips when electrons begin pairing up in the orbitals of the p subshell (recall Hund's rule).
- Generally, the trend for first ionisation energy decreases as you start a new shell because the valence electrons are further from the nucleus (recall atomic radii and electronegativity).

Similar to questions about electronegativity, when answering questions explaining the difference in the first ionisation energy between two elements, it is important that you refer to other trends discussed so far. Worked example 1C–7 guides you through these steps again.

Worked example 1C–7: Comparing first ionisation energy

Explain whether arsenic (As) or antimony (Sb) has the higher first ionisation energy.

Solution

Step 1	Define the property/trend.	First ionisation energy is the energy required to remove the outermost electron from a gaseous atom.
Step 2	Describe any similarities between the two elements (including reference to other trends).	Both arsenic and antimony are in group 15 and therefore have the same core charge (+5).
Step 3	Make a statement about the key difference (including reference to other trends).	Antimony is in period 5, whereas arsenic is in period 4, meaning it has one fewer shell of electrons and therefore a smaller atomic radius. This means there is a greater electrostatic force of attraction between the valence electrons for arsenic, so it has a higher electronegativity compared with antimony.
Step 4	Relate back to the question.	For this reason, the first ionisation energy of arsenic is higher than that of antimony.

Check-in questions – Set 4

1 Using Worked example 1C–7 as a guide, compare the first ionisation energies for magnesium and phosphorus. Explain which value is greater and why.

Metallic character and reactivity

LINK **3A** STRUCTURE AND PROPERTIES OF METALS

As you will learn in Chapter 3A, some of the general properties of metals, which we can refer to as metallic character, include:

- the ability to conduct electricity and heat
- relatively high melting and boiling points
- few electrons (one to three) in their valence shell, so an ability to lose these electrons.

LINK **1B** ELECTRON CONFIGURATIONS

When moving across a period from left to right, you know from Chapter 1B that elements change from metals to metalloids and then to non-metals. Therefore, it makes sense that **metallic character** decreases. Moving down a group of metals (group 1 or 2), you know that the atomic radius of a metal increases, meaning its electronegativity and first ionisation energy decrease. This means that the more easily a metal loses electrons, the more reactive it is (and the more metallic character it has). This can be seen in the reactions of the group 1 metals, sodium and potassium, which react vigorously with cold water. In each of these reactions, hydrogen gas is given off and a metal hydroxide is produced (Figure 1C–8).

Metallic character the ability of a metal to lose its valence electron and display common metallic properties

Figure 1C–8 Group 1 metals, such as sodium (left) and potassium (right), reacting with water

WORKSHEET 1C–1
EXPLAINING TRENDS IN THE PERIODIC TABLE

Non-metallic character and reactivity

Non-metals have many valence electrons compared to metals and therefore tend to want to gain (one, two or three) electrons as opposed to losing electrons.

Moving across a period from left to right, non-metallic character increases, as non-metals are located to the right of the periodic table. Moving down a group of non-metals (groups 15, 16 or 17), you will recall that the atomic radius of a non-metal increases, meaning its electronegativity and ionisation energy decreases. This means at the top of that group, a non-metal that gains electrons more easily will be more reactive as a result.

1C SKILLS

Summary of trends

Good revision for this section, on top of practicing your literacy skills and writing explanations of different trends and the key terms learned, is to use a summary table to represent similarities and differences in trends in the periodic table. Tables, like Table 1C–3 below, are a good way of easily comparing many different properties or features of items. Using content in this chapter, redraw Table 1C–3 in your own summary notes and complete it by filling in the gaps.

Table 1C–3 Summary of the pattern of different trends in the periodic table

Property/Trend	Change down a group	Change across a period (left to right)
Core charge		
Atomic radius		
Electronegativity		
First ionisation energy		
Metallic character		
Non-metallic character		

Summary tables like this will be important again when comparing the similarities and differences between covalent molecules, metals and ionic compounds after you have covered Chapters 2, 3 and 4, respectively.

Section 1C questions

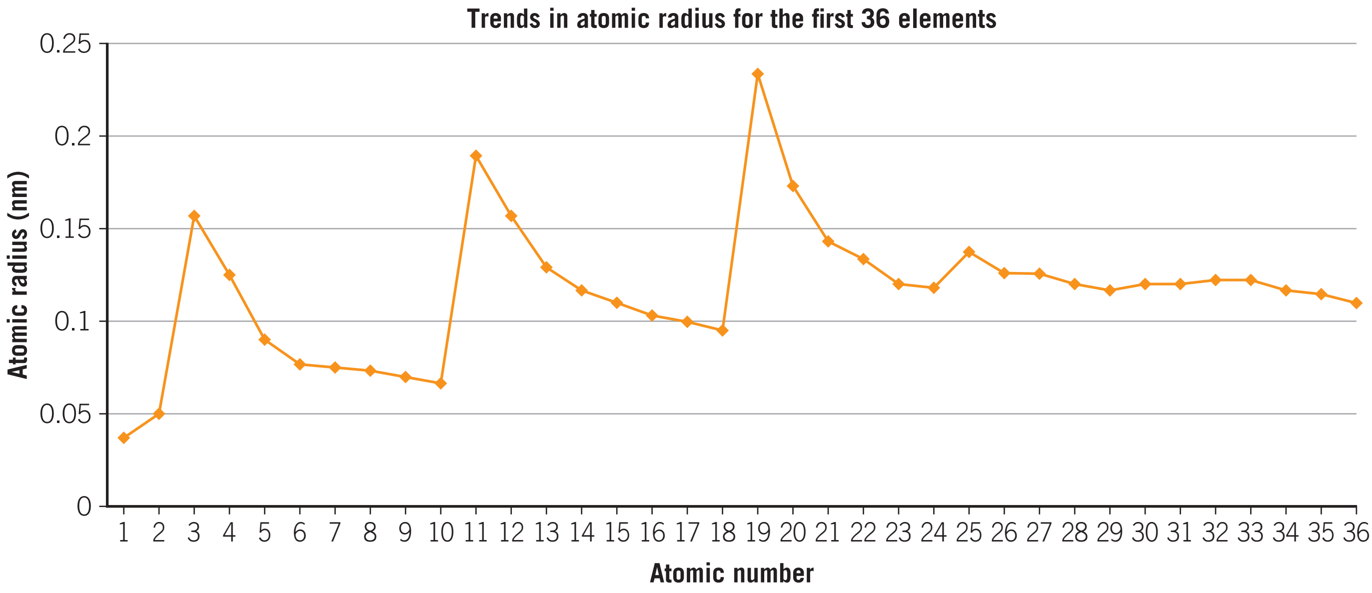
1 Compare and explain whether tin or tellurium would have a greater metallic character. Include the terms 'atomic radius', 'electronegativity' and' ionisation energy' in your answer.

2 The atomic radius is the measure of the size of an atom from the centre of the nucleus to its valence shell.

a Which numbered shell would a scientist use to determine the atomic radius of a nitrogen atom?

b Which numbered shell would be used to determine the atomic radius of a calcium ion?

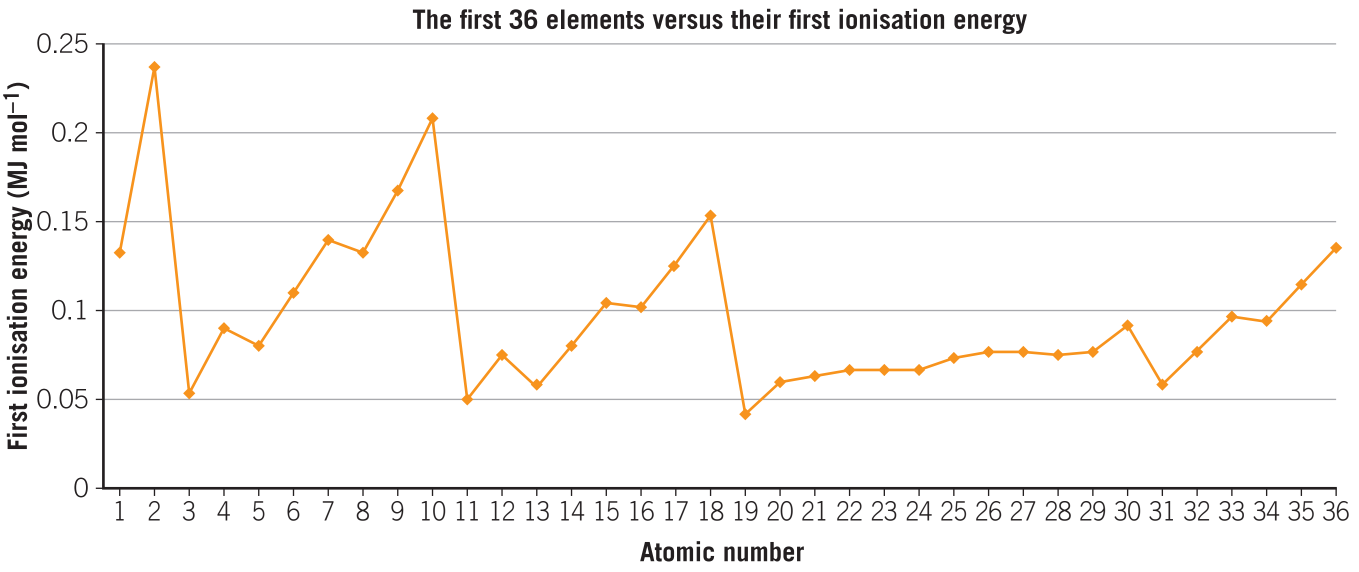
The graph below shows the atomic radii, in nanometres, for the first 36 elements on the periodic table.



c Using this data, explain the trend for atomic radius across a period.

d Can the graph be used to compare the trend in atomic radius for elements in the same group? If so, explain the trend by referring to the data.

3 The graph below shows the first ionisation energy for the first 36 elements in the periodic table.



- **a** What is the difference between ionisation energy and first ionisation energy?
- **b** For chlorine, what is the subshell of the electron that would be easiest to remove from the atom?
- **c** Explain how the electronegativity of an atom impacts its first ionisation energy. In your answer, include a description of what electronegativity is.
- **d** Referring to data in the graph, explain the trends observed in first ionisation energy both down a group and across a period.
- **e** Looking at the graph, the trend you explained in part **d** for first ionisation energy across a period is not always true when looking at two successive elements. This can be seen when looking at either beryllium and boron or nitrogen and oxygen in period 2. Using your knowledge of subshell configurations, explain why this may be the case for these elements.
- **f** Use the graph to predict the first ionisation energy of element 37, rubidium.
- **g** Can a trend in metallic or non-metallic character be predicted by looking at the values for first ionisation energy shown in the graph? Explain your answer by defining both metallic and non-metallic character.

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1D Critical elements and recycling processes

Study Design:

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

Glossary:

Boiling point
Circular economy
Critical element
Cryogenics
Eutrophication
Melting point
Ore
Slag
Smelting

ENGAGE

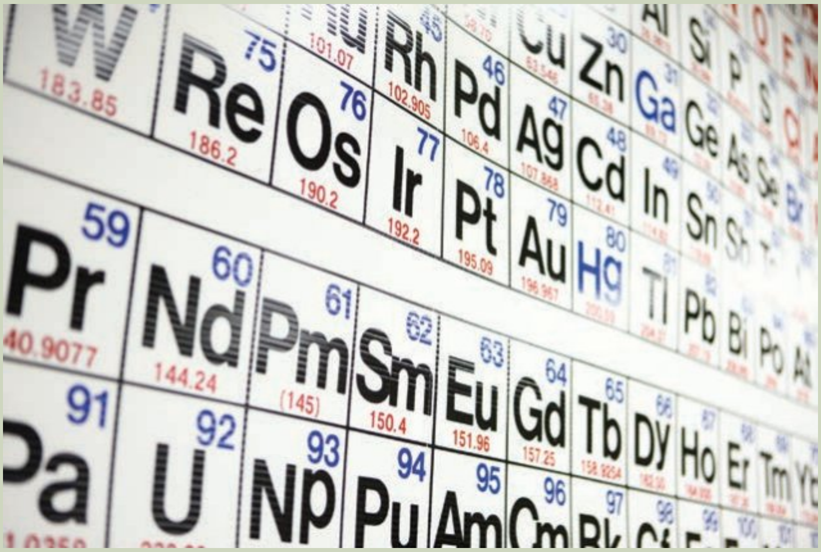
Elements essential to modern life

Over time, the discovery of new elements and an understanding of their physical and chemical properties has shaped our everyday lives. Many elements have become essential parts of our modern technologies; however, chemists are beginning to realise the use of these elements is becoming unsustainable and creating significant impacts on both our economy and environment.

The year 2019 was designated the International Year of the Periodic Table to coincide with the 150th anniversary of Mendeleev's presentation of his periodic table, marking its importance to science. There was also a large focus on not only the critical elements, but all elements, and the importance of developing sustainable recycling practices for element recovery.

There are fewer than 100 stable elements on the periodic table from which we create everything required for our 'comfortable' survival. The more recently discovered elements have all been synthesised artificially and are radioactive, meaning they decay over time. It is now not expected that any more new stable elements will be discovered for our everyday use. So, as a society, we must make the best of the elements we have and understand that these elements are finite and non-renewable.

This section will look at some examples of these critical elements and how chemists are developing and employing different recycling processes to recover elements before it is too late and they are depleted.



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EXPLAIN

What are critical elements?

Critical element a chemical element essential to the economic or national security of a nation and vulnerable to disruption of supply

Critical elements (also called technology-critical elements) are defined as those which:

- are essential to the economic or national security of a nation – for example, elements vital in the manufacture of a product or spread of a new technology – and if absent or depleted would harm the economy or defence and security capabilities
- are vulnerable to disruption of supply.

Examples of critical elements

In the examples below and on the following pages, we will look at a description of the critical elements, along with their location in the periodic table, and chemical and physical properties. Also, where applicable, the life cycle of each element or set of elements will be explored in greater depth using the following headings:

Mining and extraction
Processing (including appropriate measures to limit wastage)
Uses
Recycling

Finally, environmental issues associated with the life cycle will be covered.

Helium

Boiling point the temperature at which the vapour pressure above the liquid equals the pressure exerted by the surroundings on the liquid; temperature at which liquid boils

Melting point the temperature at which a solid melts

Helium, a colourless, odourless, non-toxic noble gas found in group 18, is a non-metal that has the lowest **boiling** and melting point of all elements. It has a **melting point** of –268°C (5.2 K) and can remain a liquid at absolute zero (0 K). Helium is the second most abundant element behind hydrogen. For this reason, it seems strange to classify this as a critical element.

Mining and extraction

Helium is both a by-product of the extraction of natural gas and formed from the radioactive decay of uranium and thorium, after which some of the helium is trapped in Earth's mantle.

Processing

The USA has the world's largest supply of helium, the National Helium Reserve in Amarillo, Texas, which holding approximately 670 million cubic metres (2019). This reservoir was created to ensure enough helium was available for US airships and the navy, not just party balloons. Globally however, there is an estimated 51.9 billion cubic metres of helium. So, why is it classified as a critical element?

UNIT 3 LINK

The first issue is when helium is sourced from natural gas, carbon dioxide is produced, which has greater impacts on our environment. The other problem lies in the fact that because helium is so light (the second lightest element), it continuously escapes Earth's atmosphere, and thus our supply is constantly diminishing. It is estimated that there is a deterioration rate of approximately 0.5% every year. However, global demand continues to increase at approximately 2% per year.

There is no simple solution to rectify the depletion rate of helium. Chemists cannot artificially produce helium or extract it from the atmosphere, and it is believed we only have 25–50 years of helium remaining. In 2018, helium was included on the US Geological Survey's (USGS) critical elements list. However, it was removed from this list in the most recent USGS survey in 2021.

Uses

Helium's low melting point and high thermal conductivity makes it very useful in **cryogenics**. It is used to cool superconducting magnets in both magnetic resonance imaging (MRI) devices and nuclear magnetic resonance (NMR) scanners. You will learn about the latter in Unit 4. Without the use of these, scientists would need to invent new modern imaging machines.

Cryogenics the production and effects of substances at very low temperatures

LINK **UNIT 4**

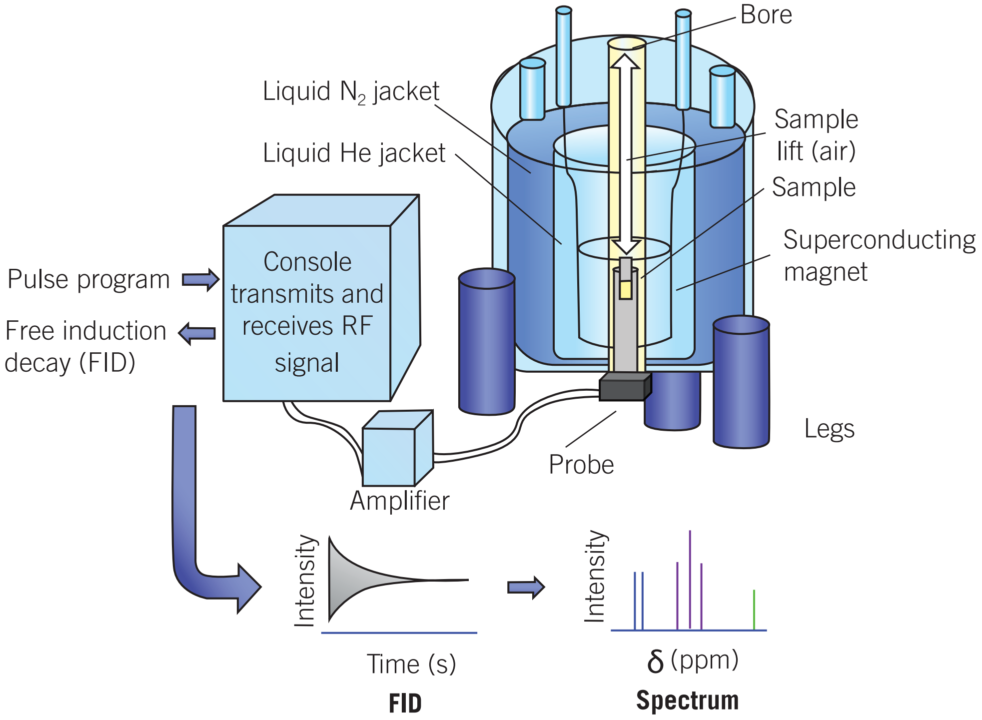

Figure 1D–1 A nuclear magnetic resonance (NMR) machine uses helium to cool its superconducting magnets.

As shown in Figure 1D–2, helium also has many other uses, including:

- semiconductor chip manufacture
- aerospace industry
- optical fibre manufacture
- nuclear power generation.

Recycling

We can recycle helium to maximise our usage of it. An example of this involves trapping the vented helium gas used in MRI devices and then re-liquefying it for re-use.

Non-recycling ways to minimise the usage of helium include:

- using alternative elements, like hydrogen, where possible
- raising the price, causing it to become too expensive to waste.

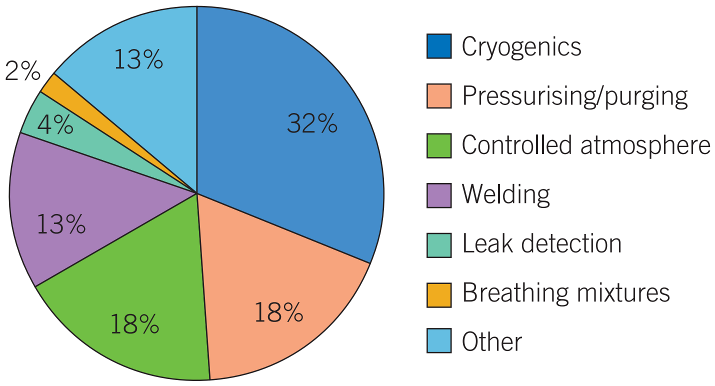

Figure 1D–2 Various uses of helium as recorded in 2016
Source: US Geological Survey (USGS)

Phosphorus

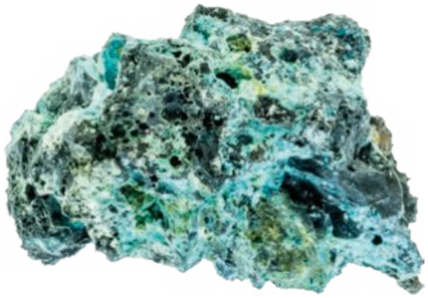

Figure 1D–3 Turquoise phosphate rock

Phosphorus, found in group 15, is a non-metal that is essential for all life. It is a key element that makes up the backbone of DNA, cell membranes, bones and teeth in living organisms, as well as a key ingredient in many fertilisers.

Mining and extraction

Phosphate rock is a general term used to describe any rock containing a high proportion of phosphate minerals. These rocks form from the mineralisation of dead sea organisms over millions of years. These sedimentary deposits become available as a result of tectonic plate movements, enabling them to be mined, mostly by surface methods. Approximately 70% of the world's phosphate rock mines are located in Morocco, and worldwide, most of the phosphate produced is used to make fertilisers for use in areas of land low in phosphate. Phosphate rock has been found at several localities in Victoria, but the only known commercial production has been from Phosphate Hill, near Mansfield.

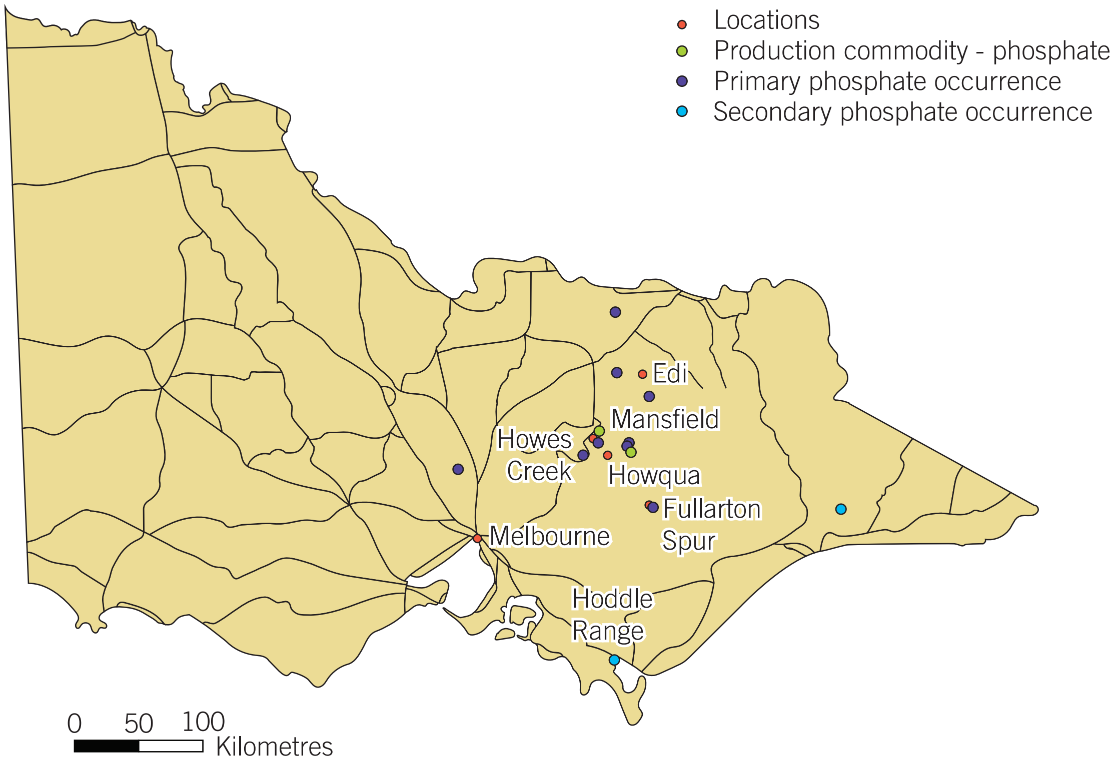

Figure 1D–4 Locations of phosphate rock in Victoria

Uses

Phosphorus, as mentioned, is more than a key element that forms part of living organisms DNA, cell membranes, bones and teeth. It is also a key ingredient in fertilisers, water softeners, soaps, detergents and insecticides.

Processing, environmental issues and potential recycling

Approximately 80% of phosphorus is lost in the supply chain from when it is mined to when it reaches you in your food, as most phosphate rock is used in food production. The large proportion of this percentage is lost to water supplies during agricultural runoff and wastewater, where any excess phosphate can cause **eutrophication**, a dense growth of harmful algae. This decreases the oxygen levels in the water that fish and other aquatic life need to survive, which can result in long-lasting imbalances to ecosystems and the loss of potential livelihood in other industries, like fisheries.

Eutrophication a dense growth of plant life as a result of high concentrations of nutrients in water, commonly due to runoff from land

It is estimated that phosphorus resources will begin to diminish by around 2080 due to the increase in global demand from developing countries for growing their own food to meet increasing population size. Currently, there are no substitutes for phosphorus in agriculture, but there are some other practices that can help to reduce this phosphorus wastage, such as:

- reduce washing of food
- compost the kitchen scraps to return phosphorus to the soil
- reduce the amount of meat and dairy in your diet
- reduce the rearing of phosphorus-dependent livestock
- innovations in industry, such as the development of:
  - time-release fertilisers by chemists
  - waste-treatment plants that are designed to recover the three million tonnes of phosphorus passed in our faeces and urine each year (some areas in Europe are already doing this).

Rare-earth elements

Rare-earth elements are the group of 17 elements in the periodic table referred to as the lanthanides, as well as yttrium (Y) and scandium (Sc). Scandium and yttrium are included as rare-earth elements because they are often found in mineral deposits with the lanthanides and have similar chemical properties. Figure 1D–5 highlights these elements.

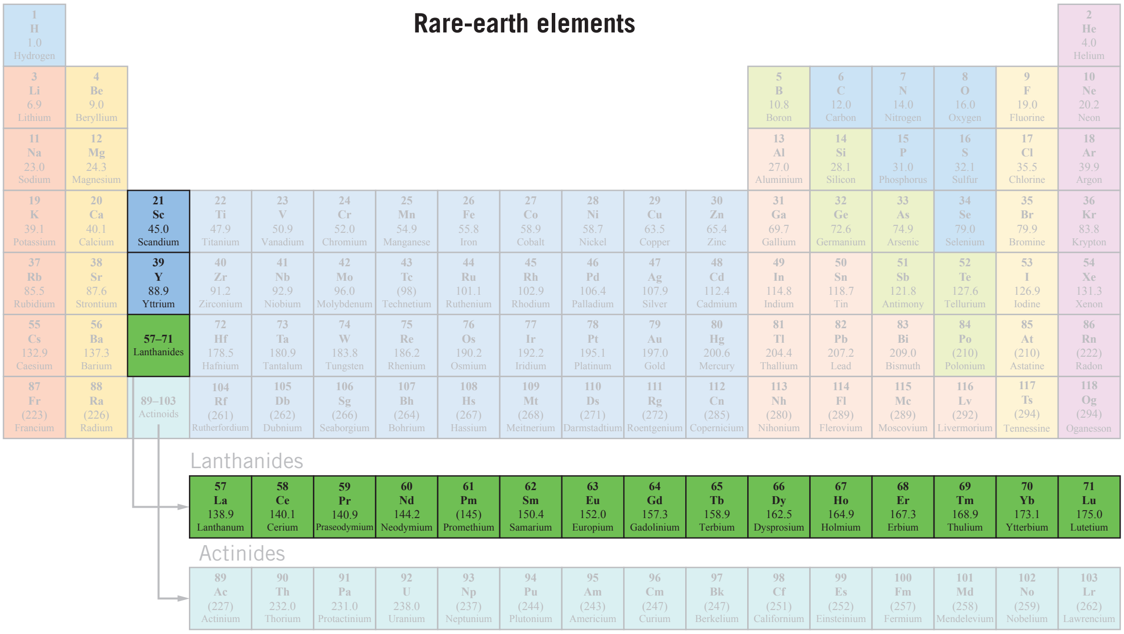

Figure 1D–5 A periodic table highlighting the location of the rare-earth elements – the lanthanides – plus scandium and yttrium

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These elements are strong reducing agents that can react with other elements to form compounds with particular properties, such as being corrosion resistant. Most rare-earth element ions are fluorescent, absorb UV light and emit visible light.

Mining and extraction

Rare-earth elements are not found in high concentrations in Earth's crust, which is why they were initially given the name 'rare-earth metals' (Figure 1D–6). These elements can further be classified into light elements (La to Sm in the lanthanides period) and heavy elements (Eu to Lu in the same period). The heavy elements are much less common and as a result are more expensive.

4A FORMATION AND NAMING OF IONIC COMPOUNDS LINK

Often these elements are found in combination with non-metals (as ionic compounds), making them difficult, and costly, to extract when mined.

Figure 1D–6 Some examples of rare-earth metal oxides: neodymium (front), samarium and lanthanum (middle left to right); and gadolinium, praseodymium and cerium (back left to right)

Processing

In the mid-1960s, production of rare-earth metals from mining and extraction in the USA increased rapidly, coinciding with the introduction of colour television (Figure 1D–7). In the mid-1980s, China started selling these elements at a very low price, which subsequently forced mining in the USA to stop as they could not make a profit. When China stopped exports in 2010, the prices of these rare-earth elements increased enormously. This then led to an increased processing of these elements in the USA again as well as in other countries, including Australia, Russia, Thailand and Malaysia.

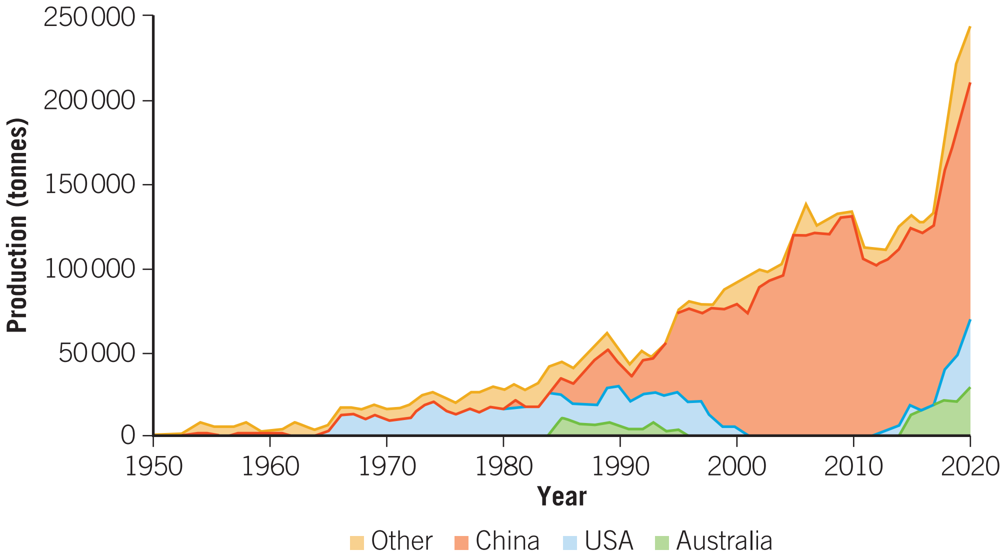

Figure 1D–7 A graph showing rare-earth element production between 1950 and 2020, published by the USGS

Today, everyone you know is likely to own a mobile phone. All these people would also probably own a computer or laptop. Thirty years ago, as hard as it is to believe, this was not the case. These devices are made with rare-earth elements along with many other elements. Additionally, the demand for longer-lasting portable electronic devices continues to increase. Car manufacturers are now well-invested in developing electric and hybrid-electric vehicles, as other issues around climate change drive this development. This will increase the demand for these rare-earth elements even further.

Uses

Rare-earth elements are used as components in recent high-tech devices, such as smart phones, digital cameras, LED televisions, and computer hard disks and monitors. More specific examples for the uses of each of these elements is shown in Table 1D–1.

Table 1D–1 Rare-earth elements and their uses

Rare-earth metal	Applications and uses
Yttrium	Ceramics and metal alloys
Lanthanum	Batteries and catalysts for refining petrol
Cerium	Chemical catalysts, glass polishing and metal alloys
Praseodymium	High-powered magnets
Neodymium	High-powered magnets
Promethium	Beta radiation source
Samarium	High-temperature magnets
Europium	Phosphors for lighting
Gadolinium	Magnetic resonance imaging (MRI)
Terbium	High-powered and high-temperature magnets
Dysprosium	High-powered and high-temperature magnets and lasers
Holmium	High-powered magnets
Erbium	Lasers and glass colourants
Thulium	Ceramic magnets (under development)
Ytterbium	Fibre optic technology and solar panels
Lutetium	PET (positron emission tomography) scanners

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Recycling and measures to limit wastage

In general, mining is a destructive process that can lead to environmental disruption, and this may become a complex issue when it impacts First Nations Australians' communities. Chemical compounds that are used to support the extraction of elements or compounds often enter the atmosphere or groundwater near mines. The processing through crushing and grinding these compounds releases dust that can have severe respiratory effects on workers and residents living in close proximity to mines.

Some deposits of rare-earth elements contain thorium and uranium, which are radioactive. For this reason, the rare-earth elements are often not recoverable and left in the residue of **ores**.

Ore
a deposit within Earth's crust that contains one or more valuable minerals, typically with a high concentration of a certain element (particularly metals)

Figure 1D–8 This image highlights the damage caused by rare-earth element mining in China.

Currently less than 1% of electronic devices are recycled. Recycling processes can result in reduced impacts. However, with the increasing demand for these elements, along with the rising mining standards outlined in many countries, production has shifted to developing countries with lower costs and relaxed environmental regulations. This has consequently resulted in greater impacts on society.

Other elements or substitutes could be used, but they are often not as effective and are very expensive.

Check-in questions – Set 1

1 Analyse and explain whether the terms 'essential element' and 'critical element' are the same.
2 Explain what happens to most of the phosphorus that is mined, and state two measures to recycle it.
3 Explain why rare-earth elements are considered to be critical elements.

Post-transition metals

Post-transition metals are the group of elements found between the right-hand side of the transition metals and the left-hand side of the metalloids, as shown in Figure 1D–9.

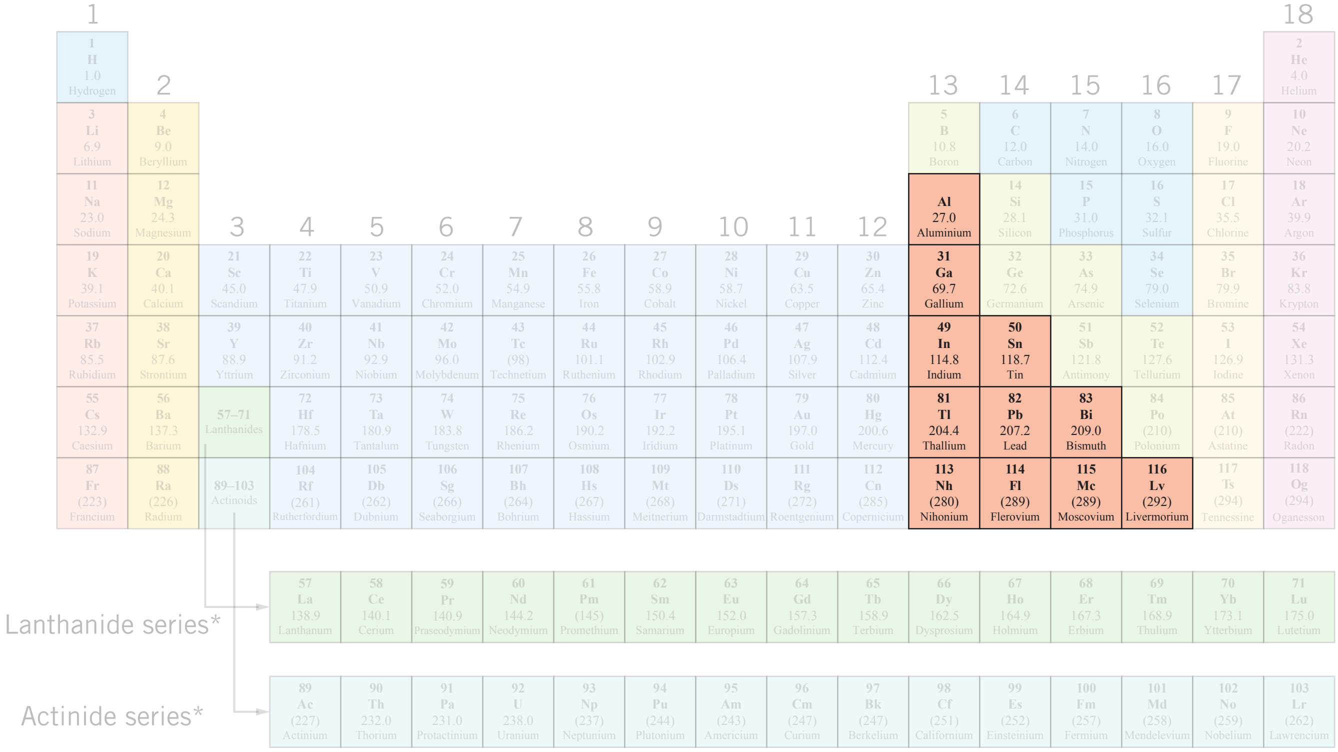


Figure 1D–9 The periodic table showing the location and elements of the post-transition metals

These metals have the physical and chemical properties listed in Table 1D–2.

Table 1D–2 Physical and chemical properties of the post-transition metals

Physical properties	Chemical properties
Soft or brittle High melting and boiling points (but lower than transition metals) High density	Ionic salts include PbI_2, PbS, PbF_2, $SnCl_2$ Covalent bonding is common between transition metals and non-metals

They tend to show less metallic nature than transition metals due to generally having smaller atomic radii with higher ionisation energies.

LINK **1C** THE PERIODIC TABLE

Uses

Post-transition metals have a variety of different uses as outlined in Table 1D–3.

Aluminium is the third-most abundant element on Earth and as such is the most abundant of these post-transition elements.

Table 1D–3 Some applications and uses of post-transition metals

Post-transition metal	Application and uses
Aluminium, Tin	Utensils, electronics and soldering
Bismuth	A key element in making some medications that soothe upset stomachs, but in Australia bismuth compounds are Schedule 4 poisons
Indium	Touch screens for electronic devices
Gallium	Semiconductors and fuel cells
Lead	Batteries

Mining, extraction and processing

A number of these post-transition metals are often produced as a by-product during the mining of other metals. For instance, the primary production of indium is from zinc mining.

Bolivia is notable as the only country in the world that mines bismuth as a primary metal, as opposed to sourcing it as a by-product.

4A FORMATION AND NAMING OF IONIC COMPOUNDS LINK

Some post-transition metals, such as lead and tin, have been in use for thousands of years. Usage of tin dates back to the Bronze Age (around 3000 BCE) when it was extracted from the ore cassiterite (SnO_2). Cassiterite remains an important source of tin today. Lead has been mined for at least 8000 years, with signs of its use in ancient Egypt and Turkey. Typically, lead was extracted from the lead sulfide, galena (PbS). Galena was applied around the eyes as a cosmetic that repelled flies and reduced the glare of the sun.

Although not a true critical element at this stage, tin is a metal that is found commonly in ores combined with high levels of sulfide (sulfur ion). The ore is first heated to 500–600°C to burn off the sulfur. This process has become more common in recent years as it has improved the amount of tin recovered from ores and from processed waste material containing tin.

Smelting
the process of heating and melting an ore to extract a metal

Slag
a silicon dioxide and metal oxide mixture that is left over as a by-product from smelting

The tin is then extracted by **smelting**, a process shown in Figure 1D–10. In this process, the tin concentrate is placed in a furnace with a source of carbon, usually in the form of coal. It is then heated to form crude tin and **slag**. This crude tin and slag mixture is then heated several times to continue to remove any further impurities as well as extract more tin from the slag, as it usually contains around 10–25% tin.

Before the tin is ready for market sale, it often undergoes further refining processes, commonly either heat treatment or electrolysis. Heat treatment is more common, firstly because it is cheaper but secondly because there is not a demand for highly pure tin, so the few impurities that remain are acceptable.

This method of metal extraction is very similar for many metal ores that are mined today.

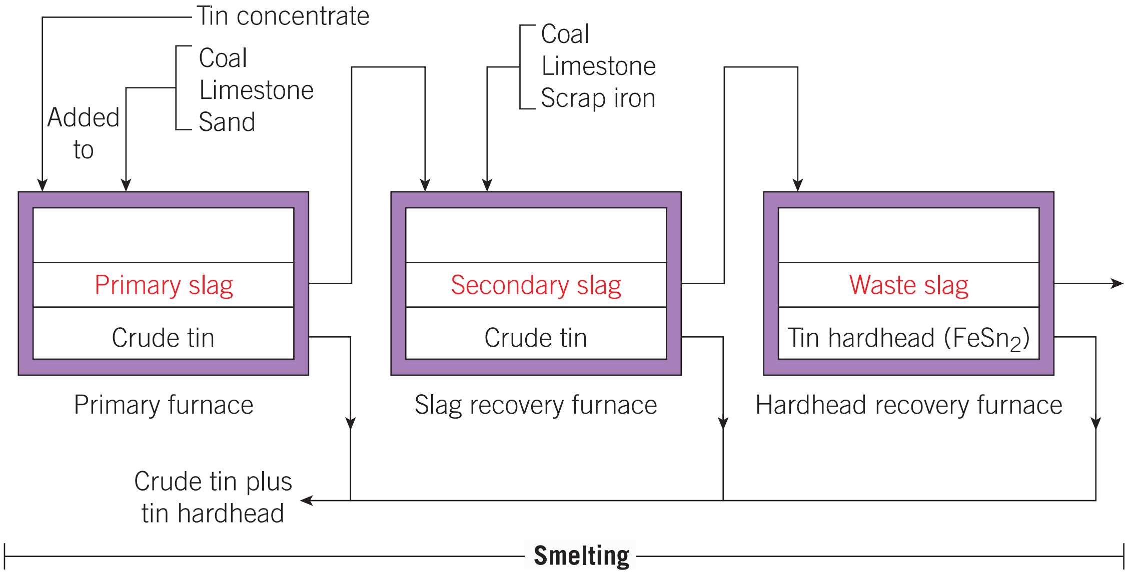


Figure 1D–10 The smelting process for obtaining tin from ores and tin concentrate

Recycling

The recycling efficiency of these post-transition metals varies greatly. Tin, for example, can be recycled to the same high quality due to its unique properties and economic value, as it is used in the manufacture of solder, cans and batteries. The recovery of tin from electronic products will be the largest source of recycled tin in our future.

Lead is another highly recyclable metal – in fact, it is 100% recyclable. It doesn't lose any of its properties in this process. Lead is mainly recycled from batteries, as 90% of scrap lead in Australia comes from lead–acid batteries in vehicles. Recycling lead uses approximately 35–40% less energy than that used in the process of obtaining primary lead.

Aluminium is similarly recyclable – approximately 75% of all aluminium ever produced is still in use today!

The recycling processes for lead and tin, like most metals, is through smelting, as shown in Figure 1D–10. However, bismuth is typically difficult to recycle, as it is used in pharmaceuticals, chemicals and paints. These are products that are either consumed or dissipated soon after usage, making recycling incredibly difficult. A global industry has emerged focusing on recovering post-transition metals that are mainly obtained as by-products, particularly as elements such as indium and gallium are used in solar and nuclear energy production. However, as mentioned earlier, bismuth can be recovered as a by-product from ores of other metals, and this method accounts for a large proportion of global bismuth production.

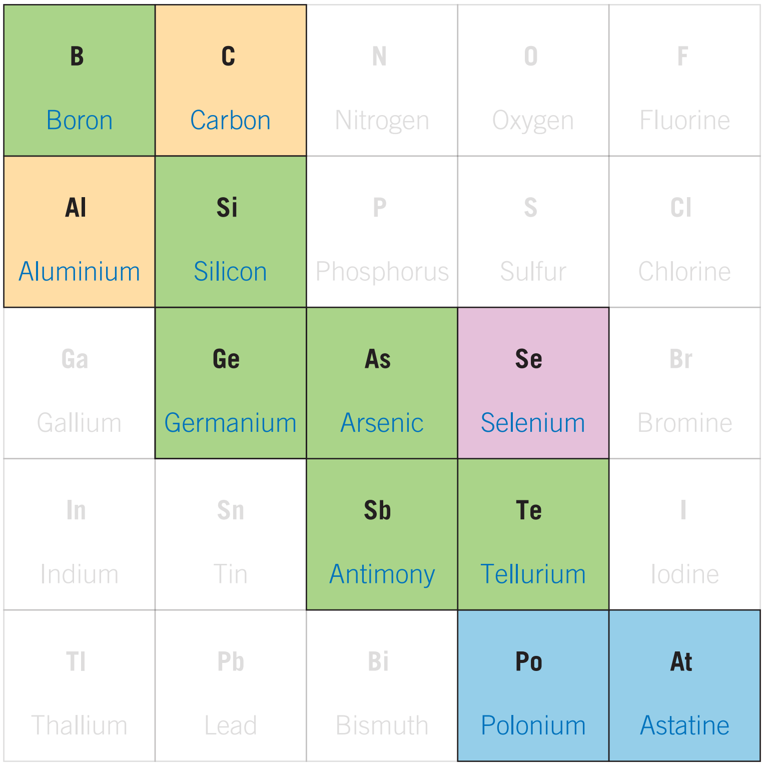
Check-in questions – Set 2

1 Outline the process of extraction of tin from its ores.
2 State which post-transition metal has the highest potential for recycling, and where most of the recycled lead comes from in Australia.
3 State the uses of gallium and indium and explain if they could be considered critical elements in the development of new technologies.

Metalloids

1A INTRODUCTION TO THE ELEMENTS LINK

You will recall from Chapter 1A that most of the elements on the periodic table can be classified as either metals or non-metals. However, there is a small group of elements, shown in Figure 1D–11, whose properties do not fit solely into one of these groups. They are classified as metalloids.



B Boron	**C** Carbon	N Nitrogen	O Oxygen	F Fluorine
Al Aluminium	**Si** Silicon	P Phosphorus	S Sulfur	Cl Chlorine
Ga Gallium	**Ge** Germanium	**As** Arsenic	**Se** Selenium	Br Bromine
In Indium	Sn Tin	**Sb** Antimony	**Te** Tellurium	I Iodine
Tl Thallium	Pb Lead	Bi Bismuth	**Po** Polonium	**At** Astatine

Figure 1D–11 Elements highlighted by how frequently they are classified as metalloids: green (most commonly), blue (irregularly), pink (less commonly), orange (rarely).

Metalloids appear physically similar to metals, as shown in Figure 1D–12, but chemically behave like non-metals in reactions. Table 1D–4 lists the key physical and chemical properties of these elements.

Metalloids currently considered critical include germanium, arsenic and tellurium.

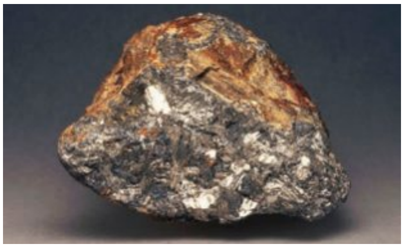


Figure 1D–12 A sample of antimony, which appears similar to a metal but behaves like a non-metal in chemical reactions

Table 1D–4 Chemical and physical properties of metalloids

Physical properties	Chemical properties
Lustrous and appear like metals	Easily form anions by gaining electrons (like non-metals), except for boron, which has three valence electrons and loses these
Solid	Multiple oxidation states
Brittle and shatter easily	Electronegativity and ionisation energies between metals and non-metals
Conduct electricity	

2A REPRESENTING COVALENT COMPOUNDS LINK

12A REDUCTION AND OXIDATION LINK

4A FORMATION AND NAMING OF IONIC COMPOUNDS LINK

Uses

Table 1D–5 highlights some of the key facts about different metalloids.

Table 1D–5 Key features and uses of the metalloids

Metalloid	Features	Uses
Boron	Forms covalent bonds, chemically behaving as a non-metal	Fuel ignition Medical applications in antiseptics and eye drops
Silicon	Brittle, crystalline solid appearing like a metal Commonly found in asteroids, moons and cosmic dust	Semiconductor in appliances, such as smart phones, televisions and general household appliances, solar panels Used in manufacturing of cement, porcelain and ceramics
Germanium	Hard and brittle Often crystallises into a diamond-like structure	Semiconductor Used in applications involving infrared optics, solar energy and metal alloys
Arsenic	Shiny grey solid appearing like a metal but forms covalent bonds with non-metals Highly toxic and carcinogenic to humans, so use has been phased out over time	Present in alloys, electronics and pesticides
Antimony	Lustrous silvery-white brittle structure that forms covalent bonds Highly reactive with halogens	Strengthens metal alloys, so useful in items such as car batteries, ammunition, cables and plumbing supplies
Tellurium	Often found in the mineral structure of other metals (e.g. gold tellurides in Western Australia)	A steel additive that can be alloyed with aluminium, copper, lead or tin to reduce corrosion Semiconductor when exposed to light

LINK **2D** DIAMOND AND GRAPHITE AS COVALENT COMPOUNDS

LINK **2A** REPRESENTING COVALENT COMPOUNDS

Mining, extraction and processing

Many metalloids, like the metals discussed in this chapter, are found naturally as oxides (element combined with oxygen) or sulfides (element present as an ionic compound with sulfur). As such, they are mined and processed in a very similar way to that explained earlier for the post-transition metals. You will learn more about ionic compounds in Chapter 4A.

LINK **4A** FORMATION AND NAMING OF IONIC COMPOUNDS

Recycling

Most germanium is either recycled or re-used in items such as optical fibre, night vision devices as broken lenses and glass, in plastic polymers such as used food containers, and in semiconductor solar cells as trimmings from substrates. A germanium recycling efficiency of about 76% was estimated to have been reached in 2000.

New optical fibres contain about 60% recycled material and 40% new material. Recovery of as much as 80% of the germanium in material sent back as scrap has been shown to be commercially feasible. However, since 2001, recycling has not been a major consideration due to a decline in optical fibre production, along with the cost of recycled material being at or above that of new material.

LINK **4C** PRECIPITATION REACTIONS

Alternatively, a process was developed in 1985 for the recovery of germanium from effluents generated in the production of optical fibres. A germanium filter cake was produced through a process that included a gas-scrubbing unit and a recirculation unit followed by precipitation and filtration. This process was reported to recover greater than 95% of the unreacted germanium in the effluent stream.

Importance of recycling going forward

As an example of society's use of elements, one of the largest global technology companies, General Electric, use 72 of the first 82 elements on the periodic table in their product lines, many of which, as you have now seen, are critical or difficult to obtain. Up to 70 of these elements can be found in a mobile phone.

For many of these critical elements, the recycling rates are very low. Outside of coming up with a better recycling program, full-scale renewable energy systems must focus on utilising elements that are less scarce as substitutes to what we currently use today. Even for many elements that seem currently abundant, such as helium or aluminium, limitations in their production and supply rates will only be experienced in decades, rather than now. However, scientists, companies and organisations realise they cannot leave this to become the problem of our future generations, so they must start taking action now.

In 2017, Apple made a pledge to create its smartphones from only recycled material. In 2021, the Tokyo Olympics organising committee manufactured gold, silver and bronze medals for which the elements were completely recycled from electronic devices. The same was done for medals awarded at the 2012 London and 2016 Rio de Janeiro Olympic Games.

It is worth noting that while we can creatively improve our recycling and re-use of critical elements, it is not always economically feasible. For many companies, this requires a major overhaul of their processing and production. A systems-thinking approach is required that considers not only science and technology, but economics and environmental impact as well. For example, in Australia, the expansion of mining for mineral ores has equated to a huge economic growth but has resulted in more consequential impacts on our environment and land use. The **circular economy** for recycling resources is explored further in Chapter 3.

Circular economy
an economy that functions on a continuous cycle with a focus on optimising the re-use of resources and materials

Check-in questions – Set 3

1 Outline the main similarities and differences between metalloids on the one hand and metals and non-metals on the other.

1D SKILLS

Environmental, economic and social issues around the mining of metals

Recycling is the treatment and processing of materials so they can be used again as new products, as opposed to throwing them away as waste. In your chemistry course, sustainability is considered from three perspectives:

- Sustainable development
- Green chemistry principles
- Movement from a linear economy to a circular economy.

This section and the recycling methods, particularly for critical elements, looks at the movement from a linear economy to a circular economy (Figure 1D–13) and the green chemistry principles behind this.

3C METAL RECYCLING LINK

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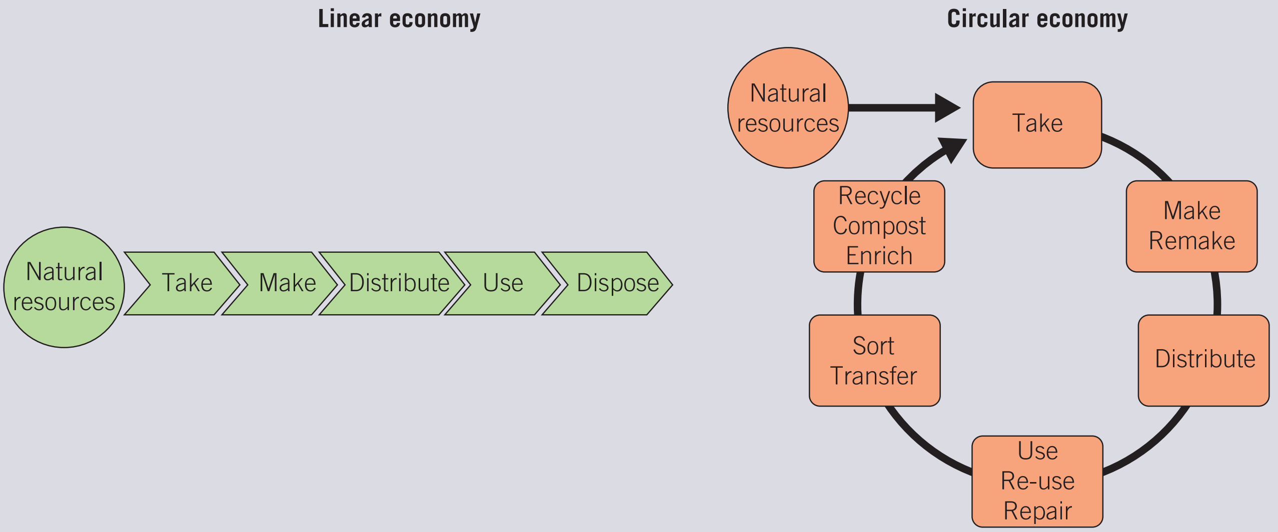

Figure 1D–13 Comparison of linear and circular economies

Additionally, recycling has many other benefits including:
- conserving natural resources
- reducing waste in landfills
- preventing pollution, therefore protecting the environment
- saving energy
- stimulating development of greener technologies
- reducing the cost of mining for raw ingredients
- supporting a community or country's economic growth
- supporting manufacturing and job opportunities
- encouraging the interaction between industry and research to develop and disseminate knowledge.

As global citizens, we need to consider how our decisions can have both positive and negative effects on our environment, economies and societies.

Table 1D–6 considers the general mining, extraction and processing looked at in this chapter and some of the implications that they have for the environment, the economy and society.

Table 1D–6 Some positive and negative environmental, economic and social implications

Environmental	Economic	Social
Loss of landscape due to mining, processing and transportation of materials	Financial benefits of mining, extraction and processing for countries, such as Australia	Land use conflicts between mining companies, agriculture and tourism providers
Pollution from extraction processes	Discrepancies in wages paid to miners compared with manufacturers or those who work in local service industries	Negotiations with Indigenous groups to ensure fair access to the resources from the land
Disposal of waste material	Cost of waste disposal (ask your school's lab technician about how much it costs to remove the different types of chemistry waste off-site)	Mining, extraction and processing are often in remote locations and the impact this has on workers being away from families
Noise pollution	Noise pollution and as a result the provision of hearing tests and ear protection	Job security for those living in remote communities where these processes occur

As part of assessments, you may be expected to analyse media articles and compare the different implications of recycling different critical elements. An example can be found in Activity 1D–1 in the Interactive Textbook.

Section 1D questions

1 Define 'critical element'.

2 List three elements in each of the following classifications.

Lanthanides	
Post-transition metals	
Metalloids	

3 State one use for each of the following critical elements.

	Critical element	Use
a	He	
b	P	
c	Eu	
d	Nd	

4 Lead is a post-transition metal that is used in lead–acid batteries, a typical component of automotive vehicles. State a physical property of lead that is required to allow its use for this application.

5 Explain the challenges associated with recycling helium, with reference to Earth's mantle degassing.

6 Mining and extraction of rare-earth elements has increased substantially since the 1950s due to demand for their use in a variety of applications. Describe the applications that require rare-earth elements in the twenty-first century.

7 Suggest one strategy to improve the recycling of rare-earth elements after their use.

8 Describe how phosphorus-containing rocks are formed over millions of years.

9 Compare the general physical properties of the post-transition metals and metalloids.

10 Tin and aluminium are post-transition metals that can be recycled many times due to their robustness. Describe the typical applications of these elements and the steps required in the physical recycling of tin and aluminium.

11 Explain whether germanium is a critical element with reference to its uses and supply. Use research to find your answer.

Chapter 1 review

Summary

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

Success criteria – I am now able to:		Linked question
1A.1	Recall what an element, isotope and ion are, and how they are different	11a☐
1A.2	Be able to write correct atomic notation for different elements, isotopes and ions	4☐
1A.3	Recall the difference between the atomic number and mass number of an element	11b☐, c☐
1A.4	Determine the atomic number and/or mass number by looking at the atomic notation of an element	11c☐
1A.5	Recall what an electron, proton and neutron are and the differences between them	1☐, 11e☐
1A.6	Determine the number of protons, neutrons and/or electrons for an element, isotope or ion from atomic notation	2☐, 3☐
1B.1	Recall Bohr's model for writing electron configurations	5☐, 8☐, 11d☐
1B.2	Recall the difference between shells, subshells and orbitals	11d☐
1B.3	Recall the types of subshells and number of orbitals in each	12c☐
1B.4	Understand the limitations of Bohr's model for electron configurations	11f☐
1B.5	Use the position of an element on the periodic table to determine its electron configuration	11d☐
1B.6	Write the electron configuration for both atoms and ions	12b☐
1C.1	Recall the different areas and overall structure of the periodic table (blocks, groups and periods)	12a☐
1C.2	Understand and be able to calculate the core charge of an atom or ion	6☐
1C.3	Explain how core charge and the number of electron shells relates to the attraction of the valence electrons to the nucleus	12d☐
1C.4	Explain the difference in atomic radius when comparing elements down a group and across the same period	7☐
1C.5	Recall the definitions of electronegativity, first ionisation energy, reactivity, and metallic and non-metallic character	13a☐, g☐
1C.6	Use the core charge and atomic radius of an element to explain trends in electronegativity, first ionisation energy, reactivity, and metallic and non-metallic character both down a group and across a period	3g☐, 8☐, 9☐, 13b☐, 13c☐, 13d☐

	Success criteria – I am now able to:	Linked question
1D.1	Explain examples of critical elements, including helium, phosphorus, rare-earth elements, post-transition metals and metalloids, and why they are classified as such	10☐, 14☐
1D.2	Recall examples of the mining, extraction and processing of critical elements	14☐
1D.3	Explain the different uses of critical elements	10☐, 14☐
1D.4	Outline examples of current recycling processes for recovering critical elements	14☐
1D.5	Explain the importance of element recovery with reference to environmental, economic and/or social implications	14☐

Multiple-choice questions

1 Which of the following statements about the proton is **false**?
- **A** Protons have a positive charge.
- **B** Protons are located in shells around the nucleus of an atom.
- **C** Protons have a mass of one atomic mass unit (amu).
- **D** The mass of a proton is very similar to the mass of a neutron.

2 How many neutrons does the atom $^{25}_{12}\mathrm{Mg}^{2+}$ contain?
- **A** 10
- **B** 12
- **C** 13
- **D** 25

3 How many electrons does the atom $^{25}_{12}\mathrm{Mg}^{2+}$ contain?
- **A** 10
- **B** 12
- **C** 13
- **D** 25

4 An atom has 9 protons, 10 neutrons and 10 electrons. What is the correct atomic notation for this atom?
- **A** $^{19}_{9}\mathrm{F}^{-}$
- **B** $^{20}_{9}\mathrm{F}^{-}$
- **C** $^{19}_{10}\mathrm{Ne}$
- **D** $^{20}_{10}\mathrm{Ne}$

5 What is the subshell electron configuration for a titanium atom in its ground state?
- **A** $1s^2 2p^8 3d^{12}$
- **B** $1s^2 2s^2 2p^6 3s^2 3p^6 3d^4$
- **C** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$
- **D** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4p^2$

6 What is the core charge of a titanium atom?

- **A** 0
- **B** +2
- **C** +4
- **D** +12

7 Which of the following elements has a smaller atomic radius than chlorine?

- **A** fluorine
- **B** sodium
- **C** sulfur
- **D** bromine

8 Moving across a period of the periodic table, which trend tends to decrease?

- **A** electronegativity
- **B** first ionisation energy
- **C** non-metallic character
- **D** metallic character

9 Moving down a group of the periodic table, which trend tends to increase?

- **A** electronegativity
- **B** first ionisation energy
- **C** non-metallic character
- **D** atomic radius

10 Which of these is **not** an example of a critical element?

- **A** helium
- **B** carbon
- **C** phosphorus
- **D** indium

Short-answer questions

11 A carbon atom can contain six, seven or eight neutrons.

- **a** Is this an example of different elements, different isotopes or different ions? Justify your answer. (4 marks)
- **b** What is the difference between the atomic number and the mass number of an element? (2 marks)
- **c** Write the atomic notation for each of the carbon atoms with differing numbers of neutrons. (3 marks)
- **d** Write both the shell and subshell electron configurations for a carbon atom containing six neutrons. (2 marks)
- **e** Does the electron configuration change depending on the number of neutrons in an atom? Explain. (1 mark)
- **f** State one limitation of the Bohr model of the atom. (1 mark)

12 The periodic table contains all the currently known elements.

- **a** Explain the structure of the periodic table, with reference to the electron configuration of each element's last subshell. A diagram may be useful. (4 marks)
- **b** Write the subshell electron configuration for phosphorus and for a phosphide ion. (2 marks)
- **c** How many orbitals are occupied by electrons in the 3p subshell of the phosphorus atom. Explain your answer by referring to how many electrons each orbital can hold. (1 mark)
- **d** Explain why phosphorus has a core charge of +5. (2 marks)

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13 In this chapter we explored some of the trends of the periodic table.

a Define each of the following properties and outline how it depends on the core charge and atomic radius of the element.

i electronegativity (3 marks)

ii first ionisation energy (3 marks)

iii metallic character (3 marks)

b For each of the following pairs of elements, state the element with the greatest electronegativity.

i Li and K (1 mark)

ii Cl and F (1 mark)

c Explain why an atom of caesium is larger than an atom of sodium. (3 marks)

d Using your answer to part **c** above, determine whether caesium will have a higher or lower first ionisation energy compared with sodium. (1 mark)

14 Explain why helium is a critical element, even though it is one of the most abundant elements in the universe. (4 marks)

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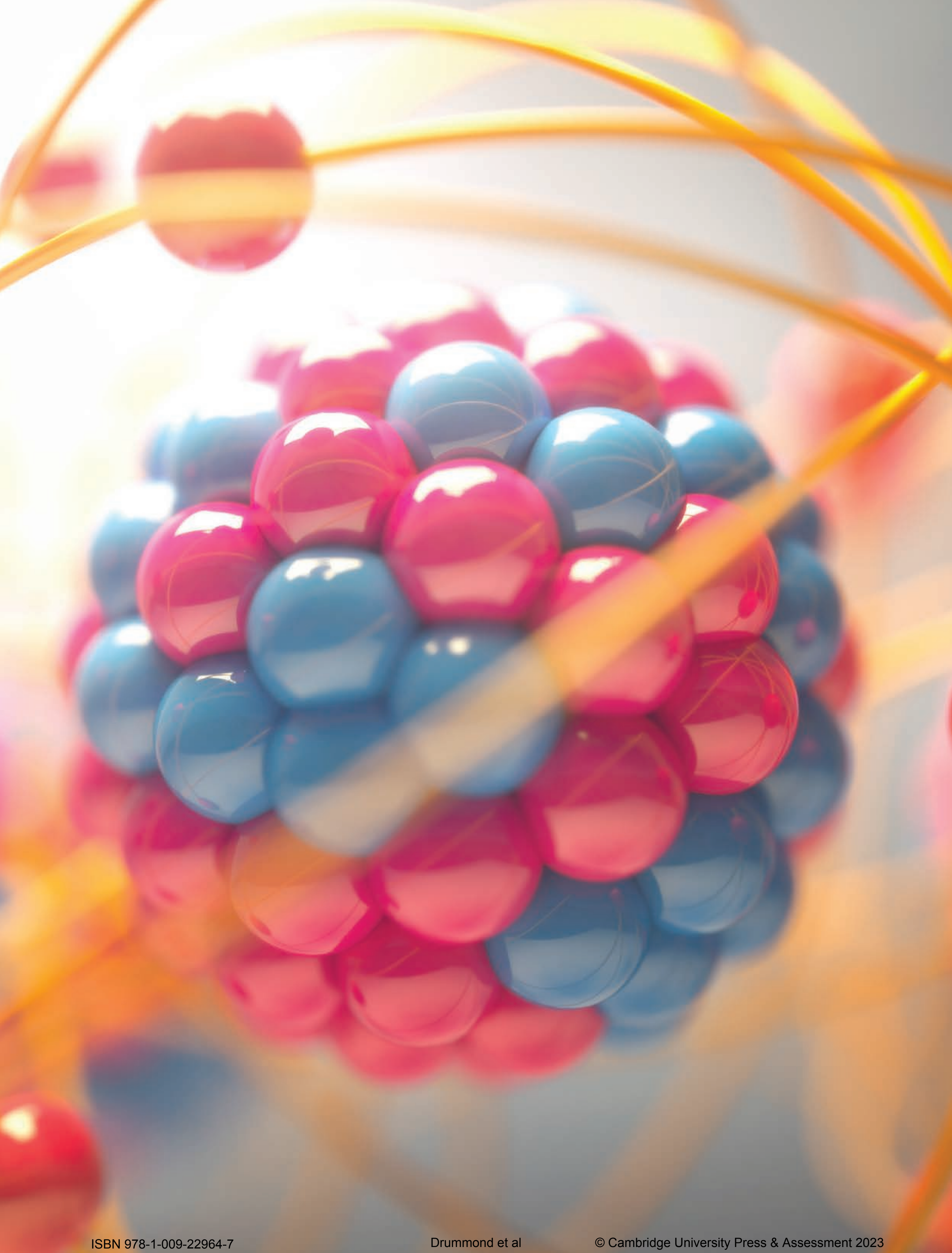

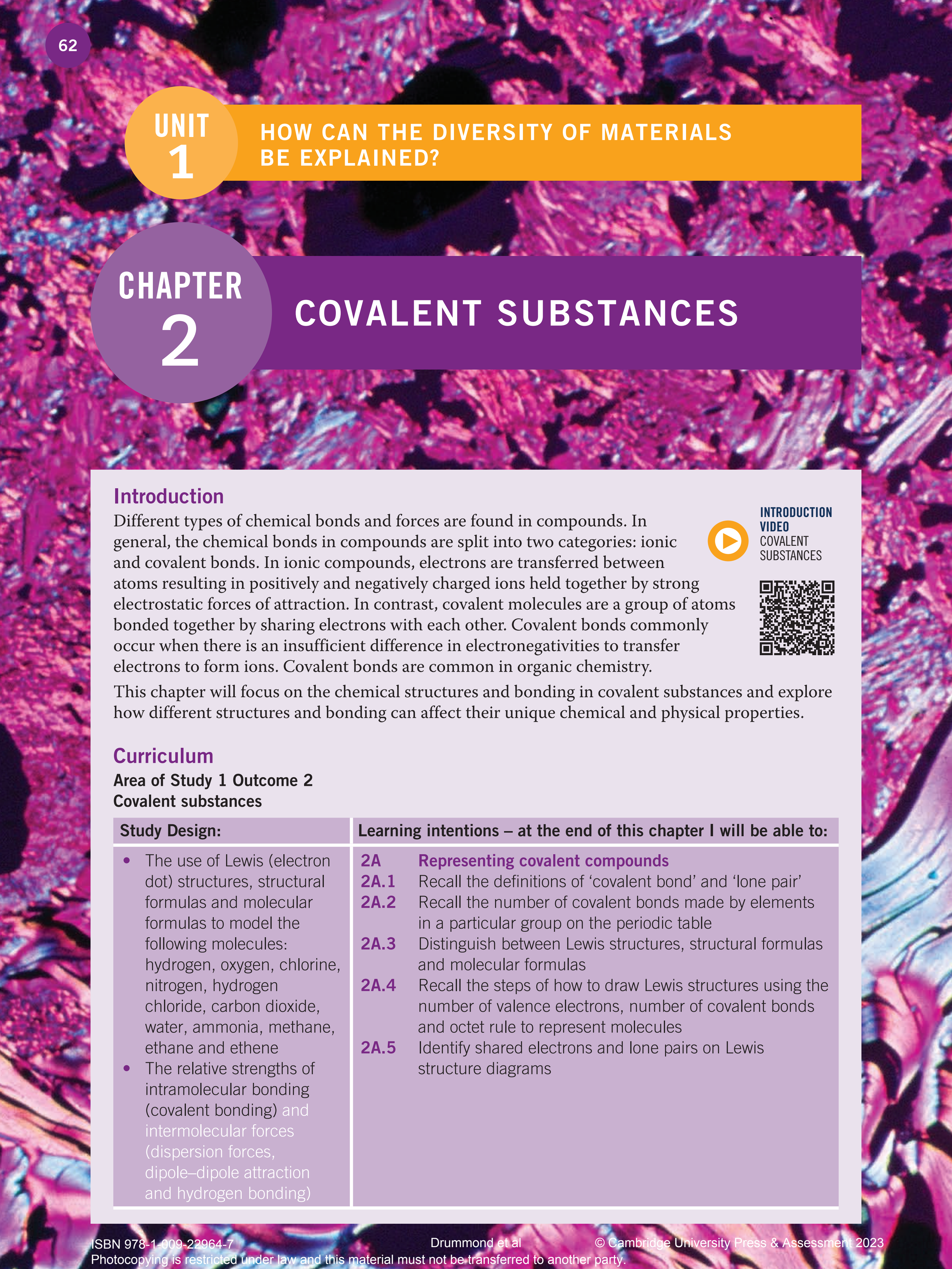

UNIT 1 HOW CAN THE DIVERSITY OF MATERIALS BE EXPLAINED?

CHAPTER 2

COVALENT SUBSTANCES

Introduction

Different types of chemical bonds and forces are found in compounds. In general, the chemical bonds in compounds are split into two categories: ionic and covalent bonds. In ionic compounds, electrons are transferred between atoms resulting in positively and negatively charged ions held together by strong electrostatic forces of attraction. In contrast, covalent molecules are a group of atoms bonded together by sharing electrons with each other. Covalent bonds commonly occur when there is an insufficient difference in electronegativities to transfer electrons to form ions. Covalent bonds are common in organic chemistry.

INTRODUCTION VIDEO
COVALENT SUBSTANCES

This chapter will focus on the chemical structures and bonding in covalent substances and explore how different structures and bonding can affect their unique chemical and physical properties.

Curriculum

Area of Study 1 Outcome 2
Covalent substances

Study Design:	Learning intentions – at the end of this chapter I will be able to:
• 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 • The relative strengths of intramolecular bonding (covalent bonding) and intermolecular forces (dispersion forces, dipole–dipole attraction and hydrogen bonding)	**2A Representing covalent compounds** **2A.1** Recall the definitions of ‘covalent bond’ and ‘lone pair’ **2A.2** Recall the number of covalent bonds made by elements in a particular group on the periodic table **2A.3** Distinguish between Lewis structures, structural formulas and molecular formulas **2A.4** Recall the steps of how to draw Lewis structures using the number of valence electrons, number of covalent bonds and octet rule to represent molecules **2A.5** Identify shared electrons and lone pairs on Lewis structure diagrams

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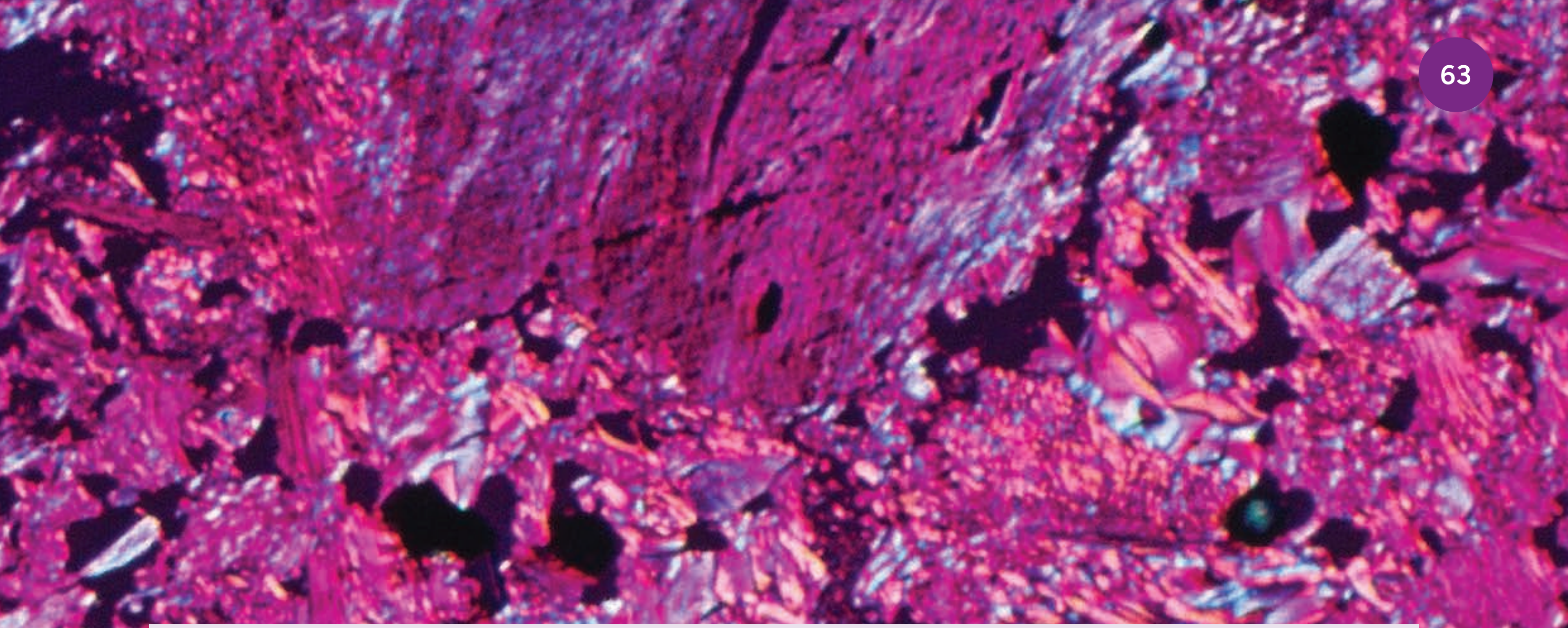

Study Design:	Learning intentions – at the end of this chapter I will be able to:
- Shapes of molecules (linear, bent, pyramidal and tetrahedral, excluding bond angles) as determined by the repulsion of electron pairs according to valence shell electron pair repulsion (VSEPR) theory - Polar and non-polar character with reference to the shape of a molecule	**2B Molecular shapes and polarity**
	2B.1 Define ‘electron domain’
	2B.2 Recall the four shapes of molecules according to the valence shell electron pair repulsion (VSEPR) theory
	2B.3 Determine two factors that affect the shape of molecules
	2B.4 Draw the shape of molecules according to the VSEPR theory and display arrows (direction) of dipoles, including partial charges where required
	2B.5 Explain the differences between non-polar and polar molecules in relation to the arrangement of atoms in molecules
	2B.6 Determine the polarity of molecules based on the polarity of the covalent bonds and molecular geometry
- 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	**2C Properties of covalent compounds determined by intermolecular forces**
	2C.1 Describe the differences between intramolecular and intermolecular bonding (forces)
	2C.2 Recall the three types of intermolecular forces
	2C.3 Determine which intermolecular forces exist between different covalent molecules based on the polarity of the covalent bonds and molecular geometry
	2C.4 Explain the relationship between the strength of intermolecular forces and a molecule's melting and boiling points
	2C.5 Rationalise the non-conduction of electricity by covalent molecules
- The structure and bonding of diamond and graphite that explain their properties (including heat conductivity, electrical conductivity and hardness) and their suitability for diverse applications	**2D Diamond and graphite as covalent compounds**
	2D.1 Understand and illustrate the crystal structure and bonding of both diamond and graphite
	2D.2 Explain the difference in physical properties of diamond and graphite: appearance, heat conductivity, electrical conductivity and hardness
	2D.3 Recall the applications of both diamond and graphite in society

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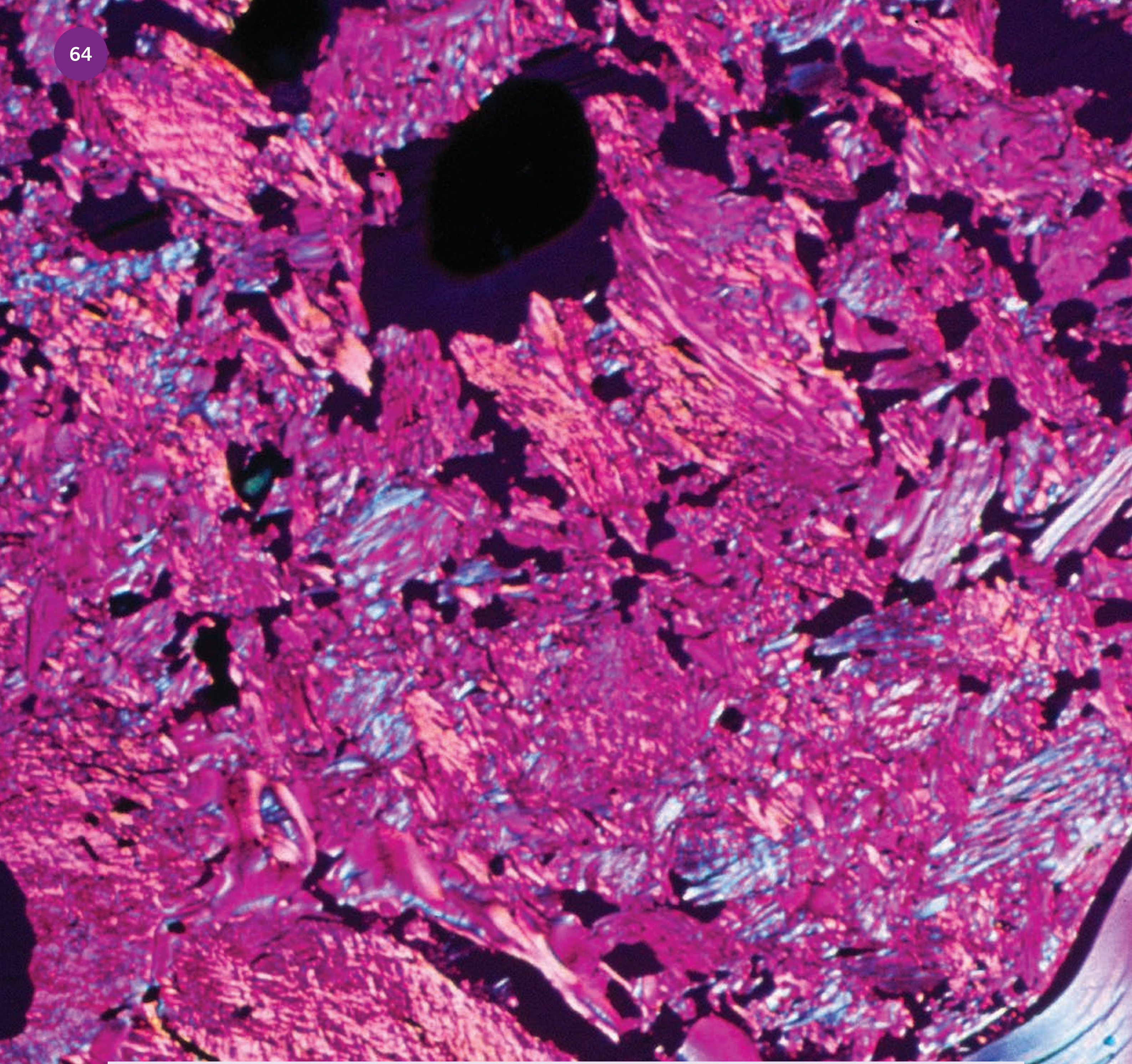
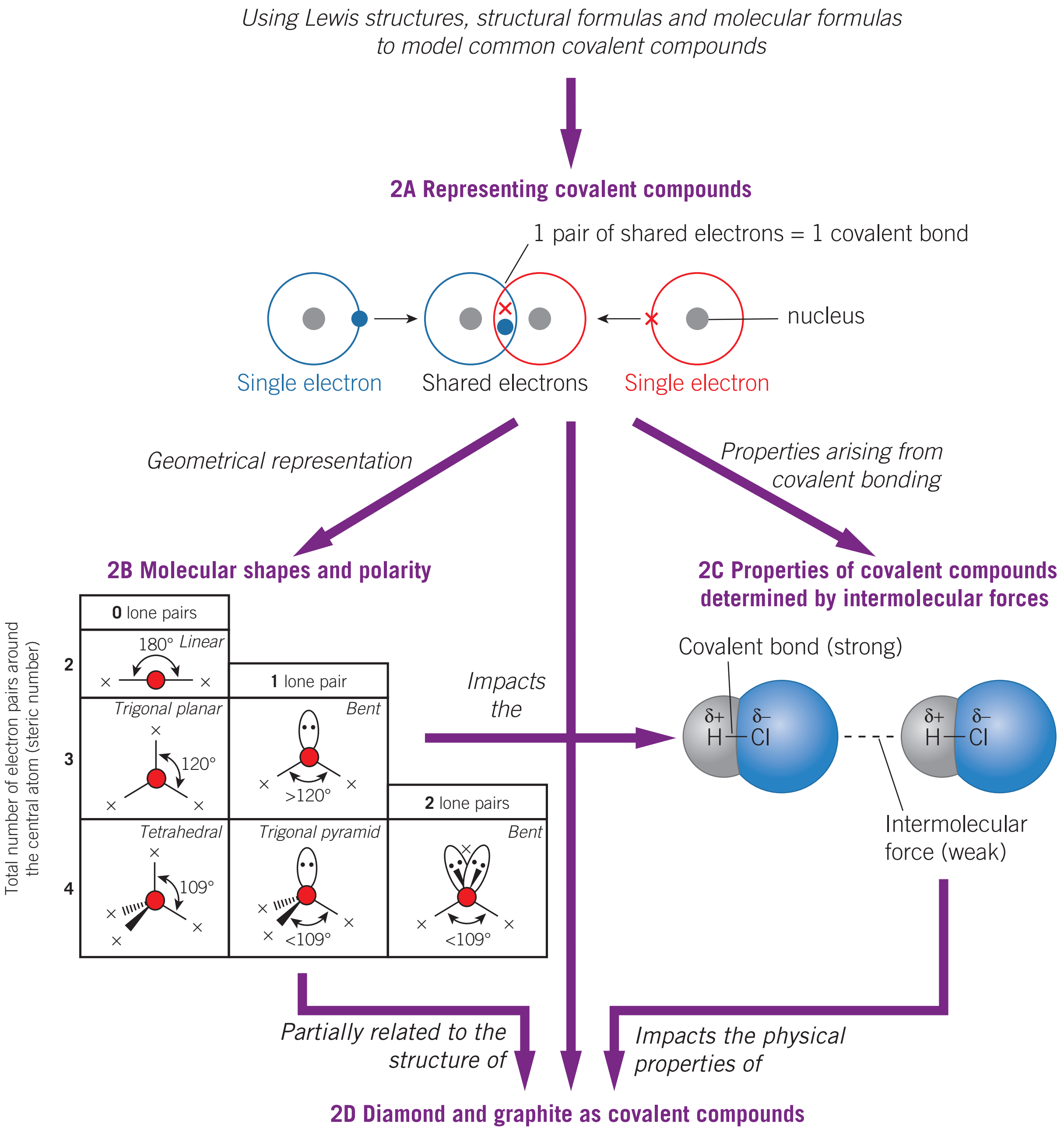


Glossary

- Allotrope
- Bond length
- Bond strength
- Covalent bond
- Dipole
- Dipole–dipole attraction
- Dispersion force
- Electron domain
- Electronegativity
- Hydrogen bonding
- Instantaneous dipole
- Intermolecular force
- Lewis structure
- Molecular formula
- Non-polar
- Octet rule
- Polar
- Polarity
- Structural formula
- Tetrahedral structure
- Valence electron
- Valence shell electron pair repulsion (VSEPR) theory

Concept map



See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

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2A Representing covalent compounds

Study Design:

- 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
- The relative strengths of intramolecular bonding (covalent bonding) and intermolecular forces (dispersion forces, dipole–dipole attraction and hydrogen bonding)

Glossary:

Bond length
Bond strength
Covalent bond
Electronegativity
Lewis structure
Molecular formula
Octet rule
Structural formula
Valence electron

ENGAGE

Lewis's legacy

Gilbert Newton Lewis was an American chemist best known for his work on covalent bonding and electron pairs. In the early 1900s, chemists realised that the electrons orbiting the nucleus had a special arrangement. Lewis's work focused on developing a theory based on the electron arrangement on an atom's outermost or 'valence' shell.

In 1902, Lewis began explaining the concept of valence electrons to his students by depicting atoms constructed from concentric series of cubes with electrons located at each corner, as shown in Figure 2A–1. The so-called 'cubical atom' expressed Lewis's idea that chemical bonds are formed from electron transfer to provide each atom with a complete set of eight outer electrons (an octet). In 1916, Lewis published his now-classic paper on chemical bonding, which suggested that a chemical bond is a pair of electrons shared by two atoms. Lewis was nominated for the Nobel Prize in Chemistry 41 times to recognise his significant work. Controversially, Lewis was never awarded the Nobel Prize in Chemistry. Nonetheless, Lewis acted as a mentor for numerous Nobel laureates at the University of California, Berkeley.

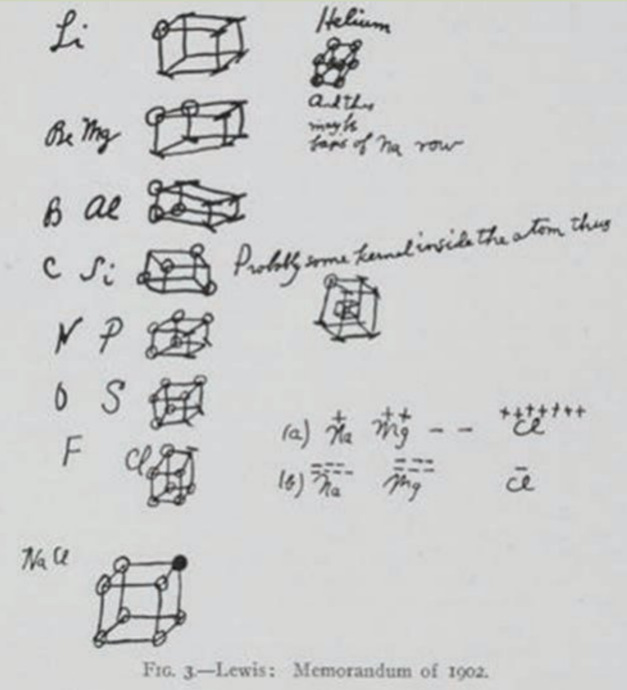


Figure 2A–1 Gilbert Lewis's 'cubical atoms' drawing from 1902 expressed his idea that atoms were composed of a concentric series of cubes with electrons at each corner.

EXPLAIN

Representing the structure of covalent compounds

A covalent compound is a molecule formed by covalent bonds, sharing one or more pairs of **valence electrons** between atoms. The sharing of electrons allows atoms to achieve the equivalent of a full valence shell, resulting in a stable electron configuration. **Covalent bonds** are formed when the difference of **electronegativities** between the atoms is too low to attract electrons from, or to donate electrons to, another atom to form ions. Shared electrons between atoms are called bonding electrons. Each pair of shared electrons (bonding electrons) equals one covalent bond, as represented in Figure 2A–2. In general, the bonding between the atoms of non-metallic elements is covalent.

Covalent compounds are represented by molecular formulas, structural formulas, and Lewis structures. This section will explore using these representations to model various covalent molecules and the relative strength of their intramolecular bonding (covalent bonding).

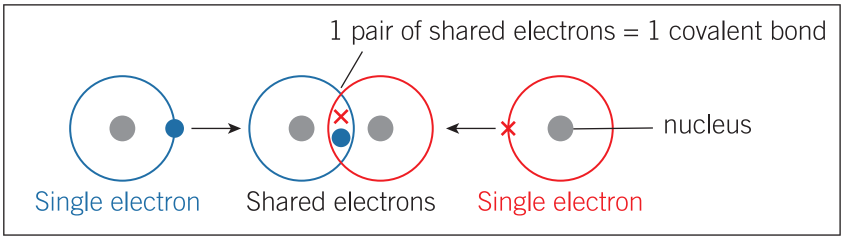


Figure 2A–2 Sharing of electrons between two non-metallic atoms in covalent compounds, indicating that every pair of shared electrons equals one covalent bond

Valence electron
an electron in the outermost shell of an atom

Covalent bond
a shared pair of electrons between two atoms within a molecule; *co* means together, *valent* means valence electrons

Electronegativity
the ability of an atom to attract shared electrons to itself. This is affected by an atom's core charge and the distance electrons are from the nucleus (atomic radius)

LINK **1A** INTRODUCTION TO THE ELEMENTS

LINK **3A** STRUCTURE AND PROPERTIES OF METALS

LINK **4A** FORMATION AND NAMING OF IONIC COMPOUNDS

Lewis structures

Lewis structures can be used for drawing any covalently bonded molecules, showing the bonding between atoms within a molecule and how the valence electrons are arranged around the outer shell of each atom in a molecule. Electrons are shown as dots or crosses in orbit around an atom. A straight line between the atoms represents pairs of electrons that are shared between the atoms. The number of electrons represented by the Lewis structure is the total number of valence electrons on individual atoms. Thus, Lewis structures are graphical representations of a molecule's valence shell electrons (lone pairs, unpaired electrons, or single/double/triple bonds).

Lewis structure
a representation of the electron arrangement in a molecule in which the valence shell electrons are represented by dots or crosses

NOTE

Non-valence electrons are not presented in Lewis structures.

VIDEO 2A–1 EXPLAINING LEWIS DIAGRAMS

Each atom aims to achieve a full outer shell of eight electrons for stable electron arrangements. Lewis referred to this as the **octet rule**. An octet of valence electrons can be achieved by either an atom's own electrons (noble gases, group 18) or some electrons being shared with other atoms. While many covalent compounds follow the octet rule, several exceptions do not follow the rule. For example, a hydrogen atom has too few electrons to form an octet. Thus, it can only hold a maximum of two electrons in its valence shell to attain a stable configuration by sharing one electron with another atom.

Octet rule
a rule to explain that atoms bond to other atoms in a way that results in each atom containing eight electrons in its valence shell

1C THE PERIODIC TABLE LINK

Figure 2A–3 summarises the pattern of valence electrons and covalent bonding illustrated by Lewis structures observed in the periodic table.

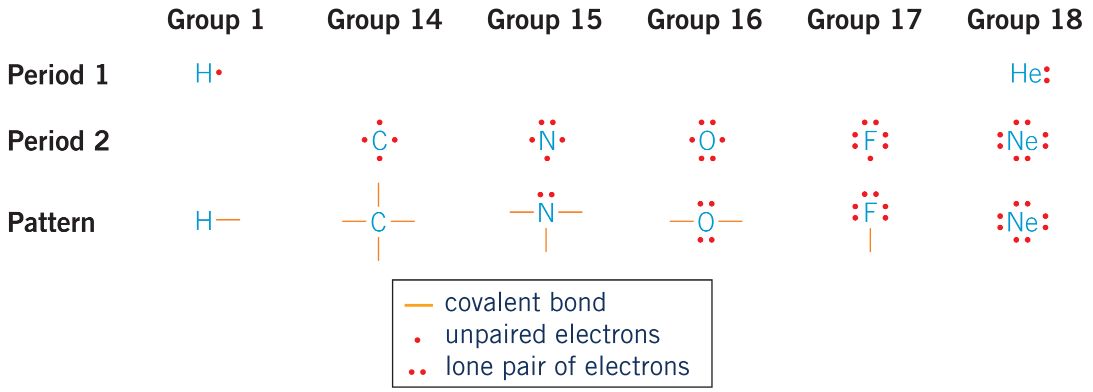


Figure 2A–3 The pattern of covalent bonding is based on the position of elements in the periodic table.

NOTE

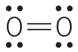
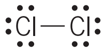
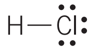
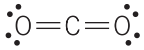
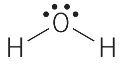
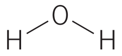
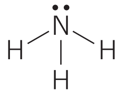
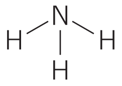
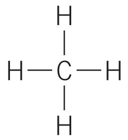
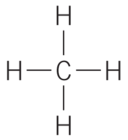
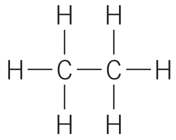
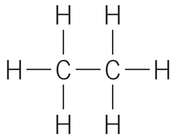
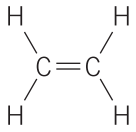
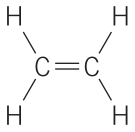
When an element is missing one electron in its valence shell, it forms one covalent bond, whereas if it is missing two electrons in its valence shell, it forms two covalent bonds, and so on.

The concepts of Lewis structures can be used to represent the covalent bonding in molecules. Table 2A–1 shows Lewis structures, structural formulas and molecular formulas to model various covalent molecules. Lewis structures represent the total number of valence electrons on each atom. **Structural formulas** represent the arrangement of atoms in a molecule and the chemical bonds between them. Note the absence of non-bonding electrons (lone pairs) in these representations. **Molecular formulas**, also known as chemical formulas, are an alphanumerical representation of the elements and the number of atoms of each element in the molecule.

Structural formula a representation of how the individual atoms of the compound may be arranged in three-dimensional space

Molecular formula the chemical formula that gives the total number of atoms of each element within in a compound

Table 2A–1 The Lewis structures, structural formulas and molecular formulas of various covalent molecules

Name	Lewis structure	Structural formula	Molecular formula
Hydrogen	H—H	H—H	H_2
Oxygen		O=O	O_2
Chlorine		Cl—Cl	Cl_2
Nitrogen	:N≡N:	N≡N	N_2
Hydrogen chloride		H—Cl	HCl
Carbon dioxide		O=C=O	CO_2
Water			H_2O
Ammonia			NH_3
Methane			CH_4
Ethane			C_2H_6
Ethene			C_2H_4

For example, hydrogen molecules, H_2, are formed from two hydrogen atoms sharing an electron pair in a single covalent bond. For oxygen molecules, O_2, two oxygen atoms share two pairs of electrons, wherein each oxygen atom requires two electrons to fill its valence shell (Figure 2A–4). Nitrogen molecules, N_2, contain triple covalent bonds because each nitrogen atom requires three more valence electrons to fulfil the octet rule.

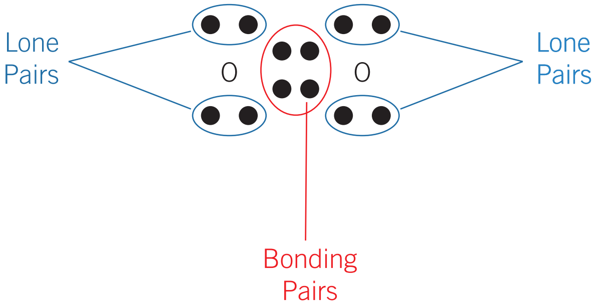


Figure 2A–4 The Lewis structure for an oxygen molecule, highlighting the shared (bonding) pair of electrons between oxygen atoms and lone pairs of electrons around each oxygen atom

Each atom in hydrogen chloride, HCl, needs one more electron to complete its valence shell. The hydrogen atom has only one valence electron, whereas chlorine has seven valence electrons. Therefore, they are bonded by a single covalent bond between them as shown in Figure 2A–5.

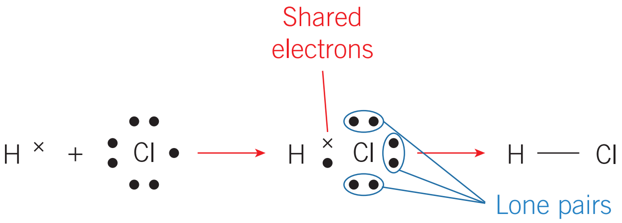


Figure 2A–5 The formation of the Lewis structure and structural formula for HCl showing shared pair of electrons (covalent bond) and non-bonding (lone) pairs of electrons around the chlorine atom

In carbon dioxide, CO_2, each oxygen atom requires two electrons to fill its valence shell and must share two of its valence electrons with the carbon atom. This means that there are four shared electrons between each carbon and oxygen atom (or two pairs of shared electrons), which are two covalent bonds between each carbon and oxygen pairing, as shown in Figure 2A–6.

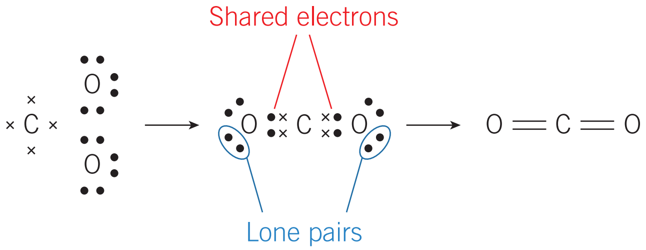


Figure 2A–6 The formation of the Lewis structure and structural formula for CO_2 showing shared pair of electrons (covalent bond) and non-bonding (lone) pairs of electrons around the oxygen atom

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A water molecule consists of two hydrogen atoms covalently bonded to the same oxygen atom. Each oxygen and hydrogen atom pairing shares one pair of electrons (or a single covalent bond).

In an ammonia molecule (NH_3), one nitrogen atom is covalently bonded to three hydrogen atoms. Each nitrogen atom requires three more electrons to fulfil the octet rule. Hence, three hydrogen atoms share one electron each with the nitrogen atom, wherein a single covalent bond connects each of the pairs of atoms. Both molecules are represented in Figure 2A–7.

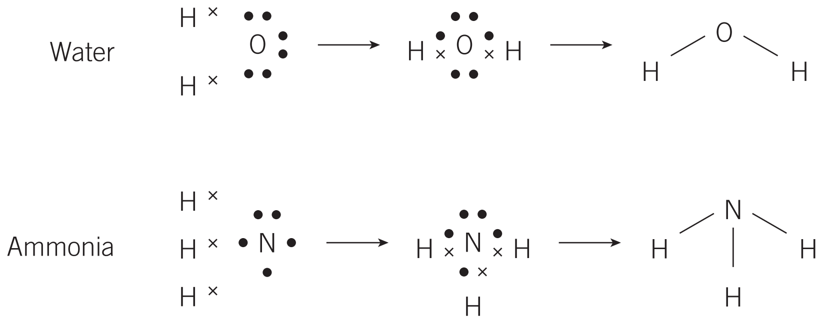

Figure 2A–7 The formation of the Lewis structure and structural formulas for both H_2O and NH_3

Methane (CH_4), ethane (C_2H_6) and ethene (C_2H_4) are represented in Figure 2A–8. Methane contains one pair of shared electrons between the carbon atom and each hydrogen atom. The carbon atom is bonded to the four hydrogen atoms by four single covalent bonds. In an ethane molecule, seven single covalent bonds are formed: six single covalent bonds between carbon and hydrogen, and one covalent bond between the two carbon atoms. An ethene molecule, which has two carbon atoms and four hydrogen atoms, has a double covalent bond between the two carbon atoms and a single covalent bond between each carbon and hydrogen pairing. Therefore, there a total of six covalent bonds in ethene.

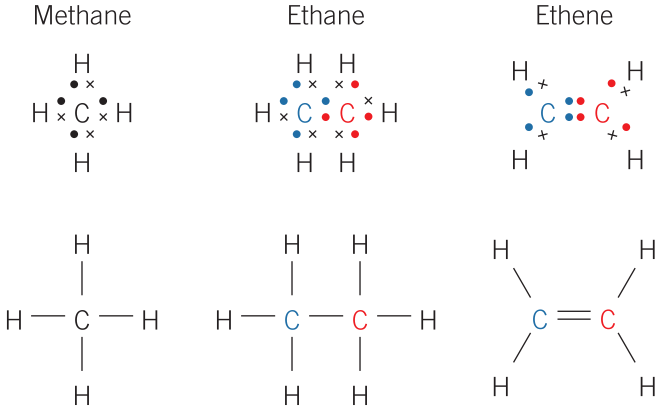

Figure 2A–8 The formation of the Lewis structures (top row) and structural formulas (bottom row) for methane (CH_4), ethane (C_2H_6) and ethene (C_2H_4)

WORKSHEET 2A–1 DRAWING LEWIS STRUCTURES DOC

2B MOLECULAR SHAPES AND POLARITY LINK

The geometric shape of some of these molecules will be covered in Chapter 2B.

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Check-in questions – Set 1

1 Define the following terms: Lewis structure, structural formula, molecular formula.

2 Provide the number of covalent bonds that atoms of the following elements want to form with other atoms, based on their position in the periodic table: hydrogen, oxygen, chlorine, carbon and nitrogen. Also, provide the total number of valence electrons for each molecule.

3 Provide the molecular formula, and draw the Lewis and structural formulas for hydrogen chloride.

4 Draw the Lewis structure for ethene (C_2H_4).

Relative strengths of intramolecular (covalent) bonds

Covalent and ionic bonds have a relative bond strength similar in magnitude. However, these are generally stronger than intermolecular forces.

One of the factors that dictate the strength of covalent bonds is bond length.

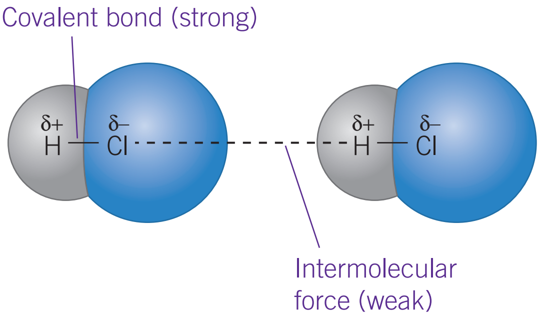

Figure 2A–9 A comparison between a covalent bond (intramolecular bond between atoms within the molecule) and an intermolecular force (represented by the dotted line between different molecules)

Bond strength

The **bond strength** between two bonded atoms is a measure of the energy required to break apart one mole of covalently bonded molecules in a gaseous state. This energy is measured in kilojoules per mole (kJ mol^{-1}).

For example, the bond strength of hydrogen, H_2, is 436 kJ mol^{-1} (at 25°C and 1 atm). This means that during the formation of H_2, 436 kJ of energy is released per mole of the molecule formed. Therefore, this also means that 436 kJ mol^{-1} of energy would be required to break one mole of H_2 molecule apart into separate H atoms.

Bond strength the amount of energy required to break apart one mole of covalently bonded molecules in the gas state, measured in kJ mol^{-1}

Bond length

The separation of atoms in non-metallic molecules with stable electron arrangements (generally fulfilled by the octet rule) requires energy. The stronger the bonds between atoms, the greater the energy needed to separate the atoms. The amount of energy needed can be determined by the bond length between atoms.

Referring to the example of hydrogen, H_2, if the two hydrogen atoms are too close, they will repel each other due to the positively charged nuclei. However, if they are too far apart, there will not be enough attraction to share the bonding electrons. Therefore, there is an optimum distance between nuclei which results in the maximum stability for the molecule. This is called the **bond length** and is typically measured in the unit angstroms (Å), equal to 10^{-10} m. In H_2, the bond length between the two H atoms is 0.74 Å.

Bond length the optimal distance between the nuclei of two atoms within a molecule allowing them to not repel each other but still share bonding electrons

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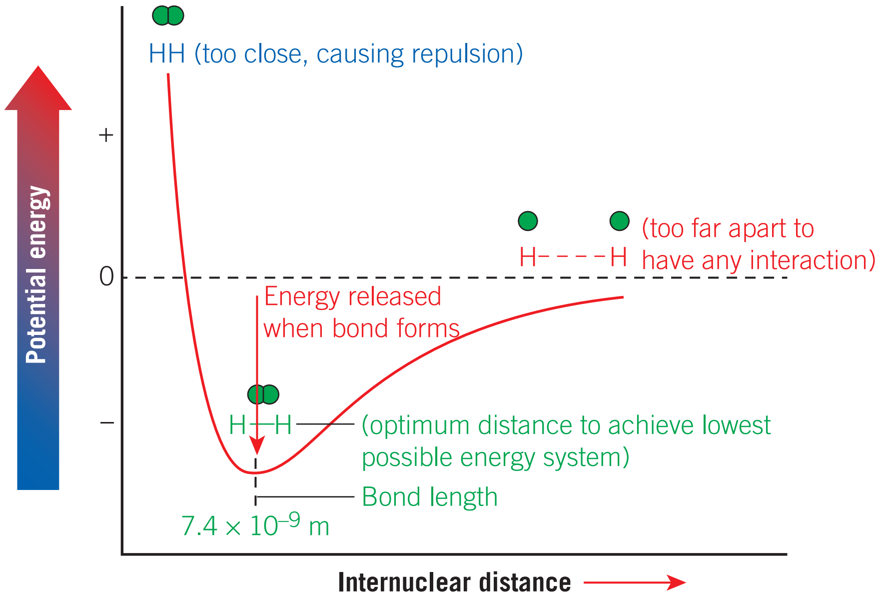


Figure 2A–10 The optimal bond length, or distance, between the two nuclei in a H_2 molecule

The bond length of triple bonds between like atoms, such as in a nitrogen molecule, N_2, is shorter than double bonds between like atoms, such as in an oxygen molecule, O_2. It requires 945 kJ to break a mole of N≡N bonds, but only 498 kJ to break a mole of O=O bonds. The shorter the bond length, the stronger the bond. Thus, double bonds are shorter in length and stronger than single bonds, and triple bonds are shorter in length and stronger than double bonds. The C=O double bond in carbon dioxide is longer and weaker than the C≡O triple bond in carbon monoxide.

For example, the bond length for a carbon-to-carbon single bond, like in ethane, is 1.54 Å, and the average bond energy is 346 kJ mol^{-1}. The bond length for a carbon-to-carbon double bond, like in ethene, is 1.34 Å, and the average bond energy is 602 kJ mol^{-1}.

Although there are four C–H bonds in molecules such as methane (CH_4), they do not require the same energy to break. After the first C–H bond is broken (approximately 439 kJ mol^{-1}), the remaining C–H bonds are easier to break. Therefore, the average bond strength value is 415 kJ mol^{-1}.

Check-in questions – Set 2

1 What is the relationship between bond length and bond strength?
2 **Challenge:** Draw the Lewis structure for hydrogen cyanide (HCN).

VIDEO 2A–2
SKILLS: USING REPRESENTATIONS TO EXPLAIN PROPERTIES

2A SKILLS

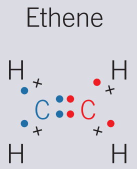
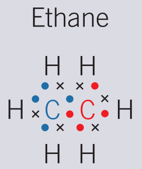
Using representations to explain properties
In many assessment and exam-style questions, you will be asked to draw structures (or analyse diagrammatical representations provided) and explain a particular property by referencing a specific aspect of these representations.

In this section, we have looked at Lewis structures and discussed the idea of bond strength for the covalent bonds within molecules. A question you may see and a response for it are given as an example on the following page.

Question:

Draw the Lewis structure for ethane (C_2H_6) and ethene (C_2H_4). Based on the Lewis structure drawings, order the bond energy between two carbons in each molecule from highest to lowest and justify your answer.

Answer:

Step 1	Draw the Lewis structure for ethane (C_2H_6) and ethene (C_2H_4).	 Ethene  Ethane
Step 2	Order the bond energy between two carbons in each molecule from highest to lowest.	Ethene (C_2H_4) > Ethane (C_2H_6)
Step 3	Justify your answer.	In this instance, the ethene molecule contains a double bond between the two carbon atoms and would have higher bond energy than the single bond between the two carbon atoms found in ethane. Therefore, the order of bond energy from highest to lowest is ethene and ethane.

Section 2A questions

1 Outline the procedure for illustrating a Lewis structure.

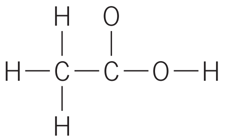
2 Select which covalent molecule in each pair is expected to have higher bond energy.

- a i C≡C ii C=O
- b i C=C ii C≡O

3 Draw the Lewis structure for water (H_2O).

4 Identify the period and group for carbon. From this, determine how many electrons are needed for it to form an octet (a similar arrangement to a noble gas). How many hydrogen atoms would carbon bond to in creating a stable molecule?

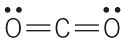
5 Complete the Lewis structure of acetic acid by adding bonds or lone pairs where required.

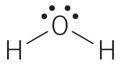


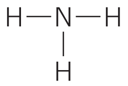
6 Draw the Lewis structures (electron dot diagram) for CH_4 and NH_3. Include all shared and non-bonding electrons in your diagrams.

7 Draw the Lewis structures and molecular formulas for hydrogen, oxygen, chlorine and nitrogen molecules. Explain which one of the these molecules does not follow the octet rule.

8 State whether each of the following Lewis structures is correct or incorrect. Justify your answer.

a 

b 

c 

2B Molecular shapes and polarity

Study Design:

- 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 a molecule

Glossary:

Dipole
Electron domain
Non-polar
Polar
Polarity
Valence shell electron pair repulsion (VSEPR) theory

ENGAGE

Seeing is believing

In 2013, a group of chemists and physicists at the University of California, Berkeley, captured the first images of molecules before and after a reaction. Using an atomic force microscope, the team at Berkeley developed an advanced technique that could take images that showed the bonds between atoms. These images clearly portrayed a change in molecular structure after a reaction. Previously, scientists inferred the structure of molecules based on molecular structure diagrams. However, as seen in the image below, we can observe that the changes in the chemical bonds are like those that have been illustrated in molecular structure diagrams for decades.

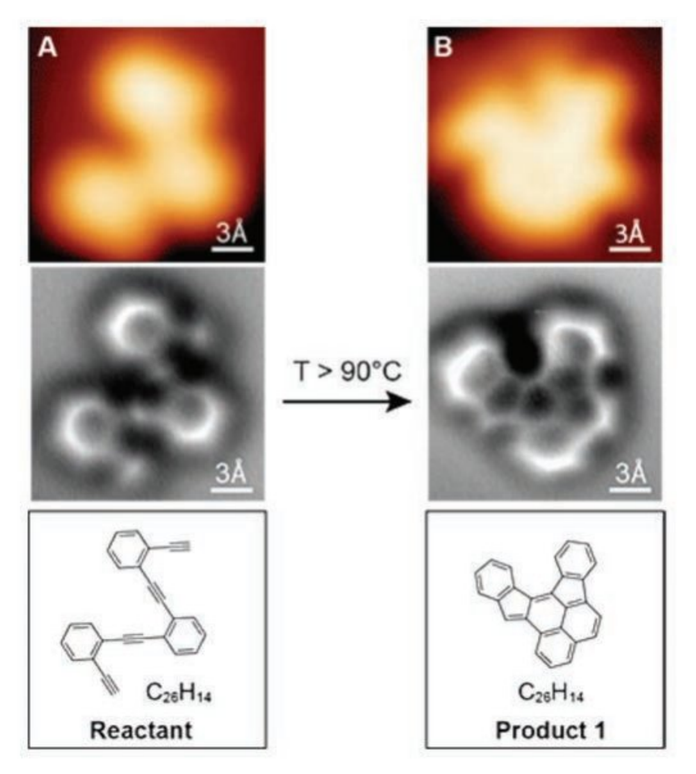


Figure 2B–1 These atomic microscopy images of a molecule before a reaction (left) and after a reaction (right) verify their classic structural molecular diagrams.

This section will explore how VSEPR theory can be used to determine the shape and polarity of covalent molecules.

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EXPLAIN

Molecular shape

Molecular shapes influence both the physical properties and chemical properties, including polarity, of a molecule. As such, in covalent molecules, the geometric arrangements of atoms in the molecule must be recognised. The **valence shell electron pair repulsion (VSEPR) theory** is widely used to predict the spatial orientations of atoms within molecules to describe the correct three-dimensional shapes of the covalent molecules.

Electron domains

The number of electron domains around a central atom determines the geometric shape of molecules. **Electron domains** are the negatively charged areas of either shared (bonding) electrons or lone pairs (non-bonding) electrons around an atom. As electrons all carry a negative charge, this means that they repel and try to get as far away as possible from each other. Depending on the number of electron domains, the angle between them is different, which influences the overall shape of molecules. A summary of this can be seen in Figure 2B–2. Note that the molecular shape is dictated by:

- the number of electron domains
- the number of lone pairs around the central atom in the molecule.

VIDEO 2B–1
SHAPES OF
MOLECULES

Valence shell electron pair repulsion (VSEPR) theory a model used to predict the geometric shape of molecules based on the number of valence shell electron pairs surrounding the central atoms in molecules

Electron domain the negatively charged area around an atom

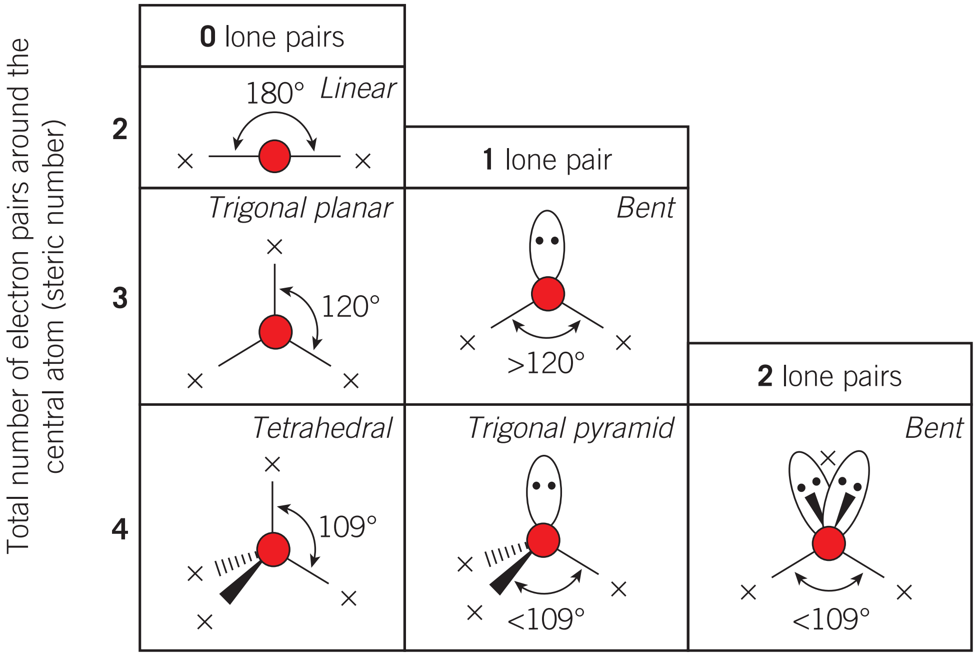


Figure 2B–2 The different shapes of molecules depend on the number of electron domains surrounding the central atom.

Valence shell electron pair repulsion theory

The valence shell electron pair repulsion (VSEPR) theory grew out of Lewis's theory. While Lewis structures present the connectivity of atoms within covalent molecules in two-dimensional diagrams, the VSEPR model is used to predict the three-dimensional molecular geometry based on the number of paired electrons associated with the central atoms of the molecules. This provides accurate shapes of covalent molecules.

When the atoms are placed near one another in the molecule, the shared valence electrons within the covalent bonds and lone-pair electrons tend to repel each other to automatically organise the arrangement into a shape that minimises repulsions among them. Consequently, the overall stability of the molecule can be improved by decreasing the energy of the molecule.

In this chapter, the VSEPR theory describes five shapes of molecules: linear, bent, trigonal planar, pyramidal and tetrahedral. These are represented in Figure 2B–3.

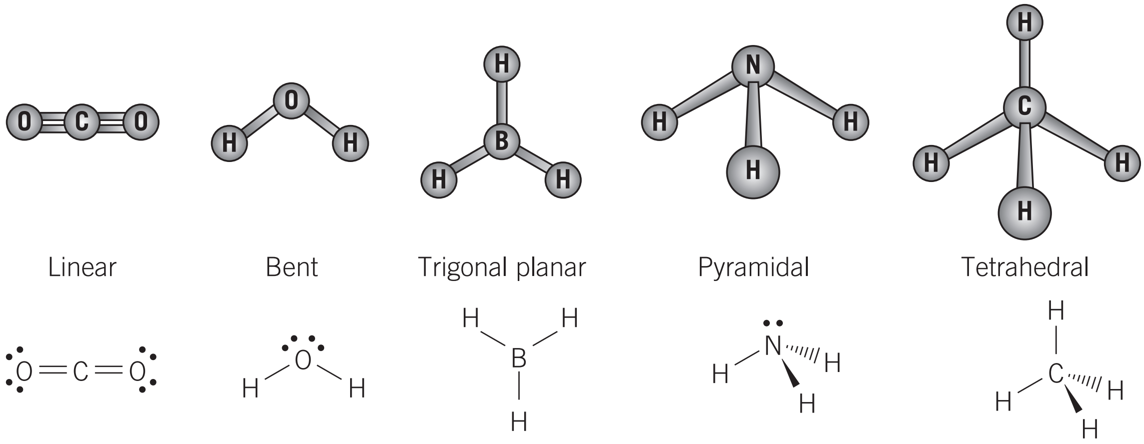


Figure 2B–3 Five shapes of molecules: linear, bent, trigonal planar, pyramidal and tetrahedral as determined by repulsion of electron pairs according to VSEPR theory (top). Examples of covalent molecules exhibiting these shapes (bottom).

NOTE

You are not required to know about trigonal planar molecular geometry in this Chemistry course.

VSEPR structures for molecules with pyramidal or tetrahedral geometry are drawn with solid lines, dashed wedged lines and solid wedges representing covalent bonds, as shown in Figure 2B–3. While solid straight lines in VSEPR models represent bonds in the plane of the paper, dashed wedged lines represent bonds that extend behind the plane (or away from the viewer), and solid wedges represent bonds that extend out of the plane (or towards the viewer). The geometry of covalent molecules is compared in Table 2B–1.

Table 2B–1 Molecular geometry comparisons for covalent molecules

	Linear	Bent	Pyramidal	Tetrahedral
Number of atoms involved	2 or 3	3	4	5
Number of lone pairs of electrons around central atom	0, 1, 2 or 3	1 or 2	1	0
Description	Two atoms bonded to each other or two atoms bonded to a central atom are placed at a bond angle of 180°	Non-linear arrangements where the central atom has one or two lone pairs of electrons	Based on the tetrahedral shape, the central atom with one lone pair of electrons at the apex is attached to three identical atoms at the base	The three-dimensional shape in which a central atom is bonded to four other atoms at the corners of the tetrahedron, without any lone pairs on the central atom

Check-in questions – Set 1

1 What is the main difference between Lewis structures and the valence shell electron pair repulsion (VSEPR) model?
2 Apart from linear, name three other molecular geometries according to VSEPR theory.
3 What are two factors that determine molecular shape?

Molecular polarity

The molecular geometry of a molecule affects its **polarity**. Although differences in the electronegativities of the atoms can determine the polarity of the bond, the shape of the molecule based on Lewis structures, and VSEPR theory determines whether the molecule is **polar** or **non-polar**.

For this section, you will need to recall electronegativity, which you learned about in Chapter 1C. Remember that electronegativity is the ability of an atom to attract shared electrons to itself. The most electronegative elements are located in the top right-hand corner of the periodic table, excluding the noble gases in group 18, as these elements are non-metals that want to gain electrons to fill their valence shell.

LINK **1C** THE PERIODIC TABLE

Polar bonds

Consider hydrogen chloride, HCl. Both hydrogen and chlorine, as you learned in Chapter 2A, require one electron to complete their respective valence shells. Therefore, they each share one of their own electrons in a single covalent bond.

LINK **2A** REPRESENTING COVALENT COMPOUNDS

Chlorine has an electronegativity of 3.16, whereas hydrogen has an electronegativity of 2.2. Therefore, the electrons will be more attracted to the chlorine atom and so spend more time closer to chlorine. As there is more negative charge closer to chlorine, a partial negative charge arises. This is represented by δ^- (delta negative). The hydrogen therefore does not have the electrons spending as much time on its side of the molecule and so has a partial positive charge. This is represented by δ^+ (delta positive).

Considering the difference in electronegativity and the non-symmetrical nature of this molecule, permanent partial negative and partial positive charges exist. The molecule is said to have a permanent **dipole**, as represented in Figure 2B–4.

Figure 2B–4 Lewis structure and structural formula for hydrogen chloride, showing a polar bond and permanent dipoles in the molecule

NOTE

The electronegativities of hydrogen and chlorine are written below the atoms. As the electrons spend more time closer to the chlorine atom, this is represented on the valence structure by an arrow next to the polar bond pointing in the direction of the more electronegative atom.

Polarity
the distribution of electric charge between two atoms joined by a bond

Polar
a molecule with two distinct poles of partial electric charge, a partial positive charge on one side of the molecule and a partial negative charge on the other side

Non-polar
a molecule with no permanent net dipole

Dipole
a bond or molecule where the ends have opposite partial charges, one positive and one negative

Polar molecules

Polar molecules are non-symmetrical. This means they have either different electronegativities among the bonded atoms or lone pair(s) of electrons on the central atom in the molecule (Figure 2B–5). A simple example is hydrogen fluoride (HF). The fluorine atom has a higher electronegativity than the hydrogen atom. This difference in electronegativities leads to an imbalance in electron sharing between atoms (with electrons pulled closer to the atom with the higher electronegativity value), creating a permanent dipole.

Another common example is water (H_2O). Water is a bent and non-symmetrical molecule, having two lone pairs of electrons on the central oxygen atom. The individual dipoles created by each oxygen–hydrogen polar bond point towards the oxygen atom and cannot cancel each other out, indicating a permanent net (overall) dipole. Here, the oxygen side of the bent molecule has a partial negative charge and the hydrogen side a partial positive charge.

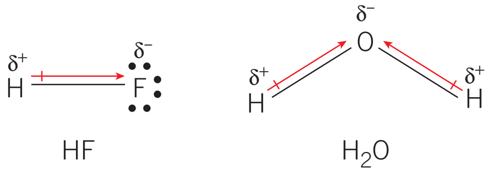
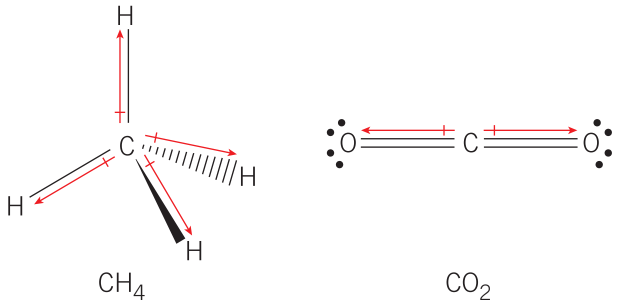


Figure 2B–5 Examples of polar and non-symmetrical molecules, HF, hydrogen fluoride (left) and H_2O, water (right), displaying that all dipoles do not cancel each other

Non-polar molecules

Non-polar molecules are symmetrical. This means all sides are bonded to the same atom with no unshared pairs of electrons (Figure 2B–6). For instance, methane (CH_4) is a non-polar molecule with tetrahedral geometry. All four carbon–hydrogen bonds are arranged symmetrically, and the equal magnitude of dipole moments cancel each other out.

Another common example is carbon dioxide (CO_2). CO_2 is a linear and symmetrical molecule. Two polar bonds are arranged symmetrically, pointing outward from the central carbon atom to each oxygen atom. As shown in Figure 2B–6, the equal strengths of the dipole moments cancel out each other, resulting in no overall polarity.



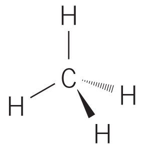
WORKSHEET 2B–1
POLAR AND NON-POLAR

Figure 2B–6 Examples of non-polar and symmetrical molecules, CH_4, methane (left) and CO_2, carbon dioxide (right), displaying that all dipoles are pointed outwards and cancel out each other

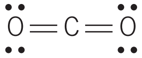
Check-in questions – Set 2

1 What is the main difference between non-polar and polar molecules in terms of the arrangement of atoms in molecules?

2 Label the molecules CH_4 and CO_2 as polar or non-polar and explain the reasons based on their geometric structures. Point out dipoles (arrows) of the individual bonds in their structures. (Label partial charges δ^+ and δ^- if it is necessary.)

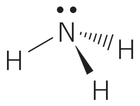


CH_4



CO_2

3 Label the following molecule, ammonia (NH_3), as polar or non-polar and explain the reason by pointing out dipoles (arrows) of the individual bonds.



2B SKILLS

VIDEO 2B–2
SKILLS: DETERMINING THE SHAPE AND POLARITY OF MOLECULES

Determining the shape and polarity of molecules

There are a number of steps you can follow and important items to recall when asked to determine the shape and/or polarity of a covalent molecule. These are explained in the example question and answer below.

Question:

Determine the shape and polarity of an ammonia molecule, NH_3. If partial charges are present, include these in a diagram.

Answer:

Step 1	Determine the central atom in the molecule.	Nitrogen, as it can make three covalent bonds (located in group 15), compared to each hydrogen making only one (located in group 1)
Step 2	Determine the number of electron domains around the central atom.	Four – three pairs of bonding electrons and one pair of non-bonding electrons (lone pair)
Step 3	Determine the shape.	Pyramidal
Step 4	Determine if the molecule is symmetrical or non-symmetrical.	Non-symmetrical
Step 5	Determine if the molecule is polar or non-polar. (Remember symmetrical = non-polar and non-symmetrical = polar)	Polar

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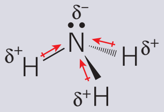
Step 6	Determine which atom the electrons spend more time closer to, then write the partial negative charge symbol on this side of the molecule. (This is shown in the Step 7 diagram below.)	Nitrogen
Step 7	Draw the molecule.	

Section 2B questions

1 In each case below, answer ‘true’ if the statement is correct or ‘false’ if it is incorrect.

- **a** The molecular shape is dictated by the number of electron domains and the number of lone pairs of electrons around the central atom.
- **b** The VSEPR model is used to predict two-dimensional molecular geometry based on the number of paired electrons associated with the central atoms of the molecules.
- **c** Non-polar molecules are non-symmetrical wherein all sides are bonded to the same atom with shared pairs of electrons.
- **d** Polar molecules are non-symmetrical, and they have either different electronegativities among the bonded atoms or lone pairs of electrons on the central atom in molecules.

2 Determine the molecular geometry of the following molecules using the VSEPR theory and state the polarity of each molecule: hydrogen sulfide (H_2S), tetrafluoromethane (CF_4), nitrogen trichloride (NCl_3), carbon disulfide (CS_2).

3 The VSEPR theory is a model used to predict molecular geometry. Explain how the VSEPR theory achieves this and provide an example.

4 Draw the molecular geometry of the following molecules using the VSEPR theory and provide the name of each.

- **a** CO_2
- **b** NH_3
- **c** $CHCl_3$

5 What is the molecular polarity of H_2, NH_3 and CH_4?

6 State whether each of the following molecules is polar or non-polar. Where the molecule is polar, draw it and show partial charges.

- **a** C_2H_6
- **b** Cl_2
- **c** HCl

7 Draw the molecular geometry of $CHCl_3$ (trichloromethane) and CCl_4 (tetrachloromethane) using the VSEPR theory and name the geometry. Label each molecule as polar or non-polar.

8 What is the expected shape of a molecule with one lone pair of electrons and three identical atoms around the central atom?

9 Draw and label the geometric arrangement of bonds in a molecule of methane.

2C Properties of covalent compounds determined by intermolecular forces

Study Design:

- 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

Glossary:

Dipole–dipole attraction
Dispersion force
Hydrogen bonding
Instantaneous dipole
Intermolecular force

ENGAGE

Cooking oil smoke points

The physical properties of covalent molecules are dictated by a variety of factors, such as the strength of intermolecular forces and chemical structure. An example of this is in the smoke point and chemical structure of various cooking oils. Cooking oils are comprised of a combination of saturated, monounsaturated and polyunsaturated fats. For example, oils such as coconut oil and palm kernel oil are mainly composed of saturated medium-chain fatty acids with some polyunsaturated fats and monounsaturated fats. Other cooking oils, such as sunflower oil, contain a lot less saturated medium-chain fatty acids and a lot more polyunsaturated long-chain fatty acids.

Why is it important to know the composition of these oils?

Knowing the medium-chain fatty acid and long-chain fatty acid content of cooking oil helps to explain the smoke point of the cooking oil. The smoke point of cooking oil is a temperature where the cooking oil begins to emit smoke due to the breakage of chemical bonds within the cooking oil. Sunflower oil has a higher long-chain fatty acid content compared with coconut oil and palm kernel oil. Consequently, sunflower oil has a higher smoke point than coconut oil and palm kernel oil because polyunsaturated long-chain fatty acids contain more intermolecular forces between the chains compared to saturated medium-chain fatty acids. Thus, more energy is required to break the chemical bonds in sunflower oil, which results in a higher smoke point.

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EXPLAIN

VIDEO 2C–1
UNDERSTANDING DIFFERENT INTERMOLECULAR FORCES

Intermolecular forces

Recall that covalent bonds are a strong type of force within molecules. They are an example of **intramolecular** forces, forces that act between the atoms making up the molecule.

However, there are also forces that keep separate molecules together. These are referred to as **intermolecular forces**. Intermolecular forces exist between molecules, indicating the forces of attraction or repulsion between neighbouring atoms or other particles (atoms, ions, or molecules). These forces are essential in determining a discrete molecular substance's physical properties, including melting point, boiling point, density and electrical conductivity.

Intermolecular force
an attraction or force between separate, or discrete, molecules

These intermolecular forces vary depending on the type of molecules, but there are three main types: dispersion forces, dipole–dipole attraction and hydrogen bonding. In this section, we will explore those three main types of intermolecular forces and investigate the physical properties of the molecules.

Dispersion forces

Within every molecule, there are electrons present, which are constantly orbiting atoms very quickly. For two atoms that have equal electronegativities, it is fair to assume that the electrons between the atoms would be equally shared and located directly in the middle; however, at any instant in time, the electrons making up a covalent bond (shared pair of electrons) could be closer to one side of the molecule than the other.

In Figure 2C–1, at this instant, the shared pair of electrons are located closer to the right-hand side of the molecule, so it has a very small partially negative charge. This is represented as $\delta\delta^-$ (delta negative). Therefore, the other side has a very small partially positive charge, represented as $\delta\delta^+$ (delta positive).

$$^{\delta\delta^+}\mathrm{H}\ \ {:}\mathrm{H}^{\delta\delta^-}$$

Figure 2C–1 Examples of dispersion forces in hydrogen, H_2

For this instant, the molecule has a negative and positive pole (dipole) which is referred to as an **instantaneous dipole**. If another H_2 molecule is near this instantaneous dipole created by the original H_2 molecule, its shared electrons will get repelled by the partially negative side of the dipole. Therefore, the hydrogen atom from the incoming H_2 molecule that is closer to the negative instantaneous dipole gains a very small partially positive charge, $\Delta\delta^+$.

These very weak attractions, as represented by the green arrows in Figure 2C–2, are known as **dispersion forces**.

Instantaneous dipole
when a dipole or opposite charges are created randomly and very quickly within a molecule

Dispersion force
a weak intermolecular force present in all molecules created between instantaneous dipoles.
Also known as London dispersion forces or van der Waals forces

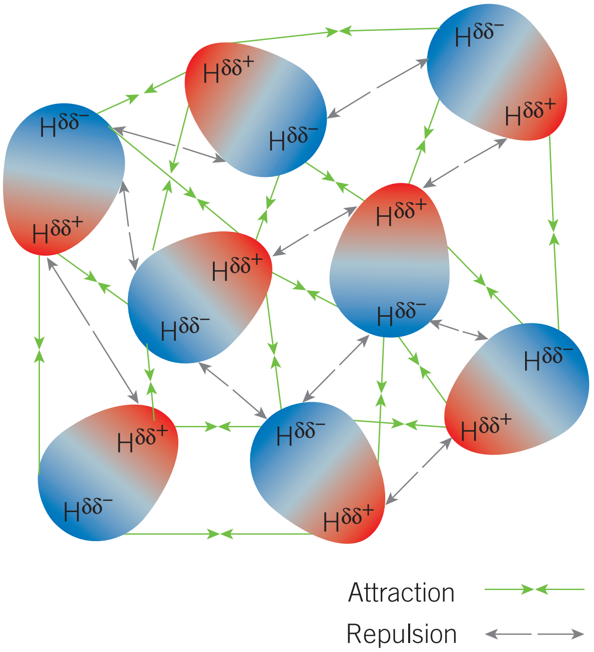


Figure 2C–2 Orientation of H_2 molecules as a result of the instantaneous dipoles created

As instantaneous dipoles can be created in all molecules regardless of whether they are polar or non-polar, dispersion forces are always present. However, they are only significant between non-polar molecules where they are the only type of intermolecular force present.

Dipole–dipole attraction

This intermolecular force exists only between polar molecules. When two or more molecules with permanent dipoles are near each other, the partially positive end of one molecule is attracted to the partially negative end of another molecule.

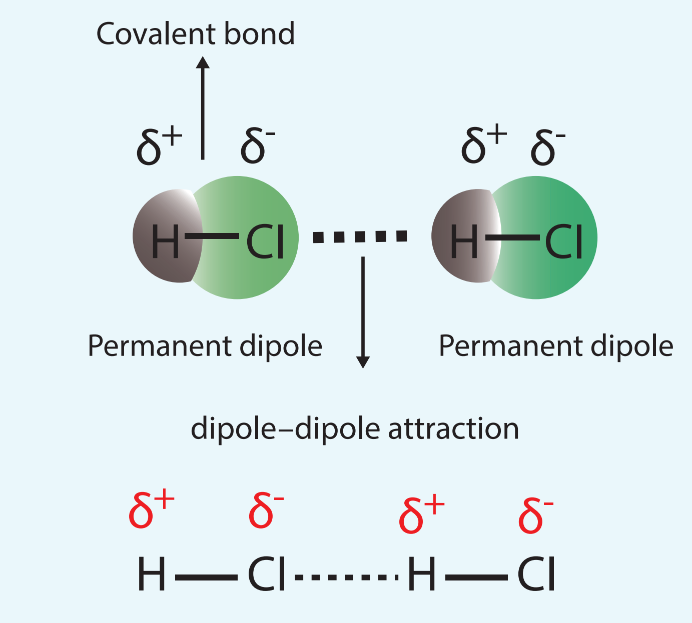


Figure 2C–3 Dipole–dipole attraction represented as a dotted line between two hydrogen chloride molecules. The partially negative chlorine atom in the molecule on the left has an attraction towards the partially positive hydrogen atom from the molecule on the right.

An example of **dipole–dipole attraction** occurs in hydrogen chloride, HCl, as shown in Figure 2C–3. Hydrogen and chlorine both require one electron to complete their respective valence shells. Therefore, they each share one of their electrons in a single covalent bond. Chlorine has an electronegativity of 3.16, whereas hydrogen has an electronegativity of 2.2. Therefore, the electrons are more attracted to the chlorine atom, so they spend more time closer to chlorine. As there is more negative charge closer to the chlorine, a partial negative charge occurs. Therefore, the electrons spend less time on the hydrogen's side, so it has a partial positive change. These opposite poles can be attracted to each other, which is what results in dipole–dipole attraction.

Dipole–dipole attraction an intermolecular force present between polar molecules that have permanent dipoles

Hydrogen bonding a stronger type of dipole–dipole attraction. Occurs between molecules that contain either an N–H, O–H or F–H covalent bond

Hydrogen bonding

When hydrogen (electronegativity of 2.20) is bonded to nitrogen (electronegativity of 3.04), fluorine (electronegativity of 3.98), or oxygen (electronegativity of 3.44), the significant difference in electronegativity between these atoms and hydrogen causes a very polar bond to form. Therefore, the intermolecular force between the partially positive hydrogen of one molecule and the nitrogen, fluorine or oxygen atom of another molecule is a stonger type of dipole–dipole attraction known as **hydrogen bonding**.

Figure 2C–4 represents the hydrogen bonding present between several molecules.

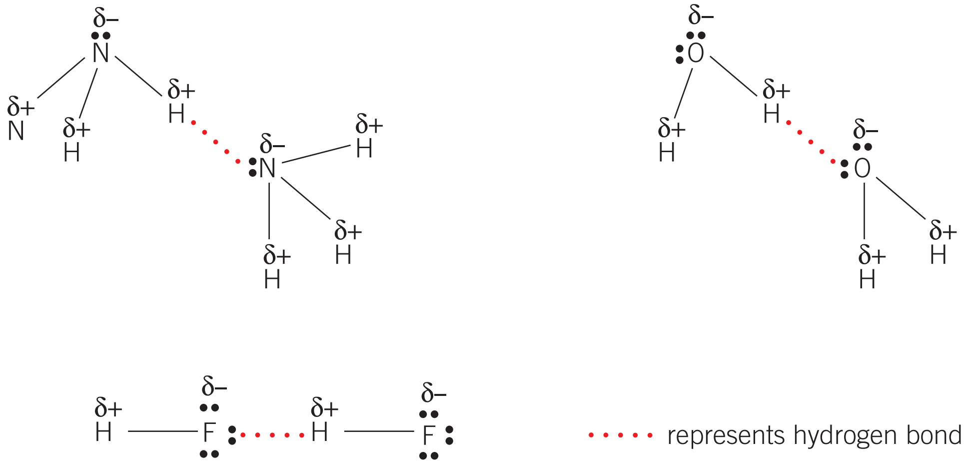
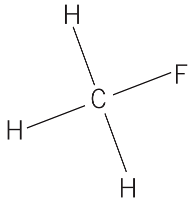


Figure 2C–4 Hydrogen bonding represented as a red dotted line between two molecules containing an O–H, N–H or F–H covalent (intramolecular) bond

A molecule must have an O–H, an N–H or an F–H covalent (intramolecular) bond within its molecule for it to be able to form a hydrogen bond between itself and another molecule. A molecule such as fluoromethane (CH_3F) would not be able to hydrogen bond with another CH_3F molecule, as there is no direct H–F bond within these molecules, as shown in Figure 2C–5.



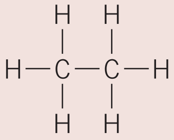
WORKSHEET 2C–1
IDENTIFYING INTERMOLECULAR FORCES

Figure 2C–5 Fluoromethane does not have hydrogen bonding as an intermolecular force as it does not contain an H–F bond. It would, however, demonstrate dipole–dipole attractions between molecules.

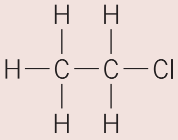
Worked example 2C–1: Strength of intermolecular forces

VIDEO
WORKED EXAMPLE 2C–1

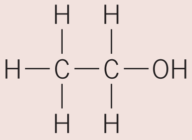
a What type of intermolecular force(s) would occur between the molecules of the following pure substances?


$\mathrm{H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H}$

Ethane


$\mathrm{H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-Cl}$

Chloroethane


$\mathrm{H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-OH}$

Ethanol

b List the molecules in order of increasing strength of the intermolecular forces acting between them in the pure substances, from weakest to strongest, and explain your response.

Solution

Step 1	Identify the differences.	Ethane is a non-polar molecule and contains instantaneous dipoles, creating weak intermolecular forces between the molecules, called dispersion forces. Chloroethane and ethanol are both polar molecules. Ethanol contains an O–H bond so is capable of hydrogen bonding, while chloroethane creates the dipole–dipole attraction between molecules. Hydrogen bonding is the stronger form of dipole–dipole attraction.
Step 2	Refer back to the question.	The order of intermolecular forces in increasing strength is dispersion forces < dipole–dipole attraction < hydrogen bonding. Therefore, the order of the molecules in terms of their strength of intermolecular forces is ethane < chloroethane < ethanol.

Check-in questions – Set 1

1 What is the main difference between intramolecular and intermolecular bonding (forces)?
2 Which is stronger – intramolecular or intermolecular forces?
3 List the three main intermolecular forces in order of increasing strength. What are the differences between them?

Physical properties of molecular substances

The physical properties of covalent molecules largely depend on the strength of the intermolecular forces acting between them.

Next, we will observe the physical properties of molecular substances, including melting points, boiling points and electrical conductivity.

Melting points/boiling points

Melting and boiling points are a measure of how strong the attractive forces (intermolecular forces) are between the molecules in the pure substance. As the intermolecular forces between molecules increases, more energy is needed to disrupt them, which means that the melting and boiling points also increase.

Molecules with hydrogen bonding between them show the highest boiling and melting points compared to those with dipole–dipole attractions or dispersion forces; whereas dispersion forces, the weakest intermolecular forces, would have the smallest effect on melting and boiling points. This is summarised in Figure 2C–6.

The strength of intermolecular forces

Hydrogen bonding > dipole–dipole attraction > dispersion forces

←

Boiling/melting points increase with increasing strength of intermolecular force

Figure 2C–6 The order of intermolecular force strength from hydrogen bonding (highest bond strength) to dipole–dipole attraction to dispersion forces (lowest bond strength). Consequently, the melting point/boiling points of molecules increase as the strength of intermolecular forces increases.

If the molecules have the same types of intermolecular forces, larger molecules generally have higher boiling and melting points compared to smaller molecules. For example, non-polar methane molecules have dispersion forces. However, when the chain length of carbon atoms increases, like in non-polar ethane, the boiling and melting points also increase. This is because longer linear chains experience more intermolecular forces, in this case dispersion forces, which need more energy to disrupt the structures of the molecules.

Electrical conductivity

3A STRUCTURE AND PROPERTIES OF METALS LINK

4B STRUCTURE AND PROPERTIES OF IONIC COMPOUNDS LINK

11B STRONG AND WEAK ACIDS AND BASES LINK

Covalent molecules are electrical non-conductors (electrical insulators). The covalent bond is formed by sharing electrons between atoms. Due to the sharing of electrons, covalent molecules do not have any freely moving charged particles (free ions or delocalised electrons as you will learn in Chapter 3A and 4B) to carry and move the electrical charges. Therefore, covalent molecules do not conduct electricity in either pure solid or liquid forms.

NOTE

In Unit 2, one standard test is to measure the conductivity of a strong acid and a weak acid of the same concentration. The weak acid is usually an organic acid, like acetic acid, a covalent molecule that dissolves in water to produce ions so that the solution will conduct electricity. Therefore, there are a number of examples of covalent molecules dissolving in water and the solution being a conductor. However, some might argue that these reactions represent a chemical reaction and not a solvation process.

Worked example 2C–2: Comparing boiling points

VIDEO WORKED EXAMPLE 2C–2

Determine whether HF or HCl has the higher boiling point.

Solution

Step 1	Identify the differences.	Both molecules are polar; however, HF contains an F–H bond and so is capable of hydrogen bonding.
Step 2	Explain the effect of any differences.	The hydrogen bonding between HF molecules is stronger than the dipole–dipole attractions present between HCl molecules.
Step 3	Relate back to the question.	Therefore, more energy is required to overcome the intermolecular forces between HF molecules, so it has the higher boiling point.

Check-in questions – Set 2

1 Which intermolecular force would have the strongest effect on melting and boiling points?
2 Which intermolecular force would have the weakest effect on melting and boiling points?
3 Do covalent molecules in a pure substance conduct electricity? Explain.

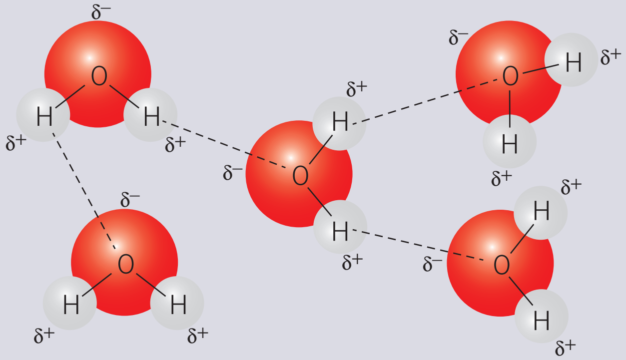
2C SKILLS

The importance of water and hydrogen bonding for future chapters

In Unit 2 you will further explore properties and behaviours of some of the covalent molecules in this chapter. Firstly, you will look at water, carbon dioxide and methane with respect to their abilities to absorb infrared radiation, thereby contributing to greenhouse effects. Consequently, you will learn how to measure the volume and mass of gases using stoichiometry and the ideal gas equation. You will also determine the solubility of substances, including many of these covalent molecules, in a given mass or volume of solvent (usually water).

LINK **15B** GAS STOICHIOMETRY AND THE GREENHOUSE EFFECT

LINK **13A** MEASURES OF SOLUBILITY



A key focus of Area of Study 1 in Unit 2 is ‘How do chemicals interact with water?’ In this chapter, you learned that water has a bent geometric shape, containing permanent dipoles, and the strongest type of intermolecular force between molecules, called hydrogen bonding. It is this intermolecular force that allows a single water molecule to form hydrogen bonds with four other water molecules.

Water is therefore a very effective solvent as it can form either hydrogen bonds or dipole–dipole attractions (and dispersion forces) with other polar molecules. Take sugars, such as glucose, for example. The oxygen–hydrogen bond of a water molecule can form a hydrogen bond with the oxygen–hydrogen bond of a glucose molecule. The same can occur with alcohols.

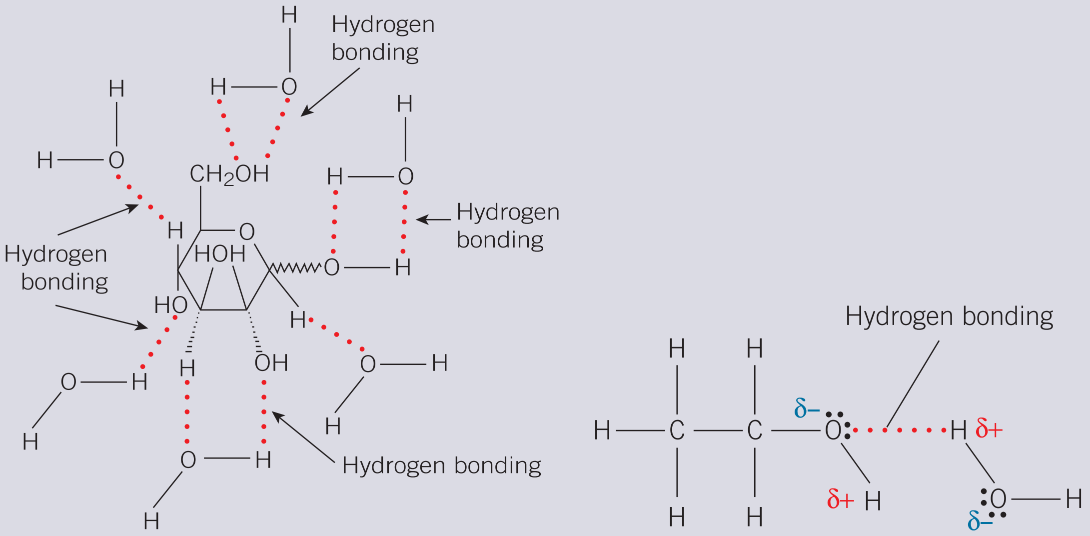


Figure 2C–7 Hydrogen bonding between water molecules and either glucose (left) or ethanol (right) makes water a very effective solvent.

Other substances, such as ionic compounds (Chapter 4C), organic compounds (Chapter 7B) and acids and bases (Chapter 11A) can also interact with water, forming aqueous solutions.

Much of the content you will learn in your chemistry studies is based on the knowledge you have established here. It is important that you are able to build connections between new and prior knowledge throughout your chemistry course.

LINK **4C** PRECIPITATION REACTIONS

LINK **7B** HYDROCARBONS

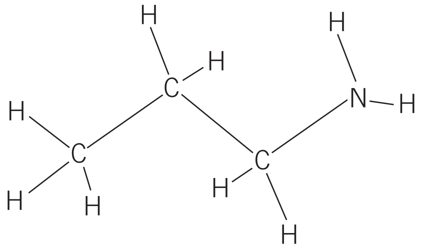
LINK **11A** INTRODUCTION TO ACIDS AND BASES

Section 2C questions

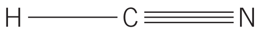
1 Name a type of intermolecular force that is found in all molecules. Explain your response.

2 Which of these molecules are capable of hydrogen bonding with the same type of molecule?

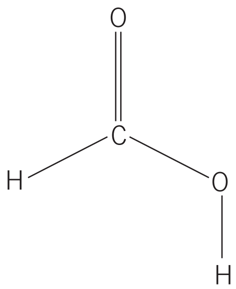
a



b



c



3 a What type of intermolecular force(s) would occur between molecules with the following formula?

   i Br_2

   ii PF_3

   iii H_2O

b Which of the above would you expect to have the highest boiling point and why?

4 Do low boiling points indicate strong intermolecular forces? Explain your answer.

5 Would CH_4 or C_2H_6 have a higher boiling point?

6 Does CH_4 or $CHCl_3$ have a higher boiling point?

7 Nitrogen and phosphorous are both elements in group 15. They both form three single covalent bonds with three hydrogen atoms. The result is ammonia (NH_3) and phosphine (PH_3).

a Draw the geometric shape of both ammonia and phosphine and identify the general name given to that shape. In your answer include any partial charges that may be present.

b The electronegativity of phosphorus is 2.19 and the electronegativity of hydrogen is 2.20. These values are very similar, suggesting that neither atom would have a substantially stronger attraction for the shared electrons. However, phosphine is a polar molecule. Explain why this is the case.

c Explain why the boiling point of ammonia (NH_3, 240 K) is greater than the boiling point of phosphine (PH_3, 185 K).

8 What type of intermolecular force is present in alkanes, such as methane and ethane?

9 Order the following compounds by the strength of their intermolecular forces, from strongest to weakest. Justify your responses.

CH_4, HCl, C_2H_6, H_2O

2D Diamond and graphite as covalent compounds

Study Design:
The structure and bonding of diamond and graphite that explain their properties (including heat conductivity, electrical conductivity and hardness) and their suitability for diverse applications

Glossary:
Allotrope
Tetrahedral structure

ENGAGE

Ancient carbon crystals

Both diamond and graphite are completely made up of carbon atoms. While both display covalent bonding and high melting/boiling points, there are significant differences between the two structures.

Diamonds have been traded for thousands of years. As far back as the first century AD, they were widely considered to be the most precious stone in the world. Their value is mainly derived from their scarcity, unique appearance and challenging mining conditions.

Diamonds mined from underground are extremely old–possibly billions of years old or as old as Earth itself. Not only do diamonds take a very long time to form, but diamonds also require an extreme amount of heat and pressure to form. To start the diamond forming process, almost 511.5 billion kilograms per square metre of pressure and a temperature of around 1200°C are required. These extreme conditions allow the carbon atoms to arrange into a tetrahedral formation and contribute to the diamond's brilliant and coveted appearance.

In the absence of such extreme conditions, graphite can be formed instead of a diamond. Graphite, or *graphein* in Ancient Greek (meaning 'to write/draw'), has a completely different set of applications compared with diamond. From 3000 BC to the present age, graphite has been used for applications in art, such as drawing and decorating ceramic pottery. Nicholas-Jacques Conte, a scientist in the army of Napoleon Bonaparte, invented the modern pencil lead using graphite and received a patent for his invention in 1795. An interesting application of graphite is its use as a crucible for storing molten steel because of graphite's high melting point and inertness.

EXPLAIN

Giant covalent molecules: diamond and graphite

Allotrope
a different structural form of the same chemical element

WORKSHEET 2D–1
COMPOUNDS OF CARBON

Diamond and graphite are made entirely out of carbon. They are known as **allotropes** of carbon. Although both diamond and graphite are chemically the same, they have entirely different atomic arrangements.

The carbon atoms in diamond and graphite are bonded to each other by strong covalent bonds, albeit with different arrangements, which lead to different crystal structures. Their distinct crystal structures also contribute to their unique physical properties. These unique physical properties are desirable for different sets of applications.

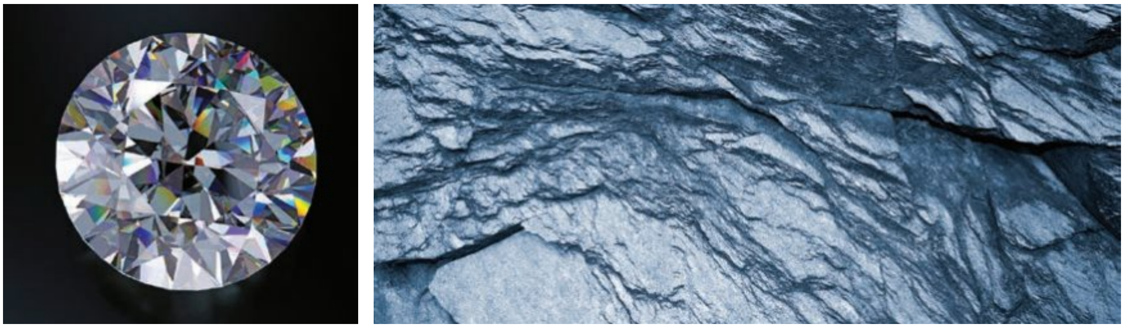


Figure 2D–1 An example of a diamond crystal (left) and graphite schist (right).

In this section, we will explore the structure and bonding of both diamond and graphite. In addition, we will investigate their unique physical properties and how these properties can be utilised in different types of applications in our lives.

Structure and bonding of diamond and graphite

Diamond and graphite are allotropes of pure carbon with giant covalent network structures. They have identical chemical compositions. However, the entirely different atomic arrangements allow them to form their distinct crystal structures.

Structure and bonding of diamond

In a diamond, all four valence electrons on each carbon atom are bonded to four neighbouring carbon atoms by covalent bonds. This tightly bonded arrangement allows the diamond to form an extremely rigid **tetrahedral structure**. Consequently, an extremely stable three-dimensional network is formed within a diamond's solid crystal structure.

Tetrahedral structure
a structure where the central atom is bonded to four other atoms that are located at the corners of a tetrahedron

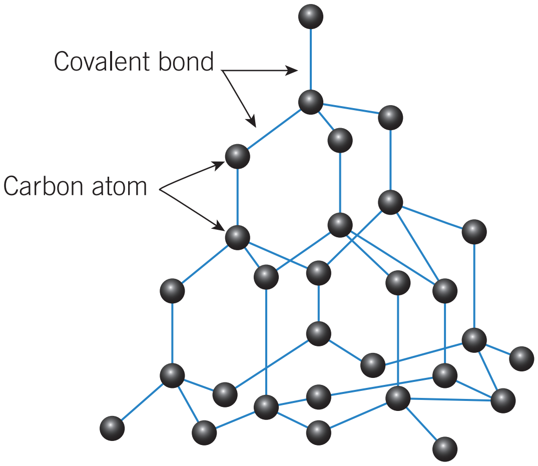


Figure 2D–2 Giant covalent network lattice of diamond indicating the covalent bond between carbon atoms and their tetrahedral arrangement

Structure and bonding of graphite

The carbon atoms in graphite are arranged in layers. Each carbon atom is covalently bonded to three other neighbouring carbon atoms leaving one free electron. Therefore, all carbon atoms are held tightly in one layer, forming a six-membered hexagonal ring structure. The delocalised electron on each carbon atom moves freely within the whole sheet of carbon atoms in one layer. In addition, each planar hexagonal carbon layer is held together by weak intermolecular forces. The distance between carbon atoms within the same layer is shorter than the distance between two layers, indicating that two different forces exist in a three-dimensional graphite structure.

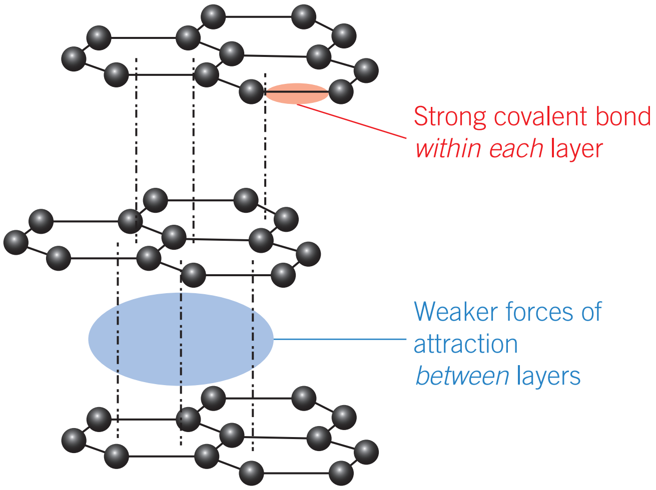


Figure 2D–3 Giant covalent layer lattice of graphite indicating the strong covalent bond within each layer and weaker forces between layers.

Physical properties of diamond and graphite

Different crystal structures in diamond and graphite contribute to their unique physical properties, which are distinct from each other. The main differences are summarised in Table 2D–1.

Table 2D–1 Comparison of physical properties of diamond and graphite

Physical property	Diamond	Graphite
Appearance	Transparent	Dark black opaque
Hardness	Hardest naturally occurring substance on Earth	Soft substance
Heat/thermal conductivity	Excellent conductor of heat	Moderate conductor of heat
Electrical conductivity	No electrical conductivity	Excellent electrical conductivity due to free moving delocalised electron from each C atom

Applications of diamond and graphite

Diamond and graphite have a wide range of industrial applications according to their unique physical properties. Several examples of their applications are outlined in Tables 2D–2 and 2D–3.

Table 2D–2 Applications of diamond

Application	Explanation based on structure and physical properties
May be incorporated into the cutting end of drill bits used for oil well drilling	Involves cutting through various hard substances in Earth's crust, so drill bits used for drilling applications must be robust. Due to its infinite tetrahedral structure arrangement, diamond has a highly strong and rigid 3D structure, making diamond the hardest substance on Earth and the best material to be used to cut through other hard materials in Earth's crust.
Found in the burrs of a dental drill	Due to excellent heat conductivity, the diamonds on the burrs can heat up quickly and consequently drill into a tooth more efficiently.

VIDEO 2D–1
ONE ELEMENT, TWO COMPOUNDS

Table 2D–3 Applications of graphite

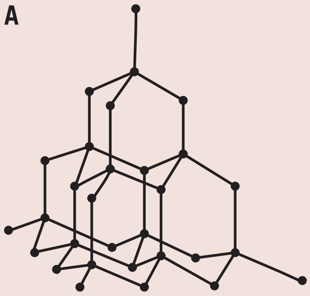
Application	Explanation based on structure and physical properties
Battery anode in lithium-ion batteries	When a lithium-ion battery is charged, lithium ions flow from the cathode to the anode through a buffer system and vice versa when the battery is discharged. The charging and discharging cycle is enabled by delocalised electrons in graphite that act as a reversible host for lithium cations.
Solid lubricant material	The bonding forces between the layers in graphite are weak. Thus, they can easily slip over each other, which is referred to as having a low shear strength. For this reason, graphite is a good candidate for lubrication and as a pencil core (easy to write with and erase).

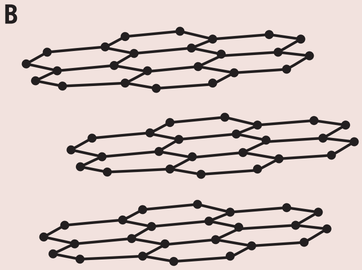
Worked example 2D–1: Structure and properties of molecules

Two different structures of carbon, A and B, are shown below. One of them can conduct electricity and the other has incredibly high heat/thermal conductivity.

a Name each material.

b Explain how the unique structures of each material allow them to better conduct either heat or electricity.





Solution:

a (A) is diamond and (B) is graphite.

b First, explain the structural differences related to bonding of each.

Diamond (A) has a tetrahedral crystal structure in which each carbon atom is covalently bonded to four other carbon atoms without any free delocalised electrons on carbons. This allows the diamond to be a strong and rigid three-dimensional network structure.

In graphite (B), each carbon atom is covalently bonded to three other neighbouring carbon atoms with one delocalised electron on each carbon, forming a six-membered hexagonal ring structure. Each planar hexagonal carbon layer is held together by weaker attractive forces to form the whole graphite crystal structure.

Second, use your answer about bonding to explain how their unique structures can allow them to better conduct either heat or electricity.

Diamond has a rigid tetrahedral crystal structure, in which all carbons are uniformly and closely bonded together. Thus, diamond can serve as a good heat conductor.

All carbons in the graphite's structure have their own delocalised electrons. These delocalised electrons can move freely between layers, allowing graphite to conduct electricity; however, in diamond, there are no free moving delocalised (non-bonding) electrons available to conduct the electricity.

Check-in questions – Set 1

1 Which element is found both in diamond and graphite?

2 With reference to graphite and diamond, which is the better conductor of:
  a heat?
  b electricity?

3 Selecting between diamond and graphite, which is the hardest naturally occurring substance?

4 What are examples of applications for graphite and diamond according to their unique physical properties? Provide at least two examples of each.

2D SKILLS

VIDEO 2D–2
SKILLS: USING VENN DIAGRAMS IN CHEMISTRY

Using Venn diagrams in chemistry

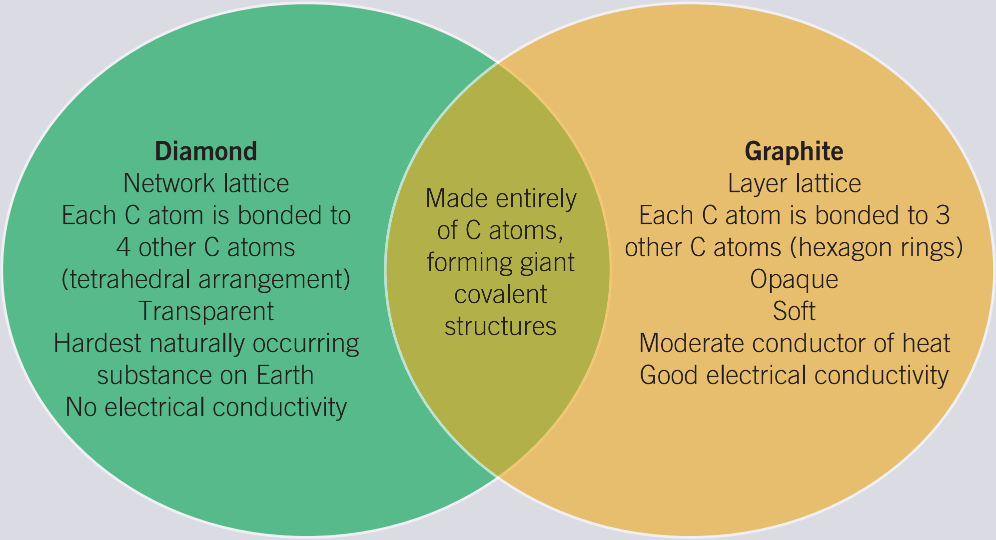
You may well have come across Venn diagrams in your mathematics studies, but did you know they are very useful in comparing structures or properties of different structures in chemistry as well?

If you are not familiar with a Venn diagram, it is an illustration that uses slightly overlapping circles to show the relationships among items. Things that are common between items are included in the overlapping region of the circles, while those that are distinct are in the stand-alone regions of each circle.

Venn diagrams have many other benefits, including:

- visually organising information to see the commonalities and differences between items
- comparing multiple sets of data.

Venn diagrams may be a useful pictorial revision tool to recall the similarities and differences for diamond and graphite learned in this chapter.



Venn diagrams can range in their level of complexity. Once you have learned about metals and ionic compounds in Chapters 3 and 4, respectively, you might like to revise either the structures or the properties of the three types of bonding (including covalent bonding from this chapter) using a three-circled Venn diagram.

LINK **CHAPTER 3**
CHAPTER 4

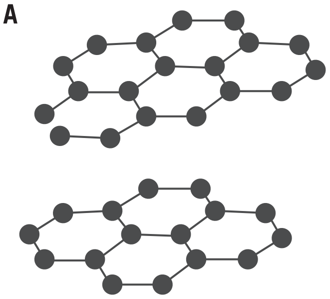
Section 2D questions

1 What is the molecular geometry of diamond and graphite? (Referring back to Figure 2B–3 earlier in this chapter may be of use here.)

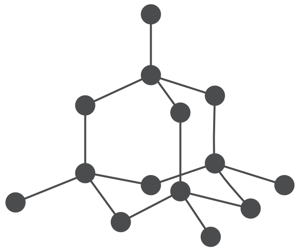
2 a Name the structures, A and B, shown below.

  b What are the main structural similarities between graphite and diamond? (Hint: You might like to use the Venn diagram in 2D Skills to assist you.)

A



B



3 Which is the better heat conductor – diamond or graphite? Refer to the covalent bonding present in each to explain your answer.

4 Why does graphite conduct electricity while diamond does not?

5 Why is diamond the hardest naturally occurring substance on Earth?

6 Provide an example of an application of diamond and explain how diamond's physical properties make it desirable for this application. Also, explain why graphite would be a poor candidate for this application.

7 Graphite is commonly used as an anode material in lithium-ion batteries. How do the physical properties of graphite enable it to act as a good anode material in lithium-ion batteries?

8 Graphite is commonly used as a solid lubricant material. What physical property does graphite possess that makes it a desirable material for this application?

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Chapter 2 review

Summary

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

Success criteria – I am now able to:		Linked question
2A.1	Recall the definitions of ‘covalent bond’ and ‘lone pair’	11☐
2A.2	Recall the number of covalent bonds made by elements in a particular group on the periodic table	3☐
2A.3	Distinguish between Lewis structures, structural formulas and molecular formulas	1☐
2A.4	Recall the steps of how to draw Lewis structures using the number of valence electrons, number of covalent bonds and octet rule to represent molecules	1☐, 2☐, 17☐
2A.5	Identify shared electrons and lone pairs on Lewis structure diagrams	11☐
2B.1	Define ‘electron domain’	17☐
2B.2	Recall the four shapes of molecules according to the valence shell electron pair repulsion (VSEPR) theory	4☐, 9☐, 12☐
2B.3	Determine two factors that affect the shape of molecules	4☐, 9☐, 12☐
2B.4	Draw the shape of molecules according to the VSEPR theory and display arrows (direction) of dipoles, including partial charges where required	4☐, 9☐, 12☐
2B.5	Explain the differences between non-polar and polar molecules in relation to the arrangement of atoms in molecules	5☐, 9☐, 13☐
2B.6	Determine the polarity of molecules based on the polarity of the covalent bonds and molecular geometry	5☐, 9☐, 13☐
2C.1	Describe the differences between intramolecular and intermolecular bonding (forces)	6☐
2C.2	Recall the three types of intermolecular forces	6☐, 7☐, 8☐, 9☐
2C.3	Determine which intermolecular forces exist between different covalent molecules, based on the polarity of the covalent bonds and molecular geometry	7☐, 8☐, 9☐
2C.4	Explain the relationship between the strength of intermolecular forces and a molecule’s melting and boiling points	14☐, 15☐
2C.5	Rationalise the non-conduction of electricity by covalent molecules	9☐
2D.1	Understand and illustrate the crystal structure and bonding of both diamond and graphite	10☐
2D.2	Explain the difference in physical properties of diamond and graphite: appearance, heat conductivity, electrical conductivity and hardness	16☐
2D.3	Recall the applications of both diamond and graphite in society	16☐

Multiple-choice questions

1 Which of the following is illustrated in a Lewis structure?

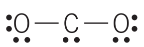
A all the electrons in an element

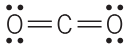
B the valence protons in an element

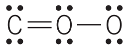
C the valence electrons in an element

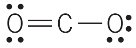
D the electrons in the last two shells in an element

2 Which of the following depicts the correct Lewis structures for CO_2?

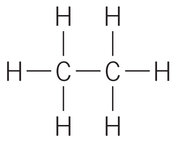
A 

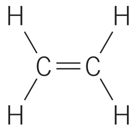
B 

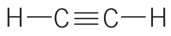
C 

D 

3 Which of the following descriptions is true regarding these molecules? (Note: There may be more than one correct answer.)


A


B


C

A The longest and strongest bond between two carbon atoms is found in molecule A.

B The shortest and weakest bond between two carbon atoms is found in molecule C.

C The shortest and strongest bond between two carbon atoms is found in molecule C.

D The longest and weakest bond between two carbon atoms is found in molecule A.

4 Based on VSEPR theory, predict the shape of the H_2S molecule.

A bent

B tetrahedral

C linear

D trigonal pyramidal

5 Which of the following descriptions is true regarding these molecules?

NH_3, H_2O, C_2H_6

A non-symmetrical polar molecule, symmetrical polar molecule, symmetrical non-polar molecule

B non-symmetrical polar molecule, non-symmetrical polar molecule, symmetrical non-polar molecule

C symmetrical polar molecule, non-symmetrical polar molecule, non-symmetrical non-polar molecule

D symmetrical non-polar molecule, non-symmetrical polar molecule, symmetrical non-polar molecule

6 What is the correct order of forces from weakest to strongest?

A covalent bonding, hydrogen bonding, dispersion forces, dipole–dipole attraction

B hydrogen bonding, dipole–dipole attraction, dispersion forces, covalent bonding

C dispersion forces, covalent bonding, dipole–dipole attraction, hydrogen bonding

D dispersion forces, dipole–dipole attraction, hydrogen bonding, covalent bonding

7 Which combination of molecules is capable of hydrogen bonding? (Note: There may be more than one correct answer.)

A H_2O and NH_3

B H_2 and H_2O

C O_2 and H_2

D NH_3 and HF

8 Which of the following molecules interact primarily through dispersion forces?

A HBr

B CH_4

C CH_3F

D C_2H_5OH

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9 Which of the following statements about water (H_2O) and methane (CH_4) molecules is incorrect?

- **A** Hydrogen bonding between water molecules is stronger than the dispersion forces present between methane molecules, therefore water has a higher boiling point than methane.
- **B** Water is a good conductor of electricity but methane is not.
- **C** Water is a polar molecule and non-symmetrical. Methane is a non-polar molecule and symmetrical.
- **D** According to VSEPR theory, methane has a tetrahedral structure and water has a bent structure.

10 Diamond and graphite are both composed of the same element – carbon. Choose the **most** accurate statement below.

- **A** They have the same arrangement of carbon atoms within their structures.
- **B** Graphite is an organic material and diamond is an inorganic material.
- **C** They have the same physical and chemical properties.
- **D** They have different three-dimensional crystalline structures, resulting in different physical properties.

Short-answer questions

11 Draw the Lewis structures for Cl_2, O_2 and N_2. Then label all covalent bonds and lone pairs of electrons. (4 marks)

12 Based on VSEPR theory, draw the geometric shape of each molecule below and provide its general geometrical shape name.

- **a** NH_3
- **b** CH_4 (4 marks)

13 Which of the following molecules have permanent dipoles? Draw each molecule, and show partial charges if the molecule is polar.

- **a** O_2
- **b** CCl_4
- **c** H_2O
- **d** HCl (4 marks)

14 Does CH_4 or C_2H_4 have a higher boiling point? Explain your answer. (3 marks)

15 List the following substances in order of increasing boiling point. Explain your answer.

C_2H_6, HCl, C_2H_5OH (4 marks)

16 The objects below are commonly made using either a pure metal or an alloy.

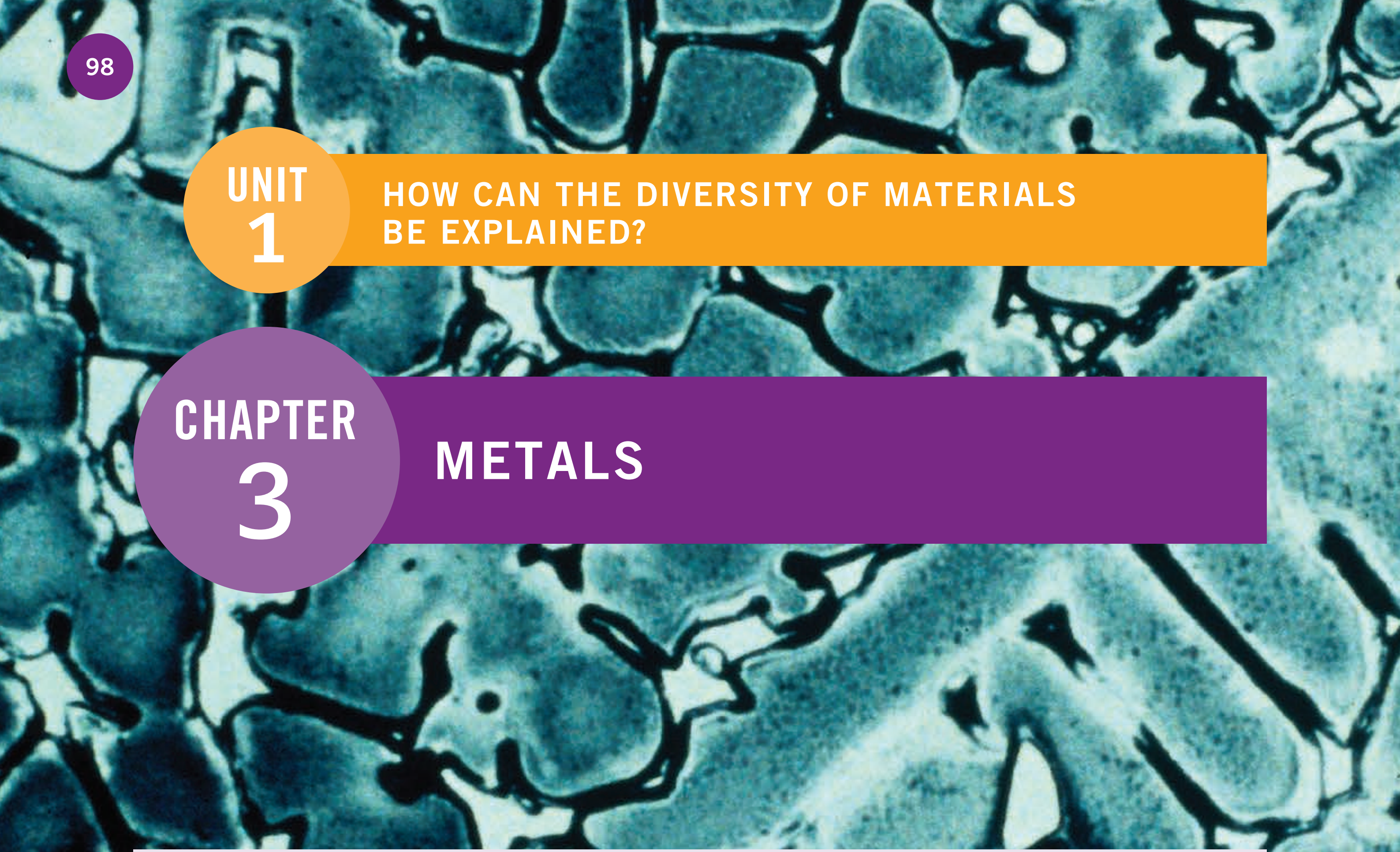
- **i** battery anodes
- **ii** concrete cutter blade
- **iii** dental drill
- **iv** dry lubricants

- **a** Group the above objects according to the material (diamond or graphite) that is (could be) used instead to impart desirable properties on its structure. (2 marks)
- **b** Explain why diamond or graphite would be desirable for making each of the above objects. The structural shape of diamond and graphite should be discussed based on VSEPR theory. (4 marks)

17 The Lewis structure of H_2 is shown below.

H : H

- **a** Define 'electron domain.' (1 mark)
- **b** How many electron domains are associated with each hydrogen atom within this hydrogen molecule? (1 mark)
- **c** Explain why hydrogen is an exception to the octet rule. (2 marks)



UNIT 1
HOW CAN THE DIVERSITY OF MATERIALS BE EXPLAINED?

CHAPTER 3
METALS

Introduction

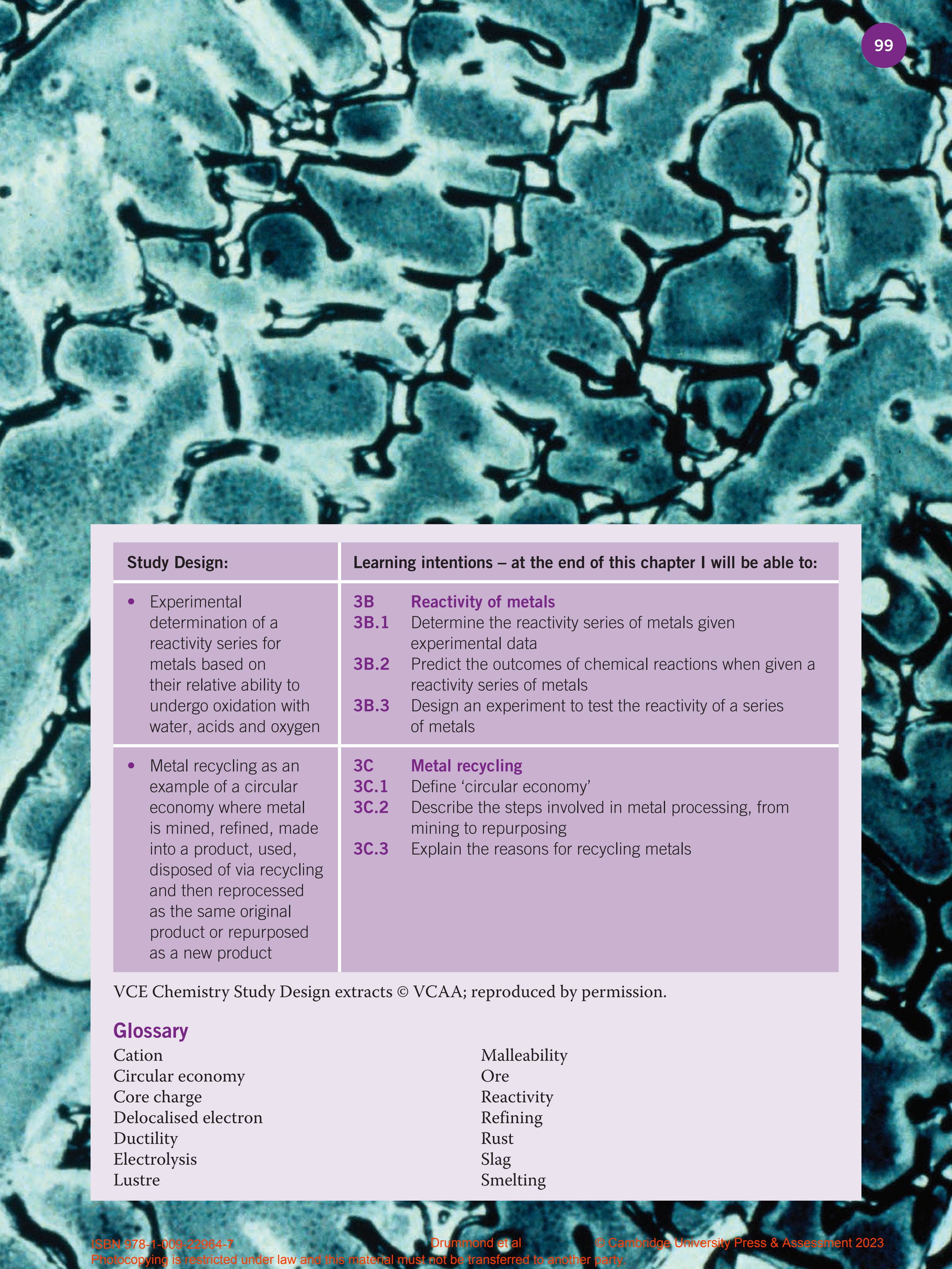
INTRODUCTION VIDEO METALS

We encounter metals all the time in our everyday life: soft drink cans are made of aluminium, coins are composed of copper and nickel, car batteries contain lead and jewellery is often produced from gold and silver. How can metals be used for so many types of items with such different functions and various shapes and structures? All in all, there are more than 90 different metals found on the periodic table. Considering the large number of metals, they do possess some unique characteristics. However, they also share some common properties. In this chapter, you will learn about those properties, as well as the different reactions they undergo and how metal recycling is being used to increase sustainability and shift society towards a circular economy.

Curriculum

Area of Study 1 Outcome 1
Reactions of metals

Study Design:	Learning intentions – at the end of this chapter I will be able to:
• 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 Structure and properties of metals** **3A.1** Describe the metallic bonding model **3A.2** Draw a labelled diagram of a metallic lattice **3A.3** Using the metallic bonding model, explain common properties of metals



Study Design:	Learning intentions – at the end of this chapter I will be able to:
- Experimental determination of a reactivity series for metals based on their relative ability to undergo oxidation with water, acids and oxygen	**3B Reactivity of metals** **3B.1** Determine the reactivity series of metals given experimental data **3B.2** Predict the outcomes of chemical reactions when given a reactivity series of metals **3B.3** Design an experiment to test the reactivity of a series of metals
- 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	**3C Metal recycling** **3C.1** Define ‘circular economy’ **3C.2** Describe the steps involved in metal processing, from mining to repurposing **3C.3** Explain the reasons for recycling metals

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Glossary

Cation
Circular economy
Core charge
Delocalised electron
Ductility
Electrolysis
Lustre
Malleability
Ore
Reactivity
Refining
Rust
Slag
Smelting

Concept map

The common properties of metals are explained by the nature of metallic bonding and existence of metallic crystals

3A Structure and properties of metals

PROPERTIES OF METALS


High melting point


Malleable


Good conductor of heat

Ductile

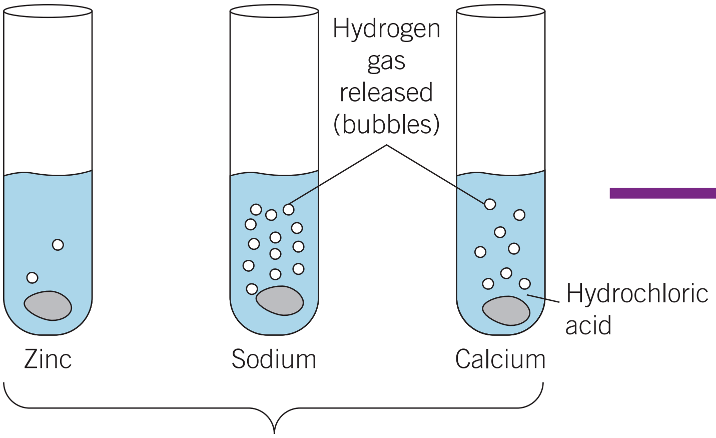

Lustrous


Good conductor of electricity


Experimentally determining reactivity series for different metals based on the metal's relative ability to undergo oxidation

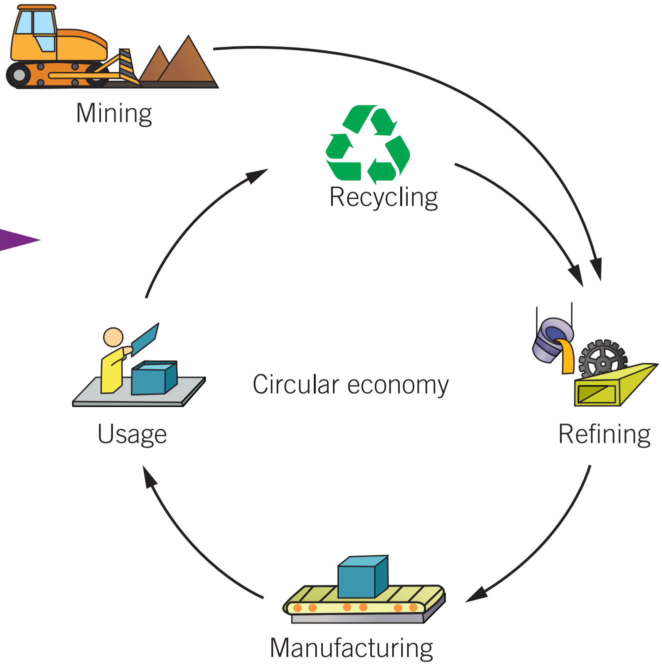
Metal recycling and the manufacturing process as an example of a circular economy

3B Reactivity of metals



In the same volume and concentration of hydrochloric acid

3C Metal recycling



See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

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3A Structure and properties of metals

Study Design:
The common properties of metals (lustre, malleability, ductility, heat conductivity and electrical conductivity) with reference to the nature of metallic bonding and the existence of metallic crystals

Glossary:
Cation
Delocalised electron
Ductility
Lustre
Malleability

ENGAGE

One of the most recognisable superheroes from the Marvel movie franchise is Iron Man. You may think that the name comes from the fact that his suit is made out of iron; however, that is not the case. The name 'Iron Man' is in reference to his strength and determination. The suit, on the other hand, is made from a gold and titanium alloy (this is mentioned by Tony Stark in one of the movies). What do iron, gold and titanium all have in common? They are all metals. In this chapter, you will learn more about metals, their properties, their structure and some common reactions.

Properties of metals

Metals have a variety of different uses in society. Their ability to be used for such a wide range of purposes has a lot to do with their properties. For example, metals are used for electrical wiring due to their **ductility** and the fact that they are good conductors of electricity. Metals are also regularly used for jewellery due to their **lustre** and **malleability**. Similarly, their high melting points and ability to conduct heat make them well suited to be used for cookware.

Ductility the ability of a material to be drawn out into a long wire without breaking

Lustre the appearance of a surface in reflected light. Lustrous metals appear shiny

Malleability the ability of a material to be distorted by an applied force without breaking

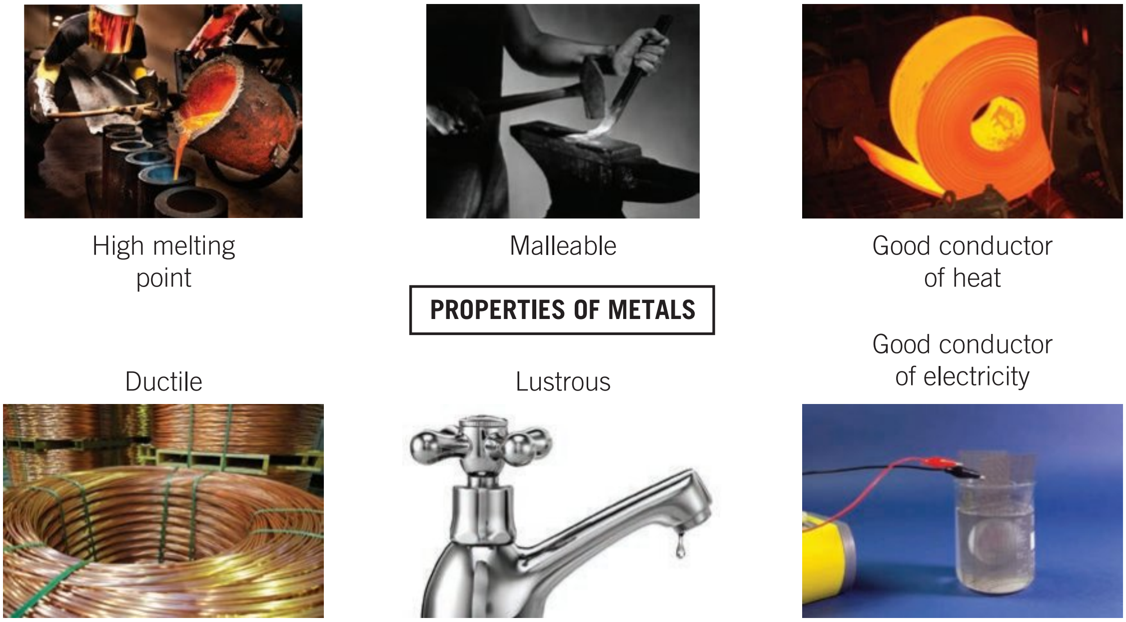


Figure 3A–1 Some of the general properties of metals

2A REPRESENTING COVALENT COMPOUNDS LINK

4A FORMATION AND NAMING OF IONIC COMPOUNDS LINK

1B ELECTRON CONFIGURATIONS LINK

Cation
an atom that has lost electron(s) from its valence shell, resulting in an overall positive charge

Delocalised electron
an electron that is not associated with any particular atom and is free to move

Metallic bonding

One of the key concepts that you will learn about in Unit 1 is bonding. There are a number of different bonding models that can be used to explain how elements form compounds. In this section, we are going to focus on metallic bonding.

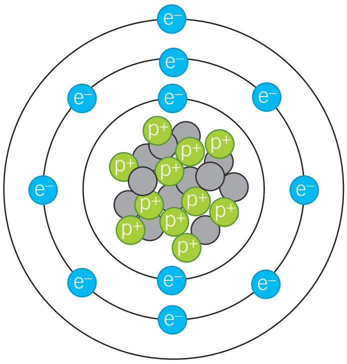


Figure 3A–2 A representation of a sodium atom with positively charged protons (green) and neutral neutrons (grey) in the nucleus surrounded by 11 negatively charged electrons (blue)

Generally, when we think about the structure of an atom, we picture a nucleus containing neutrons and positively charged protons that are surrounded by orbiting negatively charged electrons, such as the example shown for sodium in Figure 3A–2. The electrons in the outer shell are known as valence electrons – the diagram shows that sodium has a single valence electron. All atoms prefer to have a stable outer-shell electron configuration, and for metals this means losing their valence shell electrons. After a metal has lost these electrons, it now has a charge imbalance due to having more protons than electrons. This species is now an ion and, more specifically, we call an ion with a positive charge a **cation**.

In a metallic substance, this process leads to the production of many cations and free electrons. These free electrons are not associated with any particular atom, so they can move around 'freely' in the metal. The electrostatic force of attraction of these negatively charged free electrons to the positively charged cations is the basis for metallic bonding and holds everything together in a regular structure called a crystal lattice (Figure 3A–3). For these reasons, you may often hear metallic bonding referred to as an array (or lattice) of positive ions in a 'sea' of **delocalised electrons**. This is known as the metallic bonding model.

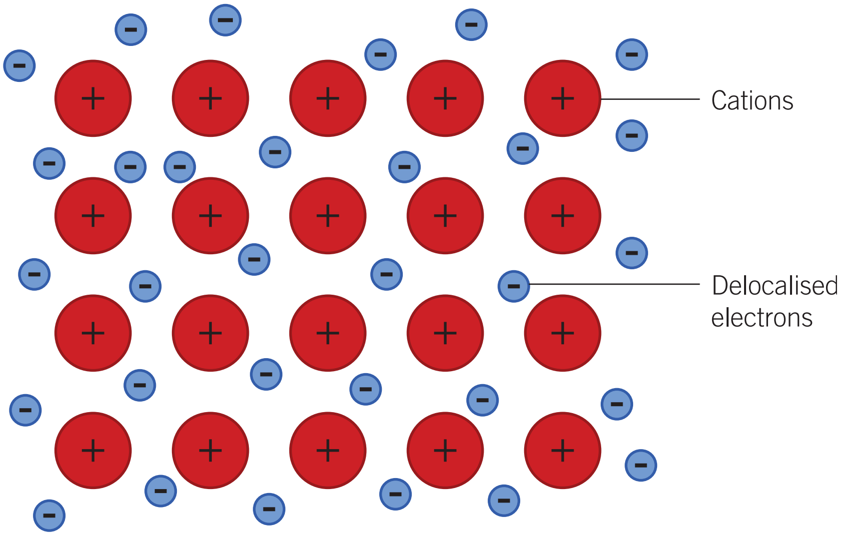


Figure 3A–3 Metallic bonding occurs by having a sea of delocalised electrons moving among positively charged metal ions (cations). This diagram shows just one layer of metal ions.

Check-in questions – Set 1

1 What term is given to the property of metals that allows them to be formed into a wire?
2 Other than your answer to Question **1**, list three other common properties of metals.
3 Define 'metallic bond'.
4 Why do metal atoms form cations and not anions (negatively charged ions)?

Properties of metals explained

We can use this metallic crystal lattice model to explain the common properties of metals that were mentioned earlier in this section. This is a good chance to start to build a skill that you will use a lot throughout VCE Chemistry, which is applying your understanding of general concepts to explain specific examples. Let's discuss each property separately.

VIDEO 3A–1 EXPLAINING THE PROPERTIES OF METALS

Metals are malleable

One of the key concepts of the metallic bonding model is that the electrostatic force of attraction is between cations and delocalised electrons. This means that the electrons aren't associated with any particular cation. So, when a force is applied to the metal, such as when it is hit with a hammer, the cations and the sea of delocalised electrons are able to shift and still maintain their attraction to each other. Thus, the metal is able to be bent and twisted without breaking. This is the physical property of malleability.

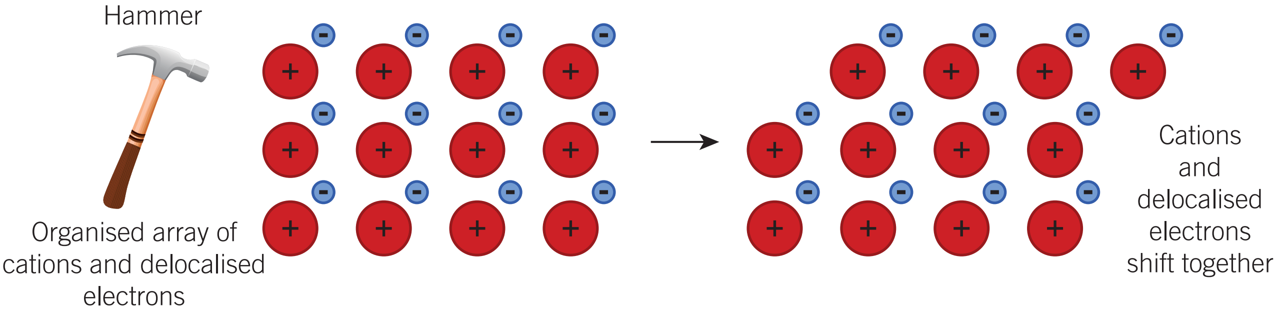


Figure 3A–4 A representation of the malleability of a metal

Metals are ductile

Ductility refers to the ability to draw a metal out into a long wire. The explanation for why metals are ductile is essentially the same as for why they are malleable (Figure 3A–4). Therefore, the description for malleability is equally applicable here.

Metals are good conductors of heat

The delocalised electrons found within a metal are able to move freely throughout the lattice structure. If heat is applied to one part of the metal, the electrons present will absorb this thermal energy and transform it into kinetic energy. These electrons will then begin to move faster around the lattice, going from the 'hot' end to the 'cool' end, as shown in Figure 3A–5. In this way, the electrons are able to quickly transfer the thermal energy throughout the metal, which is what makes metals such good conductors of heat.

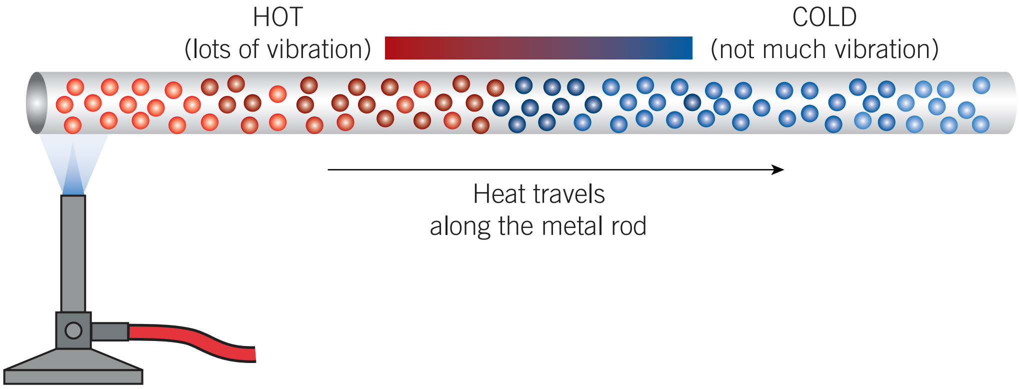


Figure 3A–5 The conduction of heat by a metal being heated by a Bunsen burner

Metals are good conductors of electricity

1A INTRODUCTION TO THE ELEMENTS LINK

To understand this property, we first need to understand that electricity is the movement of charged particles through a substance. Importantly, when it comes to metals, we know that the lattice contains a sea of freely moving, delocalised electrons. Recall that electrons are negatively charged subatomic particles. Therefore, these free electrons are able to carry an electrical current throughout the metal, which is what makes them such excellent conductors.

NOTE

Group 1 metals, and mercury, have low melting points, which is atypical of other metals.

Metals have high melting points

When a substance melts, it changes state from a solid to a liquid. For metals, this occurs when they reach a temperature that causes the cations to no longer be held within their regular lattice structure. The electrostatic forces of attraction between the cations and the delocalised electrons, which hold the lattice together in metals, are very strong. This means that a large amount of heat energy is required to disrupt those electrostatic forces of attraction, which is why metals have high melting points.

Metals are lustrous

Free-moving electrons, particularly those found on the surface of the substance, can reflect light. As the metal lattice has many delocalised electrons moving freely, they are able to readily reflect light. The reflection of the light from the electrons in the metal lattice, when observed by the human eye, makes the metal appear shiny.

WORKSHEET 3A–1 EXPLORING THE PROPERTIES OF METALS

NOTE

While all of these properties can be explained by referring to the metallic bonding model shown in Figure 3A–3, other properties cannot, meaning this model has some limitations.

Some of the properties or specifics that cannot be explained include the following:

- actual melting point (specifically group 1 metals, which have a lower melting point than other metals)
- difference in electrical conductivity between metals
- magnetism of metals such as cobalt, iron and nickel
- varying densities between metals
- rapid movement of delocalised electrons through the structure.

3A SKILLS

Using labelled diagrams to support answers

You have probably heard the saying, 'a picture paints a thousand words'. Essentially, this means that sometimes the best and easiest way to explain something is to illustrate it. This can also apply to chemistry, where an annotated diagram with a few key written points can be a quick and effective way of answering a question.

Remember these key points when using a diagram to support your answer.

Key points to remember:

The following points will help you to identify the key elements.

1. Plan what you want your diagram to show.
2. Label your diagram with the key components.
3. Support your diagram with text. The text should reference the diagram, highlight key aspects and explain anything that the diagram isn't able to show.

Let's look at how we would do this for the following question.

Question:

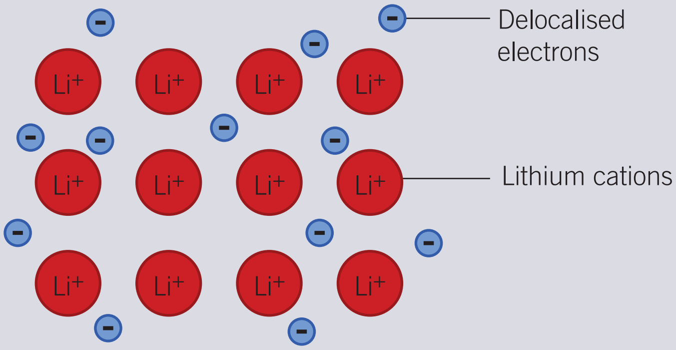
Use the metallic bonding model to explain the structure of lithium.

Comments:

1 The diagram shows a regular lattice of lithium cations in a sea of delocalised electrons. Lithium has one valence electron, so there will be equal numbers of lithium cations and electrons.
2 The components of the diagram, the lithium cations and the electrons, are labelled.
3 The text explains the arrangement of the lithium cations in a regular lattice structure. The valance electrons are delocalised and free to move. There is an electrostatic force of attraction between the cations and the delocalised electrons.

Considering all of this, your final response should look like this.

Answer:



The lithium cations are arranged in a regular lattice structure. The valance electrons are delocalised and free to move. There is an electrostatic force of attraction between the cations and the delocalised electrons.

Section 3A questions

1 Using the metallic bonding model, explain why metals are good conductors of electricity.
2 Provide an example of a property of metals that cannot be explained using the metallic bonding model.
3 Explain why it is unlikely for oxygen to form a metallic bond.
4 Draw a fully labelled diagram of a metallic lattice for aluminium.
5 Copper is commonly used for a variety of different applications. For each application below, state and explain in detail the property that makes it suitable for this use.
   a the base of a frypan
   b electrical wiring
   c water pipes

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3B Reactivity of metals

Study Design:
Experimental determination of a reactivity series for metals based on their relative ability to undergo oxidation with water, acids and oxygen

Glossary:
Core charge
Reactivity
Rust

ENGAGE

Figure 3B–1 shows a photo of two iconic Melbourne landmarks: the dome of Flinders Street Station and the flagpole support at the very top of the Royal Exhibition Building. Looking at these photos, which building do you think is older?

Figure 3B–1 The domes of Flinders Street Station (left) and the Royal Exhibition Building (right) provide an interesting insight into the reactivity of metals.

If you answered Flinders Street Station, you would be incorrect. The construction on Flinders Street Station finished in 1910, whereas the Royal Exhibition Building was completed in 1880. However, you would have been justified in assuming the opposite. We can see the original colour of the dome at Flinders Street Station poking through; however, it has mostly changed to a green colour now. The top of the Royal Exhibition Building dome, though, is still brilliantly golden in colour and looks brand new. The reason for this comes from the metals that they were constructed from. The Flinders Street Station dome is made from copper, whereas the ball from which the flag protrudes at the top of the Royal Exhibition Building dome is made from gold. While the copper dome has reacted over time with the environment to cause the change in colour that can be observed, the gold dome has not reacted and looks identical to how it originally was. It is this difference in reactivity of metals that we will be exploring in this section.

EXPLAIN

Reactivity in metals

1C THE PERIODIC TABLE LINK

In Chapter 1C, you learned that elements in the periodic table are ordered according to atomic number. As you move across the periodic table, the atomic number (the number of protons in the nucleus) increases by one – hydrogen has one proton, helium has two protons, and so on.

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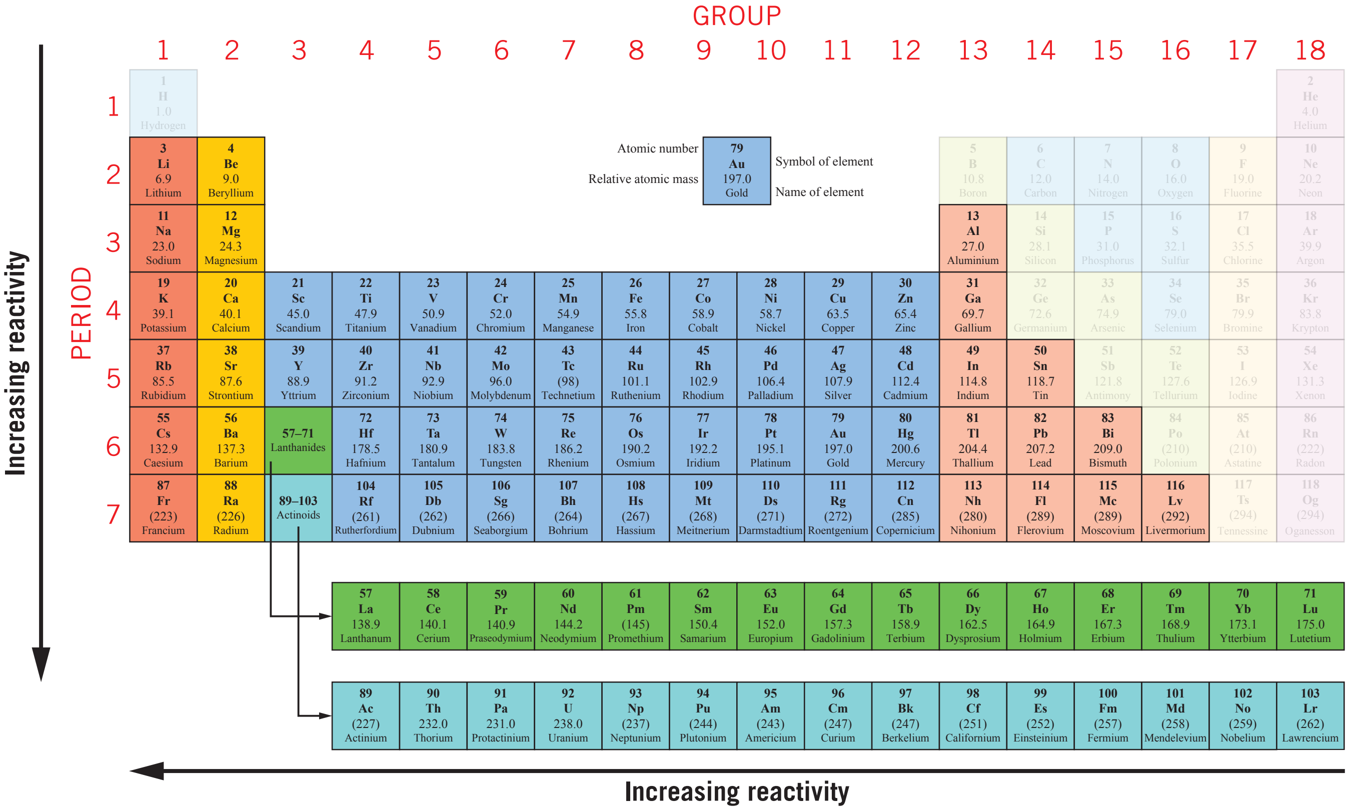
However, other patterns and trends exist within the periodic table. In fact, there are a number of properties of elements that increase or decrease when going across a period (rows) or down a group (columns). One of these general trends has to do with the reactivity of metals.

LINK **1C THE PERIODIC TABLE**



PERIOD	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1.0 Hydrogen																	2 He 4.0 Helium
2	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
3	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
4	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
5	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
6	55 Cs 132.9 Caesium	56 Ba 137.3 Barium	57–71 Lanthanides	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
7	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

Figure 3B–2 Periodic table indicating the general trend of increasing reactivity of metals

When metals react, they lose electrons. Therefore, the **reactivity** of a metal can be thought of as how readily a metal loses an electron. The more easily a metal loses an electron, the more reactive it is.

Reactivity how likely a chemical is to take part in a chemical reaction

As a general rule, the reactivity of metals increases going down a group. Looking at group 1, this means that potassium (K) is more reactive than sodium (Na), and caesium (Cs) is more reactive than potassium. This can be explained by thinking about the position of the valence electrons in relation to the nucleus of the atom. As we go down the group, the valence shell is further away from the nucleus (Figure 3B–3). Recall from Chapter 1C that this can also be related to the atomic radius and electronegativity. This means that the valence electrons contained within that shell have less electrostatic attraction to the nucleus and are therefore more easily removed. This ability to easily lose electrons is what explains the increased reactivity of metals going down a group.

LINK **1C THE PERIODIC TABLE**

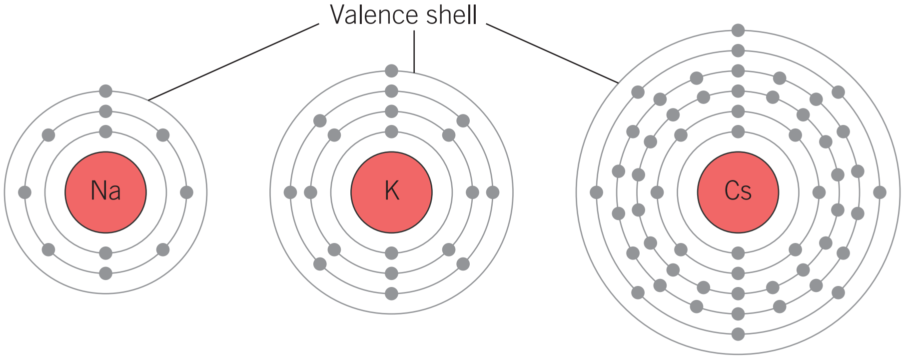


Figure 3B–3 As we move down a group, the number of electron shells in an atom increases. Therefore, there is less electrostatic attraction between the nucleus and the valence electrons. The weaker this attraction, the more easily the electrons can be removed.

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As we move across a period from left to right, the reactivity of metals decreases. This means that sodium (Na) is more reactive than magnesium (Mg), which is more reactive than aluminium (Al). As with the trend going down a group, this is based on how easily the metal will lose an electron.

1C THE PERIODIC TABLE **LINK**

Core charge the electrostatic force of attraction between valence electrons and the nucleus. It can be calculated by subtracting the number of non-valence electrons from the number of protons

To understand this, let's look at the electron configurations of Na, Mg and Al (Figure 3B–4). As you can see, unlike when going down a group, the valence shell for all of these atoms is the same (third shell). The thing that does change as we move across a period is that the **core charge** increases. An increased core change means an atom will less readily lose electrons and will therefore be less reactive.

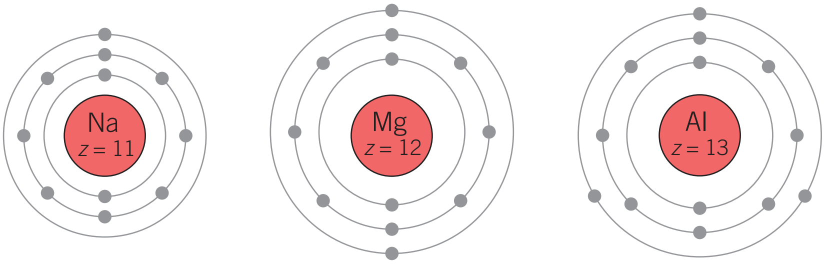


Figure 3B–4 As we move across a period, the core charge increases. Therefore, the electrostatic attraction between the nucleus and the valence electrons increases. The stronger this attraction, the more difficult it is to remove the electrons, and the less reactive the atom is.

NOTE

The letters in brackets following each reactant and product in chemical equations represents the states for those substances. There are four types of states, which are:

(s) = solid

(l) = liquid

(g) = gas

(aq) = aqueous

This will be explored more in 3B Skills and in Chapter 4C.

It is possible to determine the relative reactivities of metals through conducting experiments. We will now look at some common reactions that metals undergo and think about how they may be analysed to work out which metal is more reactive.

Reactions of metals and water

Metals react with water according to the general equation:

metal (s) + water (l) → metal hydroxide (aq) + hydrogen gas (g)

We can look at a specific example of this using sodium:

sodium (s) + water (l) → sodium hydroxide (aq) + hydrogen gas (g)

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

The reactivity of metals can be assessed based on the temperature of the water they react with:

- cold water – only reactive metals will react
- hot water – some less reactive metals that don't react in cold water will react.

NOTE

Unreactive metals won't react in either cold water or hot water.

4C PRECIPITATION REACTIONS **LINK**

VIDEO 3B–1 REACTIVITY OF METALS

Let's look at some example qualitative experimental data in Table 3B–1 that could be generated from experiments such as these. We will look at the relative reactivity of six metals: potassium, copper, zinc, magnesium, sodium and gold.

Table 3B–1 Experimental results for the reaction of different metals with cold and hot water

Water temperature	Reaction	No reaction
Cold water	Potassium Sodium	Magnesium Zinc Copper Gold
Hot water	Potassium Sodium Magnesium Zinc	Copper Gold

Based on these results, we can see that potassium and sodium are the two most reactive metals in this series, as these are the only two that react in cold water. It is important to note that we can't determine which of these two is more reactive based on these results. We can see that two of the metals, magnesium and zinc, do not react in cold water but do react with hot water. This makes them less reactive than potassium and sodium but more reactive than copper and gold, which do not react with either cold or hot water.

Therefore, we can construct a relative reactivity series that looks like this:

Reactions of metals and acid

Metals react with acids according to the general equation:

metal (s) + acid (aq) → salt (aq) + hydrogen gas (g)

Thus, we can look at a specific example of this using sodium and hydrochloric acid:

sodium (s) + hydrochloric acid (aq) → sodium chloride (aq) + hydrogen gas (g)

$$2Na(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2(g)$$

LINK **11E** NEUTRALISATION REACTIONS

For reactions of metals with acids, we can use the production of hydrogen gas to monitor the relative reactivity of the metals. As hydrogen gas is released, the solution will begin to bubble. Therefore, we can look at how vigorous the bubbling is to get a qualitative idea of the relative reactivity of the metals involved. Figure 3B–5 shows an example of an experiment that may be performed to test this.

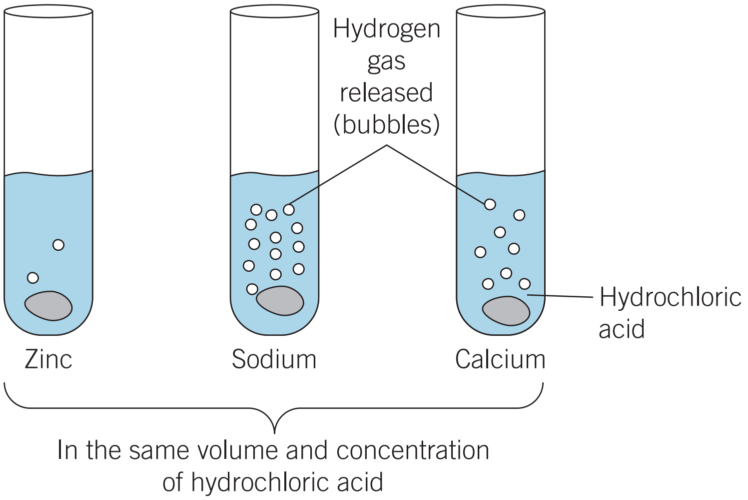
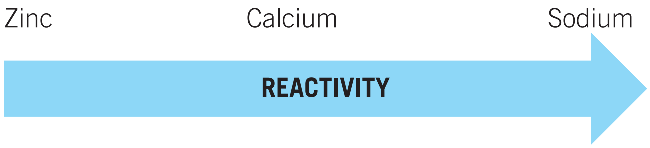


Figure 3B–5 Reactions involving zinc, sodium and calcium with hydrochloric acid. The rate of bubbling can provide an indication of the relative reactivity of the metals.

Figure 3B–5 illustrates that there is a big difference in the rate of bubbles produced between the reactions with different metals. This observation is due to the different reactivities of the metals. Sodium, which bubbles most vigourously, is the most reactive of the three metals tested. Zinc, which produces almost no bubbles, is the least reactive of these three metals. From this data, we can again devise a simple reactivity series for these metals:



Reactions of metals and oxygen

Metals react with oxygen gas (O_2) according to the general equation:

$$\text{metal (s)} + \text{oxygen gas (g)} \rightarrow \text{metal oxide (s)}$$

Thus, we can look at a specific example of this using iron:

$$\text{iron (s)} + \text{oxygen gas (g)} \rightarrow \text{iron(III) oxide (s)}$$

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

Rust
the product of corrosion, which involves iron reacting with oxygen and moisture in the air, converting the iron into iron oxides

The product of this reaction is iron(III) oxide. Although you may not have heard of it before, you have almost definitely seen it. We more commonly refer to the oxides of iron as **rust** (Figure 3B–6).

Figure 3B–6 The flaky red-brown coating on this iron frying pan is rust.

It is not quite as easy to monitor experiments using the reaction of metals with oxygen to determine the relative reactivity of metals. However, one method involves holding a small piece of the metal over a Bunsen burner. As the reaction takes place, a glow will appear in the flame. The brighter and larger this glow, the more reactive the metal. As an example, magnesium will glow so brightly within the Bunsen burner flame that you shouldn't look directly at it for too long. Zinc, on the other hand, will barely glow at all. Based on this, we could conclude that magnesium is a more reactive metal than zinc.

WORKSHEET 3B–1
DETERMINING THE REACTIVITY OF METALS USING EXPERIMENTAL DATA

By performing these experiments with many metals, a fuller reactivity series can be obtained, as shown in Table 3B–2.

Table 3B–2 Reactivity of metals with water, acids and oxygen

Metal	Symbol		Reactivity with water	Reactivity with acids	Reactivity with oxygen
Caesium	Cs	Most reactive	**Cs, Fr, Rb, K** Violent reaction with cold water **Na, Li, Ba, Ra, Sr, Ca** Moderate reaction with cold water Strong reaction with hot water	**Cs, Fr, Rb, K, Na, Li, Ba, Ra, Sr** Reacts explosively with dilute acids to give hydrogen gas **Ca** Reacts less vigorously	**Cs, Fr, Rb, K, Na, Li, Ba, Ra, Sr, Ca** Reacts with oxygen at ordinary temperature to form oxides
Francium	Fr				
Rubidium	Rb				
Potassium	K				
Sodium	Na				
Lithium	Li				
Barium	Ba				
Radium	Ra				
Strontium	Sr				
Calcium	Ca				
Magnesium	Mg	Moderately reactive	**Mg** Slowly reacts with cold water Strong reaction with hot water **Al, Ti, Mn, Zn, Fe** Moderate reaction with hot water **Ni, Sn, Pb** Does not react with water	**Mg** Reacts vigorously with dilute acids to give hydrogen gas **Al, Ti, Mn, Zn, Fe, Ni, Sn** Reacts moderately with dilute acids to give hydrogen gas **Pb** Reacts with concentrated acids to give hydrogen gas	**Mg, Ti, Mn, Zn, Fe, Ni, Sn, Pb** Reacts with oxygen on heating to form oxides
Aluminium	Al				
Titanium	Ti				
Manganese	Mn				
Zinc	Zn				
Chromium	Cr				
Iron	Fe				
Cadmium	Cd				
Cobalt	Co				
Nickel	Ni				
Tin	Sn				
Lead	Pb				
Hydrogen	H		Non-metal included for comparison		
Antimony	Sb	Least reactive	**Sb, Bi, Cu, W, Hg, Ag, Au, Pt** Does not react with water	**Sb, Bi, Cu, W, Hg, Ag, Au, Pt** Does not react with dilute acids	**Sb, Bi, Cu, W, Hg, Ag, Au, Pt** Does not react, even at high temperatures
Bismuth	Bi				
Copper	Cu				
Tungsten	W				
Mercury	Hg				
Silver	Ag				
Gold	Au				
Platinum	Pt				

3B SKILLS

Writing chemical equations

As you have seen in this section, writing equations for chemical reactions is an important part of this course. When writing chemical equations, there are a number of key things that you need to remember.

1. Reactants are found on the left of the equation; products are on the right.
2. You need to include coefficients to make a balanced equation.
3. You should always include states (that is, solid (s), liquid (l), gas (g) or aqueous (aq)).

VIDEO 3B–2
SKILLS: WRITING CHEMICAL EQUATIONS

Let's work through a question that would require us to apply these skills.

Question:

A piece of magnesium reacts with a solution of hydrochloric acid (HCl) to produce a solution of magnesium chloride ($MgCl_2$) and hydrogen gas. Write the chemical equation for this reaction.

Answer:

1 We are told that our reactants are magnesium and hydrochloric acid, so we will write these on the left-hand side of our equation. We are also told that our products are magnesium chloride and hydrogen gas, so these will go on the right-hand side of our equation. The left- and right-hand sides of the equation are separated by an arrow pointing from left to right:

$$\mathrm{Mg + HCl \rightarrow MgCl_2 + H_2}$$

2 This equation is currently unbalanced. We have one chlorine on the left and two on the right, as well as one hydrogen on the left and two on the right. We must balance the equation by including the appropriate coefficient(s) in front of reactants and/or products:

$$\mathrm{Mg + 2HCl \rightarrow MgCl_2 + H_2}$$

3 We also need to include states in this reaction. The question tells us that we have a piece of magnesium, which indicates it is a solid. Similarly, we are told that we have solutions of hydrochloric acid and magnesium chloride, which we will assume means they are both aqueous. Lastly, we are told that hydrogen is also produced as a gas:

$$\mathrm{Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)}$$

As you progress through your chemistry studies, you will get better at picking up these little hints found within worded questions, which will help you to write chemical equations. This book contains other Skills sections that will help you to determine the formulas for compounds and balancing equations.

Section 3B questions

1 When a metal reacts, does it gain or lose electrons?

2 What is one way that you could measure the relative reactivity of metals through their reactions with water? Explain how you would interpret the results to develop a reactivity series.

3 The following qualitative observations were made by scientists when adding metals to a solution of hydrochloric acid.

Metal	Observation
Sodium	Rapid bubbling, and the solid is consumed
Magnesium	Bubbling occurs, but more slowly than observed with sodium metal
Copper	No visible reaction
Iron	Bubbles appear very slowly

  a Use this information to order the metals in terms of increasing reactivity.

  b Describe how this determination of reactivity in part **a** was made.

4 Define 'reactivity series' as it relates to metals.

3C Metal recycling

Study Design:
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

Glossary:
Circular economy
Electrolysis
Ore
Refining
Slag
Smelting

ENGAGE

Did you know that Australia is one of the leading producers in the world for iron ore, lithium, gold, lead, aluminium ore (also known as bauxite) and zinc? In fact, the mining industry is so prominent in Australia that it is the largest single contributor to our economy. The most recent report by the Australian Bureau of Statistics showed that the gross domestic product (GDP) of the mining industry was $202 billion in a 12-month period, which equates to 10.4% of the national economy! Despite the economic importance of this industry, it can also be associated with negative environmental impacts. Considering this, there is also a lot of work being done to find ways to make the mining process more sustainable. In this section, we will explore this sustainability in more detail with a particular focus on the recycling of metals.

EXPLAIN

While we are surrounded by metals in our everyday life, we don't often think about the process that has happened to create our cookware, bathroom fittings, electronic equipment or our cars. As you can see in Figure 3C–1, there are a number of steps that occur before and after these products are made.

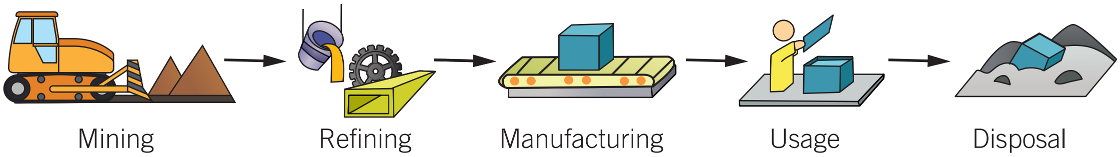


Figure 3C–1 The mining of metals for the production of goods has often been modelled as a linear economy, where the metals are mined, refined, manufactured into products, used and disposed of.

Mining of metals

Metals are found in rocks within Earth's crust. Depending on the location, different mining strategies have to be used. For metal deposits found close to the surface, quite simple techniques can be implemented. An example you may be familiar with is the panning technique used to search riverbeds for gold (Figure 3C–2).

Figure 3C–2 Some metals, for example gold, can sometimes be found through simple mining techniques such as panning in rivers.

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For other metals that are just below the surface, an open-cut mine can be used. In open-cut mining, the soil layer is removed and then the underlying rock layer is blasted (Figure 3C–3), such that it can then be easily removed by trucks. This may help uncover the resource of interest or to produce rocks that can be further refined.

For metals that are further below the surface, underground mining is required. In this situation, a shaft is drilled into the ground to the appropriate depth for where the metal has been located. From this point, the mine is then dug out horizontally so that the metal can be accessed and then transported back to the surface.

Figure 3C–3 Mining for metals can involve both the use of open-cut mines (left) and deeper underground mines (right).

The mining process has several negative impacts on the environment. These include the clearing of land to be used for a mine, which often involves deforestation.

It may also involve the loss of access to Country and its resources for First Nations Australians and the loss of cultural and sacred ancestral Indigenous sites. For example, in the Northern Territory, mining on Indigenous land contributes greater than 80% of the Territory's mineral value. One of these sites, the Newmont Tanami Granites Mine complex produces approximately 14 175 kilograms of gold per year. Mining complexes and newer developments are only possible in this region as a result of the *Aboriginal Land Rights (NT) Act 1976*. Part IV of this Act ensures that ancestral Indigenous land owners still have meaningful control over their land, while still allowing mining and energy interests to be set up with clear prior-informed decision-making. This ensures current Indigenous communities living in these mining areas can make mutually beneficial agreements with mining organisations. This often includes environmental protection for sacred sites but also financial compensation and employment or training opportunities. Nevertheless, there are often complex and difficult issues to resolve between mining companies and traditional owners of the land.

Ore
a deposit within Earth's crust that contains one or more valuable minerals, typically with a high concentration of a certain element (particularly metals)

Environmental impacts also include a loss of plant and animal life, decreasing the biodiversity of the area. As well as this, mining also results in significant levels of air and water pollution in the surrounding area. It is important to keep these factors in mind, as they form some of the key reasons why metal recycling (which we will discuss later in the chapter) is so critical.

3B REACTIVITY OF METALS LINK

The method for extracting metals depends largely on the metal's reactivity and position within the reactivity series.

Remember that unreactive metals, such as gold, can be found in the Earth's crust, not combined with other elements. However, when metals are mined from the ground, they are often not found as pure metals. Instead, they are found as **ores**. At this stage in the process,

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they need to be treated to obtain pure versions of the metal. This is referred to as extraction. The extraction process can involve different mechanisms, such as heating with other substances or undergoing various chemical reactions. See the Chapter 3C Skills section for more information about these. Another extraction process commonly used is smelting.

Smelting involves separating metals from other impurities by melting them at very high temperatures. It can have a number of negative impacts on the environment, particularly due to the emissions that are released during the process. These emissions can include sulfuric acid mist, which can enter the atmosphere and lead to the production of acid rain. Similarly, the water that is used to cool the ore after smelting must be disposed of. While some mechanisms are in place to treat this wastewater before disposal, if it is done incorrectly, it can end up contaminating water systems with toxic and hazardous compounds. Smelting also results in the production of **slag**, which contains several elements that can cause pollution to soil and water when released into the environment.

Smelting
the process of heating and melting an ore to extract a metal

Slag
a silicon dioxide and metal oxide mixture that is left over as a by-product from smelting

Refining
the elimination of impurities from metals to obtain highly pure versions of the metal

Refining of metal

The term **refining** refers to increasing the purity of the metal, through the elimination of other impurities.

One technique used in the process, discussed in the Chapter 3C Skills section, is called **electrolysis**. You will also look at this technique more in Chapter 12D. Other techniques include:

- distillation, the process of vaporising a substance and allowing it to condense in a separate location. This is good for metals with lower boiling points, such as mercury or zinc.
- liquation, where the impure metal is melted rather than vaporised, like in distillation. This is good for metals with low melting points, such as lead and tin.
- chromatographic methods, where different components of a mixture are adsorbed at different levels of a column. Chromatography is explored more in Chapter 5B.

LINK **12D** APPLICATIONS OF REDOX REACTIONS IN SOCIETY

LINK **5B** CHROMATOGRAPHY

Manufacturing: making metals into a product

Once the purified metal is obtained from the refining process, it is then ready for manufacturing. During the manufacturing process, the metals are used to make a large number of the products that we use in our everyday lives. The manufacturing of metallic products takes advantage of the properties of metals that were discussed earlier in this chapter. In particular, their malleability and ductility allow them to be hammered out into large thin sheets, which can then be cut into any required shape. As well as this, they can be melted and placed into a mould, where they are allowed to solidify into the desired dimensions.

LINK **3A** STRUCTURE AND PROPERTIES OF METALS

Moving to a circular economy

In Figure 3C–1, we looked at the linear economy that exists for the production of metallic consumer goods, including the stages of mining, refining, manufacturing, usage of the products and disposal. While this process is well established, it has a few drawbacks, including the following:

- Many of the steps have negative environmental consequences. In particular, mining leads to deforestation, loss of habitats for plants and animals, and air and water pollution.
- All of the metals that are mined are present in finite quantities, and many of these metals are also quite rare, such as gold and copper.
- The mining and refining processes require a large energy input, which further depletes non-renewable sources of energy.

VIDEO 3C–1 COMPARISON OF LINEAR AND CIRCULAR ECONOMIES

Circular economy an economy that functions on a continuous cycle with a focus on optimising the re-use of resources and materials

1D CRITICAL ELEMENTS AND RECYCLING PROCESSES LINK

For these reasons, it has become the preferred option to adopt a **circular economy** method for metal products (Figure 3C–4). As you can see, the main difference in the circular economy model is the introduction of metal recycling. When recycling is used, it bypasses the need for mining and extraction (and in some cases refining as well). By not requiring these steps, the process becomes much more environmentally friendly, is more energy and cost-efficient and helps preserve the limited supply of metals.

While recycling requires specific processes for different metals, many of the steps are the same. Generally, the strategy involves separating the metal from any other components (such as glass or plastic), melting at a high temperature and then placing it in a mould or letting it solidify as a large sheet. At this point, the metal is available to re-use and can be reprocessed into the same product or repurposed as an entirely new product.

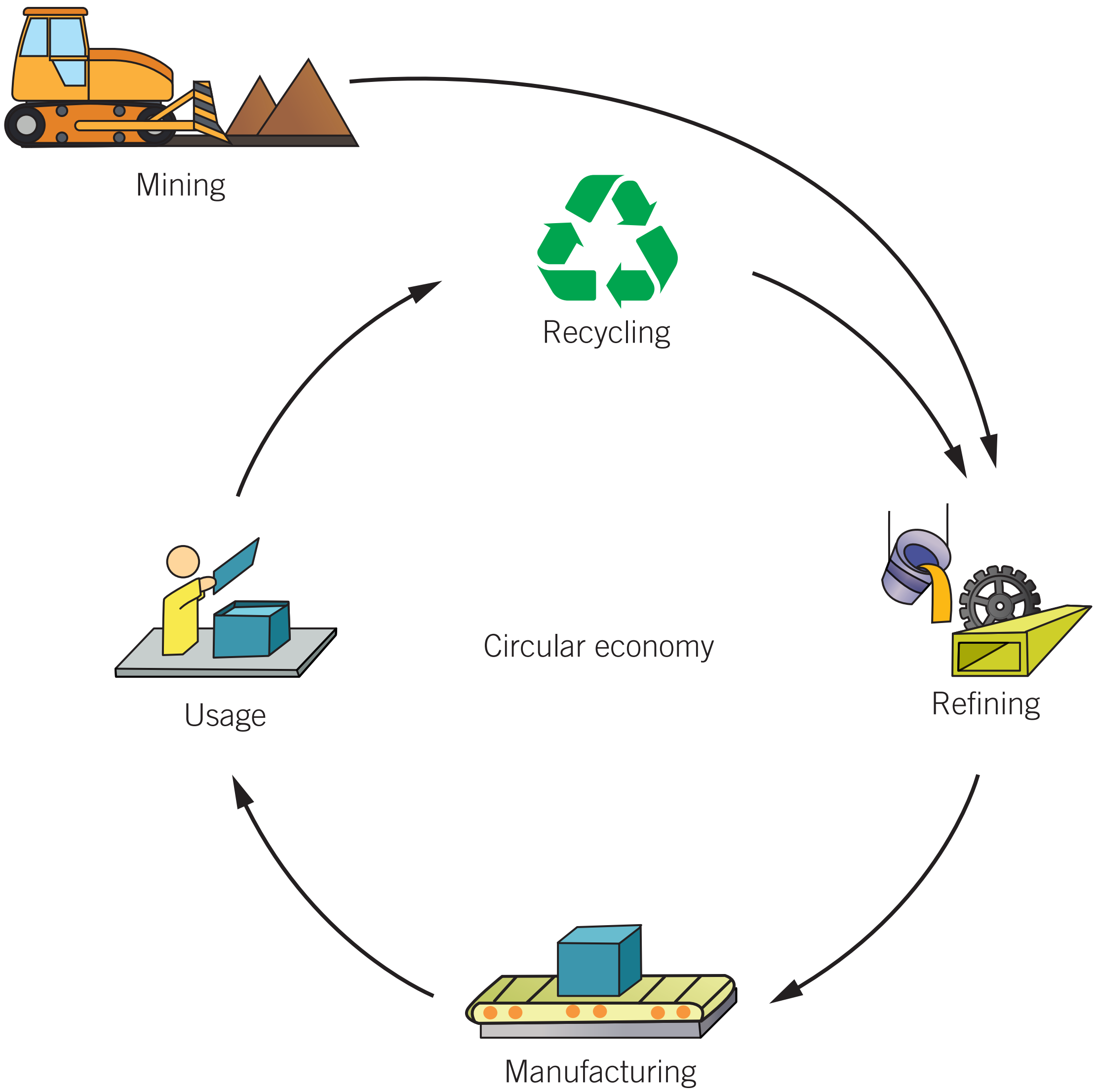

Figure 3C–4 The general steps involved in the circular economy process for metals

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3C SKILLS

VIDEO 3C–2 SKILLS: CHEMICAL REACTIONS INVOLVED IN THE EXTRACTION OF METALS

The chemical reactions involved in the extraction of metals

As mentioned in this section, the extraction of metals from their compounds is largely dependent on their position in the reactivity series. A reactivity series for some metals is shown in Figure 3C–5.

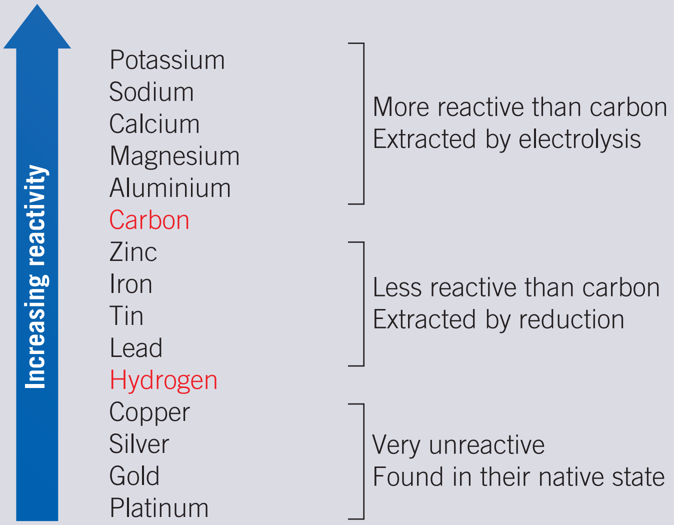

Figure 3C–5 The reactivity series of metals. This series also shows the different extraction methods used depending on a metal's position in the series relative to hydrogen and carbon.

The least reactive metals, such as platinum, gold and silver, are found as uncombined elements in Earth's crust.

Metals more reactive than hydrogen but less reactive than carbon, such as iron and zinc, can be extracted by reacting (reducing) the metal ore with carbon. Reduction helps to remove traces of oxygen from the ore, as the carbon is more reactive than iron and can displace the iron from its oxide (oxygen). For iron, this is done by reacting iron(III) oxide with carbon monoxide at very high temperatures in a blast furnace, resulting in the following reaction:

$$\text{iron(III) oxide} + \text{carbon monoxide} \rightarrow \text{iron} + \text{carbon dioxide}$$

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$$

As a result, molten (liquid) iron flows to the bottom of the furnace where it can be collected, and the carbon dioxide gas is released as a waste gas from the top of the furnace. While this process is inexpensive, it does result in large amounts of carbon dioxide emissions, which contribute greatly to greenhouse gas emissions. It is therefore environmentally unfriendly.

LINK **UNIT 3**

LINK **4A** FORMATION AND NAMING OF IONIC COMPOUNDS

Metals more reactive than carbon, such as aluminium, sodium and potassium, need to be extracted using **electrolysis**. It cannot be done by reacting with carbon as the carbon is not reactive enough to displace the metal. Electrolysis is much more expensive and uses a lot of electricity. As you will learn in Chapter 4A, aluminium oxide is an ionic compound. It is insoluble in water and has a high melting point. As such, a lot of energy would be required to break these bonds, which is expensive. Therefore, powdered aluminium ore is first dissolved in cryolite, which allows it to melt at a lower temperature.

Electrolysis the process of forcing a redox reaction to occur by passing an electric current through an electrolyte

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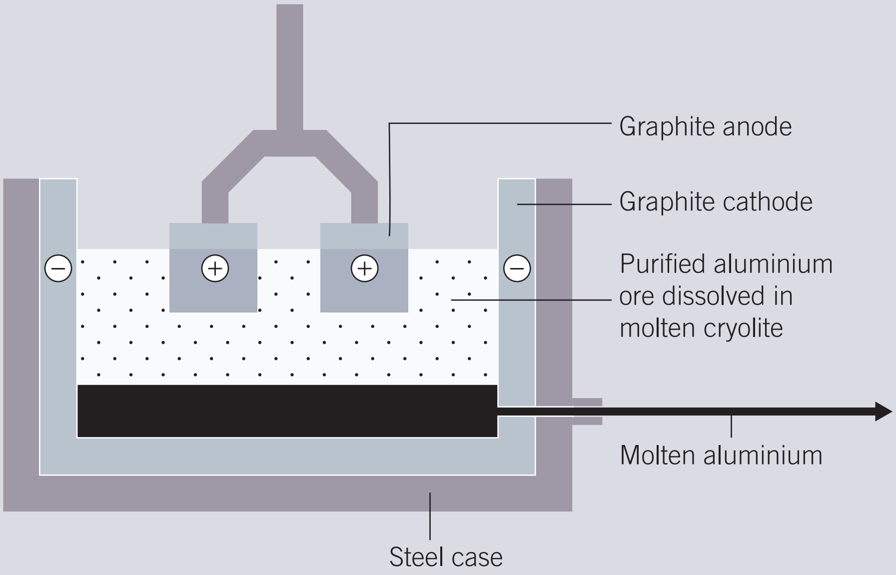

Figure 3C–6 The electrolysis of aluminium ore into molten aluminium

2D DIAMOND AND GRAPHITE AS COVALENT COMPOUNDS LINK

You will recall from Chapter 2D that graphite is a large, layered carbon compound. Graphite is used as the anode and cathode due to its very high melting point and ability to conduct electricity thanks to its delocalised electrons. The reactions occurring at each electrode are:

At anode (oxidation)	At cathode (reduction)
$2O^{2-}(aq) \rightarrow O_2(g) + 4e^-$	$Al^{3+}(aq) + 3e^- \rightarrow Al(l)$

12A REDUCTION AND OXIDATION LINK

The reduction reaction occurring at the cathode results in molten aluminium being produced.

12C METAL DISPLACEMENT REACTIONS LINK

These displacement reactions and oxidation/reduction (redox) reactions will be explored in more detail in Chapters 12A and 12C. It will provide a good reflection back to the knowledge of metals you have developed here.

Section 3C questions

1. State two types of mines and explain what determines which one is used.
2. Extraction is an important part of the metal production process.
   a. What is the purpose of extracting a metal?
   b. Name and explain one technique used to extract metals.
3. List three environmental consequences of mining.

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Chapter 3 review

Summary

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

Success criteria – I am now able to:		Linked question
3A.1	Describe the metallic bonding model	2☐
3A.2	Draw a labelled diagram of a metallic lattice	11☐
3A.3	Using the metallic bonding model, explain common properties of metals	1☐, 3☐, 6☐
3B.1	Determine the reactivity series of metals given experimental data	13☐
3B.2	Predict the outcomes of chemical reactions when given a reactivity series of metals	4☐, 10☐
3B.3	Design an experiment to test the reactivity of a series of metals	8☐, 12☐
3C.1	Define ‘circular economy’	9☐, 14☐
3C.2	Describe the steps involved in metal processing, from mining to repurposing	5☐, 14☐
3C.3	Explain the reasons for recycling metals	7☐, 14☐

Multiple-choice questions

1 Which of the following is **not** a common property of metals?
- **A** ductile
- **B** low melting point
- **C** lustrous
- **D** malleable

2 Which of the following describes the force that holds together a metallic lattice?
- **A** electrostatic attraction between cations and anions
- **B** sharing of electrons between neighbouring atoms
- **C** electrostatic attraction between neighbouring cations
- **D** electrostatic attraction between cations and delocalised electrons

3 Metals are good conductors of electricity because they
- **A** are malleable.
- **B** have a high melting point.
- **C** have freely moving cations.
- **D** have freely moving electrons.

4 As a general rule, the reactivity of metals

A increases down a group.
B decreases down a group.
C is the same for all metals in the same group.
D mostly depends on the other reagent involved in the reaction.

5 Which of the following is **not** a step associated with refining of metals?

A distillation
B reacting with chemicals
C washing with chemicals
D extraction of metals from the ground

6 Copper is able to be manufactured into long, thin wires. This is mostly due to its

A lustre.
B high ductility.
C high melting point.
D high electrical conductivity.

7 Which of the following is **not** a negative environmental impact of mining?

A increased deforestation
B increased air pollution
C increased water pollution
D increased biodiversity

8 Magnesium and calcium are added to separate test tubes with sulfuric acid. One way to tell which metal is most reactive would be to measure the

A colour change of the solution over time.
B volume of gas produced over time.
C starting mass of the magnesium and calcium.
D starting concentration of the sulfuric acid.

9 Which of the following steps is present in a circular but not a linear economy of metal products?

A mining
B refining
C manufacturing
D recycling

10 If a metal reacts with water, the gas most likely to be produced is

A hydrogen.
B nitrogen.
C oxygen.
D carbon dioxide.

Short-answer questions

11 Sketch a diagram of the metallic lattice for aluminium. Ensure that your diagram is fully labelled. (4 marks)

12 Describe an appropriate experiment that could be conducted to determine the relative reactivity of lithium, beryllium and ytterbium with water. In your response, refer to variables that are controlled, changed and measured. (5 marks)

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13 A scientist has samples of four unidentified substances. She labels them X1, X2, X3 and X4. As a first step in identifying these samples, she determines some properties of each. She performs five tests on each of the samples. The results are presented below.

Element	Ductility	Melting point	Electrical conductivity	Heat conductivity	Malleability
X1	Moderate	High	High	High	Moderate
X2	Low	Low	Low	Moderate	High
X3	Moderate	Moderate	Moderate	High	High
X4	High	High	High	Moderate	High

a Which substance would be the best choice to be used as an electrical wire? Why? (2 marks)

b Which substance is the least likely to be a metal? Why? (3 marks)

c Substance X3 is being considered for the production of saucepans. What two properties of X3 would make it suitable for this application? (2 marks)

d What extra piece of information would you want to know about substance X3 before you decided to start using it to make saucepans? (2 marks)

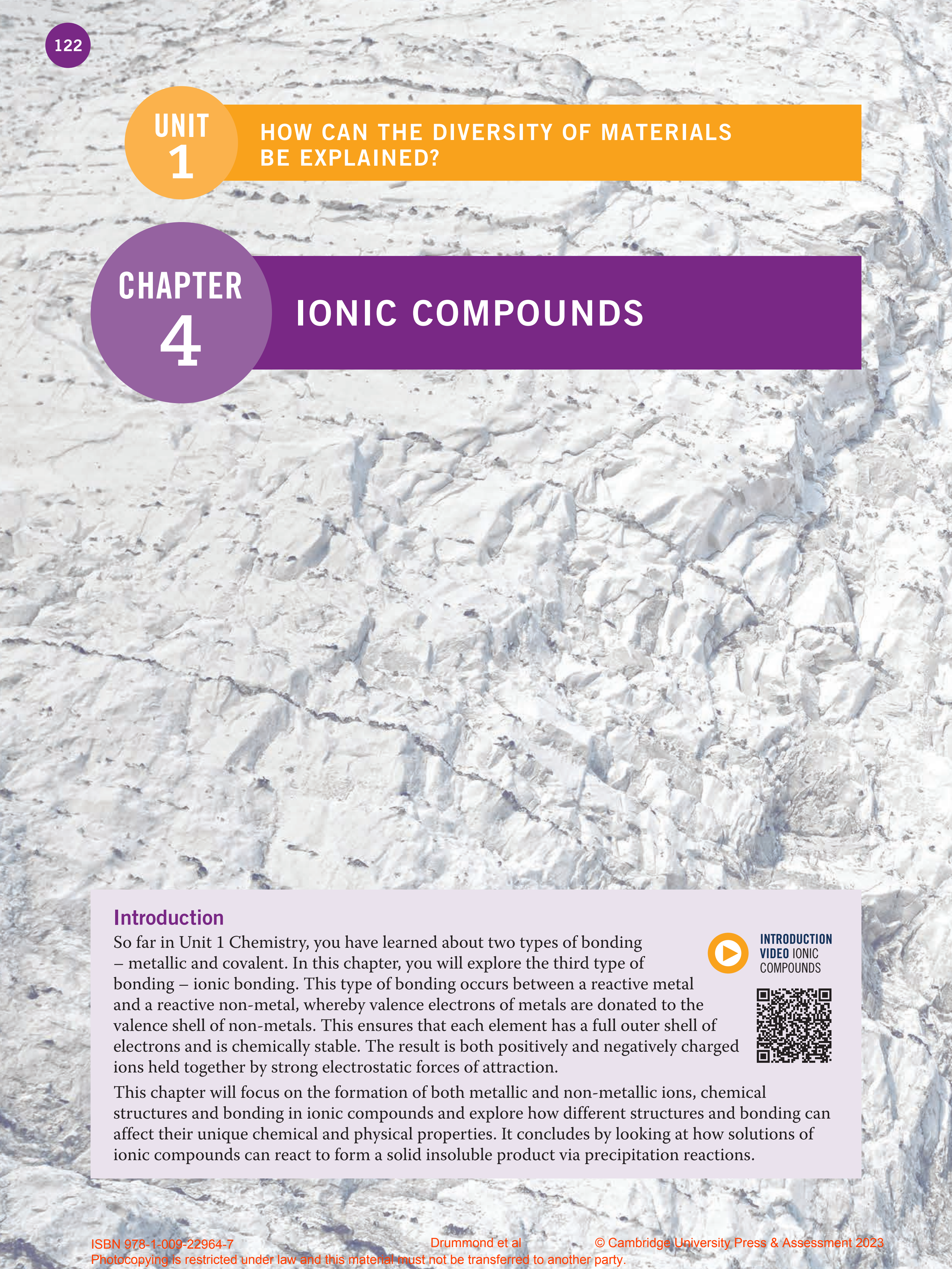
14 Bauxite is an ore that is rich in aluminium. A new bauxite deposit has been discovered in Queensland, just 10 metres below the surface.

a Outline the steps that would be necessary to go from this discovery through to the production of an aluminium soft-drink can. (Hint: You should use the Skills section in Chapter 3C to assist with aspects of your response here.) (3 marks)

b Define ‘circular economy’. (1 mark)

c You consume a soft drink from an aluminium can. State the next step that should occur here to ensure that the circular economy model is applied to aluminium production. (1 mark)

d Explain two environmental benefits of a circular economy model. (4 marks)



UNIT 1

HOW CAN THE DIVERSITY OF MATERIALS BE EXPLAINED?

CHAPTER 4

IONIC COMPOUNDS

Introduction

So far in Unit 1 Chemistry, you have learned about two types of bonding – metallic and covalent. In this chapter, you will explore the third type of bonding – ionic bonding. This type of bonding occurs between a reactive metal and a reactive non-metal, whereby valence electrons of metals are donated to the valence shell of non-metals. This ensures that each element has a full outer shell of electrons and is chemically stable. The result is both positively and negatively charged ions held together by strong electrostatic forces of attraction.

INTRODUCTION VIDEO IONIC COMPOUNDS

This chapter will focus on the formation of both metallic and non-metallic ions, chemical structures and bonding in ionic compounds and explore how different structures and bonding can affect their unique chemical and physical properties. It concludes by looking at how solutions of ionic compounds can react to form a solid insoluble product via precipitation reactions.

Curriculum

Area of Study 1 Outcome 1
Reactions of ionic compounds

Study Design:	Learning intentions – at the end of this chapter I will be able to:
• 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 • 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-})	**4A** **Formation and naming of ionic compounds** **4A.1** Recall and apply the idea that an ion is charged (anion and cation), meaning it has a differing number of protons and electrons **4A.2** Describe how ions are formed, with regard to electrons being gained and lost and how this allows an atom to achieve a more stable electron configuration (full outer shell) **4A.3** Explain how ionic compounds are made up of positive and negative ions but the overall charge is neutral **4A.4** Recall that the positive ion in an ionic compound is written and named first in the formula **4A.5** Determine the name and ionic formula of ionic compounds using the periodic table **4A.6** Determine the charge on an ion given an ionic formula **4A.7** Determine the charge of a main group ion (using group number) **4A.8** Write the formula of ionic compounds involving polyatomic ions and transition metals **4A.9** Determine the charge on a transition metal when given an ionic formula
• 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	**4B** **Structure and properties of ionic compounds** **4B.1** Recall that the size of ions and the charge on ions affects the melting point and boiling point of an ionic compound **4B.2** Recall that generally ionic compounds are hard, brittle and can conduct in a molten liquid state and in aqueous solution, but not a solid state **4B.3** Use the ionic bonding model to explain the difference in melting point of two different ionic compounds (must have one different ion) **4B.4** Compare and contrast properties of two different ionic compounds
• 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)	**4C** **Precipitation reactions** **4C.1** Recall that a precipitation reaction involves mixing solutions of two different ionic compounds **4C.2** Recall the solubility rules (SNAPE, BICS, CHPS) to determine if an ionic compound formed will be solid – that is, not in solution **4C.3** Recall what an ionic equation is and how it differs from a full precipitation reaction **4C.4** Write and balance chemical equations, including states, for a precipitation reaction **4C.5** Write and balance ionic equations for precipitation reactions

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Glossary

- Anion
- Brittle
- Cation
- Dissolution
- Double displacement reaction
- Electrostatic force of attraction
- Hardness
- Insoluble
- Ion
- Ionic bond
- Ionic compound
- Ion–dipole attraction
- Ionic equation
- Lattice
- Polyatomic ion
- Precipitate
- Precipitation reaction
- Solubility
- Soluble
- Solute
- Solvent
- Spectator ion
- Valence shell

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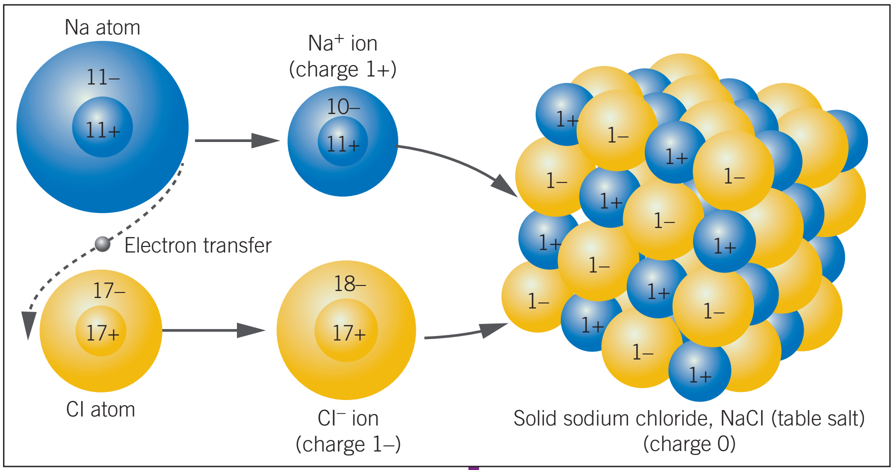
Concept map

Atoms can gain or lose electrons to become more stable.

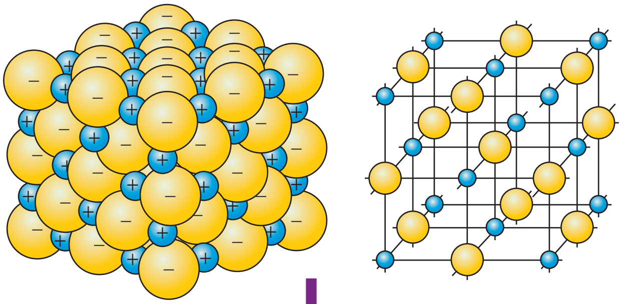
Ionic compounds form by the transfer of electrons from metals to non-metals.

4A Formation and naming of ionic compounds



Common properties of ionic compounds are explained by the ionic bonding model and formation of crystal structure.

4B Structure and properties of ionic compounds



Soluble ionic compounds can react in solution to form new insoluble ionic compounds, which can be predicted with solubility tables.

4C Precipitation reactions

See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

4A Formation and naming of ionic compounds

Study Design:

- 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
- 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-})

Glossary:

Anion
Cation
Electrostatic force of attraction
Ion
Ionic compound
Lattice
Polyatomic ion
Solute
Solvent
Valence shell

ENGAGE

The discovery of ions

In 1830, Michael Faraday provided the first known theory for the discovery of ions. He was the first to use the term *ion* (meaning 'wanderer' in Greek) and the terms *anion* and *cation*. His experiments uncovered that particular substances could conduct electricity when dissolved in water.

However, these ideas were not fully explained until much later, in 1884, by Svante August Arrhenius. His research and experimentation showed that substances (**solutes**) broke up into charged ions when placed in a **solvent**, regardless of whether an electrical current was applied to the solution.

Solute
the solid, liquid or gas that dissolves in a solvent

Solvent
the substance that the solute dissolves in

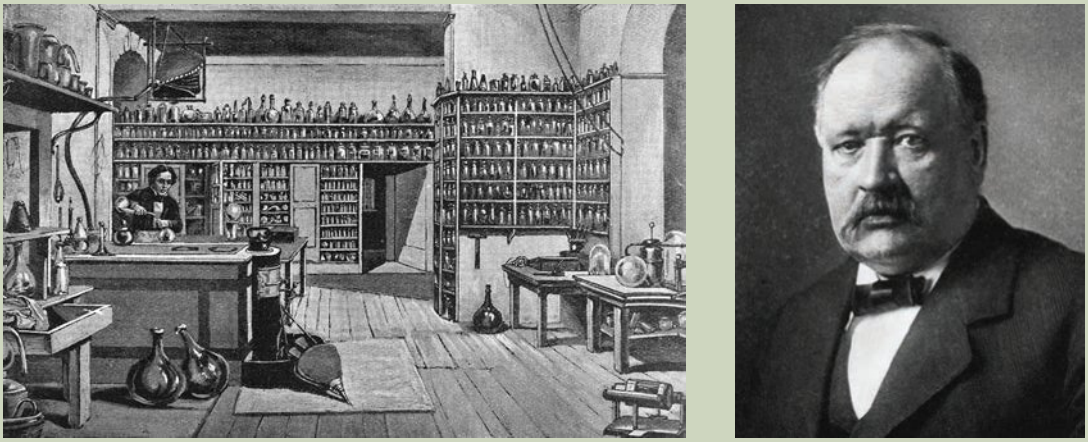


Figure 4A–1 Michael Faraday (1791–1867) in his laboratory (left) and Svante Arrhenius (1859–1927) (right)

EXPLAIN

What is an ion?

1A INTRODUCTION TO THE ELEMENTS LINK

As you will recall from Chapter 1A, changing the number of electrons in an atom does not change its atomic number or the element it is. However, it results in a change to its chemical properties. It is now referred to as an ion.

Ions are atoms that have either lost or gained electrons to become more stable. These electrons are lost from the outermost shell, or **valence shell**, of the atom, as these electrons are the most weakly attracted to the protons within the nucleus.

Ions are denoted by a ‘+’ or a ‘–’ sign as a superscript after the chemical symbol. A ‘+’ indicates the atom has become more positive (lost negatively charged electrons); whereas a ‘–’ indicates the atom has become more negative (gained negatively charged electrons). For example, Mg^{2+} denotes a magnesium atom that has donated two electrons to another atom and is now carrying an overall 2+ charge. Magnesium has 12 protons in the nucleus, but as it has given away two electrons it now only has 10 electrons. Atoms tend to lose or gain electrons according to how many electrons they possess in their valence shell.

Ion
a positively or negatively charged atom that has either lost or gained electron(s)

Valence shell
the outermost shell of an atom

Cation
an atom that has lost electron(s) from its valence shell, resulting in an overall positive charge

Cations

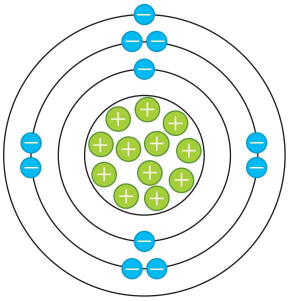
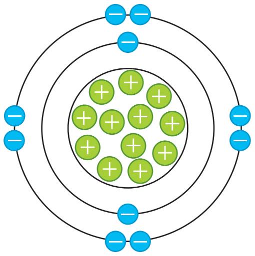
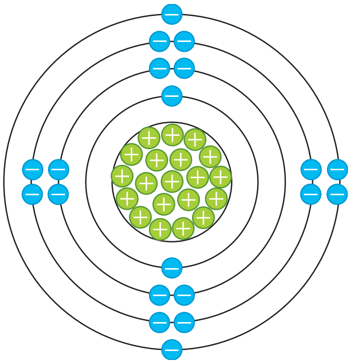
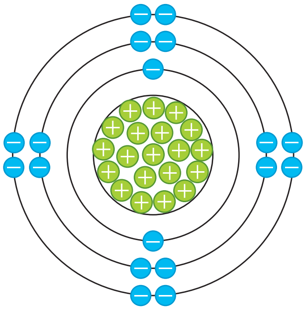
When atoms lose electrons, the result is that there are now more protons (positive charge) than electrons (negative charge), and therefore an overall positive charge exists. These are referred to as **cations**. You can remember this through the saying,

‘Cats have paws, so cat-ions are paws-itive.’

Metals lose electrons to form cations, as you learned in Chapter 3A. These are the metals with one, two or three valence electrons (group 1, group 2 and group 13, respectively) or some of the transition metals that you will see later can lose varying numbers of electrons. Table 4A–1 shows the result of atoms drawn using the Bohr model of losing electrons.

LINK **3A** STRUCTURE AND PROPERTIES OF METALS

Table 4A–1 Formation of positive ions (cations)

Starting atom	Result from loss of electron(s)	Explanation and result
Sodium atom  Electron configuration: 2,8,1	Sodium ion (Na^+)  Electron configuration: 2,8	Sodium has one electron in its valence shell, so it is easier to lose one electron than gain seven. It now has one less electron compared to the number of protons. Therefore, it has a ‘+’ charge with two full shells of electrons.
Calcium atom  Electron configuration: 2,8,8,2	Calcium ion (Ca^{2+})  Electron configuration: 2,8,8	Calcium has two electrons in its valence shell, so it is easier to lose two electrons than gain six. It now has two less electrons compared to the number of protons. Therefore, it has a ‘2+’ charge with three full shells of electrons.

VIDEO 4A–1 FORMATION OF POSITIVE IONS

Anions

Anion
an atom that has gained electron(s), filling its valence shell and resulting in an overall negative charge

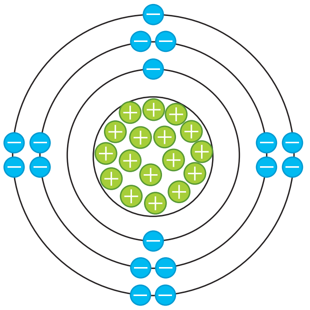
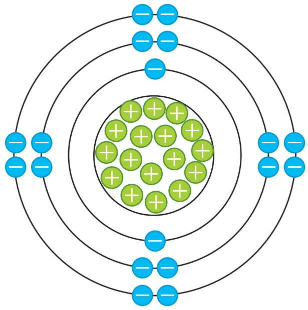
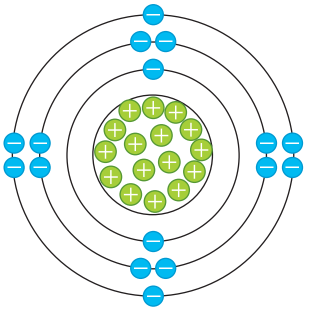
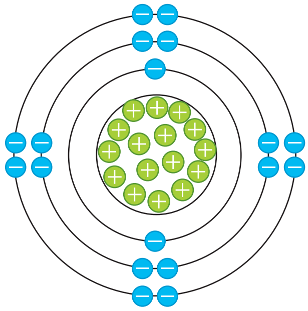
When atoms gain electrons, the result is that there are now more electrons (negative charge) compared to protons (positive charge), and therefore an overall negative charge exists. These are referred to as **anions**. You can remember this through realising the word ‘anions’ looks similar to ‘onions’, and the saying,

‘Onions make you cry, so anions are negative.’

The non-metals in group 15, group 16 and group 17 have five, six or seven electrons in their valence shell, respectively, so they gain electrons to form anions. Table 4A–2 shows atoms drawn using the Bohr model and the result of them gaining electrons.

VIDEO 4A–2
FORMATION OF NEGATIVE IONS

Table 4A–2 Formation of negative ions (anions)

Starting atom	Result from gain of electron(s)	Explanation and result
Chlor**ine** atom  Electron configuration: 2,8,7	Chlor**ide** ion (Cl^-)  Electron configuration: 2,8,8	Chlorine has seven electrons in its valence shell, so it is easier to gain one electron than lose seven. It now has one extra electron compared to the number of protons. Therefore, it has a ‘–’ charge with three full shells of electrons.
Sulf**ur** atom  Electron configuration: 2,8,6	Sulf**ide** ion (S^{2-})  Electron configuration: 2,8,8	Sulfur has six electrons in its valence shell, so it is easier to gain two electrons than lose six. It now has two extra electrons compared to the number of protons. Therefore, it has a ‘2–’ charge with three full shells of electrons.

To determine if an atom will gain or lose electrons, you look for what will result in the movement of the least number of electrons to achieve a full valence shell.

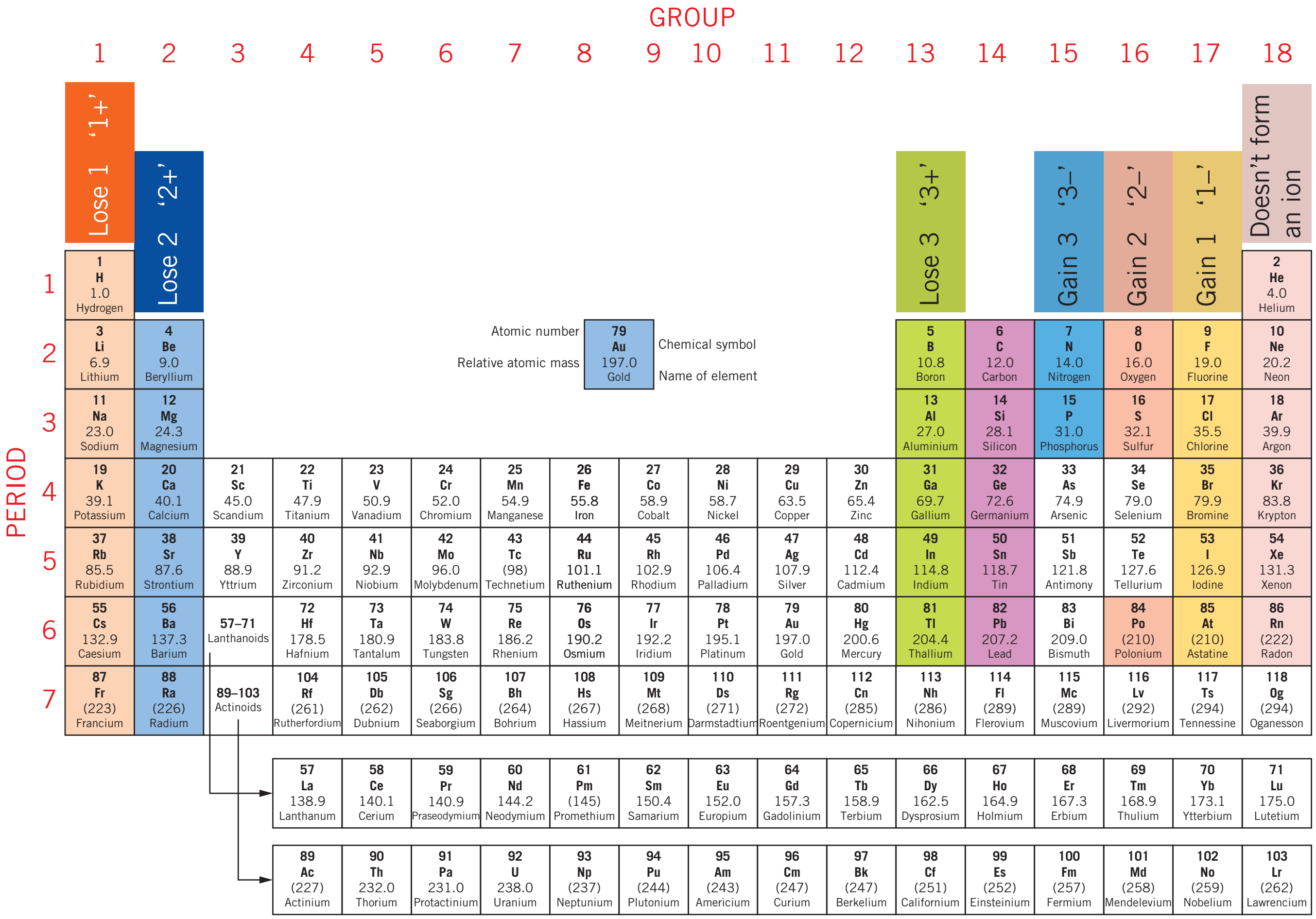
NOTE

The ending of the name in non-metals when they lose electrons and form negative anions changes. The suffix of the element’s name is replaced with ‘ide’. In the examples above, for chlorine, ‘-ine’ is replaced by ‘-ide’ and for sulfur, ‘-ur’ is replaced by ‘-ide’.

Elements that do not form ions

There is one group of elements, the noble gases in group 18 of the periodic table, shown in Figure 4A–2, that do not form ions. These elements already have a full valence shell of electrons and are therefore stable and usually unreactive.

A summary of the number of electrons lost or gained and the resulting ion charges formed is also shown in Figure 4A–2.



PERIOD	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
	Lose 1 '1+'	Lose 2 '2+'											Lose 3 '3+'		Gain 3 '3–'	Gain 2 '2–'	Gain 1 '1–'	Doesn't form an ion
1	1 H 1.0 Hydrogen																	2 He 4.0 Helium
2	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
3	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
4	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
5	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
6	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
7	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 (286) Nihonium	114 Fl (289) Flerovium	115 Mc (289) Muscovium	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

Figure 4A–2 Periodic table indicating the main group elements and the charges of the ions formed by these element's atoms when they lose or gain electrons. Note that both groups 14 and 18 (noble gases) typically do not form ions.

The other main group of elements that you will notice does not have an associated ion charge is group 14. The atoms in this group, consisting of elements such as carbon and silicon, contain four electrons in their valence shell. In this case, the element's atoms could either lose or gain four electrons. But, in most instances, the non-metallic elements tend to share electrons and form covalent compounds, as you learned in Chapter 2A.

Forming ionic compounds

Ionic compounds are substances composed of both positive and negative ions. This occurs when a metal atom loses its electron(s), transferring them to the non-metal atom, which gains the electron(s). The result is a positive metal cation and a negative non-metal anion. These ions then arrange themselves into an overall neutral three-dimensional **lattice** structure of alternating cations and anions held together by strong **electrostatic forces of attraction**, as shown in Figure 4A–3, with sodium and chlorine atoms forming ions that are the basis of sodium chloride. You will recall from Chapter 3A that an electrostatic force of attraction was also established between metal cations and the sea of delocalised electrons in the metallic bonding model. The ionic bonding model will be explored further in Chapter 4B.

LINK **2A** REPRESENTING COVALENT COMPOUNDS

LINK **3A** STRUCTURE AND PROPERTIES OF METALS

LINK **4B** STRUCTURE AND PROPERTIES OF IONIC COMPOUNDS

Ionic compound
a compound or molecule formed when positive cations and negative anions chemically combine

Lattice
an organised arrangement of atoms or ions into fixed positions in a 3D structure

Electrostatic force of attraction
the attraction between positively and negatively charged particles that are not directly in contact

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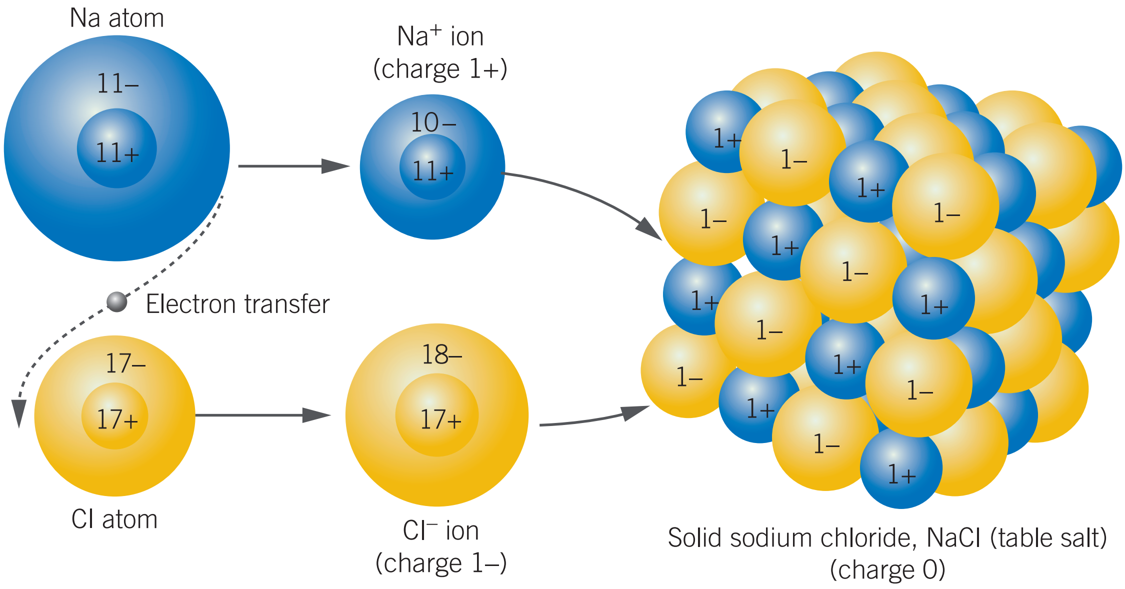

Figure 4A–3 Sodium chloride, or table salt, formed from the combination of alternating sodium and chloride ions

Electron transfer diagrams

Electron transfer diagrams are used to show the path that electrons take when they are lost from a metal atom and gained by a non-metal atom during ionic bond formation, as shown with sodium and sulfur in Figure 4A–4.

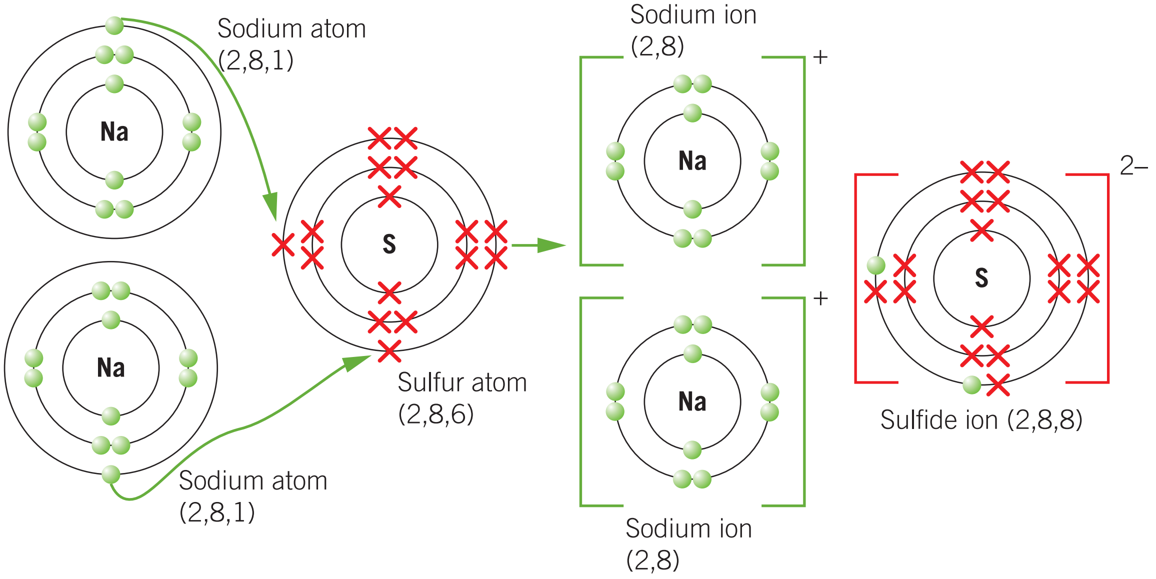

Figure 4A–4 The transfer of electrons between sodium atoms and a sulfur atom. Note that as a sodium atom can only lose one electron but sulfur needs to gain two electrons to obtain a full valence shell, two sodium atoms are required.

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In Figure 4A–4, sulfur requires two electrons for it to achieve a full outer shell. However, a single sodium atom can only lose (or donate) one electron. Therefore, to balance the overall charge in the formation of this ionic compound, two sodium atoms are required, with each donating a single electron to the sulfur atom, as shown by the arrows. This results in:

- each sodium atom forming a 1+ sodium ion
- sulfur becoming a 2– sulfide ion (note that the suffix of the anion has changed to '-ide')
- the ionic compound formed having two sodium ions for every one sulfide ion (Figure 4A–4).

As you work towards being able to determine the formula of an ionic compound, these electron transfer diagrams, while time-consuming to draw, are an excellent way to determine how many cations and anions are required to form a neutral ionic compound as a result of the movement of electrons.

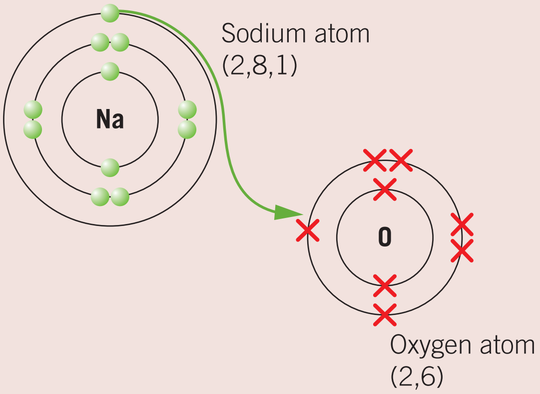
Worked example 4A–1: Drawing electron transfer diagrams

VIDEO
WORKED EXAMPLE 4A–1

Question

Draw an electron transfer diagram to show what would happen when sodium and oxygen atoms react together.

Solution

Step 1	Starting with the metal atom (sodium), recall the group and period it is in.	Group 1 (one electron in its valence shell) Period 3 (will have three shells of electrons)
Step 2	Recall what group and period the non-metal atom (oxygen) is in.	Group 16 (six electrons in its valence shell) Period 2 (will have two shells of electrons)
Step 3	Draw the starting atoms for sodium and oxygen and indicate the transfer of electrons from metal (sodium) to non-metal (oxygen) using an arrow.	
Step 4	Determine if more than one metal or non-metal is needed to result in a full valence shell of electrons in the non-metal.	Two sodium atoms are needed, each transferring one electron to the oxygen atom, which requires two electrons.

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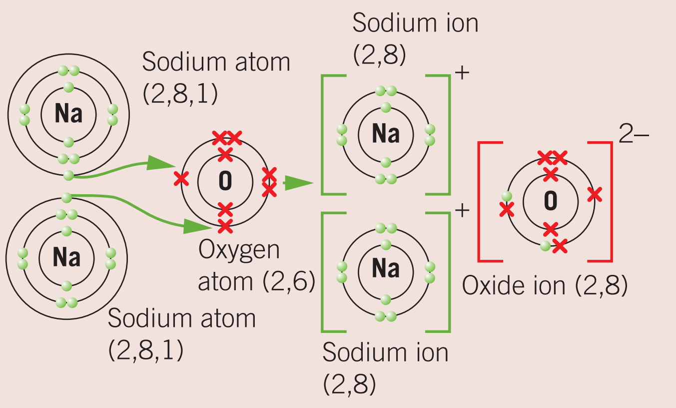
Step 5	Draw the resulting ions formed in the ionic compound. (Note that the electron configuration of both sodium and oxygen ions is now the same as a neon atom, a group 18 noble gas.)	

NOTE

In this worked example, electrons are shown as dots in the metal atom/cation and as crosses in the non-metal atom/anion. This allows us to easily visualise the transfer of electrons from metal to non-metal.

Check-in questions – Set 1

1 Explain the difference between an anion and a cation, defining these key terms in your answer.
2 Draw an electron transfer diagram to show what would happen when atoms of aluminium and oxygen react together.

Naming and writing ionic compounds

Naming ionic compounds

3A STRUCTURE AND PROPERTIES OF METALS LINK

2A REPRESENTING COVALENT COMPOUNDS LINK

1C THE PERIODIC TABLE LINK

Being able to look at the chemical formula of an ionic compound and name it is an important aspect of developing your chemistry knowledge for later topics. It is crucial that you can, firstly, look at the formula of a substance or compound and determine if it is metallic (Chapter 3A), ionic (this chapter) or covalent (Chapter 2A). For a substance to be ionic, it must contain both a reactive metal and a non-metal. Secondly, you will also need to be able to use the periodic table to determine the position of an element and whether it is a metal or a non-metal (Chapter 1C).

There are rules that must be followed for naming ionic compounds. These are outlined in Table 4A–3.

Table 4A–3 Rules for naming ionic compounds

Rule	Description	Examples
1	The name of the metal ion (cation) is placed first in the name of the compound. Its name does not differ from the name of the element.	Lithium Magnesium Aluminium
2	The name of the non-metal ion (anion) is placed second in the name of the compound. Its name is similar to that of the element but the suffix is changed to '-ide'.	Chor~~ine~~ → Chlor**ide** Nitr~~ogen~~ → Nitr**ide** Phosph~~orus~~ → Phosph**ide**
3	Place the name of the metal and altered name of the non-metal together with a space between them.	Lithium chloride Magnesium nitride Aluminium phosphide

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Writing ionic formulas

The ability to write the correct formula for ionic compounds is more important than naming them because these formulas are used in balanced chemical equations, precipitation reactions – as you will learn in Chapter 4C – as well as many other aspects of chemistry across Units 2–4.

LINK **4C** PRECIPITATION REACTIONS

To do this, you will need to become familiar with using the periodic table, determining the location (group number) of elements, how many valence electrons they will lose or gain and consequently what the charge on the resultant ion will be.

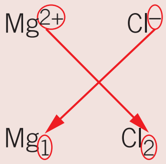
You learned earlier how to use electron transfer diagrams to determine the formula of a neutral ionic compound based on the transfer of electrons from the metal atom to the non-metal atom. Worked example 4A–2 explores another method that can be used to assist you when writing the formula for an ionic compound.

Worked example 4A–2: Writing ionic formulas

VIDEO WORKED EXAMPLE 4A–2

Write the ionic compound formula for magnesium chloride.

Solution

Step 1	Write the chemical symbol for each element.	Mg Cl
Step 2	Using the periodic table, determine the group number of each element to determine their respective ion charges (written as superscripts).	Mg is in group 2 Cl is in group 17 Mg^{2+} Cl^{-}
Step 3	Swap the coefficients from the superscript ion charges to subscripts on the other element and remove the charges (+ and –).	
Step 4	If there is just one of a particular ion present in the formula, omit the subscript ‘1’. (This can be done for the Mg in this example.)	$MgCl_2$

NOTE

The limitation with using this model is that for ionic compounds where the charge of the metal ion and the non-metal ion are the same, the coefficients when written as subscripts after each element can be simplified.

For example, in magnesium oxide, Mg has a 2+ charge and O has a 2– charge, but the formula for the ionic compound is written as MgO, not Mg_2O_2.

Check-in questions – Set 2

1 Name the following ionic compounds based on their formulas.
   - a Na_2S
   - b MgO
   - c Li_3N

2 Write the chemical formulas for the following ionic compounds.
   - a calcium fluoride
   - b potassium bromide
   - c magnesium phosphide

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Naming and writing chemical formula for ionic compounds with transition metals

1C THE PERIODIC TABLE LINK

Transition metals, located in the 'd' block of the periodic table, are mostly lustrous, hard and strong substances with high melting and boiling points, as you learned in Chapter 1C. Most transition metals can lose a varying number of electrons to form a cation. This has to do with their subshell electron configurations and the relative energies of the 4s and 3d subshells.

1B ELECTRON CONFIGURATIONS LINK

For example, copper, which is an exception to the statement above, has the subshell electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$. Therefore, as a metal, copper loses electrons. In this instance, it would lose the single valence electron from the 4s subshell, therefore forming a '1+' cation charge. However, the most common ion for copper is the '2+' cation, where it loses two 4s electrons and the 3d subshell contains 9 electrons. The 4s subshell is of lower energy than the 3d subshell when the orbitals are empty, but as orbitals begin to be occupied, the relative energies change. This is particularly apparent with the copper and chromium exceptions.

12A REDUCTION AND OXIDATION LINK

The charge of the ion for a transition metal must be specified when naming the ionic compound. The charge number is a number whose magnitude is the ionic charge. It is written in brackets immediately after the name of an ion, in Roman numerals, without a space (Table 4A–4). No charge number is used following the name of an ion that only has one oxidation state. Different charge numbers correspond to different oxidation states and will be covered in Chapter 12A.

Table 4A–4 Names and formulas for oxides of copper

Chemical name	Chemical formula
Copper(I) oxide	Cu_2O
Copper(II) oxide	CuO

In other examples, you could be given the formula of an ionic compound and asked to determine the charge on the transition metal. This can be done through knowing the charge of the anion (by knowing which group the non-metal appears in on the periodic table or how many valence electrons it has) and the idea that there is a neutral charge for the ionic compound. This is explored in Worked example 4A–3 below.

Worked example 4A–3: Determining charge from ionic formulas

VIDEO WORKED EXAMPLE 4A–3

Determine the charge of the transition metal ion in chromium oxide, Cr_2O_3.

Solution

Step 1	Determine the charge of the anion.	Cr_2O_3 Oxygen is in group 16 and so has six valence electrons. Oxide is a 2– ion.
Step 2	Determine the total charge of the non-metal part of the compound dependent on the number of non-metal atoms present.	Three oxide ions means there is a 3 × 2– = 6– total charge.

Step 3	The total charge of the metal cations present must be the same as the total charge of the anions present to form a neutral compound.	The charge of the two chromium ions must add to 6+ so that it is equal to overall charge provided by oxygen ions present in this neutral formula. 6+ = 2 × (?+)
Step 4	Divide the total charge of cations by the number of metal atoms present to determine the cation charge.	6+ divided by 2 (the number of chromium ions) = Cr^{3+}

Check-in questions – Set 3

1 Iron can exist with two different ionic charges, ‘2+’ and ‘3+’. Write the formula for each of these when combined with chlorine and write the appropriate name to distinguish between the two.

2 Determine the charge of the zinc ion in the formula Zn_3P_2.

Polyatomic ion an ion that consists of more than one atom covalently bonded together to form a charged compound; *poly* means ‘many’

Naming and writing chemical formula for ionic compounds with polyatomic ions

Polyatomic ions consist of more than a single atom covalently bonded together to form a charged compound. When different elements are present, the atoms combine in a fixed ratio determined by the number of covalent bonds a non-metal atom can make, as you learned in Chapter 2A.

LINK **2B** REPRESENTING COVALENT COMPOUNDS

However, occasionally a particular molecule loses one or more H^+ ions, so the electron(s) are ‘left behind’ on the central atom of the molecule, producing a negative charge, indicating that the total number of electrons isn’t equal to the total number of protons.

These ions are particularly important when discussing precipitation reactions in Chapter 4C and acid–base chemistry in Chapter 11A.

LINK **4C** PRECIPITATION REACTIONS

LINK **11A** INTRODUCTION TO ACIDS AND BASES

Common examples of polyatomic ions that you will need to be familiar with are shown in Table 4A–5.

Table 4A–5 Common polyatomic ions

Name	Formula (must include charge)
Ammonium	NH_4^+
Hydroxide	OH^-
Nitrate	NO_3^-
Hydrogen carbonate	HCO_3^-
Carbonate	CO_3^{2-}
Sulfate	SO_4^{2-}
Phosphate	PO_4^{3-}

Importantly, these ions behave as a single entity. So, where there is more than one of these ions present in a chemical formula, brackets must be placed around the entire polyatomic ion and a subscript outside of the brackets, which is used to indicate the number of ions in the neutral ionic compound. Worked examples 4A–4 and 4A–5 demonstrate this.

DOC **WORKSHEET 4A–1** IONIC COMPOUNDS

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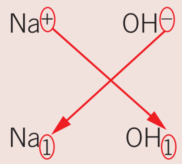
Worked example 4A–4: Writing ionic formulas containing a hydroxide ion

Write the formula for sodium hydroxide, where a single polyatomic ion is present in the formula.

Solution

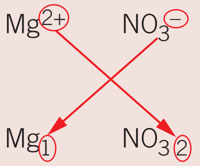
Step 1	Write the symbol for the cation and anion. Use superscripts to represent ion charges.	Na^+ OH^-
Step 2	Swap the superscript coefficient for ion charges to subscripts and remove the charges (+ and –).	
Step 3	If there is just one of a particular ion present in the formula, omit the subscript '1'. (This can be done for both the Na and the OH in this example.)	NaOH

Worked example 4A–5: Writing ionic formulas containing a nitrate ion

VIDEO
WORKED
EXAMPLE 4A–5

Write the formula for magnesium nitrate, where a single type of polyatomic ion is present in the formula.

Solution

Step 1	Write the symbol for the cation and anion. Use superscripts to represent ion charges.	Mg^{2+} NO_3^-
Step 2	Swap the superscript coefficient for ion charges to subscripts and remove the charges (+ and –).	
Step 3	As there is also more than one nitrate ion, brackets must be placed around the entire nitrate formula first before adding the subscript.	Mg_1 $(NO_3)_2$
Step 4	If there is just one of a particular ion present in the formula, omit the subscript '1'. (This can be done for the Mg in this example. Note that there is no space between the formulas for the cation and anion in the overall formula for the compound.)	$Mg(NO_3)_2$

4A SKILLS

Writing the chemical formulas for ionic compounds

As covered in this section, when writing the chemical formula for ionic compounds, there are rules that must be followed. These are outlined below and on the following page.

1. Write the symbol for the positively charged ion first.
2. Use subscripts to indicate the number of each ion in the formula, making sure to write the subscript after the ion they refer to.
3. If there is just one ion present in the formula, omit the subscript '1'.
4. Do not include the charges on the ions in the balanced formula.

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5 For polyatomic ions, if more than one of these is present, the formula for the polyatomic ion must be enclosed in round brackets with the subscript after the final bracket.
6 The formulas for the cation and anion are written directly after each other in the overall neutral formula with no space between them.

Section 4A questions

1 Define the following terms.
  a ionic compound
  b ion
  c cation
  d anion
2 Copy the following table into your workbook and complete it.

Element	Subshell electron configuration	Number of valence electrons	What it does to become stable	Charge after gain/loss	Chemical symbol for the ion
N					
F					
Na					
Ca					
S					

3 Using your knowledge established from Question **2** above and the charges of the ions you identified, what formula will each of the following compounds have?
  a sodium sulfide
  b calcium fluoride
4 Draw a complete transfer diagram for the following combinations of elements showing both the electron configurations for the atom and ion forms of the element. Include the resulting charge on each ion.

  (Hint: You must figure out how many of each element to use before you start to draw or as you draw, so using a pencil is a good idea.)

  a potassium and nitrogen
  b magnesium and oxygen
  c calcium and phosphorus
5 What is included after the name of a transition metal to denote the charge of its ion? Give an appropriate example in your response.
6 Describe the difference between a monoatomic ion and a polyatomic ion.
7 What is the charge on the ion formed by non-metallic elements that are present in the following groups of the periodic table? Explain how you arrived at your answer for each.
  a group 2
  b group 18
  c group 15
  d group 14
8 Write the formula for the following ionic compounds.
  a lithium bromide
  b calcium chloride
  c magnesium sulfide
  d aluminium hydrogen carbonate
  e ammonium sulfate
  f nickel(II) phosphide
  g vanadium(III) nitrate
9 Name the ionic compounds with the following formulas.
  a $NaHCO_3$
  b Pb_3N_4
  c CoOH
  d Li_3PO_4
  e $Fe(NO_3)_3$
  f $(NH_4)_2CO_3$
  g $FeBr_2$
  h CrF_3

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4B Structure and properties of ionic compounds

Study Design:
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

Glossary:
Brittle
Hardness
Ion–dipole attraction
Ionic bond

ENGAGE

Using ionic compounds to de-ice roads

In snowy environments, moving around can be slow and dangerous. Over the years, many solutions have been put forth to help remove ice and snow on roads: sand, table salt, vinegar, coffee beans and even kitty litter! However, one of the most effective methods is using the ionic compound calcium chloride ($CaCl_2$). This is because when calcium chloride dissolves into its ions Ca^{2+} and Cl^- in water, it lowers the freezing point of water to –29°C, making ice harder to form, which keeps roads safer for vehicles. Calcium chloride can also act as a thinner in fabric softener or as a heating agent in some heating pads.

This section will look at the structure and properties of ionic compounds that can enable such a wide range of uses.

Figure 4B–1 Calcium chloride is spread on an ice-covered road to lower the freezing point of water.

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EXPLAIN

Ionic compound bonding model

Similar to what you saw with the metallic bonding model in Chapter 3A, the ionic bonding model consists of a three-dimensional crystal lattice. Here the ions are also held in fixed positions in an organised arrangement. However, unlike metals, this lattice is composed of alternating positive cations and negative anions, as shown in Figure 4B–2.

LINK **3A** STRUCTURE AND PROPERTIES OF METALS

VIDEO 4B–1 COMPARING PROPERTIES OF IONIC COMPOUNDS

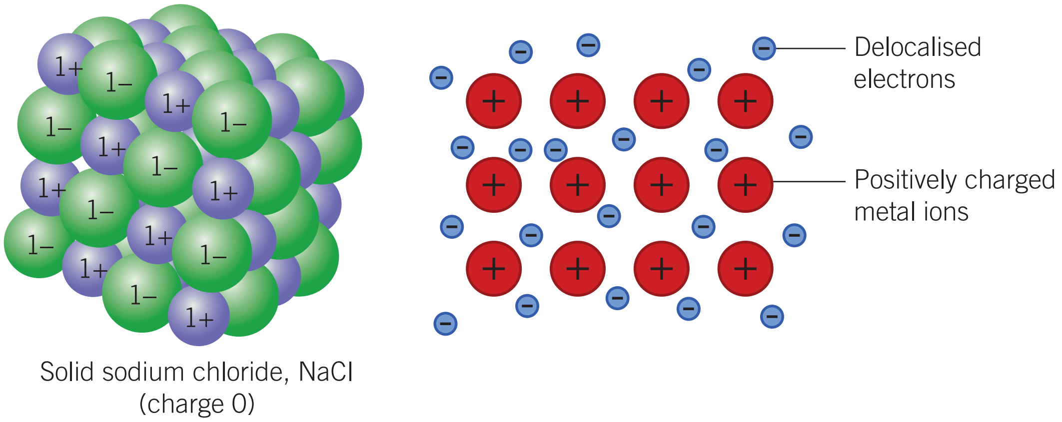

Figure 4B–2 Lattice structure in an ionic compound (left) consisting of alternating positive and negative ions, compared to that of a metallic lattice (right) consisting of positive ions surrounded by a sea of delocalised electrons

Like in the metallic bonding model, there is a strong electrostatic force of attraction between the positive and negative charges, as shown in Figure 4B–2. However, this attraction is between the cations and anions in the lattice, as there are no free moving electrons (electrons have been completely donated by metals to non-metals). This electrostatic force of attraction between cations and anions is referred to as an **ionic bond**.

Ionic bond the electrostatic force of attraction between cations and anions in a lattice

This ionic bonding model can be used to predict the general properties of ionic compounds.

Properties of ionic compounds

High melting and boiling point

Recall from Chapter 2A that the bond strength between two bonded atoms is a measure of the energy required to break apart 1 mole of covalently bonded molecules in a gaseous state. This energy is measured in kilojoules per mole (kJ mol^{-1}) units.

LINK **2A** REPRESENTING COVALENT COMPOUNDS

For ionic compounds, this 'intermolecular force' is the electrostatic force of attraction between cations and anions. When heat, or thermal energy, is applied to a solid ionic compound, this causes ions within the lattice to gain this energy and begin to vibrate (Figure 4B–3). The ions then transfer this energy to other ions in the solid lattice, eventually causing enough disruptions where the ions are able to move.

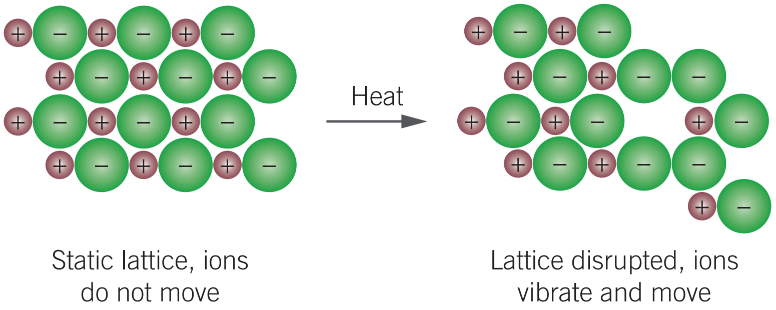

Figure 4B–3 This general structure of an ionic compound shows the electrostatic force of attraction between ions being disrupted as thermal energy is applied.

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Almost all ionic compounds have high melting and boiling points, but there are slight differences between each. This is due to two factors outlined in Table 4B–1.

Table 4B–1 Factors that affect the melting and boiling point of ionic compounds

Factor	Explanation of effect
Size of ions	The bigger the ions, the more spread out the charges are, which leads to less electrostatic force of attraction between the ions holding the lattice together.
Charge of ions	The greater the charge of the ion, the greater the electrostatic force of attraction.

4A FORMATION AND NAMING OF IONIC COMPOUNDS
1C THE PERIODIC TABLE
LINK

You should be able to use your understanding of the charges of ions from Chapter 4A and knowledge of atomic radii from Chapter 1C to determine the difference between two ionic compounds, where either the anion or cation is altered. Worked examples 4B–1 and 4B–2 guide you through how to structure this.

Worked example 4B–1: Comparing melting points

VIDEO WORKED EXAMPLE 4B–1

Determine whether lithium fluoride or sodium fluoride would have the higher melting point. Compare each structure in justifying your answer.

Solution

Step 1	Identify the difference.	Lithium and sodium both have an ion charge of 1+ (group 1 metals). However, lithium is in period 2 and sodium is in period 3. Therefore, lithium ions are smaller.
Step 2	Explain the effect of the difference.	Ions in lithium fluoride can position closer together and will have a stronger electrostatic force of attraction.
Step 3	Link back to the question.	This results in a higher melting point for lithium fluoride than sodium fluoride as more energy is required to disrupt its ionic bond.

Worked example 4B–2: Comparing boiling points

Determine whether CaF_2 or KF would have the higher boiling point. Compare each structure in justifying your answer.

Solution

Step 1	Identify the difference.	Calcium and potassium are both in period 4, so will have similar size ions; however, potassium is in group 1 (ion charge of 1+) and calcium is in group 2 (ion charge of 2+). Therefore, calcium has a greater charge.
Step 2	Explain the effect of the difference.	The electrostatic force of attraction between calcium and fluoride ions will be greater than potassium and fluoride ions.
Step 3	Link back to the question.	This results in more energy required to overcome this attraction holding the ions in the lattice, so CaF_2 has a higher boiling point.

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Check-in questions – Set 1

1 Predict whether aluminium nitride (AlN) or aluminium phosphide (AlP) would have the higher melting point. Explain.

Hardness and brittleness

Ionic compounds as solids are recognised by their **hardness**, which is due to the strong electrostatic forces of attraction between their cations and anions. To shift ions in the lattice, this attraction needs to be overcome.

However, if enough force is applied and these layers of alternating cations and anions do shift, this results in like charges lining up, anions with anions, cations with cations. This causes the like charges to repel each other and the crystal lattice at those points to shatter. Hence, the compound is **brittle**. This property is different to that of metals, which are malleable and ductile, as opposed to being brittle. Recall the structure of the metallic bonding model. It consists of only cations in an organised crystal lattice surrounded by a sea of negative delocalised electrons. In metals, if a force is applied, the cations move. However, the delocalised electrons floating between these ions can also move and the electrostatic force of attraction is maintained. Therefore, the structure does not break and is not brittle. Figure 4B–4 highlights this difference between the two bonding models.

Hardness
a measure of the resistance of a material to abrasion or surface indentation

Brittle
hard and rigid but easily broken. A result of the same charged ions in ionic compounds coming together, causing them to repel and the crystal lattice to shatter/break

LINK **3A** STRUCTURE AND PROPERTIES OF METALS

LINK **3B** REACTIVITY OF METALS

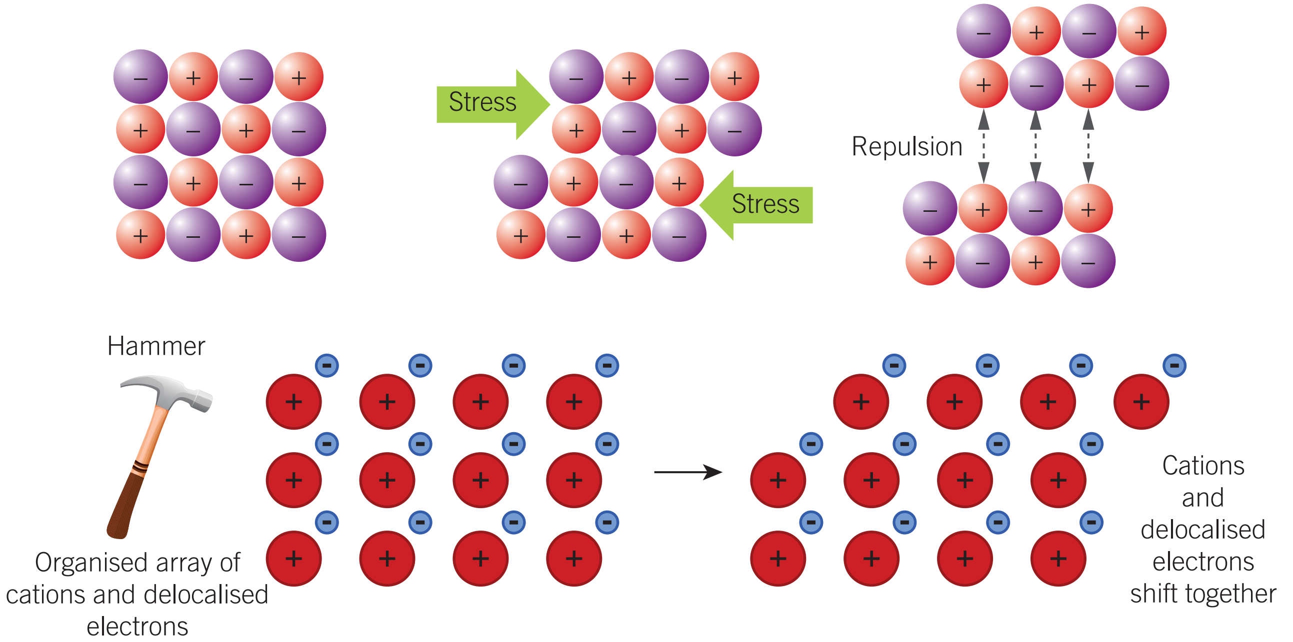

Figure 4B–4 Difference between the ionic bonding model (top) and metallic bonding model (bottom) when a force (stress) is applied. Note that ionic compounds are brittle, whereas metals are malleable and ductile.

Electrical conductivity

For any substance to conduct electricity, it requires free-moving charged particles. For ionic compounds, you need to consider two things:

- the types of charged particles present
- the arrangement of these charged particles.

The state of an ionic compound, whether it is a solid or a liquid, will determine its ability to conduct electricity.

Solid state

In the solid state, the cations and anions in ionic compounds are in fixed positions in the lattice structure. Therefore, they are not free to move, as shown in Figure 4B–5. So, ionic compounds do not conduct electricity as a solid.

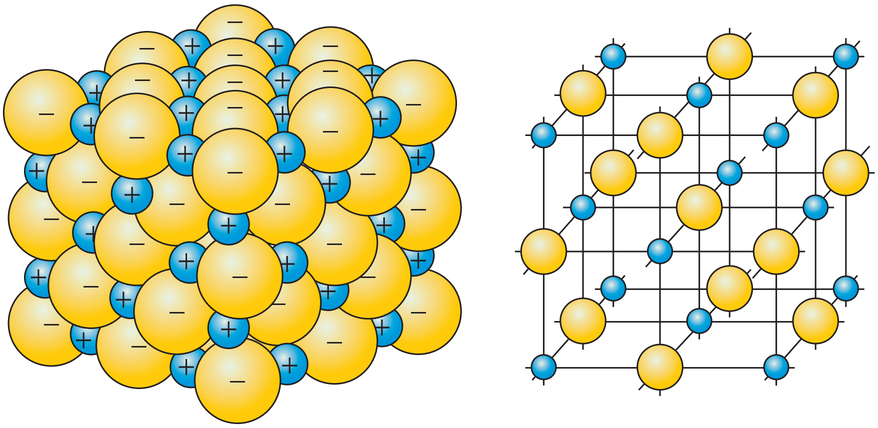

Figure 4B–5 These two images show an ionic compound in the solid state and highlight the fixed positions of both anions and cations in the lattice structure.

Molten and aqueous state

In the molten (melted, or liquid, or fused) state, as a result of heating the ionic compound to disrupt the electrostatic force of attraction (ionic bond) between cations and anions, the ions become free-moving and can conduct electricity, as shown in Figure 4B–6.

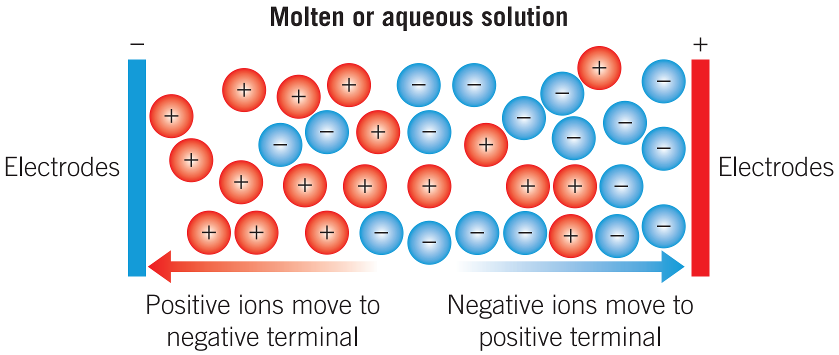

Figure 4B–6 When electrodes are placed in either a molten sample or aqueous solution of an ionic compound, the charged ions will move to the oppositely charged electrodes, giving the compound the ability to conduct electricity.

2B MOLECULAR SHAPES AND POLARITY LINK

Another way that ionic compounds can conduct electricity is if they are in an aqueous state – in other words, dissolved in water. As you learned in Chapter 2B, water is a polar bent-shaped molecule. Recall that this means one side of the non-symmetrical molecule has a partially positive charge (hydrogen atoms) and the other side has a partially negative charge (oxygen, as this is the more electronegative element).

As a result, when the solid ionic compound is placed in water, the partial positive charge of the hydrogen atoms are attracted to the negative anions. Conversely, the partial negative charge of the oxygen atoms are attracted to the positive cations. This draws the outer ions in the lattice away from each other, and the electrostatic force of attraction between the cations and anions is replaced by the **ion–dipole attraction** between the ions and water molecules. This is shown in Figure 4B–7.

Ion–dipole attraction
the bond formed between an ion and the oppositely charge end of a polar molecule, such as water

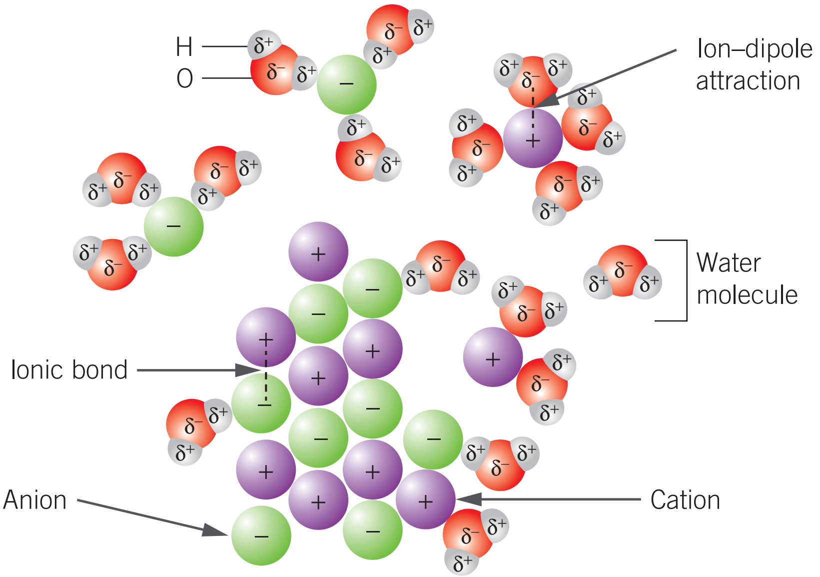

Figure 4B–7 Ionic compound dissolving in water

You will explore the solubility of different ionic compounds more in Chapter 4C, and this concept is also related to their ability to conduct electricity.

WORKSHEET 4B–1
IONIC PROPERTY PREDICTION

LINK **4C** PRECIPITATION REACTIONS

Check-in questions – Set 2

1 Explain the difference between the electrical conductivity of ionic compounds in the solid and molten (liquid) state.

4B SKILLS

VIDEO 4B–2
SKILLS: HOW TO ANSWER 'EXPLAIN' QUESTIONS

How to answer 'explain' questions

Command terms are important components of questions, as they indicate what it is that you need to do to attain full marks.

Explain is an example of a command term used regularly in assessment situations. To explain means to give a detailed account of causes, reasons or mechanisms or to essentially say why something happened. Consider the following question:

Explain why the element iron is malleable but iron oxide is brittle.

Your answer is best set out in dot points (which are acceptable to use in assessments). An acronym you might find useful when answering these types of questions is **DER**, which stands for:

Define – include key and relevant definitions in your answer for key content words.

Explain – give a detailed account of causes, reasons or mechanisms relevant to the question.

Relate – relate your explanation to the information or context of the question so an assessor or teacher can clearly see the link.

So, you would answer the question on the previous page as follows:

Answer:

Iron is a metal, which means it is composed of a lattice of positive iron cations surrounded by a 'sea' of delocalised electrons. However, iron oxide is an ionic compound, which means it is composed of alternating positive cations and negative anions in fixed positions in a lattice.

When a force is applied to a metal lattice, the layers of cations slide over each other. However, the delocalised electrons also move with the cations, maintaining an electrostatic force of attraction. Alternatively, when a force is applied to an ionic lattice, the layers of ions slide over each other, resulting in like-charged ions that are next to each other, which repel, causing the structure to shatter at these points. (Note that you could also use a diagram, similar to the one represented in Figure 4B–4, to support your answer.)

Therefore, iron is malleable, whereas iron oxide is brittle.

Section 4B questions

1 What is the difference between cations and anions?
2 Describe the ionic bonding model.
3 Similar to some of the limitations that were presented for the metallic bonding model in Chapter 3A, identify some potential limitations of the ionic bonding model.
4 Draw a labelled diagram of a three-dimensional sodium chloride crystal lattice.
5 Use the ionic bonding model to explain the following properties of ionic compounds.
  a Potassium bromide melts at over 700°C.
  b Sodium chloride is a hard and brittle solid at room temperature.
  c Calcium nitrate is a poor conductor of electricity in the solid state.
  d Calcium nitrate conducts electricity in the molten state.
6 The melting point of MgO is more than 3.5 times greater than the melting point of NaCl. Suggest a reason why this might be the case.
7 The melting point of NaCl is 801°C, slightly higher than the melting point of KCl at 770°C. Explain this difference in melting point.
8 Explain why NaBr is a hard and brittle solid.
9 Describe what happens when an ionic compound is dissolved in water and why this aqueous solution conducts electricity.

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4C Precipitation reactions

Study Design:
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)

Glossary:

Dissolution
Double displacement reaction
Insoluble
Ionic equation
Precipitate
Precipitation reaction
Solubility
Soluble
Spectator ion

ENGAGE

Using ionic liquids to remove water contaminants

In developing countries, access to clean drinking water is a huge problem. Areas of the community may only have access to contaminated water from rivers, streams and lakes. Traditional water filtration methods use separate filters to help remove specific impurities from contaminated water. This can make water filtration systems costly.

However, in 2017, scientists in Germany and Spain designed a new material that can remove many different types of contaminants. This supported ionic liquid phase (SILP) material is an ionic liquid, formed by melting an ionic compound at room temperature. This ionic liquid is then adsorbed onto a solid substrate, which can then be altered through further chemical modifications. The SILP is essentially transition metal atoms connected with oxygen atoms forming a three-dimensional lattice. Specifically, for their new filtration material, they chose polyoxotungstate anions and tetra-alkylammonium cations.

The anions were able to trap heavy metal ions present in the water, including chromium, cobalt, copper, lead and nickel. They also removed the water-soluble dye called blue trityl commonly used in textile manufacturing in these countries. The cations, however, provide an antimicrobial property, inhibiting the growth and reproduction of *E. coli* bacteria.

It is hoped that this ionic liquid shows the way for the development of future chemically designed filtration systems that help to remove contaminants from water in developing countries and following chemical spills or natural disasters.

Figure 4C–1 New filtration material will potentially revolutionise the cleaning up of chemical spills.

LINK **1D** CRITICAL ELEMENTS AND RECYCLING PROCESSES

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EXPLAIN

Recap of the solubility of ionic compounds

4B STRUCTURE AND PROPERTIES OF IONIC COMPOUNDS LINK

When dissolved in water, ionic compounds are able to conduct electricity as the ions separate and become mobile charge carriers in the aqueous solution, as considered in Chapter 4B. However, not all ionic compounds are soluble in water. In this section, you will use the solubility rules to determine the solubility of an ionic compound, as well as write full equations and ionic equations to demonstrate that aqueous solutions of ionic compounds may react when mixed to produce an insoluble ionic compound product.

Solubility of ionic compounds

13B PREDICTING SOLUBILITY LINK

All ionic compounds will dissolve in water to a certain extent at different temperatures, as you will learn in Chapter 13B. In this section, we will consider the **solubility** of ionic compounds at 25°C for consistency. Usually, compounds that have a solubility of greater than 1 g/100 g of water would be considered completely **soluble**. Similarly Compounds that have a solubility of less than 0.1 g/100 g of water are considered **insoluble**, meaning there will be solid present in a mixture with water.

Solubility the ability of a solute to dissolve in a solvent

Soluble when a solute is able to dissolve in a solvent

Insoluble when a solute has limited solubility in a solvent

Solubility is a relative scale, rather than a fixed notion of soluble versus insoluble, and compounds that have a solubility between 0.1 g and 1 g/100 g of water would be considered partially or slightly soluble. The observation here would be that some solid has dissolved, but some solid remains in a mixture with water.

How ionic compounds dissolve

As you learned in Chapter 4A, ionic compounds have negative and positive ions. When an ionic compound dissolves in water, the bent-shaped polar molecules are attracted to each of the ions in the following way.

- The partial positive end of water (hydrogen atoms) is attracted to the negative anions.
- The partial negative end of water (oxygen atom) is attracted to the positive cations.

4A FORMATION AND NAMING OF IONIC COMPOUNDS LINK

The attraction between the ions in the lattice and water molecules as shown in Figure 4C–2 is referred to as an ion–dipole attraction. For an ionic compound to be soluble, this attraction must be stronger than the hydrogen bonding between water molecules and the electrostatic force of attraction between cations and anions in the solid lattice.

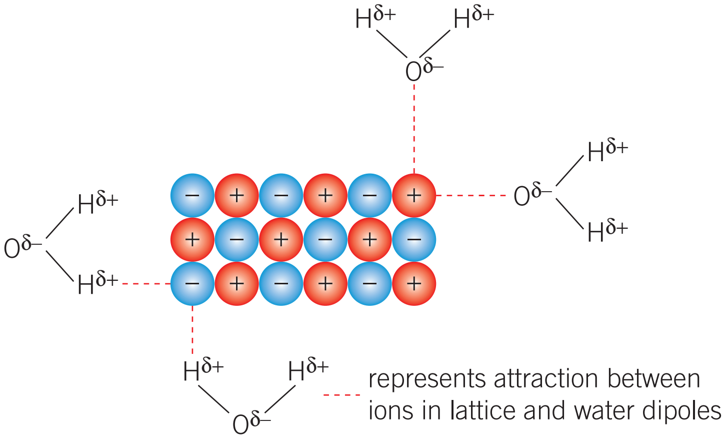

Figure 4C–2 Attraction between polar water molecules (H_2O) and the ions in the lattice of an ionic compound.

The process of an ionic compound dissolving in a solvent, such as water, is called **dissolution**. This is a two-step process.

1. Dissociation: the separation of a molecule or ion pair into two or more molecules or free ions.
2. Solvation: a stabilising interaction of a solute and the solvent. These interactions can involve electrostatic forces of attraction, dispersion forces or hydrogen bond formation.

Dissolution the process in which molecules or ions become separated and mobile when surrounded by solvent molecules. For ionic compounds, they are split into ions that then form an electrostatic force of attraction with the solvent

This is represented in Figure 4C–3.

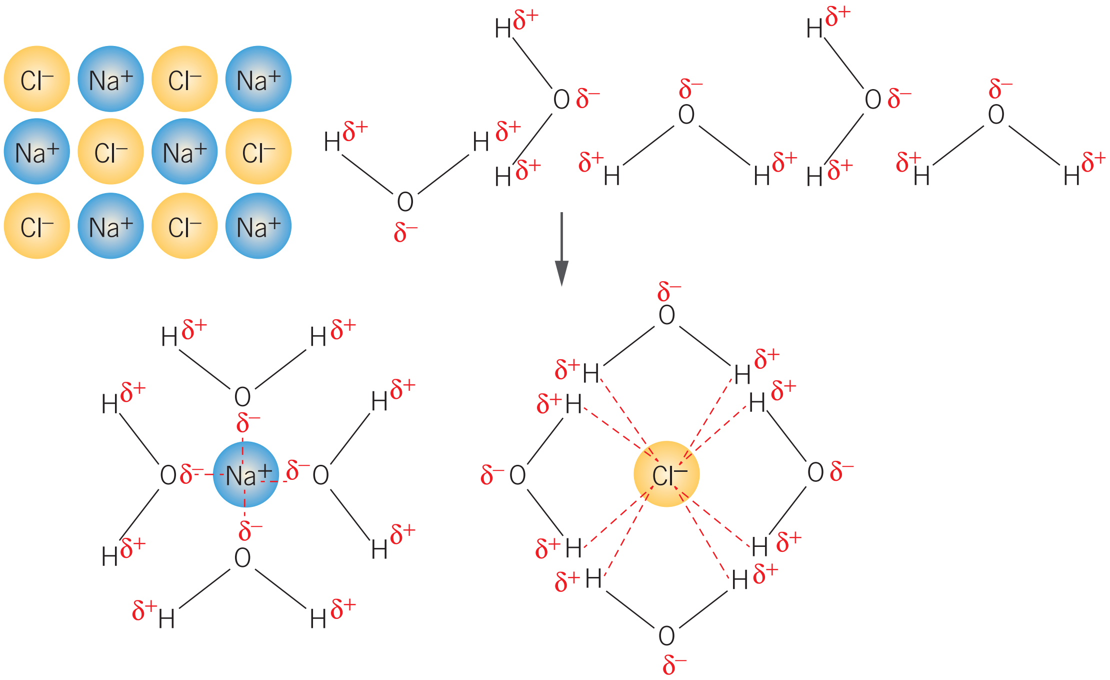

Figure 4C–3 The process of dissolution for solid sodium chloride when mixed with water

The solubility table

The solubility of ionic compounds in water is strongly related to the identity of the cations and anions that compose the ionic compound. This relationship can be summarised effectively using a solubility table, as depicted in Table 4C–1.

VIDEO 4C–1
INTERPRETING THE SOLUBILITY TABLE

Table 4C–1 Solubilities of common ionic compounds and their exceptions

Soluble in water	Exceptions – insoluble	Exceptions – slightly insoluble
All nitrates (NO_3^-)	No exceptions	No exceptions
All ammonium salts (NH_4^+)	No exceptions	No exceptions
All sodium (Na^+) and potassium (K^+) salts	No exceptions	No exceptions
All ethanoates (CH_3COO^-)	No exceptions	No exceptions
Most chlorides (Cl^-), bromides (Br^-) and iodides (I^-)	AgCl, AgBr, AgI, PbI_2	$PbCl_2$, $PbBr_2$
Most sulfates (SO_4^{2-})	$BaSO_4$, $PbSO_4$, $SrSO_4$	Ag_2SO_4, $CaSO_4$
Insoluble in water	**Exceptions – soluble**	**Exceptions – slightly insoluble**
Most carbonates (CO_3^{2-})	Na_2CO_3, K_2CO_3, $(NH_4)_2CO_3$	No exceptions
Most sulfides (S^{2-})	Na_2S, K_2S, $(NH_4)_2S$	No exceptions
Most phosphates (PO_4^{3-})	Na_3PO_4, K_3PO_4, $(NH_4)_3PO_4$	No exceptions
Most hydroxides (OH^-)	NaOH, KOH, $Ba(OH)_2$, NH_4OH	$Ca(OH)_2$, $Sr(OH)_2$

In lieu of a solubility table, there are some general rules that can be considered when deducing the relative solubility of an ionic compound based upon the rules suggested in the solubility table.

Solubility rules

There are some ions that are generally always soluble in water, and therefore their ionic compounds will be soluble in water. The ions in the list below can be considered 'always soluble'. An easy way to remember this information is to use the initials in a short mnemonic, **SNAPE**:

- **S**odium (Na^+)
- **N**itrate (NO_3^-)
- **A**mmonium (NH_4^+)
- **P**otassium (K^+)
- **E**thanoate (CH_3COO^-).

Ionic compounds that contain the following halides are also considered soluble: bromide (Br^-), iodide (I^-) and chloride (Cl^-). However, these ions have exceptions and when bonded to silver (Ag^+) and lead (Pb^{2+}) ions in ionic compounds, they are insoluble. Moreover, most sulfate (SO_4^{2-}) containing compounds are soluble in water except when bonded to the cations of silver (Ag^+) and lead (Pb^{2+}) as well but also when bonded to strontium (Sr^{2+}), barium (Ba^{2+}) and calcium (Ca^{2+}).

This information can be remembered by using short mnemonic too, **BICS,** which denotes those ions that can be considered *generally* soluble with some exceptions:

- **B**romide (Br^-)
- **I**odide (I^-)
- **C**hloride (Cl^-)
- **S**ulfate (SO_4^{2-}).

Finally, there are those ionic compounds, as listed below, that are considered insoluble in water when not bonded to a SNAPE ion (as these are always soluble!). This information can be combined into another short mnemonic, **CHPS**, which denotes those ions that can be considered *generally* insoluble with some exceptions:

- **C**arbonate (CO_3^{2-})
- **H**ydroxides (OH^-), except when bonded to SNAPE ions, or barium (Ba^{2+}), calcium (Ca^{2+}) and strontium (Sr^{2+})
- **P**hosphate (PO_4^{3-})
- **S**ulfide (S^{2-}).

When considering whether an ionic compound is soluble or insoluble in aqueous solutions, the general solubility rules can be summarised by **SNAPE BICS CHPS**. However, do not forget the exceptions!

Check-in questions – Set 1

1 Use the solubility rules to determine whether the following ionic compounds are soluble or insoluble in water.

	Compound	Solubility
a	$Ba(OH)_2$	
b	$CaCl_2$	
c	Na_3PO_4	
d	Ammonium carbonate	
e	Magnesium sulfide	
f	Silver bromide	

WORKSHEET 4C–1 SOLUBILITY RULES

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Precipitates

Water-soluble ionic compounds are able to react in solution to form new ionic compounds, in which cations and anions exchange their respective ionic counterparts. For example, a solution of sodium chloride and silver nitrate can be combined to form a mixture of silver chloride solid and a solution of sodium nitrate. In this case, the Na^+ and Ag^+ cations have swapped their Cl^- and NO_3^- anion counterparts as shown in the equation:

$$\text{sodium chloride} + \text{silver nitrate} \rightarrow \text{sodium nitrate} + \text{silver chloride}$$

$$NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$$

VIDEO 4C–2 UNDERSTANDING PRECIPITATION REACTIONS

This type of reaction is typically known as a **double displacement reaction**, as two sets of ions have been displaced. The general form is shown in Figure 4C–4.

Double displacement reaction a chemical reaction in which ionic compounds swap their cations and anions

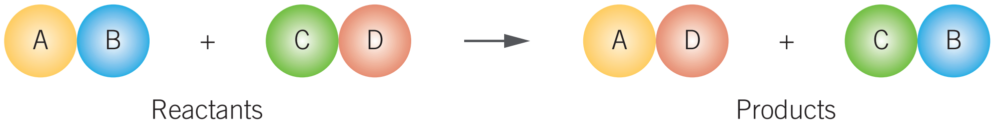

Figure 4C–4 Pictorial representation of a double displacement reaction

NOTE

To provide more information about a double displacement reaction equation, states are expected to be used to denote whether the reactants and products are in a solid (s), liquid (l), gas (g) or aqueous (aq) state. These state symbols are written next to the reactant or product species. Thus, for the example above, the equation would read:

$$NaCl(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)$$

Using the solubility rules, sodium chloride, silver nitrate and sodium nitrate are soluble in water and the state symbol (aq) is used. Silver chloride is insoluble in water and will form a solid in the resulting mixture, and is therefore given the state symbol (s).

Precipitation reactions

A double displacement reaction that produces a solid from two aqueous solution reactants is called a **precipitation reaction**. When writing chemical equations for precipitation reactions, a knowledge of the solubility rules is important. Thus, all resulting products should be considered as to whether they would be soluble or insoluble in water to determine a possible **precipitate** and the appropriate state symbols.

Precipitation reaction a chemical reaction in which two or more solutions of soluble salts are mixed and one or more of the products formed is an insoluble solid – known as a precipitate

Precipitate the solid formed in a reaction in which two or more solutions of soluble salts are mixed

Worked example 4C–1: Writing full equations for precipitation reactions

An aqueous solution of K_2CO_3 was added to an aqueous solution of $CuCl_2$. Write a balanced full equation for the reaction, including states.

Solution

Step 1	Write ionic formulas for all reactants and products, remembering to swap the ionic counterparts for the products formed.	$K_2CO_3 + CuCl_2 \rightarrow KCl + CuCO_3$
Step 2	Use solubility rules to write state symbols. As it is a double displacement reaction, the cations will swap anion counterparts to form KCl and $CuCO_3$. All are soluble as per the solubility rules except copper(II) carbonate, which is insoluble and thus has a solid state symbol (s).	$K_2CO_3(aq) + CuCl_2(aq) \rightarrow KCl(aq) + CuCO_3(s)$
Step 3	Balance the equation. The products side requires two potassium and chloride ions, and therefore KCl is given the coefficient 2.	$K_2CO_3(aq) + CuCl_2(aq) \rightarrow 2KCl(aq) + CuCO_3(s)$

WORKSHEET 4C–2 PRECIPITATION REACTIONS

VIDEO WORKED EXAMPLE 4C–1

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Check-in questions – Set 2

1 Write a balanced full equation with state symbols for the following precipitation reactions.
  a An aqueous solution of Na_3PO_4 is added to an aqueous solution of $Ca(CH_3COO)_2$.
  b An aqueous solution of barium hydroxide is added to an aqueous solution of nickel(II) bromide.

Ionic equations

When an ionic compound dissolves in water, the ions dissociate and become free mobile ions in the solvent. When ions combine in solution to form an insoluble compound, the solid starts to precipitate out of the solution.

Spectator ion an ion that does not take part in the precipitation of the solid but remains soluble in solution

Some ions do not participate in the formation of the precipitate. These ions are called **spectator ions**. Take the reaction of sodium chloride and silver nitrate as an example again.

$$NaCl(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)$$

In an aqueous solution, the ions are free and mobile, so the equation can more accurately be shown as a total ionic equation:

$$Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \rightarrow Na^+(aq) + NO_3^-(aq) + AgCl(s)$$

In the more accurate representation below, the Na^+ and NO_3^- ions are present as free mobile ions at the start of the reaction and are also present as free mobile ions at the end of the reaction. They have not participated in the reaction. This is expected if you recall the SNAPE solubility rule. These ions are the spectator ions and should not be considered as part of the chemical reaction that takes place.

$$\cancel{Na^+(aq)} + Cl^-(aq) + Ag^+(aq) + \cancel{NO_3^-(aq)} \rightarrow \cancel{Na^+(aq)} + \cancel{NO_3^-(aq)} + AgCl(s)$$

Ionic equation a balanced equation that shows only the chemical species that participate in a reaction

Therefore, an **ionic equation** shows only the species that participate in the reaction:

$$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$$

A precipitation reaction equation can be written in full form showing the identity of full ionic compounds, but it can more accurately be written as an ionic equation that removes spectator ions that do not participate in the reaction.

Worked example 4C–2: Writing ionic equations for precipitation reactions

VIDEO
WORKED
EXAMPLE 4C–2

An aqueous solution of K_2CO_3 was added to an aqueous solution of $CuCl_2$. Write a balanced ionic equation for the reaction, including states.

Solution

Step 1	Write the full balanced equation with state symbols. Write the equation in full, with states, as a starting point to convert to an ionic equation.	$K_2CO_3(aq) + CuCl_2(aq) \rightarrow 2KCl(aq) + CuCO_3(s)$
Step 2	Write the total ionic equation. Show all aqueous ionic compounds as their free mobile ions. Remember to include the coefficient to ensure the correct number of ions are present. Keep any solids in full bonded form.	$2K^+(aq) + CO_3^{2-}(aq) + Cu^{2+}(aq) + 2Cl^-(aq) \rightarrow 2K^+(aq) + 2Cl^-(aq) + CuCO_3(s)$
Step 3	Write the ionic equation. Remove the ions that are the same on both sides of the equation (spectator ions).	$CO_3^{2-}(aq) + Cu^{2+}(aq) \rightarrow CuCO_3(s)$

Solutions where no precipitate forms

In some cases, two solutions of ionic compounds can be mixed with no precipitate being formed. The question is then, did a double displacement chemical reaction occur?

An ionic equation most accurately demonstrates what is happening in the solution.

For example, if a solution of sodium chloride and a solution of potassium bromide are mixed, the full balanced equation, following a double displacement reaction, would be:

$$\mathrm{NaCl(aq) + KBr(aq) \rightarrow NaBr(aq) + KCl(aq)}$$

Writing this equation as a total ionic equation, which shows all aqueous ions as free ions, would be:

$$\mathrm{Na^+(aq) + Cl^-(aq) + K^+(aq) + Br^-(aq) \rightarrow Na^+(aq) + Br^-(aq) + K^+(aq) + Cl^-(aq)}$$

The total ionic equation demonstrates that all free ions in solution at the start of the reaction are the same as at the end of the reaction. They are all spectator ions! Has a chemical reaction occurred? As there are no new bonds formed, the answer is: no reaction has occurred here.

Worked example 4C–3: Writing ionic equations for spectator ions

An aqueous solution of ammonium chloride was added to an aqueous solution of potassium nitrate. Write a balanced ionic equation for the reaction, including states.

Solution

Step 1	Write the full balanced equation with state symbols as a starting point to convert to an ionic equation.	$\mathrm{NH_4Cl(aq) + KNO_3(aq) \rightarrow NH_4NO_3(aq) + KCl(aq)}$
Step 2	Write the ionic equation. All species are ions dissolved in water, so no reaction has taken place.	No reaction

Check-in questions – Set 3

1 Write a balanced ionic equation, with state symbols, for the following precipitation reactions.
  a $\mathrm{K_2SO_4(aq)}$ and $\mathrm{BaBr_2(aq)}$ are mixed.
  b An aqueous solution of copper(II) ethanoate is added to an aqueous solution of nickel(II) nitrate.

Summary

In this section you have explored the solubility of ionic compounds as well as the effect of solubility on combinations of aqueous solutions of ionic compounds. Reactions between ionic compounds that produce a precipitate can be most accurately described by an ionic equation with states. Combinations that do not produce a precipitate, where all ions are spectator ions, have no reaction occurring.

4C SKILLS

VIDEO 4C–3
SKILLS:
MEMORY TOOLS

Memory tools

As you will see throughout this year, some concepts require you to remember a lot of specific information about how different elements, ions or substances react. It can be very helpful to have strategies for recalling this information – these are referred to as memory tools. One such device that can be handy is the creation of a mnemonic. A mnemonic is a pattern of letters or associations that can help you remember something. The list of ions that are always soluble, slightly soluble (with exceptions) and insoluble (with exceptions), as you saw in this section, can be converted into a mnemonic. If you look back to the these, you will see we used the mnemonics below.

Always soluble (no exceptions)	Soluble (with exceptions)	Insoluble (with exceptions)
S for sodium N for nitrate A for ammonium P for potassium E for ethanoate	B for bromide I for iodide C for chloride S for sulfate	C for carbonate H for hydroxide P for phosphate S for sulfide

The difficult thing to recall here is that the chemical symbols for elements are not always the same letter(s) as the beginning of their English name. For example, potassium has the chemical symbol ‘K’. However, in the context of these mnemonics, it is represented as a ‘P’ (as part of SNAPE). It is important that you do not confuse this for phosphide (not included here) or even phosphate (the ‘P’ in CHPS).

In most reactions, it is enough to remember just those ions under the SNAPE mnemonic. As most of you will be familiar with the *Harry Potter* books and/or films, this should be easy to recall by knowing who Professor Snape is.

If you find other information in Units 1 and 2, or Units 3 and 4 next year, difficult to remember, try developing a mnemonic to help you commit it to memory.

Section 4C questions

1 Define these terms: solubility, soluble, insoluble.

2 Recall the state symbols used to denote solid, liquid, gas and aqueous states in chemical reactions.

3 Use your knowledge of the solubility rules to predict whether the following ionic compounds will be soluble or insoluble in water.

	Salt	Solubility
a	Ammonium phosphate	
b	Zinc hydroxide	
c	Iron(III) nitrate	
d	Lanthanum(III) carbonate	
e	Barium sulfate	
f	Sodium sulfide	

4 Balance the following precipitation reaction equations.
   - a $Pb(NO_3)_2(aq) + NaCl(aq) \rightarrow PbCl_2(s) + NaNO_3(aq)$
   - b $Al_2(SO_4)_3(aq) + (NH_4)_3PO_4(aq) \rightarrow AlPO_4(s) + (NH_4)_2SO_4(aq)$

5 Write fully balanced equations, including states, for the following precipitation reactions.
   - a $KOH(aq) + Cu(NO_3)_2(aq)$
   - b $Ba(OH)_2(aq) + Na_2S(aq)$
   - c $(NH_4)_2CO_3(aq) + Sn(CH_3COO)_4(aq)$
   - d $AlCl_3(aq) + Pb(NO_3)_2(aq)$

6 Write balanced ionic equations, including states, for the following precipitation reactions.
   - a $FeBr_3(aq) + AgNO_3(aq)$
   - b $NiSO_4(aq) + Ba(CH_3COO)_2(aq)$
   - c $Na_3PO_4(aq) + ZnI_2(aq)$
   - d $Mn(NO_3)_4(aq) + K_2CO_3(aq)$
   - e $NaI(aq) + K_2SO_4(aq)$

7 Four aqueous solutions of sodium chloride, potassium nitrate, ammonium carbonate and copper(II) ethanoate are all mixed together. Identify the name and chemical formula for the insoluble precipitate formed.

8 An aqueous solution of iron(III) iodide is mixed with an aqueous solution of ammonium carbonate. Write an ionic equation, including states, for the precipitation reaction that occurs.

9 An aqueous solution of barium hydroxide is added to an aqueous solution of nickel(II) sulfate. Write ionic equations, including states, for the precipitation reactions that occur.

10 The precipitation of sodium urate in human body joints is a common cause of gout. While sodium ions are generally considered soluble, this is an example of an insoluble ionic compound made from Na^+ and $C_5H_3N_4O_3^-$ ions. Write an ionic equation, with state symbols, for the formation of insoluble sodium urate from sodium and urate ions.

Chapter 4 review

Summary

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

Success criteria – I am now able to:		Linked question
4A.1	Recall and apply the idea that an ion is charged (anion and cation), meaning it has a differing number of protons and electrons	1☐
4A.2	Describe how ions are formed, with regard to electrons being gained and lost and how this allows an atom to achieve a more stable electron configuration (full outer shell)	2☐, 11☐, 12☐
4A.3	Explain how ionic compounds are made up of positive and negative ions but the overall charge is neutral	11☐, 13☐
4A.4	Recall that the positive ion in an ionic compound is written and named first in the formula	3☐
4A.5	Determine the name and ionic formula of ionic compounds using the periodic table	5☐, 6☐
4A.6	Determine the charge on an ion given an ionic formula	14☐
4A.7	Determine the charge of a main group ion (using group number)	12☐
4A.8	Write the formula of ionic compounds involving polyatomic ions and transition metals	15☐
4A.9	Determine the charge on a transition metal when given an ionic formula	14☐
4B.1	Recall that the size of ions and the charge on ions affects the melting point and boiling point of an ionic compound	7☐
4B.2	Recall that generally ionic compounds are hard, brittle and can conduct in a molten liquid state and in aqueous solution, but not a solid state	8☐, 9☐, 17☐
4B.3	Use the ionic bonding model to explain the difference in melting point of two different ionic compounds (must have one different ion)	16☐
4B.4	Compare and contrast properties of two different ionic compounds	16☐, 17☐
4C.1	Recall that a precipitation reaction involves mixing solutions of two different ionic compounds	18☐
4C.2	Recall the solubility rules (SNAPE, BICS, CHPS) to determine if an ionic compound formed will be solid – that is, not in solution	10☐
4C.3	Recall what an ionic equation is and how it differs from a full precipitation reaction	19b☐
4C.4	Write and balance chemical equations, including states, for a precipitation reaction	19a☐
4C.5	Write and balance ionic equations for precipitation reactions	19b☐

Multiple-choice questions

1 An anion is an atom that has
  A lost one or more electrons.
  B gained one or more electrons.
  C lost one or more protons.
  D gained one or more protons.

2 Naturally occurring elements in which of the following groups on the periodic table are **most** likely to form a cation?
  A group 13
  B group 15
  C group 16
  D group 18

3 Which of the following ionic compound formulas is written correctly?
  A $ClNa$
  B P_2Ca_3
  C Na_2O
  D BrK

4 Which of the following compounds contains a polyatomic ion?
  A $NaCl$
  B $NiSO_4$
  C CaF_2
  D Cr_2O_3

5 If calcium metal and fluorine gas are reacted together, what is the correct name for the resulting compound?
  A calcium fluorine
  B fluorine calcide
  C calcium fluoride
  D calcium fluorate

6 If calcium metal and fluorine gas are reacted together, what is the correct formula of the resulting compound?
  A CaF
  B FCa
  C Ca_2F
  D CaF_2

7 Which of the following ionic compounds has the highest melting point?
  A NaF
  B $NaCl$
  C $NaBr$
  D NaI

8 Ionic compounds can conduct electricity
  A in either a solid or a molten state.
  B only if dissolved in water.
  C in a molten state or when dissolved in water.
  D in any state.

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9 What property of solid ionic compounds causes them to shatter when stress is applied?
- **A** brittleness
- **B** strong electrostatic attractions
- **C** weak electrostatic attractions
- **D** malleability

10 Which of the following ionic compounds would be in a solid state when placed in water?
- **A** Na_2S
- **B** $MgCO_3$
- **C** KI
- **D** $NaCO_3$

Short-answer questions

11 Draw an electron transfer diagram to show what would happen when atoms of the following elements are reacted together.
- **a** calcium and fluorine (4 marks)
- **b** magnesium and sulfur (4 marks)

12 Copy and complete the following table. (5 marks)

Element	Subshell electron configuration	Number of valence electrons	What it does to become stable	Charge after gain/loss of electrons	Chemical symbol for the ion
Br					
Al					
Ne					
K					
Cl					

13 Explain how the overall charge of an ionic compound is neutral. (2 marks)

14 Write the names of the following ionic compounds. (6 marks)
- **a** Hg_2S
- **b** $Cr_2(CO_3)_3$
- **c** $NiCl_3$
- **d** SrF_2
- **e** Li_3N
- **f** MgO

15 Write the formulas for these ionic compounds. (6 marks)
- **a** chromium(IV) oxide
- **b** chromium(III) oxide
- **c** magnesium chloride
- **d** nickel(II) sulfate
- **e** iron(III) phosphate
- **f** lead(IV) hydroxide

16 Explain why lithium oxide has a higher melting point than lithium sulfide. (3 marks)

17 Using the ionic bonding model, explain why ionic compounds can be hard as solids but are highly soluble (dissolve readily) in water. (3 marks)

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18 Explain why adding sodium chloride to water is not an example of a precipitation reaction. (2 marks)

19 Consider the chemical reaction that occurs when aqueous solutions of silver nitrate and potassium chloride are mixed.

- **a** Write a fully balanced equation for the precipitation reaction, including states. (3 marks)
- **b** Write a balanced ionic equation, including states, for this reaction. (2 marks)

UNIT 1 HOW CAN THE DIVERSITY OF MATERIALS BE EXPLAINED?

CHAPTER 5 SEPARATING AND IDENTIFYING COMPOUNDS IN MIXTURES

Introduction

INTRODUCTION VIDEO
SEPARATING AND IDENTIFYING COMPOUNDS IN MIXTURES

Have you ever heard of an athlete being banned for using performance-enhancing drugs and wondered how they were caught? Blood and urine samples are taken, but these fluids are filled with thousands of different compounds, which often makes it very difficult to be able to identify a single illegal molecule. However, scientists are able to use a technique, known as chromatography, to initially separate out all of these different substances. Therefore, instead of trying to find the needle in a haystack, they are trying to find a needle in many little piles. As you can imagine, this makes things much easier! While this is a slightly simplified explanation of how drug testing works, it does help explain how chromatography can be incredibly useful for many real-world applications. In this chapter, you are going to recall the properties of compounds that make it possible for them to be separated from a mixture and how chromatography uses those properties to enable the identification of certain molecules.

Curriculum

Area of Study 1 Outcome 1
Separation and identification of the compounds of mixtures

Study Design:	Learning intentions – at the end of this chapter I will be able to:	
• 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	**5A**	**Polarity**
	5A.1	Define ‘solute’ and ‘solvent’
	5A.2	Recall what makes a solute or solvent polar or non-polar
	5A.3	Describe the process of a solute dissolving in a solvent
	5A.4	Explain why polar solutes are soluble in polar solvents and non-polar solutes are soluble in non-polar solvents
	5A.5	Determine the relative solubility of solutes in a given solvent

Study Design:	Learning intentions – at the end of this chapter I will be able to:
- Experimental application of chromatography as a technique to determine the composition and purity of different types of substances, including calculation of R_f values	**5B Chromatography** **5B.1** Describe the principle of thin-layer chromatography (TLC) **5B.2** Describe the process involved in performing a TLC experiment **5B.3** Determine the qualitative purity of a substance from TLC data **5B.4** Determine the composition of a substance or mixture from TLC data **5B.5** Describe a TLC experiment that could be used to determine the purity and/or composition of a substance or mixture **5B.6** Predict the results of a TLC experiment for a known set of compounds **5B.7** Calculate the R_f value of a given substance from experimental data

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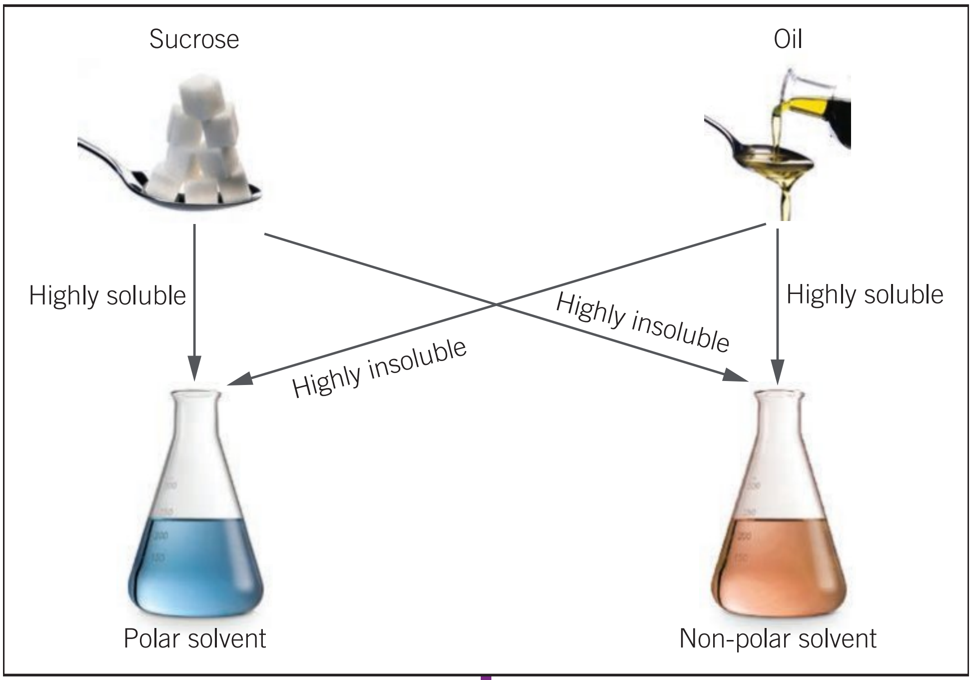
Glossary

Adsorption
Capillary action
Desorption
Mobile phase
Non-polar
Polar
Retention factor (R_f)
Solute
Solution
Solvent
Solvent front
Stationary phase
Thin-layer chromatography

Concept map

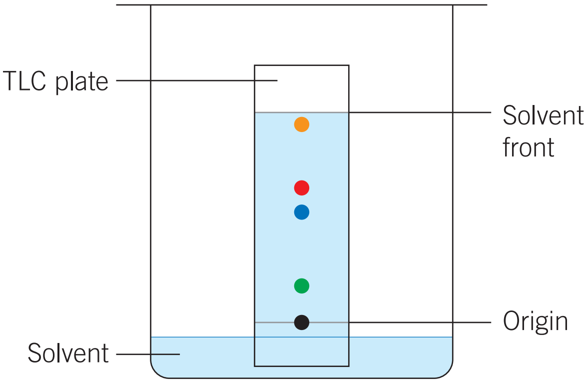
Polarity of a compound affects its properties, including its ability to dissolve in polar and non-polar solvents

5A Polarity



Thin-layer chromatography uses solubility of compounds in different solvents to experimentally separate samples to determine their composition and purity

5B Chromatography



See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

5A Polarity

Study Design:
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

Glossary:
Non-polar
Polar
Solute
Solution
Solvent

ENGAGE

Why don't oil and water mix?

If you've ever tried adding oil to water, like in Figure 5A–1, you'll know that no matter how hard you try, they just won't mix! However, take a spoonful of sugar and add that to a glass of water and give it a quick stir. What do you notice? The sugar has all disappeared! How is it possible that two substances can behave so differently when added to water? While you will learn much more about solubility in Unit 2, this section is going to explore the concept of polarity and how it influences the way things dissolve (or do not dissolve) in different liquids.

LINK **2B** MOLECULAR SHAPES AND POLARITY

LINK **13A** MEASURES OF SOLUBILITY

Figure 5A–1 While oil and water do not mix, a spoonful of sugar will dissolve easily in a glass of water.

EXPLAIN

Polarity

The concepts of electronegativity, polarity, dipole moments and intermolecular forces have all been covered in Chapter 2. These principles are important here as we begin to think about what drives interactions between solutes and solvents, which underlies the basis for how chromatography works later in this chapter. Answer the questions below and on the following page, referring to content from Chapter 2 where necessary.

LINK **2B** MOLECULAR SHAPES AND POLARITY

LINK **2C** PROPERTIES OF COVALENT COMPOUNDS DETERMINED BY INTERMOLECULAR FORCES

Check-in questions – Set 1

1 Define 'electronegativity'.

2 Which of the following bonds are polar? For those that are polar, determine the correct partial charges (δ^+ and δ^-) on each element.

- a $C{=}O$
- b $C{-}C$
- c $N{-}H$

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3 Explain how a molecule can be non-polar, even if it contains polar bonds.

4 State whether each molecule below is polar or non-polar.
  - a NH_3
  - b CCl_4
  - c H_2O
  - d CO_2

5 List the three types of intermolecular forces (hydrogen bonding, dispersion forces, dipole–dipole attractions) in increasing order of strength.

6 Draw the Lewis structures of two water molecules, clearly labelling the hydrogen bonding present between them.

Solute
the solid, liquid or gas that dissolves in a solvent

Solvent
the substance that the solute dissolves in

Solution
a mixture of a solute dissolved in a solvent

Solubility

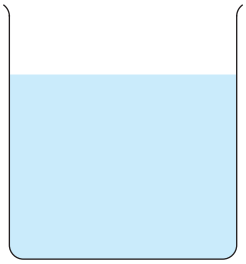
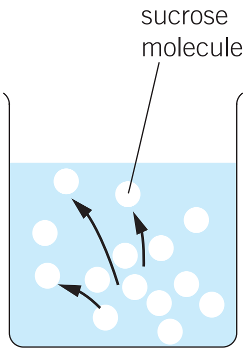
The polarity of a compound has a direct impact on its ability to dissolve in another substance. In these instances, the solid, liquid or gas that dissolves is known as the **solute**, and the substance that it dissolves in is known as the **solvent**. If we know the polarity of our solute and solvent, we can predict whether a **solution** can be formed. To do this, we must first understand what is happening at the atomic level when a substance dissolves in another.

Earlier in this chapter, the example of sugar dissolving in a glass of water was mentioned. Let's continue with that situation but look deeper into what is occurring.

VIDEO 5A–1
THE SOLUBILITY OF A SUGAR CUBE

Firstly, it is important to realise that a grain of sugar is not made up of a single sugar molecule. In fact, every grain of sugar contains trillions of sugar molecules. Sugar is actually a complex molecule called sucrose, the structure of which is shown in Figure 5A–2. In Table 5A–1, the step-by-step process of what is happening as the sugar cube dissolves is outlined.

Table 5A–1 Step-by-step process of what is happening as a sugar cube dissolves

	Observation	Explanation	Diagram
Step 1	A sugar cube and a beaker of water are present next to each other.	The polar sucrose molecules in the cube are attracted to each other through hydrogen bonds (and dispersion forces), which hold the molecules together in the crystal lattice. The polar water molecules are also attracted to each other by hydrogen bonds (and dispersion forces).	 Beaker of water  Sugar
Step 2	When the sugar cube is placed in water and the solution is stirred, the sugar cube begins to get smaller.	There is attraction between the polar water molecules and the polar sucrose molecules. This attraction (hydrogen bonding) is stronger than that of a sucrose molecule to another sucrose molecule, which leads to them being pulled apart from each other.	

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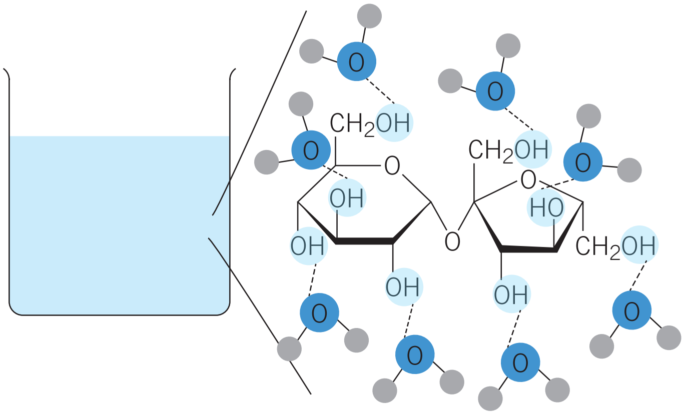
	Observation	Explanation	Diagram
Step 3	A clear solution of water is left in the beaker. No solid sugar grains can be seen.	All the sucrose molecules are now surrounded by water molecules and forming hydrogen bonds with them.	 Dashed lines = hydrogen bonds

It should be noted that this explains what happens when a covalent molecule dissolves in water; however, the process is slightly different if the solute is an ionic compound. In this instance, the ionic compound would completely separate (dissociate) into individual ions at step 2, rather than individual molecules. As an example, when placed in water, crystals of table salt (or sodium chloride) would break apart into positively charged sodium cations and negatively charged chloride anions. This was shown in Figure 4C–3 in Chapter 4.

LINK **4C** PRECIPITATION REACTIONS

If we come back to our example of the sugar cube – why is dissolving possible? To understand this, we must look at the structure of sucrose (Figure 5A–2). All of the oxygen–hydrogen bonds (highlighted in blue) are polar. It can be seen that there are eight of these in a sucrose molecule, which makes the overall compound highly polar. These polar components of the sucrose molecule are able to form hydrogen bonds with water molecules. It is for this reason that sucrose is very soluble in water.

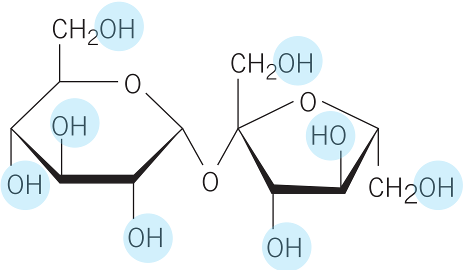


Figure 5A–2 A sucrose molecule

Solubility is often talked about in the general terms that **polar** compounds are soluble in polar solvents and **non-polar** compounds are soluble in non-polar solvents. Similarly, we could say that non-polar compounds are insoluble in polar solvents and polar compounds are insoluble in non-polar solvents (Figure 5A–3). Overall, this is true; however, the reality is slightly more complicated. The situation can be explained through looking at Figure 5A–4. In this figure, we can see that polarity is a continuous spectrum going from highly non-polar through to highly polar. This will become really important in the next section on chromatography in which we discuss how we separate compounds.

Polar
a molecule with two distinct poles of partial electric charge, a partial positive charge on one side of the molecule and a partial negative charge on the other side

Non-polar
a molecule with no permanent net dipole

LINK **5B** CHROMATOGRAPHY

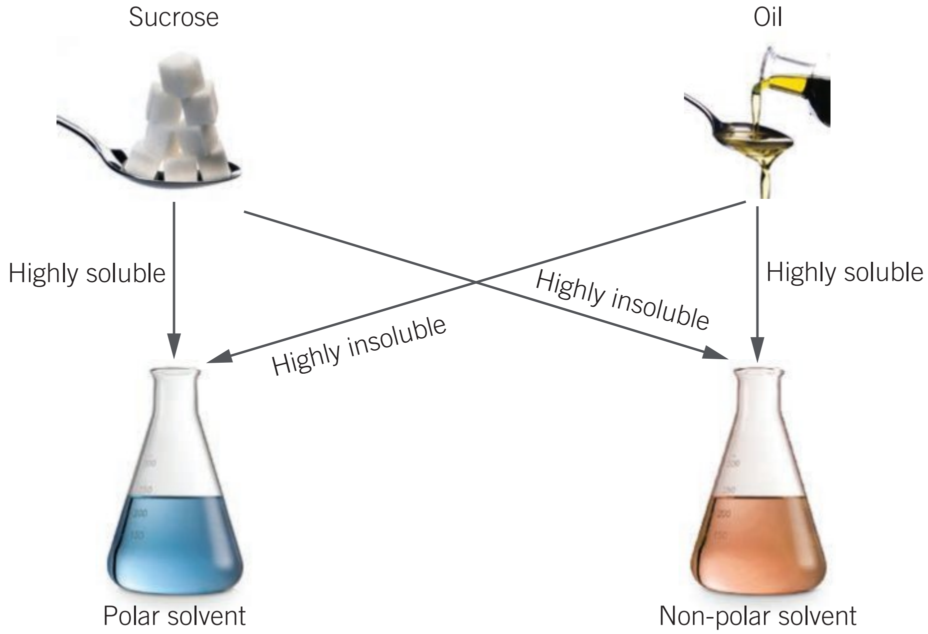


Figure 5A–3 Generally, polar solutes, such as sucrose, will dissolve in polar solvents. Non-polar solutes, such as oil, will dissolve in non-polar solvents.

WORKSHEET 5A–1 POLARITY AND SOLUBILITY

13A MEASURES OF SOLUBILITY LINK

Figure 5A–4 considers three compounds, which are named compounds A, B and C. Compound C is highly polar and compound A is highly non-polar, whereas compound B is moderately polar. If three beakers of water were taken and we dissolve increasing amounts of the three compounds at the same time, we would find that compound A would be the first one to reach its solubility limit (that is, if any more of the solid compound is added it won't dissolve), whereas compound C would be the last. This concept of solubility will be discussed in more detail in Chapter 13A.

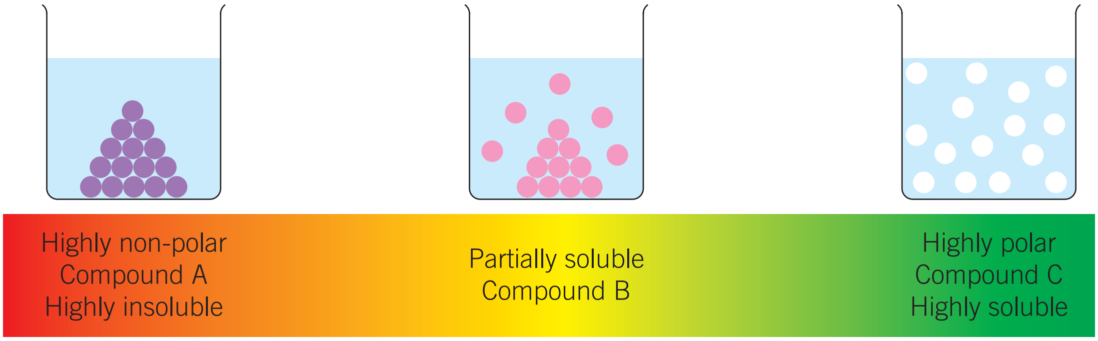


Figure 5A–4 Polarity and solubility occur across a spectrum. Two polar substances, such as compound B and compound C, will both dissolve in a polar solvent, but by differing amounts.

5A SKILLS

VIDEO 5A–2 SKILLS: MAKING CONNECTIONS BETWEEN TOPICS

Making connections between topics

In VCE Chemistry, you are expected to be able to apply your knowledge by making connections between different areas of the Study Design. The questions that test your knowledge in this way require higher-order thinking skills, and answering them correctly will be the difference between an average mark and a very good mark. You will have noticed by now that throughout each chapter there are icons in the margin that show a link between the content you are currently reading and concepts you have already read about in previous sections (or will read about in later sections). These links are designed to help you make connections between key concepts and take your understanding to a deeper level.

Consider this section in relation to the following question as an example.

Question:

Explain why sucrose is soluble in water.

Answer:

Worded answer	Other chapters this links to
Water molecules are polar, non-symmetrical, bent-shaped molecules containing partial negative and partial positive charges due to differences in electronegativity between O and H atoms.	Chapter 2B, where the idea of the shapes of molecules was discussed, with water molecules having a non-symmetrical bent shape.
Electronegativity is the strength of attraction of electrons to the positively charged nucleus of an atom.	Chapter 1C defines this term as shown.
This difference in electronegativity is due to the increased core charge, which is the number of protons minus the number of inner-shell electrons.	Chapter 1C would allow you to explain the difference in electronegativity between the O and H atoms in water as a result of core charge, as defined.
Oxygen's position in the periodic table is group 16, whereas H is in group 1. Even though they are both non-metals, the greater number of protons compared to inner-shell electrons means that oxygen has a higher core charge.	Chapter 1B covers the electron configurations, determining position on the periodic table and then relating this to Chapter 1C to use these configurations to determine their position based on period and group.
As a water molecule is therefore a polar molecule due to containing an O–H bond, it can form hydrogen bonds with other molecules, such as another water molecule or sucrose.	Chapter 2C is where hydrogen bonding is first introduced as a stronger type of intermolecular force existing between two molecules.
A sucrose molecule is a polar molecule, due to its non-symmetrical structure.	Chapter 2B extends ideas from Chapter 1C, relating the concept of electronegativity to polarity.
Sucrose contains many O–H bonds and thus will be able to form hydrogen bonding with other molecules, such as water, which also contain O–H bonds.	Chapter 2C explains that hydrogen bonding exists between a hydrogen atom and either an N, O or F atom of another molecule. The example used in this section is sucrose. Look at the structural diagram shown in Figure 5A–2. Sucrose has eight O–H bonds within a molecule, and therefore this will allow it to form a hydrogen bond with a water molecule (which also has an O–H bond).

When mixed together, the hydrogen bonds that form between a water molecule and a sucrose molecule are stronger and more favourable than the hydrogen bonds between only water molecules or between only sucrose molecules.	Chapter 5A extends your knowledge by stating that the ability for a molecule to dissolve in water is two-fold. Firstly, it needs to be polar. Secondly, the hydrogen bonding that occurs between a water molecule and another molecule, like sucrose, must be stronger or more favourable than the hydrogen bonding occurring between water molecules themselves or sucrose molecules themselves.
Sucrose is soluble in water.	Finally, link back to context of question.

Thus, you could provide a large amount of detail in your answers to questions. While this is not always required in an assessment and you may run out of time, this is good practice for your own revision to build a deeper understanding of the course, which will help in future topics.

It is important that when revising and making connections between these topics or answering them in assessment situations, you are able to use the **DER** framework. This is colour-coded, as outlined below and in the example solution above and on the previous page.

- **D**efine any key terms (shown in blue).
- **E**xplain any key ideas/concepts, similarities or differences between two molecules (shown in red).
- **R**elate back to the context of the question (shown in green).

Section 5A questions

1 The preparation of an aqueous solution involves dissolving a solute in water.

  a Define 'solute'.

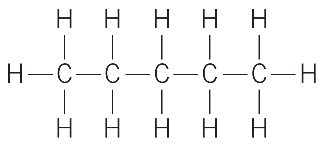
  b Define 'solvent'.

  c Sugar, a highly polar substance, is dissolved in many commercial soft drinks. Explain the process by which sugar would dissolve in a bottle of soft drink (assume it is largely water). Your answer should reference the polarity of both sugar and water.

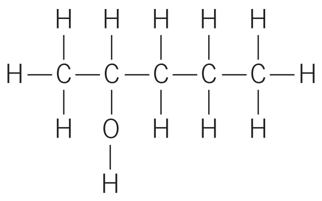
  d Carbon dioxide is also added to soft drinks to make them fizzy. Using your knowledge of the nature of carbon dioxide molecules from Chapter 2B and 2C, explain whether the sugar or carbon dioxide would dissolve more easily in the water.

2 What is the main difference that occurs when an ionic compound is dissolved in water compared to when a polar covalent compound dissolves? Use a fully labelled diagram to support your answer.

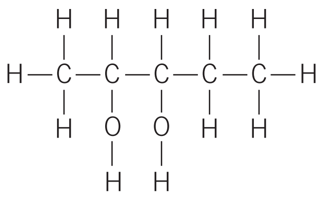
3 A chemist is performing an experiment in which three different compounds will be mixed with water. The three compounds are pentane, pentan-2-ol and pentane-2,3-diol. Their structures are shown below.



Pentane



Pentan-2-ol



Pentane-2,3-diol

  - a Copy out the structures on the previous page and circle any polar bonds that you can identify for each molecule.
  - b List the compounds in order from most soluble to least soluble in this experiment and explain why you have chosen that sequence.
  - c What type of solvent would be the most appropriate to dissolve the least polar solute identified in part **b**?

4 Sucrose is a highly polar compound that is commonly called table sugar. A tablespoon of sucrose is added to two beakers, A and B, which contain unknown solvents. After stirring for a couple of minutes, it is observed that there are still grains of sucrose present at the bottom of beaker A, but none in beaker B. Explain what these observations tell you about the relative polarity of the solvents in beakers A and B.

5 Explain why ammonia, NH_3, is soluble in water, H_2O. In your response, refer to the intermolecular forces in the solute and the solvent.

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5B Chromatography

Study Design:
Experimental application of chromatography as a technique to determine the composition and purity of different types of substances, including calculation of R_f values

Glossary:
Adsorption
Capillary action
Desorption
Mobile phase
Solvent front
Stationary phase
Retention factor (R_f)
Thin-layer chromatography

ENGAGE

Chromatography at home

Imagine you are holding a garage sale on the weekend. You spend time using coloured markers to create a really eye-catching sign to display on the street. However, five minutes after you put up the sign on the tree outside, it starts to rain. You watch on as the sign becomes wet and the letters begin to drip. While you are really sad, you also notice something interesting. Although the letter 'G' was written in red, as the colour begins to run, you observe that the line contains not just the colour red, but also blue, green and yellow. The same thing happens with all of the other coloured letters as well! This happens because that single sample of ink is actually made up of a number of different compounds that have different colours. As the water moves the ink across the paper, these colours begin to separate. Although this is a very basic example, it perfectly demonstrates the experimental process of chromatography, which you will learn more about in this section.

Stationary phase a solid, liquid or gel that remains static during a chromatography experiment

Mobile phase the liquid or gas that flows during a chromatography experiment, which transports the compounds to be separated at different rates over the stationary phase

EXPLAIN

Separation of mixtures

In chemistry, there is often a need to test samples to assess their purity, composition or properties.

The separation of compounds in mixtures by chromatography relies on a general principle of differing levels of attraction to an immobile solid surface and a solvent. The solid surface is known as the **stationary phase**, whereas the solvent is known as the **mobile phase**. When a substance interacts with the stationary phase, this is called **adsorption**. When it releases from the stationary phase and interacts with the mobile phase, it is called **desorption**. The more that a substance interacts with the stationary phase, the slower it moves through the chromatography process. On the other hand, the more a substance interacts with the mobile phase, the faster it moves through. In this chapter, we are going to focus on one chromatography technique in particular – thin-layer chromatography. You will learn more about other chromatography methods, such as high-performance liquid chromatography, in Unit 3.

Adsorption the process in which a compound interacts with the stationary phase during a chromatography experiment

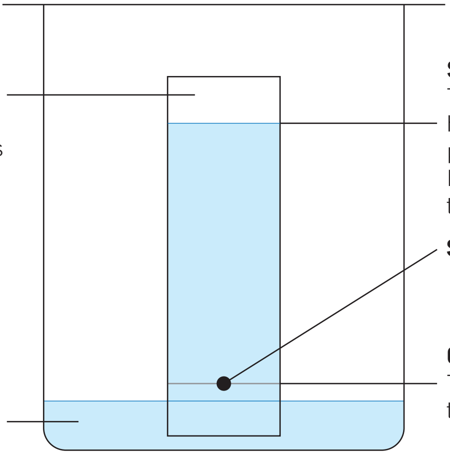
Desorption the process in which a compound interacts with the mobile phase during a chromatography experiment

UNIT 3 LINK

ISBN 978-1-009-22964-7 Drummond et al

Thin-layer chromatography

Thin-layer chromatography (TLC) is a common technique in chemistry that is used to identify compounds present in mixtures, monitor the progress of reactions and to assess them for purity. TLC is an effective technique for these sorts of analyses, as it provides results quickly, only requires a very small amount of sample and doesn't require expensive instrumentation. TLC works on a similar basis to other forms of chromatography in that it separates molecules based on their polarity. The key features of a TLC experiment are shown in Figure 5B–1.



TLC plate (stationary phase)
Typically a glass, plastic or aluminium plate that is covered with a silica or aluminium oxide gel. This gel is polar and acts as the stationary phase.

Solvent (mobile phase)
The mobile phase for the chromatography experiment. The solvent can be altered to suit the types of compounds that are being separated (e.g. using a polar or non-polar solvent).

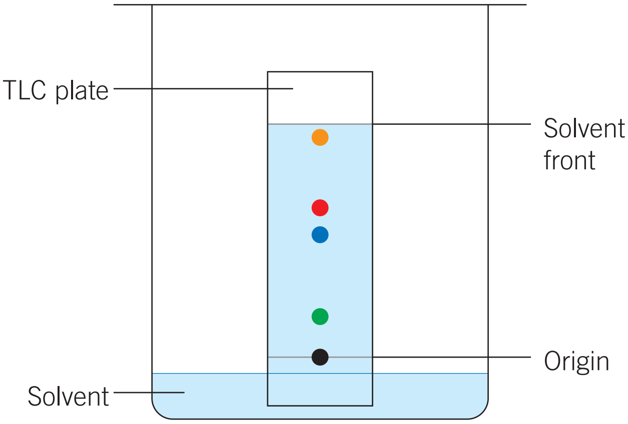
Solvent front
The distance the solvent has travelled up the TLC plate/stationary phase.
Note that no compound can travel further than the solvent.

Sample

Origin
The point on the plate where the original sample is added.

Figure 5B–1 Key features of a TLC experiment

To run a TLC experiment, the following method is used.

Step 1	The sample(s) are spotted onto the origin.	 **Figure 5B–2** A general result from a TLC experiment. In this experiment, the mobile phase is a non-polar solvent.
Step 2	The TLC plate (stationary phase) is lowered into the beaker containing the solvent. It is important that the origin on the TLC plate is not lowered below the level of the solvent. If this happens, the samples can dissolve directly into the solvent in the beaker and will not be carried up the TLC plate.	
Step 3	The experiment is left to run. During this time, the solvent (mobile phase) will move up the TLC plate through **capillary action**. It is possible to monitor the progress of the experiment by observing the solvent front progressing up the plate. When the **solvent front** nears the top of the TLC plate, the experiment is stopped by removing the plate from the beaker.	
Step 4	Once the experiment is stopped, the results are then visualised. There are several ways this can be done, but most commonly they involve placing the plate under UV light (which illuminates the spots) or reacting with a reagent that causes the spots to change colour (or sometimes the compounds themselves are coloured). It is not important for you to understand the mechanisms behind how these visualisation techniques work.	

Thin-layer chromatography an analytical technique for separating and identifying mixtures in which a silica or polar aluminium oxide gel acts as the stationary phase

VIDEO 5B–1
THIN-LAYER CHROMATOGRAPHY

Capillary action the movement of a substance due to adhesion (attractive forces between the solvent and stationary phase) and cohesion (attractive forces between molecules of the solvent)

Solvent front the distance the solvent has travelled up the TLC plate/stationary phase

It can be seen in Figure 5B–2 that the sample contained five different compounds, represented by the black, green, blue, red and orange dots. Based on this, we can answer several questions about the sample and the compounds it contains.

1 Was the sample pure?
No, this is an impure sample, as more than one spot was observed after separation by TLC. If it was a pure sample (which means it only contained one compound), we would only observe a single spot.

2 Which compound is the most polar?
As the stationary phase is polar on a TLC plate, the most polar compound will have the greatest interaction (adsorption) with this phase. This will result in it travelling the shortest distance. Therefore, we can determine that the black spot is the most polar compound in this sample.

3 Which compound is the least polar?
This will be the opposite of Question 2. The least polar molecule in the sample will have the least interaction (desorption) with the polar stationary phase on the TLC plate. This means that it will be relatively more attracted to the non-polar mobile phase and so travel the furthest distance on the plate. Therefore, we can determine that the orange dot is the least polar compound in the sample.

4 What is the order of polarity of the compounds in the sample?
Based on our answers to Questions 2 and 3, we can determine that the order of polarity of the compounds in the sample, going from most polar to least polar, is black, green, blue, red then orange.

5A POLARITY **LINK**

NOTE

Importantly, and you may already be questioning this, we cannot say whether the compounds are polar or non-polar. If you think back to Chapter 5A, it was discussed that polarity (and therefore solubility) occurs across a spectrum. For this reason, it is possible to separate only polar (or only non-polar, or a combination of the two) compounds using a TLC plate. This is why it is important that we have used the terms ‘most polar’ and ‘least polar’ in our answers to these questions, rather than ‘polar’ and ‘non-polar’. The orange spot in this example, while it has travelled the furthest and therefore interacted more than any of the other compounds with the non-polar mobile phase, could still be a polar molecule. All that we can determine is that it is less polar than the other compounds.

5 How would the separation of the components change if the mobile phase was changed to a polar solvent?
The order of the compounds on the plate wouldn’t change; however, the distances that they travel (their **retention factor, R_f**, values) may change. This is because the stationary phase used in TLC is usually highly polar, so it is very likely that it will always be more polar than the mobile phase, even if a polar solvent is used. In that case, it would still be the least polar substance in the sample that will interact the most with the mobile phase and so travel the furthest up the TLC plate. Similarly, the most polar compound will be the one that will interact the most with the stationary phase and travel the least distance up the TLC plate.

Retention factor (R_f) the distance between the origin and the spot of interest relative to the distance between the origin and the solvent front

6 What is the identity of the compounds in the sample?
This is not a question that we can answer at this stage. To be able to answer this, we need more information, such as can be obtained by using a reference sample or calculating the R_f value of the compound.

Identifying compounds on a TLC plate

LINK **6A** RELATIVE ATOMIC AND ISOTOPIC MASSES

LINK **UNIT 4**

While there are much better methods for the absolute identification of molecules in a mixture, such as mass spectrometry, which you will learn about in Chapter 6A and in more detail in Unit 4, there are ways that the presence of compounds can be predicted from TLC. One of these is through the calculation of the retention factor. Retention factor (R_f) is a relative measure of the distance travelled by a compound compared to the solvent front. To calculate the retention factor, the following equation is used:

$$R_f = \frac{\text{distance between the origin and the spot of interest}}{\text{distance between the origin and the solvent front}}$$

This can be seen more clearly in Figure 5B–3.

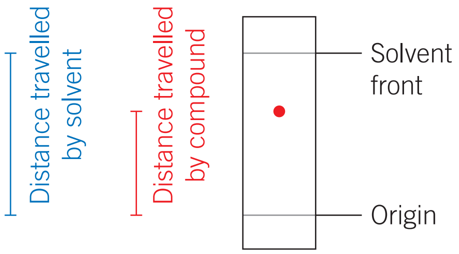


Figure 5B–3 The retention factor (R_f) can be calculated by dividing the distance travelled from the origin by the compound of interest (red spot) by the distance travelled from the origin to the solvent front.

The R_f value can provide some relative information about the properties of different spots on a TLC plate. In general, a smaller R_f value indicates that the compound has interacted more with the stationary phase. In the case of a TLC plate with a polar stationary phase and a non-polar mobile phase, it could be concluded that the molecule with the smallest R_f value would be the most polar and the compound with the largest R_f value would be the least polar (or most non-polar).

NOTE

It is important to note that you can never have an R_f value for a compound that is greater than 1, as the compound cannot travel further than the solvent front.

So how does calculating an R_f value help you identify the compound that is present in a sample? If the TLC experiment is run under the same conditions, the same compound will have the same R_f value. Therefore, you can compare the R_f values of a spot from an unknown sample and a spot from a pure sample of known identity to get information about the unknown sample. It is important to note that you cannot conclude the identity of a compound with 100% certainty using this method – it is likely that there will be more than one compound with the same R_f value. However, it is a good starting point, especially if you have some idea to begin with as to what may be present in the starting mixture.

In a similar way, the use of a standard on a TLC plate can help identify potential compounds that may be present in the mixture of interest. We can understand this by looking at Worked example 5B–1.

Worked example 5B–1: Interpreting TLC plates

VIDEO
WORKED EXAMPLE 5B–1

Recently, two new multivitamin products have become available in supermarkets. Advertising for these products suggests that they help deal with the 'winter blues' by improving energy and boosting vitamin D levels. Doctors have contacted a research lab after they had several patients using these products, which we will call Product A and Product B, to boost their vitamin D levels during winter. When the doctors performed routine blood tests to monitor the amount of vitamin D present, they didn't consistently see any increase. Figure 5B–4 represents a TLC experiment that the scientists have performed to analyse Product A and Product B.

Using this data, determine whether the multivitamins from Product A and Product B actually contain vitamin D.

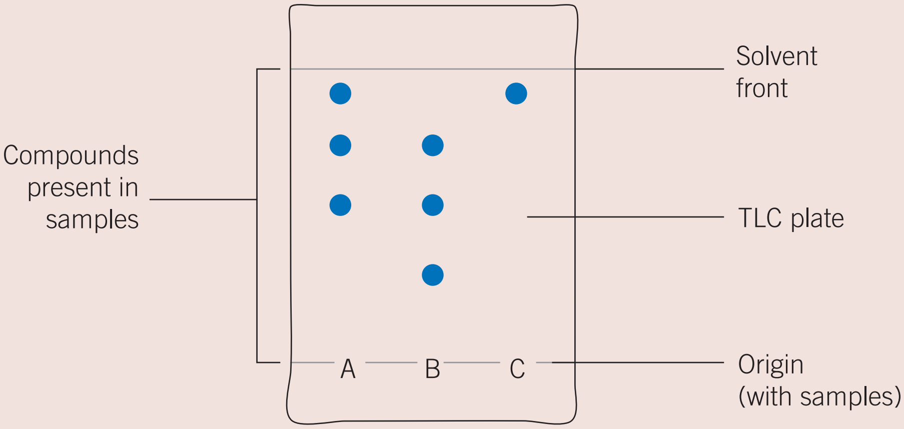


Figure 5B–4 The results of a TLC experiment for two multivitamins, Product A (A) and Product B (B), as well as a reference sample of vitamin D (C).

WORKSHEET 5B–1
ANALYSING CHROMATOGRAPHY

Solution:

It can be observed that three samples are present on the origin of the TLC plate corresponding to Product A, Product B and a vitamin D reference sample (Product C). Product A and Product B both contain multiple components, as indicated by more than one spot present in each sample. The vitamin D reference sample is pure, as shown by the presence of only a single spot in that sample.

Importantly, the question was around whether Product A and Product B contained vitamin D. To answer this, we can use the reference sample as a guide. It can be observed that Product A has a spot that has travelled the same distance along the TLC plate (and therefore will have the same R_f value) as the spot present in the reference sample. However, Product B does not have a spot at this location on the TLC plate. From this, it would be reasonable to suggest that Product A does contain vitamin D, but Product B does not.

5B SKILLS

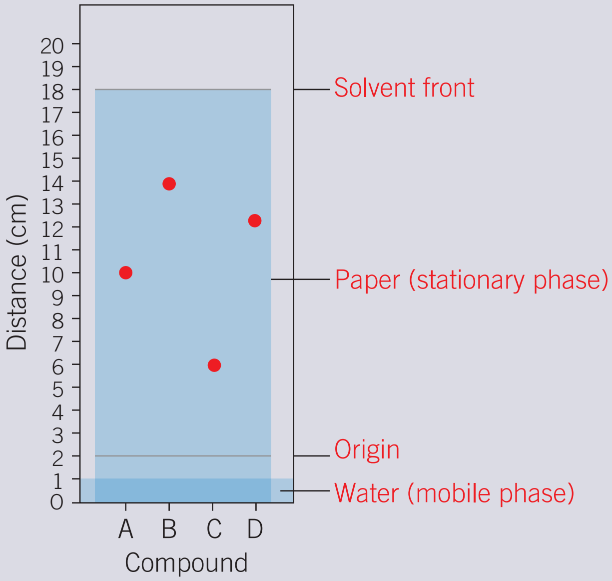
VIDEO 5B–2
SKILLS: USING KEY TERMINOLOGY

Using key terminology

When describing the movement of certain compounds or components of substances due to chromatography, it is important to use precise key terminology and not oversimplify your language. Take the following question as an example.

Question:

Explain the difference between compounds B and C in the paper chromatography below when water is used as the solvent.



Initial answer:

All compounds separate out at different distances along the paper with sample B moving the furthest.

Comments:

While this is not wrong, the language used is not specific enough for the level expected in VCE. Before providing a more detailed answer to this question, it might be good to do two things.

1. Label the key parts of the diagram (shown above in red).
2. List all the key terms that should be included in your answer (listed below in blue).
   Paper
   Polar solvent
   Origin
   Solvent front
   Stationary phase
   Water
   Mobile phase
   Adsorption
   Desorption
   Retention factor (R_f)
   Polar
   Non-polar

When it comes to writing your answer, you can use dot points if it is easier to keep it succinct, while still including the key terms that you identified.

Complete answer:
- Compounds A, B, C and D were spotted onto chromatography paper at the origin line.
- When the bottom of the paper, below the origin, is immersed into the polar solvent, the water (mobile phase) moves slowly up the paper (stationary phase).
- Compounds that are less polar (or more non-polar) experience more adsorption to the stationary phase and consequently travel less distance. This can be seen with compound C.
- Compounds that are more polar tend to experience desorption from the stationary phase and migrate with the solvent further along the mobile phase. This can be seen with compound B.
- The solvent front is how far the solvent, water, has been adsorbed along the stationary phase from the origin, measured in centimetres.
- As such, compound B will have a larger retention factor compared to compound C.

Note: If this question had further parts, you would probably guess it would ask you to measure the exact retention factor for some (or all) of these compounds. It is important to note that the origin is not always labelled on these diagrams as beginning at distance '0 cm'. Here, the origin is at a distance of 2 cm. Therefore, you need to account for this in any of your R_f calculations comparing how far the compound has moved along the stationary phase compared to the solvent front.

Section 5B questions

1 Explain how thin-layer chromatography separates the components of a mixture.

2 One of the pieces of information about a compound that can be determined from a TLC experiment is the retention factor.
  - a What is the retention factor?
  - b If a TLC experiment is performed using a polar stationary phase and a non-polar mobile phase, will a polar or non-polar compound in the sample have a higher R_f value? Explain your answer.
  - c A student performs a TLC experiment and calculates the R_f value of the compound of interest to be greater than 1. What can be concluded from this?

3 Your sample mixture contains two polar compounds. Will it be possible to separate these using a silica TLC plate? Explain why or why not.

4 Draw the general setup of a TLC experiment. Ensure that your diagram is fully labelled with all the key features.

5 Describe how the purity of a sample can be determined from a TLC experiment.

Chapter 5 review

Summary

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

Success criteria – I am now able to:		Linked question
5A.1	Define 'solute' and 'solvent'	3☐
5A.2	Recall what makes a solute or solvent polar or non-polar	
5A.3	Describe the process of a solute dissolving in a solvent	4☐
5A.4	Explain why polar solutes are soluble in polar solvents and non-polar solutes are soluble in non-polar solvents	9☐
5A.5	Determine the relative solubility of solutes in a given solvent	6☐, 13☐
5B.1	Describe the principle of thin-layer chromatography (TLC)	1☐, 7☐, 10☐
5B.2	Describe the process involved in performing a TLC experiment	5☐, 8☐, 12☐
5B.3	Determine the qualitative purity of a substance from TLC data	14☐
5B.4	Determine the composition of a substance or mixture from TLC data	11☐, 14☐
5B.5	Describe a TLC experiment that could be used to determine the purity and/or composition of a substance or mixture	11☐, 14☐
5B.6	Predict the results of a TLC experiment for a known set of compounds	2☐, 11☐, 13☐
5B.7	Calculate the R_f value of a given substance from experimental data	11☐

Multiple-choice questions

1 What property of compounds is the basis of separation in thin-layer chromatography?

A size
B shape
C polarity
D symmetry

2 A TLC experiment is performed using a polar stationary phase. Which R_f value would best correspond to the most non-polar substance in the sample?

A 0.1
B 0.3
C 0.5
D 0.7

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3 The solute in an aqueous solution is
- **A** the substance that is dissolved in water.
- **B** the liquid that the substance dissolves in.
- **C** the combination of the dissolved substance and the liquid it is dissolved in.
- **D** always a covalent molecule and always polar.

4 When a covalent substance is fully dissolved
- **A** solid particles can be observed in the solution.
- **B** solute molecules are interacting with each other.
- **C** the solution is made up of many individual ions interacting with water molecules.
- **D** all of the solute molecules are interacting with water molecules.

5 Spots on a TLC plate can be visualised using
- **A** an ultraviolet light.
- **B** R_f values.
- **C** the solvent front.
- **D** a non-polar mobile phase.

6 Which of the following will most readily dissolve in a glass of water?
- **A** carbon dioxide (CO_2)
- **B** methane (CH_4)
- **C** ammonia (NH_3)
- **D** hydrogen (H_2)

7 Which of the following pieces of information would you **not** be able to determine using data from a TLC experiment?
- **A** purity
- **B** reactivity
- **C** relative polarity
- **D** retention factors

8 To set up a TLC experiment, the sample is placed
- **A** in the solvent.
- **B** at the origin.
- **C** at the solvent front.
- **D** in the mobile phase.

9 The best explanation for why a non-polar compound will not dissolve in a polar solvent is because
- **A** it cannot break up into smaller components.
- **B** the attraction of the solute to the solvent is too strong.
- **C** the size of solute is too big to allow it to be soluble.
- **D** the attraction of the solvent molecules to each other is greater than to the solute.

10 Which of these is an advantage of TLC as an analytical method?
- **A** expensive
- **B** uses a large amount of sample
- **C** provides results quickly
- **D** requires sophisticated equipment

Short-answer questions

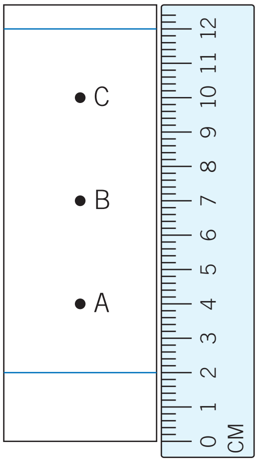
11 A chemistry student wanted to monitor a chemical reaction that was being conducted in which a single reactant was being converted to a single product. A student took a sample during the reaction and spotted it onto a TLC plate. The TLC plate that was used had a polar stationary phase and a non-polar mobile phase. The results are shown on the right.

a Calculate the R_f values for compounds A, B and C. (2 marks)

b Which compound is the most polar? (1 mark)

c Which compound would you expect to have the highest R_f value if the experiment was changed to use a polar mobile phase and a non-polar stationary phase? (2 marks)

d It is thought that compound C might be a contaminant from the laboratory. Describe how you could test this hypothesis. (2 marks)



12 You plan to conduct an experiment to analyse the samples that you've prepared. You are going to use a silica TLC plate and hexane (C_6H_{14}) as the solvent.

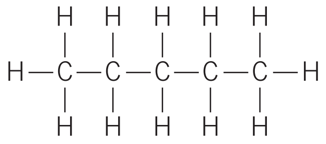
a Name the stationary phase and the mobile phase in your experiment. (2 marks)

b The steps you will perform in this experiment are listed below. Place them in the correct order for the procedure.

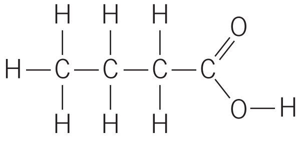
A Samples are spotted onto the origin.
B The TLC plate is analysed under UV light.
C Samples are prepared.
D The R_f values are calculated for the observed compounds.
E The TLC plate is lowered into the solvent.
F The mobile phase travels up the TLC plate as the solvent front. (3 marks)

13 The following questions relate to these chemical structures.

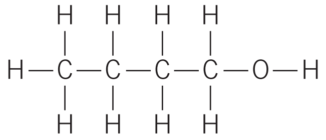
Compound A



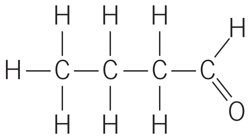
Compound B



Compound C



Compound D



a Which compound will be the most soluble in a non-polar solvent? Explain. (2 marks)

b Which compound will be the most soluble in a polar solvent? Explain. (2 marks)

c Which compound will be the least soluble in a polar solvent? Explain. (2 marks)

d You separate these compounds by TLC using a polar stationary phase and a non-polar mobile phase. Which compound will have the highest R_f value? (1 mark)

14 You have synthesised a new batch of artificial sweetener, but the company requires you to prove the purity of the product before it can be sold.

a Design and describe a thin-layer chromatography experiment that you would perform to test the purity of your artificial sweetener.

b Having conducted the experiment, you calculate the following data.

Spot #	R_f values
1	0.34
2	0.52
3	0.80

Considering that the solvent front travelled 85 mm, sketch a diagram that represents how this data would appear on the TLC plate.

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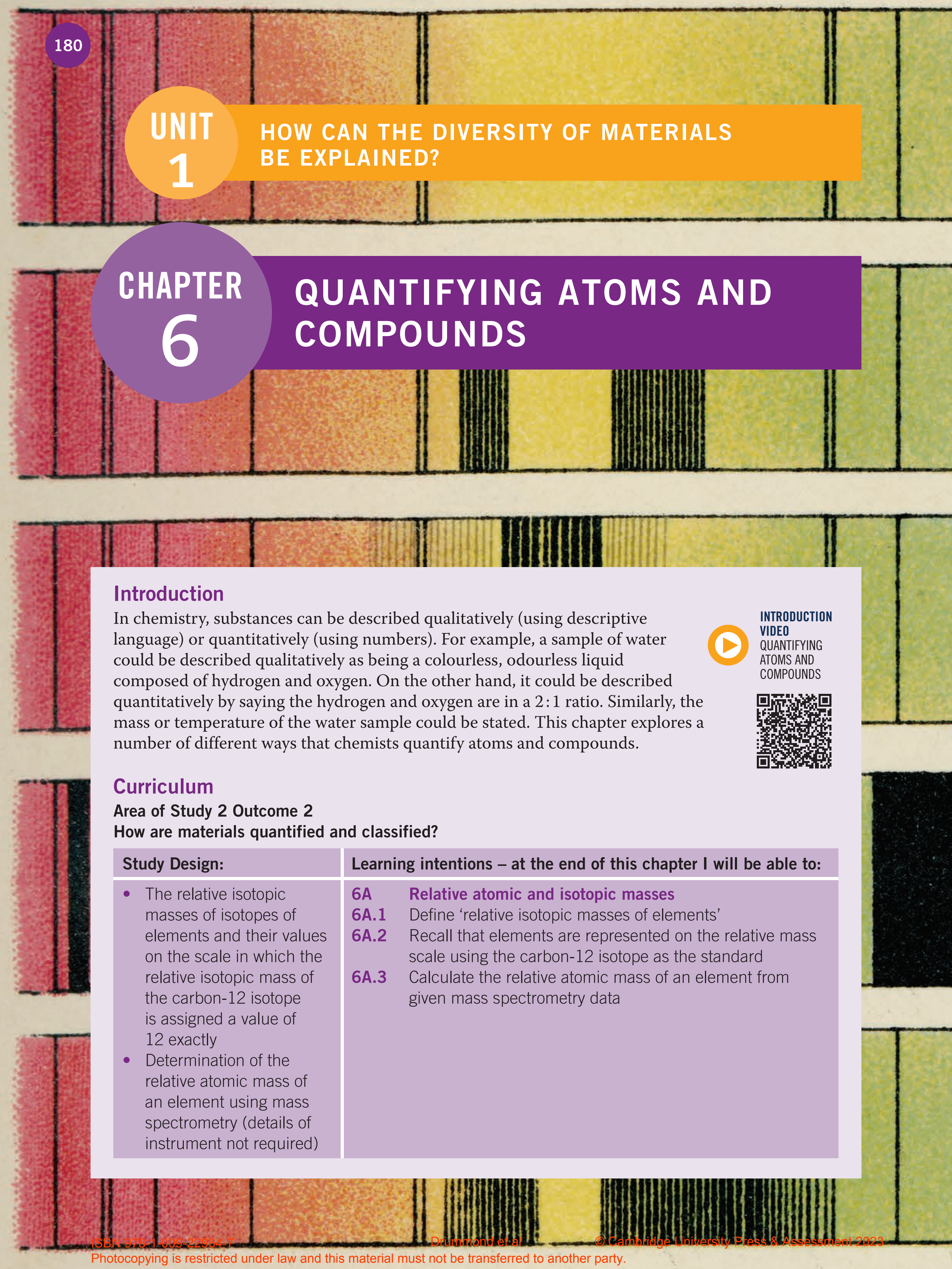


UNIT 1 HOW CAN THE DIVERSITY OF MATERIALS BE EXPLAINED?

CHAPTER 6 QUANTIFYING ATOMS AND COMPOUNDS

Introduction

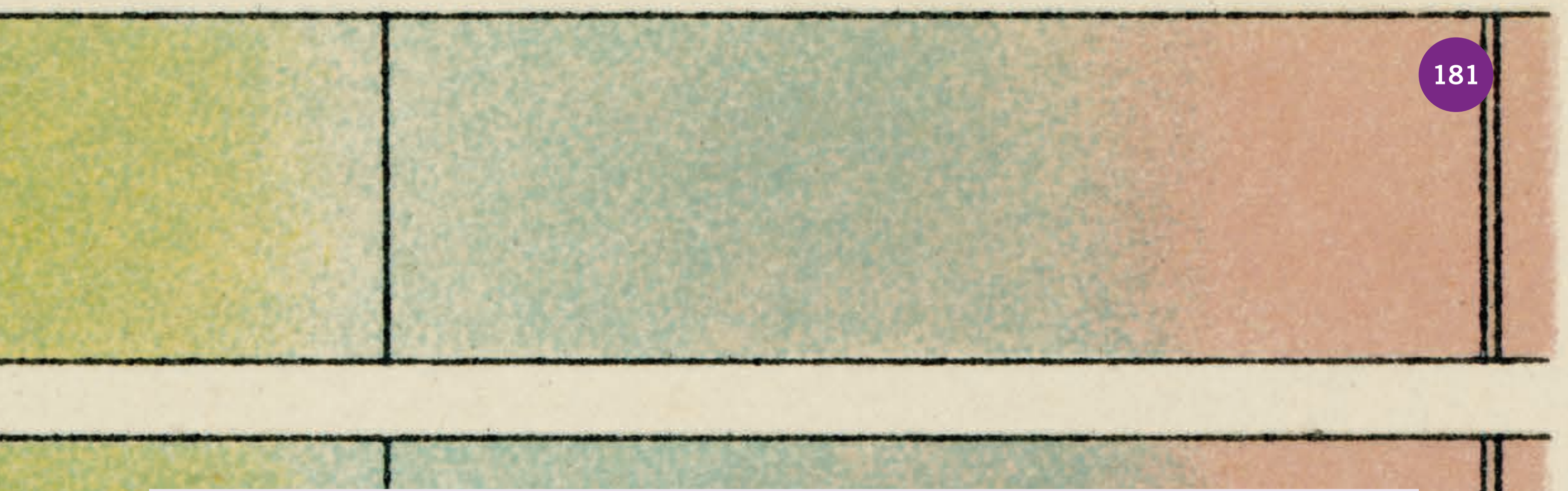
In chemistry, substances can be described qualitatively (using descriptive language) or quantitatively (using numbers). For example, a sample of water could be described qualitatively as being a colourless, odourless liquid composed of hydrogen and oxygen. On the other hand, it could be described quantitatively by saying the hydrogen and oxygen are in a 2 : 1 ratio. Similarly, the mass or temperature of the water sample could be stated. This chapter explores a number of different ways that chemists quantify atoms and compounds.

INTRODUCTION VIDEO
QUANTIFYING ATOMS AND COMPOUNDS

Curriculum

Area of Study 2 Outcome 2
How are materials quantified and classified?

Study Design:	Learning intentions – at the end of this chapter I will be able to:
• 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 Relative atomic and isotopic masses** **6A.1** Define ‘relative isotopic masses of elements’ **6A.2** Recall that elements are represented on the relative mass scale using the carbon-12 isotope as the standard **6A.3** Calculate the relative atomic mass of an element from given mass spectrometry data



Study Design:	Learning intentions – at the end of this chapter I will be able to:
• 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	**6B The mole** **6B.1** Recall Avogadro's constant **6B.2** Convert between moles and the number of atoms or molecules using Avogadro's constant **6B.3** Determine the moles of atoms or molecules in a sample of known mass **6B.4** Determine the mass of an atom or molecule given the moles **6B.5** Determine the molar mass of an unknown compound given the mass and moles
• 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	**6C Percentage composition and empirical formula** **6C.1** Determine the molar mass of a compound from its formula **6C.2** Determine the percentage composition by mass of covalent compounds **6C.3** Define 'empirical formula' **6C.4** Determine the empirical formula of a covalent compound from its percentage composition by mass **6C.5** Determine the molecular formula from an empirical formula given its molar mass

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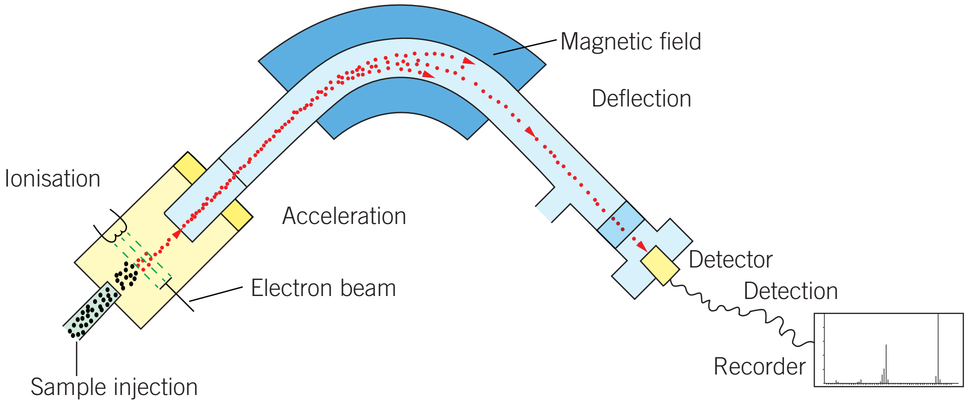
Glossary

Avogadro's constant
Cation
Empirical formula
Ionisation
Mole
Molecular formula
Percentage abundance
Relative atomic mass
Relative isotopic mass
Weighted average

Concept map

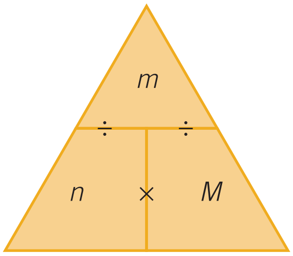
Relative mass makes calculations easier for samples made up of many isotopes and is measured with mass spectrometry

6A Relative atomic and isotopic masses



The mole is the SI unit of measurement for the amount of a substance

6B The mole



Calculating percentage composition by mass of covalent compounds and determining empirical and molecular formulas from percentage composition

6C Percentage composition and empirical formula

See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

6A

Relative atomic and isotopic masses

Study Design:

- 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)

Glossary:

Cation
Ionisation
Percentage abundance
Relative atomic mass
Relative isotopic mass
Weighted average

ENGAGE

Relative atomic mass

Imagine weighing all the members of a football team, and then calculating the average mass. You might expect that someone in the team would have the same mass as the average. However, this won't necessarily always be the case.

In fact, when we talk about elements and their average mass, this often isn't the case. If we look at Figure 6A–1, we can see that the relative atomic mass of gadolinium is 157.3. This represents the average mass of a sample of gadolinium atoms. However, no matter how long you looked, you would never find an atom of gadolinium within that sample that had this exact mass. Instead, you would find atoms with masses corresponding to the seven isotopes of gadolinium (151.9, 153.9, 154.9, 155.9, 156.9, 157.9 and 159.9). In this section, you will learn more about relative atomic mass, why it is important and how it is calculated.

64
Gd
157.3
Gadolinium

Figure 6A–1 The information for gadolinium (Gd) as seen on the periodic table

Relative atomic mass the average of the relative masses of all atoms for an element based on the isotopes for an element and their natural percentage abundances. Also known as relative atomic weight or standard atomic weight

Weighted average a calculation of the average that takes into account the frequency of a number within the data set

Relative isotopic mass the mass of an atom of an isotope relative to the mass of a carbon-12 atom, which is set at 12 exactly

EXPLAIN

Relative atomic mass (A_r) is the **weighted average** of the **relative isotopic masses** of an element relative to the mass of a carbon-12 atom, which has been determined experimentally to be 12.0. Recall from Chapter 1A that 'isotope' is the name given to atoms that have the same atomic number but different mass numbers. As both relative atomic mass and relative isotopic mass are calculated relative to carbon 12, they are unitless.

LINK **1A** INTRODUCTION TO THE ELEMENTS

The relative atomic mass becomes critical when using samples or undertaking reactions involving an element. This is because any sample of an element will be composed of a mixture of the different isotopes that exist for it. Using the mass of only one of those isotopes, or even just an average that didn't factor in the **percentage abundance** of each isotope, would lead to significant errors in calculations.

Let's look at the example of chlorine. There are two naturally occurring stable isotopes of chlorine: chlorine-35 and chlorine-37. If we were to take the average mass of these isotopes (that is, $\frac{(35+37)}{2}$), we would calculate the average mass to be 36. However, Figure 6A–2 shows us that this is not the relative atomic mass of chlorine. We can understand this better by analysing a representative sample of chlorine atoms, illustrated in Figure 6A–3.

17
Cl
35.5
Chlorine

Figure 6A–2 The information for chlorine (Cl) as displayed in the periodic table

Percentage abundance the relative proportions of the different isotopes for an element. Also called relative abundance

The key aspect to note in Figure 6A–3 is that we do not have equal amounts of chlorine-35 and chlorine-37. Indeed, if we did have the same amount of the two isotopes, then the average of 36 would be equal to the relative atomic mass. Instead, this is where the importance of the weighted average becomes apparent.

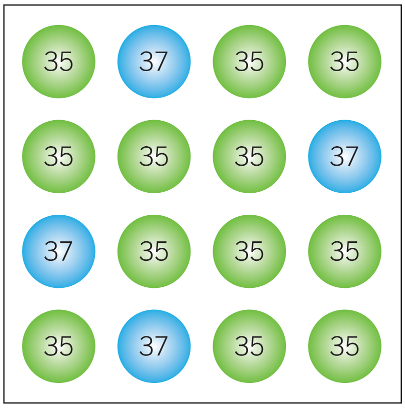


Figure 6A–3 A sample containing atoms of chlorine. In this sample, two isotopes exist corresponding to chlorine-35 and chlorine-37.

We can use the simple example in Figure 6A–3 to show how the weighted average is calculated.

1. Count how many atoms of chlorine-35 are present in the sample. (12)
2. Count how many atoms of chlorine-37 are present in the sample. (4)
3. Sum the answers to give the total number of atoms in the sample. (16)
4. Calculate the weighted average mass. To do this, take into account the number of atoms of each mass that we have and then divide by the total number of atoms.

$$\text{Weighted average mass} = \frac{(12 \times 35) + (4 \times 37)}{16} = 35.5$$

As mentioned previously, this represents a very simplified version of a sample of an element. In reality, we can't just look at a sample with the naked eye and determine the masses and abundances of the isotopes contained within it. Instead, this requires a complex piece of scientific instrumentation known as the mass spectrometer.

The mass spectrometer

What do drug testing in sport, the dating of fossils and the forensic assessment of crime scenes have in common? The answer is that all of them use mass spectrometers as part of their analysis. In essence, the role of a mass spectrometer is to separate components in complex samples based on their masses so that they can be identified.

The basic components of a mass spectrometer are shown in Figure 6A–4. It is important to note that you don't need to know the specifics of how this machine works. However, there are a couple of key features that will help in your understanding and analysis of the data produced by a mass spectrometer.

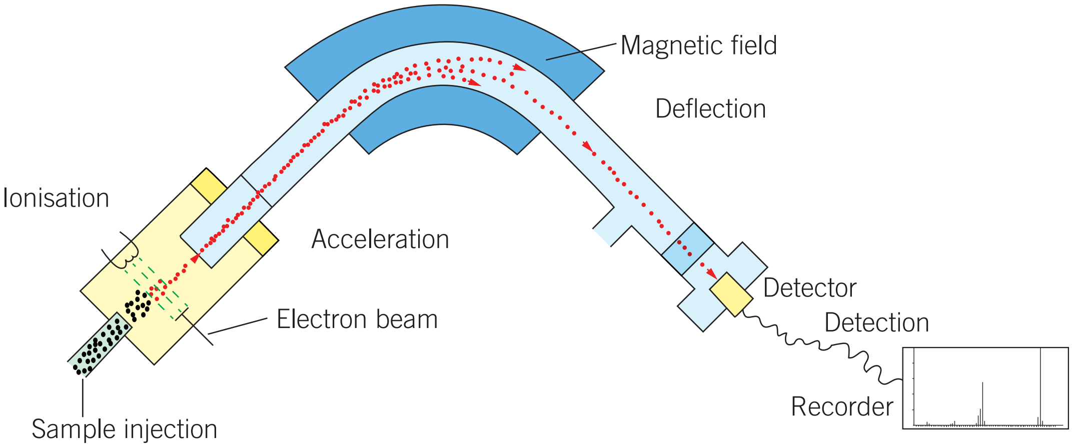


Figure 6A–4 A diagrammatic representation of a mass spectrometer, highlighting the different stages of the process, from ionisation of the sample through to detection at the end

For now, we will focus on the beginning and end of the process. As you can see in Figure 6A–4, the first step in the analysis is **ionisation** of the injected sample. In this instance, the sample is bombarded with fast-moving electrons. These electrons collide with electrons in the sample, causing them to be 'knocked off' or dissociated from their atoms. This results in the atoms in the sample becoming composed of **cations**. Depending on the number of electrons that are removed, these will 1+, 2+ or 3+ ions.

The last step in a mass spectrometer is detection. At this stage, the different components of the sample are identified and displayed on a mass spectrum, as shown in Figure 6A–5.

LINK **4A** FORMATION AND NAMING OF IONIC COMPOUNDS

Ionisation
when an atom or a compound becomes charged via the gain or loss of electrons

Cation
an atom that has lost electron(s) from its valence shell, resulting in an overall positive charge

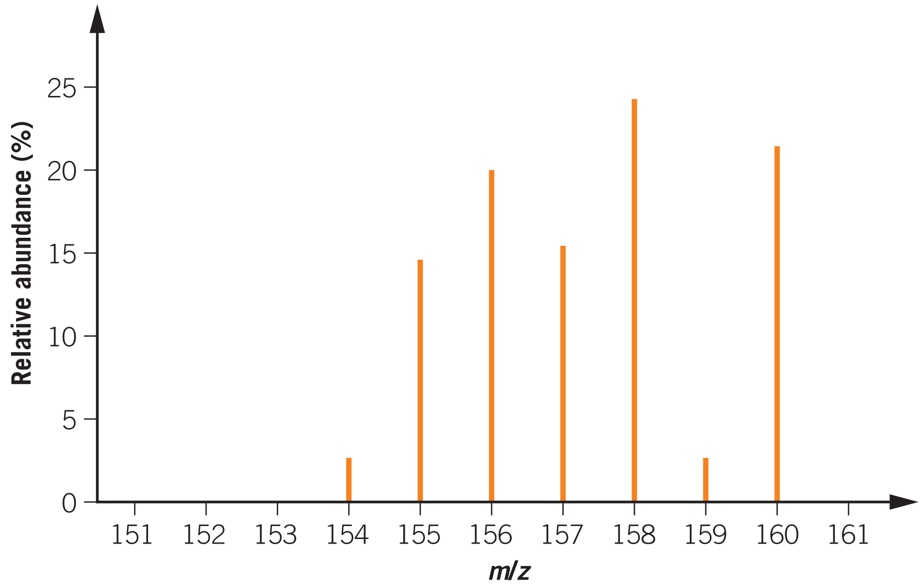


Figure 6A–5 The mass spectrum obtained from analysing a sample of gadolinium

In the example shown in Figure 6A–5, we can see the isotopes of the element gadolinium, which was discussed at the beginning of this section. There are two important aspects of this data to consider. Firstly, we can see that the *x*-axis is labelled as $\frac{m}{z}$, m/z, where m represents mass and z represents charge. This means that if a 2+ ion is detected, the value of $z = 2$ and so the m/z at which this ion is detected at is half of the actual mass (for example, a 2+ ion that appears at an m/z of 75 would have a mass of 150). For the elemental samples we will be looking at, the charge will usually be 1+. This means that the m/z will be equal to the relative isotopic mass. Secondly, the information derived from a mass spectrometer also allows us to see the percentage abundance (or relative amount) of each of the different isotopes. This can be observed from the data by examining the heights of the peaks and matching them to the corresponding value on the *y*-axis.

Check-in questions – Set 1

1 Draw the mass spectrum that you would expect to observe for the chlorine sample illustrated in Figure 6A–3.

Calculating relative atomic mass using experimental data

To summarise the content above, we can use the following pieces of information from our mass spectrometry data:

- Isotopic mass: through matching the peaks to their values on the *x*-axis
- Percentage abundance/relative abundance: through matching the height of the peaks to their values on the *y*-axis.

VIDEO 6A–1 USING MASS SPECTROMETRY DATA TO CALCULATE RAM

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Looking back at the example of chlorine described in Figure 6A–3, these were the two pieces of information that were needed to be able to calculate the relative atomic mass of the element. Therefore, mass spectrometry data can be used to complete these calculations.

As mentioned earlier, the relative atomic mass is a weighted average of the isotopic masses. Therefore, the equation to perform this calculation involves both the percentage abundance and mass of these isotopes. A general formula for this calculation is shown below:

$$A_r = \frac{(\%\ \text{Abundance}_1 \times I_{r1}) + (\%\ \text{Abundance}_2 \times I_{r2}) + \ldots + (\%\ \text{Abundance}_n \times I_{rn})}{\%\ \text{Abundance}_1 + \%\ \text{Abundance}_2 + \ldots + \%\ \text{Abundance}_n}$$

where:

A_r = The relative atomic mass
% Abundance = The relative abundance of the isotope
I_r = The relative isotopic mass of that isotope

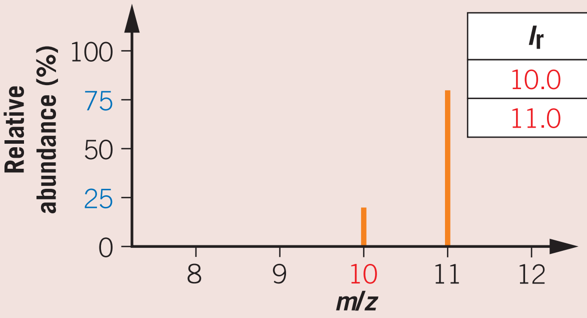
It is important to note that this formula can be applied for any number of isotopes, where the number of isotopes is represented by '*n*' in the equation above.

Using mass spectrometry data, it is possible to do this calculation either directly from the graph, or you might find that the information is converted into a table format first. We will look at both of these examples below.

Worked example 6A–1: Calculating relative atomic mass

VIDEO
WORKED
EXAMPLE 6A–1

The data below is the mass spectrum obtained for a sample of boron. Using this information, calculate the relative atomic mass of boron.



I_r	% Abundance
10.0	19.9
11.0	80.1

Solution

Step 1	Arrange your general equation based on having two isotopes.	$A_r = \frac{(\%\ \text{Abundance}_1 \times I_{r1}) + (\%\ \text{Abundance}_2 \times I_{r2})}{\%\ \text{Abundance}_1 + \%\ \text{Abundance}_2}$
Step 2	Substitute in values obtained from the data presented.	$A_r = \frac{(19.9 \times 10.0) + (80.1 \times 11.0)}{19.9 + 80.1}$ $= \frac{1080.1}{100}$
Step 3	Present your answer, remembering to check for units and significant figures.	$A_r = 10.8$

The structure of these types of questions can be varied. Sometimes the relative atomic mass is known, but the relative abundances of all the isotopes isn't. In these instances, we need to use some algebra to help solve our equation.

Worked example 6A–2: Calculating percentage abundance

Copper is composed of two stable isotopes, copper-63 and copper-65. The relative atomic mass of copper is 63.5. What is the percentage abundance of each of the two copper isotopes?

VIDEO
WORKED
EXAMPLE 6A–2

Solution

Step 1	Define the percentage abundance of the lighter isotope as *x*. As the sum of the percentages of the two isotopes must equal 100, define the percentage abundance of the heavier isotope as (100 – *x*).	% abundance of copper-63 = *x*, % abundance of copper-65 = (100 – *x*)
Step 2	Write out your equation for relative atomic mass using your new terms for the percentage abundance of each isotope.	$A_r = \frac{(x \times I_{r1}) + ((100 - x) \times I_{r2})}{100}$
Step 3	Fill in the equation with the information that is provided in the question.	$63.5 = \frac{(x \times 63) + ((100 - x) \times 65)}{100}$
Step 4	Multiply both sides of the equation by 100.	$6350 = (x \times 63) + ((100 - x) \times 65)$
Step 5	Expand the brackets on the right-hand side of the equation.	$6350 = 63x + 6500 - 65x$
Step 6	Group like terms on each side of the equation.	$6350 - 6500 = 63x - 65x$
Step 7	Solve the equation.	$-150 = -2x$ $x = 75$
Step 8	State the percentage abundance of the two isotopes.	Copper-63 = 75% Copper-65 = (100 – 75)% = 25%

WORKSHEET
6A–1
CALCULATING
RELATIVE ATOMIC
MASS

6A SKILLS

Using estimation to check answers

As you progress through the chemistry course, one skill that you should develop is an ability to be able to recognise answers that look right and those that don't. One of the best ways that you can do this is to be able to estimate the value of an answer before you begin. A good example of a situation where this can be helpful is when doing relative atomic mass calculations.

Question: Boron contains two stable isotopes, ^{10}B and ^{11}B. The abundance of the two isotopes is 20% and 80% respectively. Calculate the relative atomic mass of boron.

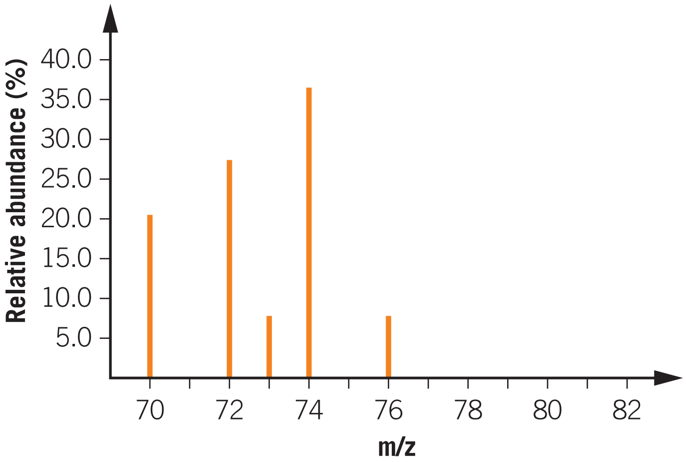
Answer: Before we start using the formula that we learned on page 186 to calculate the relative atomic mass, let's think about how we can get an estimate for what to expect in our answer. The points that we should consider are:

- The isotopic masses of our two isotopes are 10 and 11, so our relative atomic mass will fall between these two values.
- The heavier isotope is more abundant, which means that our relative atomic mass will be closer to 11 than 10 (that is, greater than 10.5).

If we were to go through the calculation steps for this, as shown in Worked example 6A–1, we would see that our relative atomic mass for boron was equal to 10.8. Imagine, though, that you had accidentally made a mistake in your calculation and the answer you generated was 10.2 or 11.8. By having done an initial estimate for your expected answer, you would hopefully realise in both these instances that you have made an error in your calculation and double-check your working out.

Section 6A questions

1 The following question relates to the mass spectrum for germanium shown below. For this mass spectrum, assume that all ions have a charge of 1+.



a How many isotopes does germanium have?

b What is the mass of the most abundant isotope?

c What is the abundance of the heaviest isotope?

d What is the relative isotopic mass of each of the two isotopes that are present in approximately equal abundance?

e Using the mass spectrum data, calculate the relative atomic mass for germanium. Show full working in your answer.

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2 An element X has three isotopes with the relative isotopic masses 23.99, 24.99 and 25.98. The relative abundance of these isotopes is 78.70%, 10.13% and 11.17%, respectively.
  - a Calculate the relative atomic mass of this element.
  - b Using your periodic table, identify element X.

3 Define 'relative isotopic mass.'

4 Define 'relative atomic mass.'

5 What unit of measurement is used for relative isotopic mass and relative atomic mass?

6 Bromine has two naturally occurring isotopes, bromine-79 and bromine-81. The bromine-79 isotope has a percentage abundance of 50.7%. Determine the percentage abundance of the bromine-81 isotope.

7 Silicon has three stable isotopes, ^{28}Si, ^{29}Si and ^{30}Si, and a relative atomic mass of 28.085. Use the information in the table below to determine the relative isotopic mass of silicon-29.

Isotope	Relative isotopic mass	Relative abundance (%)
^{28}Si	27.98	92.22
^{29}Si		4.69
^{30}Si	29.97	3.09

8 Neon has three stable isotopes, ^{20}Ne, ^{21}Ne and ^{22}Ne, and a relative atomic mass of 20.18. Use the information in the table below to determine the relative isotopic mass and relative abundance of neon-20.

Isotope	Relative isotopic mass	Relative abundance (%)
^{20}Ne		
^{21}Ne	20.99	0.27
^{22}Ne	21.99	9.25

9 Nitrogen has a relative atomic mass of 14.006. Its two naturally occurring isotopes, nitrogen-14 and nitrogen-15, have relative isotopic masses of 14.00 and 15.00, respectively.
  - a Which of the following estimates is most likely to be closest to the percentage abundance of the nitrogen-14 isotope – 10%, 50%, 90%? Explain why.
  - b Use the information provided to determine the percentage abundance of each of the two isotopes.

10 Magnesium has three stable isotopes, ^{24}Mg, ^{25}Mg and ^{26}Mg, and a relative atomic mass of 24.305. Use the information in the table below to calculate the percentage abundance of the ^{24}Mg and ^{26}Mg isotopes.

Isotope	Relative isotopic mass	Relative abundance(%)
^{24}Mg	23.985	
^{25}Mg	24.985	0.1000
^{26}Mg	25.983	

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6B The mole

Study Design:

- 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

Glossary:
Avogadro's constant
Mole

VIDEO 6B–1 WHAT IS AVOGADRO'S CONSTANT?

ENGAGE

Avogadro's constant

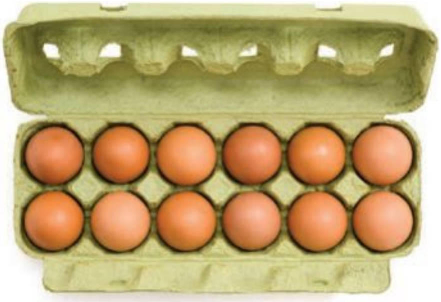
Figure 6B–1 Amedeo Avogadro (1776–1856) was an Italian scientist whose work with gases contributed to the calculation of Avogadro's constant, named for him after his death.

Amedeo Avogadro was born in 1776 and became a famous Italian scientist. His major contribution to chemistry came through his work on gases. His gas law states that if equal volumes of different gases are under the same conditions of temperature and pressure, they will contain the same number of molecules. Importantly, he made a distinction between atoms and molecules in relation to gases, which hadn't been done before. He was able to distinguish the fact that gases are composed of molecules and these molecules are made up of atoms.

However, many people that study chemistry know Avogadro's name for work that he didn't conduct. Avogadro's constant is a commonly used figure in chemistry; however, it was actually first calculated by an Austrian scientist called Johann Josef Loschmidt. Despite this, the number became known as Avogadro's constant, as a way of recognising his large contributions to the field of chemistry. In this section, you will learn a lot more about Avogadro's constant and how it is used to quantify atoms and compounds.

EXPLAIN

The mole



Thinking about the number of atoms contained within a sample of a substance is almost like thinking about how many stars are in the sky or how many grains of sand are on a beach. The answer is too many to really comprehend! However, we have ways of being able to talk about large numbers that make things easier. While not on the same scale, think about shopping for eggs in the supermarket. Would you be buying 24 eggs or two dozen eggs? Although they are describing the same quantity, it is more convenient to say

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two dozen. While it may not seem too noticeable when you only need 24 eggs, imagine if you wanted to buy 240 eggs ... doesn't it seem easier to just say 20 dozen? Thankfully, we have a similar system for being able to discuss atoms and molecules. It is known as the **mole**.

Similarly to how one dozen describes 12 eggs, one mole describes 602 214 076 000 000 000 000 000 particles (atoms, molecules or ions). This is more conveniently written as 6.02×10^{23} and is known as **Avogadro's constant (N_A)**. The relationship between the moles and Avogadro's constant is shown by the formula:

$$n = \frac{N}{N_\text{A}}$$

where:

n = The amount of moles (we express the units as mol)
N = The number of particles (atoms, molecules or ions)
N_A = 6.02×10^{23} mol^{-1} (Avogadro's constant)

This formula can be used to convert between the number of particles in a substance and the moles. This is explored in the worked examples to follow.

Mole
one mole is exactly $6.02214076 \times 10^{23}$ particles (6.02×10^{23} to three significant figures)

Avogadro's constant (N_A)
the number of particles in a mole: 6.02×10^{23} mol^{-1}

Worked example 6B–1: Calculating number of molecules

VIDEO
WORKED EXAMPLE 6B–1

If a chemist has 1.45 mol of butane gas molecules, C_4H_{10}, how many gas molecules does this chemist have?

Solution

Step 1	Identify from the question the unknown quantity you need to determine.	$N(C_4H_{10})$
Step 2	Identify the known quantities from information in the question.	$n(C_4H_{10}) = 1.45$ mol $N_A = 6.02 \times 10^{23}$ mol^{-1}
Step 3	Write the general equation needed (transposed if required) to solve for the unknown identified in Step 1.	$N = n \times N_A$
Step 4	Input values from Step 2.	$N = 1.45 \times 6.02 \times 10^{23}$
Step 5	Use a calculator to complete the calculation.	$N = 8.73 \times 10^{23}$ C_4H_{10} gas molecules

Worked example 6B–2: Calculating number of atoms

VIDEO
WORKED EXAMPLE 6B–2

Determine the number of hydrogen atoms in 1.50 mol of water molecules.

Solution

Step 1	Identify from the question the unknown quantity you need to determine.	N(H atoms in H_2O)
Step 2	Identify the known quantities from information in the question.	$n(H_2O) = 1.50$ mol $N_A = 6.02 \times 10^{23}$ mol^{-1}
Step 3	Write the general equation needed (transposed if required) to solve for the unknown identified in Step 1.	$N = n \times N_A$

Step 4	Input values from Step 2.	$N(H_2O) = 1.50 \times 6.02 \times 10^{23}$
Step 5	Use a calculator to complete the calculation.	$N = 9.03 \times 10^{23}$ H_2O molecules Note: At this step, you have calculated the number of H_2O molecules, not the number of H atoms that are in that amount of H_2O molecules.
Step 6	As there are two H atoms in every one H_2O molecule, the number of H atoms will be double the number of H_2O molecules.	$N(\text{H atoms}) = 2 \times N(H_2O)$ $= 2 \times 9.03 \times 10^{23}$ $= 1.81 \times 10^{24}$ H atoms

Molar mass

In 2018, the mole (6.02×10^{23} particles) became the SI (standard international system of units) unit of measurement for an amount of substance. Prior to this, one mole was defined as the number of carbon atoms in 0.012 kg of carbon-12. Carbon-12 is used as the reference for the relative isotopic masses so this means that a neat relationship exists between the relative atomic mass of a substance and the mass of one mole of that substance. Carbon-12 has an experimentally determined relative atomic mass of 12.0, and the mass of one mole of carbon-12 is 12.0 grams. Magnesium has a relative atomic mass of 24.3. This tells us that the mass of one mole of magnesium atoms is about twice the mass of one mole of carbon-12 atoms. This means that one mole of magnesium would have a mass of 24.3 grams. The mass of one mole of an element or compound is known as the **molar mass (M)** and has units of g mol^{-1}.

Molar mass (M) the mass of one mole of a substance measured in g mol^{-1}; symbol M

The relationship between mole, mass and molar mass is shown by the formula:

$$n = \frac{m}{M}$$

where:

n = The amount of moles (we express the units as mol)
m = The mass (in grams)
M = The molar mass (g mol^{-1})

Worked example 6B–3: Calculating molar mass

VIDEO
WORKED
EXAMPLE 6B–3

Calculate the molar mass of copper(II) carbonate, $CuCO_3$.

Solution

Step 1	Identify from the question the unknown quantity you need to determine.	$M(CuCO_3)$
Step 2	Determine the molar mass of each of the elements present in the molecular formula. Where there is more than one of an atom present, multiply its relative atomic mass by this number.	$M(\text{Cu}) = (1 \times 63.5\ \text{g mol}^{-1}) = 63.5\ \text{g mol}^{-1}$ $M(\text{C}) = (1 \times 12.0\ \text{g mol}^{-1}) = 12.0\ \text{g mol}^{-1}$ $M(\text{O}) = (3 \times 16.0\ \text{g mol}^{-1}) = 48.0\ \text{g mol}^{-1}$
Step 3	Add the molar masses for each element determined in Step 2 together.	$M(CuCO_3) = 63.5 + 12.0 + 48.0$ $= 123.5\ \text{g mol}^{-1}$

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Worked example 6B–4: Calculating mass from mol

Calculate the mass, in grams, of hydrogen peroxide, H_2O_2, in a 0.700 mol sample.

Solution

Step 1	Identify from the question the unknown quantity you need to determine.	$m(H_2O_2)$
Step 2	Identify the known quantities from information in the question.	$n(H_2O_2) = 0.700\text{ mol}$ $M(H_2O_2) = (2 \times 1.0\text{ g mol}^{-1}) + (2 \times 16.0\text{ g mol}^{-1})$ $= 34.0\text{ g mol}^{-1}$ Note: To calculate molar mass for a substance, you use the relative atomic mass of elements from the periodic table. This is then multiplied by the number of atoms of that element present in the molecular formula of that substance.
Step 3	Write the general equation needed (transposed if required) to solve for the unknown identified in Step 1.	$m = n \times M$
Step 4	Input values from Step 2.	$m(H_2O_2) = 0.700 \times 34.0$
Step 5	Use a calculator to complete the calculation.	$m(H_2O_2) = 23.8\text{ g}$

VIDEO
WORKED EXAMPLE 6B–4

WORKSHEET 6B–1
MOLE CALCULATIONS

6B SKILLS

Transposing equations

You will have seen in this section that you have a number of different equations that you will need to use throughout your chemistry studies. While these are presented to you in a standard format, you could be asked to solve for any of the terms in that formula. To be able to do this effectively, you have to be able to transpose or rearrange the equation correctly.

Let's look at the formula that shows the relationship between moles, mass and molar mass:

$$n = \frac{m}{M}$$

You could be asked three types of questions using this formula, which are:

1. Calculate the amount in moles (n).
   In this instance, you will be using the formula as it is provided above. For this reason, the equation does not need to be transposed.
2. Calculate the mass (m).
   In this instance, you will need to rearrange the equation above to solve for mass (m). To do this, you have to multiply both sides of the equation by the molar mass (M). This leaves you with the following equation:

$$m = n \times M$$

VIDEO 6B–2
SKILLS: TRANSPOSING EQUATIONS

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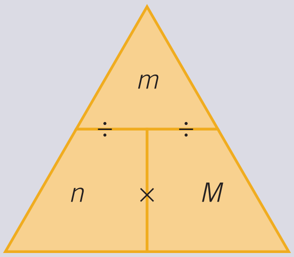
3 Calculate the molar mass (M).
In this instance, you will need to rearrange the equation to solve for molar mass (M). To do this, you have to multiply both sides of the equation by the molar mass (M) and then divide both sides of the equation by the moles. This leaves you with the following equation:

$$M = \frac{m}{n}$$

Another good way of remembering how to do this is by using the following triangle:



Section 6B questions

1 State the number and unit of measurement for Avogadro's constant.
2 Define 'mole' in chemistry.
3 Calculate the number of sodium ions in 1.777×10^{23} formula units of sodium sulfate, Na_2SO_4.
4 Determine the molar mass of the following, using a periodic table.
- **a** gold, Au
- **b** barium hydroxide, $Ba(OH)_2$
- **c** ammonium phosphate, $(NH_4)_3PO_4$
- **d** water, H_2O

5 Calculate the amount of the following molecules in moles.
- **a** 3.011×10^{24} molecules of carbon monoxide
- **b** 4.015×10^{23} molecules of nitrogen gas
- **c** Carbon dioxide if there are 7.528×10^{22} oxygen atoms

6 Calculate the amount, in mol, of the following compounds.
- **a** 4.08 g of sodium chloride, NaCl
- **b** 0.526 g of ammonia, NH_3
- **c** 58.2 g of cyclohexane, C_6H_{12}

7 Calculate the mass of the following quantities.
- **a** 2.09 moles of calcium carbonate, $CaCO_3$
- **b** 1.23×10^{-3} moles of sodium ethanoate, $NaC_2H_4O_2$
- **c** 6.44×10^{24} atoms of silver metal, Ag
- **d** 1.15×10^{23} molecules of ozone, O_3

8 Determine the number of sulfur atoms in the following quantities.
- **a** 1.00 mol of sulfur dioxide, SO_2
- **b** 0.060 mol of potassium sulfite, K_2SO_3
- **c** 0.899 g of hydrogen sulfide, H_2S
- **d** 5.55 g of sulfur, S_8

9 An 8.05 g sample of gaseous gallane, GaH_3, is known to contain 0.110 mol. Showing full working, confirm that the molar mass of gallane is 73.2 g mol^{-1}.
10 A sample of a pure unknown compound contains 2.956×10^{22} molecules in 4.88 g. Use this information to determine the molar mass of the compound.

6C Percentage composition and empirical formula

Study Design:
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

Glossary:
Empirical formula
Molecular formula

ENGAGE

Iron ore

VIDEO 6C–1 PERCENTAGE COMPOSITION

Iron is a major commodity that is required in large amounts throughout the world. For this reason, many countries with iron-rich deposits, including Australia, have iron mining industries (Figure 6C–1). Iron is not found as a pure metal within the ground, but rather as various iron ores. Iron ores are rocks and minerals that contain iron compounds. These iron ores vary in the percentage of iron that they contain. Siderite ($FeCO_3$) is a mineral that is composed of 48% iron. Magnetite (Fe_3O_4), on the other hand, contains 72% iron. The iron content in these minerals is important, as the higher the percentage, the more economically attractive the deposit is. In this section, we will look at the percentage composition of compounds, as well as understanding what an empirical and molecular formula is and how to determine them from experimental data.

Figure 6C–1 The mining and processing of iron ore is a major economic commodity in many countries around the world.

EXPLAIN

Percentage composition

The percentage composition of a compound describes the relative contribution of each element to the overall mass of the substance. Later in this section, we will look at how the percentage composition can be used to calculate the empirical formula of a compound. However, here we will focus on learning how to answer questions that ask you to calculate the percentage composition.

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VIDEO
WORKED
EXAMPLE 6C–1

Worked example 6C–1: Calculating percentage composition

What is the percentage by mass of carbon in ethanol (C_2H_5OH)?

Solution

Step 1	Calculate the molar mass of the substance.	$M(C_2H_5OH) = (2 \times 12.0\ \text{g mol}^{-1}) + (6 \times 1.0\ \text{g mol}^{-1}) + (1 \times 16.0\ \text{g mol}^{-1}) = 46.0\ \text{g mol}^{-1}$
Step 2	Divide the molar mass of the desired element (C) by the molar mass of the compound. The molar mass of carbon is multiplied by 2 because every mole of ethanol has 2 moles of carbon. Multiply by 100 to find the percentage.	$\%C = \frac{2 \times M(C)}{M(C_2H_5OH)} \times 100$ $= \frac{2 \times 12.0\ \text{g mol}^{-1}}{46.0\ \text{g mol}^{-1}} \times 100$
Step 3	Present your answer, remembering to check for units and significant figures	$\%C = 52.2\%$

VIDEO
WORKED
EXAMPLE 6C–2

Worked example 6C–2: Calculating mass of an element within a compound

What is the mass of hydrogen in a 12.0 g sample of ammonia, NH_3?

Solution

Step 1	Calculate the molar mass of the substance.	$M(NH_3) = (1 \times 14.0\ \text{g mol}^{-1}) + (3 \times 1.0\ \text{g mol}^{-1}) = 17.0\ \text{g mol}^{-1}$
Step 2	Calculate the percentage by mass of hydrogen in ammonia (similar to Step 2 in Worked example 6C–1)	$\%H = \frac{3 \times M(H)}{M(NH_3)} \times 100$ $= \frac{3 \times 1.0\ \text{g mol}^{-1}}{17.0\ \text{g mol}^{-1}} \times 100$ $= 17.6\%$
Step 3	The mass of hydrogen is 17.6% of the mass of ammonia, so calculate 17.6% of 12.0 g.	$m(H) = \frac{17.6}{100} \times 12.0\ \text{g}$ $= 2.11\ \text{g}$

Empirical formula

Empirical formula
the simplest whole-number ratio of elements within a compound

The **empirical formula** is defined as the simplest whole-number ratio of elements within a compound. For ionic compounds, this represents the ratio of elements present within the lattice. For covalent compounds, the molecular formula is the number of each atom type present in a molecule. For example, the molecular formula of carbon dioxide, CO_2, shows that each molecule has one carbon atom and two oxygen atoms. For carbon dioxide, the empirical formula is the same as the molecular formula, CO_2. However, for many covalent compounds, the empirical formula is not the same as the molecular formula.

By way of explanation, let's look at ethane. Ethane has the molecular formula C_2H_6. We can see that the ratio of carbon to hydrogen in ethane is 2 : 6, which can be simplified to 1 : 3. Therefore, the empirical formula of ethane is CH_3.

The empirical formula of compounds can be determined through a variety of different experimental techniques. We will now look at a number of examples of scenarios where you may be asked to determine the empirical formula. In each of these examples, you will be provided with slightly different starting information with which to solve the empirical formula.

Worked example 6C–3: Calculating empirical formulas from percentage composition

VIDEO WORKED EXAMPLE 6C–3

An unknown compound is composed of 40.0% carbon, 53.3% oxygen and 6.70% hydrogen. Calculate the empirical formula of this compound.

Solution

Step 1	Choose a total mass of the compound to allow you to write down the mass of each element. The calculations are easiest with 100 g of the compound (C is 40.0% of the 100 g = 40.0 g)	C = 40.0 g, O = 53.3 g and H = 6.70 g
Step 2	Calculate the moles for each element. Recall from Chapter 6B, for this step you will need to use $n = \frac{m}{M}$.	$n(\text{C}) = \frac{40.0 \text{ g}}{12.0 \text{ g mol}^{-1}} = 3.33 \text{ mol}$ $n(\text{H}) = \frac{6.70 \text{ g}}{1.00 \text{ g mol}^{-1}} = 6.70 \text{ mol}$ $n(\text{O}) = \frac{53.3 \text{ g}}{16.0 \text{ g mol}^{-1}} = 3.33 \text{ mol}$
Step 3	Divide the moles for each element by the lowest calculated mole value in Step 2.	$\text{C} = \frac{3.33}{3.33} = 1.00$ $\text{H} = \frac{6.70}{3.30} = 2.01$ $\text{O} = \frac{3.33}{3.33} = 1.00$
Step 4	If these are whole numbers, write your empirical formula. If not, multiply each element by the same value to obtain a whole number ratio.	Empirical formula = CH_2O

LINK **6B** THE MOLE

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Worked example 6C–4: Calculating empirical formulas from mass of elements

VIDEO
WORKED EXAMPLE 6C–4

You have a 9.90 g sample of an unknown substance that you know is composed of only carbon and oxygen. You find that the sample contains 2.70 g of carbon. Calculate the empirical formula of the compound.

Solution

6B THE MOLE LINK

Step 1	Determine the mass of each element in the compound.	$m(\text{C}) = 2.70$ g (as given in the question) $m(\text{O}) = 9.90\text{ g} - 2.70\text{ g} = 7.20\text{ g}$
Step 2	Calculate the amount, in moles, of each element. Recall from Chapter 6B, for this step you will need to use $n = \frac{m}{M}$.	$n(\text{C}) = \frac{2.70\text{ g}}{12.0\text{ g mol}^{-1}} = 0.225\text{ mol}$ $n(\text{O}) = \frac{7.20\text{ g}}{16.0\text{ g mol}^{-1}} = 0.450\text{ mol}$
Step 3	Divide the moles for each element by the lowest calculated mole value in Step 2.	$n(\text{C}) = \frac{0.225}{0.225} = 1$ $n(\text{O}) = \frac{0.450}{0.225} = 2$
Step 4	Write out the empirical formula.	Empirical formula = CO_2

Molecular formula

While the empirical formula of a compound is the simplest whole number ratio of elements within that compound, the **molecular formula** is defined as the chemical formula that gives the total number of atoms of each element in each molecule of that compound. If you have the empirical formula of a compound and the molar mass, it is possible to work out the molecular formula.

Molecular formula
the chemical formula that gives the total number of atoms of each element in each molecule of a compound

Worked example 6C–5: Calculating molecular formulas

VIDEO
WORKED EXAMPLE 6C–5

A compound has a molar mass of 60.0 g mol^{-1} and an empirical formula of CH_2O. What is the molecular formula for this compound?

Solution

Step 1	Determine the molar mass of the empirical formula.	$\text{M}(CH_2O) = (1 \times 12.0\text{ g mol}^{-1}) + (2 \times 1.0\text{ g mol}^{-1}) + (1 \times 16.0\text{ g mol}^{-1})$ $= 30.0\text{ g mol}^{-1}$
Step 2	Divide the molar mass of the compound by the molar mass of the empirical formula.	$\frac{60.0\text{ g mol}^{-1}}{30.0\text{ g mol}^{-1}} = 2$
Step 3	Multiply all elements in the empirical formula by the answer from Step 2 to give the molecular formula.	$CH_2O \times 2 = C_2H_4O_2$

WORKSHEET 6C–1
EMPIRICAL AND MOLECULAR FORMULAS

6C SKILLS

Significant figures

At the end of any calculation question in chemistry, you should check two things: units and significant figures. It is frustrating to have done your working out correctly and gotten the right value, only to lose marks because you haven't included (or used the wrong) units or written your answer to the correct number of significant figures. Depending on the assessment, you may lose marks every time you do this ... so these mistakes can really add up!

In terms of significant figures, there are a few rules that you should know:

1. Any digit that isn't a zero is significant (for example, 19.782 has five significant figures).
2. Any zero that is found between two significant numbers is significant (for example, 19.702 has five significant figures).
3. After the decimal point, a zero that doesn't come between two significant numbers is only significant if it comes at the end (for example, 0.78200 has five significant figures, whereas 0.00782 has three significant figures).

If you are multiplying or dividing, you consider all significant figures provided in the question. For your answer, you will use the same number of significant figures as the value in the question that contains the LEAST number of significant figures (not just after the decimal point).

Let's do an example.

Question: Multiply the following values: 54.7 × 0.02

Answer: 1.094. However, our answer should use the smallest number of significant figures in the question. Therefore, our actual answer is 1 (because 0.02 is only one significant figure).

Section 6C questions

1. Define 'empirical formula'.
2. Determine the percentage composition by mass of potassium, K, in the following compounds.
   - a potassium cyanide, KCN
   - b potassium carbonate, K_2CO_3
   - c potassium phosphate, K_3PO_4
   - d potassium permanganate, $KMnO_4$
3. Determine the percentage composition of each element in the compound caesium hexachlororhenate(IV), Cs_2ReCl_6.
4. The mineral zircon has a general chemical formula of $ZrSiO_4$ and is found in mineral sand deposits in Victoria. Use the percentage composition of oxygen in the compound to calculate the mass of oxygen in 2.36 kg of the mineral.
5. Azurite is a mineral of copper with the general formula, $Cu_3(CO_3)_2(OH)_2$. A sample of azurite contains 245 g of copper. Calculate the total mass of this sample of azurite.
6. A compound contains 78.77% tin and 21.23% oxygen by mass. Determine the empirical formula of the compound.

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7 A compound contains only the elements sodium, chromium and oxygen. The compound is known to contain 17.56% sodium and 39.69% chromium by mass. Determine the empirical formula of this compound.

8 Petalite is a mineral containing lithium, aluminium, silicon and oxygen with the percentage compositions given in the following table.

Element	Percentage composition (%)
Li	2.25
Al	8.81
Si	?
O	52.24

Determine the empirical formula of this compound.

9 A hydrocarbon is found to be 82.8% by mass carbon. Its molecular mass is 58.0 g mol^{-1}.

a Determine the empirical formula of this hydrocarbon.

b What is the molecular formula of this hydrocarbon?

10 A hydrocarbon has the empirical formula CH_2. If a 0.213 mol sample of this hydrocarbon has a mass of 17.91 grams, find the molecular formula of the hydrocarbon.

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Chapter 6 review

Summary

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

Success criteria – I am now able to:		Linked question
6A.1	Define 'relative isotopic masses of elements'	11☐
6A.2	Recall that elements are represented on the relative mass scale using the carbon-12 isotope as the standard	9☐
6A.3	Calculate the relative atomic mass of an element from given mass spectrometry data	5☐
6B.1	Recall Avogadro's constant	3☐
6B.2	Convert between moles and the number of atoms or molecules using Avogadro's constant	6☐, 12☐
6B.3	Determine the moles of atoms or molecules in a sample of known mass	13☐
6B.4	Determine the mass of an atom or molecule given the moles	13☐
6B.5	Determine the molar mass of an unknown compound given the mass and moles	2☐, 12☐
6C.1	Determine the molar mass of a compound from its formula	1☐
6C.2	Determine the percentage composition by mass of covalent compounds	8☐, 10☐
6C.3	Define 'empirical formula'	4☐
6C.4	Determine the empirical formula of a covalent compound from its percentage composition by mass	13☐
6C.5	Determine the molecular formula from an empirical formula given its molar mass	7☐

Multiple-choice questions

1 What is the molar mass of aluminium trinitrate ($Al(NO_3)_3$)?

A 89.0 g mol^{-1}
B 185.0 g mol^{-1}
C 213.0 g mol^{-1}
D 267.0 g mol^{-1}

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2 A 23.0 g sample of an unknown substance was found to contain 0.500 moles. What could the substance be?

A Na
B C_2H_5OH
C H_2O
D CO_2

3 The value of Avogadro's constant is

A 3.01×10^{22}
B 6.02×10^{22}
C 3.01×10^{23}
D 6.02×10^{23}

4 An empirical formula is

A the simplest whole number ratio of elements within a compound.
B different to the formula of an ionic compound.
C able to be simplified by dividing with a common factor.
D always the same as the molecular formula.

5 An element has two stable isotopes with the masses 68 and 72. The lighter isotope is three times more abundant than the heavier one. The relative atomic mass of this element will be approximately

A 68
B 69
C 71
D 72

6 How many oxygen atoms are present in 0.333 mole of $NaNO_3$?

A 3
B 2.00×10^{23}
C 6.02×10^{23}
D 1.80×10^{24}

7 A compound has an empirical formula of CH_2O and a molar mass of 180.0 g mol^{-1}. What is the molecular formula of this compound?

A $C_3H_6O_3$
B $C_4H_8O_4$
C $C_5H_{10}O_5$
D $C_6H_{12}O_6$

8 Which of the following contains 80.0% carbon by mass?

A CH_4
B C_2H_4
C C_2H_6
D C_2H_5OH

9 Which of the following is the standard used for the relative mass scale?

A hydrogen-1
B carbon-12
C nitrogen-14
D oxygen-16

10 What is the mass of nitrogen in a 3.60 g sample of hydrogen cyanide (HCN)?

A 1.20 g
B 1.87 g
C 3.60 g
D 14.0 g

Short-answer questions

11 You are working in a research laboratory that is particularly interested in the isotopes of magnesium. It is known that three stable isotopes exist: ^{24}Mg, ^{25}Mg and ^{26}Mg.

a Define 'isotope' and 'relative isotopic mass'. (2 marks)

b Determine the number of protons, neutrons and electrons found in the three isotopes of magnesium. (3 marks)

c A scientist at the laboratory provides you with a data table from some previous experiments that they have performed. Use this to calculate the relative atomic mass of magnesium. (2 marks)

Relative isotopic mass	Relative abundance (%)
^{24}Mg (mass = 23.985)	78.70
^{25}Mg (mass = 24.986)	10.13
^{26}Mg (mass = 25.983)	11.17

d The laboratory is involved in the production of magnesium chloride, which is sold as a supplement to treat magnesium deficiency. What is the percentage by mass of magnesium in magnesium chloride? (2 marks)

e You are provided with a sample that has been chemically analysed. You are told that it contains magnesium, sulfur and oxygen and that the percentage composition for sulfur and oxygen is 26.7% and 53.1%, respectively. What is the empirical formula of this compound? (3 marks)

f You are given 0.500 mol of the compound. How many oxygen atoms are present in the sample? (2 marks)

12 The school laboratory assistant has asked you to help out with a problem that she is having. She has only just started in the job and has found two beakers, but neither is labelled. By the notes left behind, she knows that one of the beakers contains sodium chloride and the other contains calcium carbonate. The notes also say that the beaker on the left contains twice as many formula units as the beaker on the right. There are exactly 3.01×10^{22} formula units in the beaker on the left. Having weighed the samples, she knows that both beakers contain 2.93 g of each sample. What sample is in the beaker on the left? (3 marks)

13 You have a 3.500 g sample of a substance that contains aluminium, phosphorus and oxygen. It is found that this sample contains 0.7735 g of aluminium and 1.8375 g of oxygen.

a What is the mass of phosphorus contained within this sample? (1 mark)

b What is the empirical formula of this compound? (3 marks)

c You are provided with 0.200 mol of a separate sample of this compound. What will the mass of this sample be? (1 mark)

UNIT 1
HOW CAN THE DIVERSITY OF MATERIALS BE EXPLAINED?

CHAPTER 7
ORGANIC COMPOUNDS

Introduction

Organic chemistry is the study of carbon and the diverse ways in which it can form molecular substances. Carbon is uniquely positioned on the periodic table. Sitting in group 14, it generally neither donates electrons to form cations nor accepts electrons to form anions. Instead, it generally shares its valence electrons to form covalent bonds. Its four valence electrons can form up to four bonds with surrounding atoms. Another key feature of carbon is that it can form double and triple bonds with itself or other elements. These features of carbon give rise to the diverse range of molecules it can form, as seen in Chapter 2. Carbon is also used by nature as the building block of life and is the second most abundant element by mass in the human body. The most abundant element by mass is oxygen, which exists mostly as water in our body.

INTRODUCTION VIDEO ORGANIC COMPOUNDS

In this chapter, we will explore the origins of carbon compounds used in our society. We will also investigate how carbon-based compounds are classified, represented, named and used. Finally, we will rationalise the physical and chemical properties of different organic compounds through their intramolecular and intermolecular bonding.

Curriculum

Area of Study 2 Outcome 2
How are materials quantified and classified?

Study Design:	Learning intentions – at the end of the chapter I will be able to:
• 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 and artificial hearts) and an assessment of the benefits of those products for society, as well as the health and/or environmental hazards they pose	**7A Organic chemistry in society** **7A.1** Recall that fossil fuels provide a non-renewable source of chemicals that can be used for energy generation and production of goods **7A.2** Explain how plant-based biomass can serve as an alternative to fossil fuels in the production of materials and goods **7A.3** Explain why plant-based biomass is renewable, whereas fossil fuels are non-renewable **7A.4** Analyse the impact of materials and products made from organic chemistry on society and the environment
• 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)	**7B Hydrocarbons** **7B.1** Define the characteristics of alkanes and alkenes **7B.2** Determine the molecular formula of alkanes and alkenes using their general formulas **7B.3** Represent linear and branched alkanes and alkenes using structural and semi-structural formulas **7B.4** Write systematic names for linear and branched alkanes and alkenes (including isomers) **7B.5** Connect physical properties to the intermolecular forces between molecules of hydrocarbons **7B.6** Describe the chemical properties and uses of alkanes and alkenes
• 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)	**7C Haloalkanes** **7C.1** Define the characteristics of haloalkanes **7C.2** Represent haloalkanes using structural and semi-structural formulas **7C.3** Write systematic names for haloalkanes (including isomers) **7C.4** Connect physical properties to the intermolecular forces between molecules of haloalkanes **7C.5** Describe the chemical properties and uses of haloalkanes

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Study Design:	Learning intentions – at the end of the chapter I will be able to:
- 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)	**7D** **Alcohols and carboxylic acids** **7D.1** Define the characteristics of alcohols and carboxylic acids **7D.2** Represent alcohols and carboxylic acids using structural and semi-structural formulas **7D.3** Write systematic names for alcohols and carboxylic acids (including isomers) **7D.4** Connect physical properties to the intermolecular forces between molecules of alcohols and carboxylic acids **7D.5** Describe the chemical properties and uses of alcohols and carboxylic acids

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Glossary

Addition reaction
Alcohol
Alkanol
Alkyl group
Biodegradable
Biomass
Branched alkane
Bromo-
Carbon neutral
Carboxyl group
Carboxylic acid
Chloro-
Combustion reaction
Crude oil
Dimer
Feedstock
Fluoro-
Fossil fuel
Fractional distillation
Functional group
Greenhouse gas
Haloalkane
Homologous series
Hydrocarbon
Hydroxyl group
Iodo-
Molecular weight
Non-renewable resource
Oxidant
Oxidation reaction
Parent hydrocarbon
Polymerisation
Positional isomer
Renewable
Saturated
Semi-structural formula
Straight-chain alkane
Structural formula
Structural isomer
Substitution reaction
Sustainable
Systematic nomenclature
Unsaturated

Concept map

Plant-based biomass as an alternative to fossil fuels and the impact products produced by organic chemistry have on society

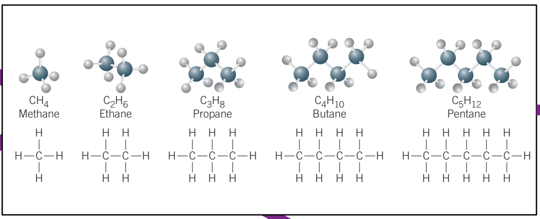
7A Organic chemistry in society

The grouping of hydrocarbon compounds into families of molecules based on properties and IUPAC nomenclature

7B Hydrocarbons

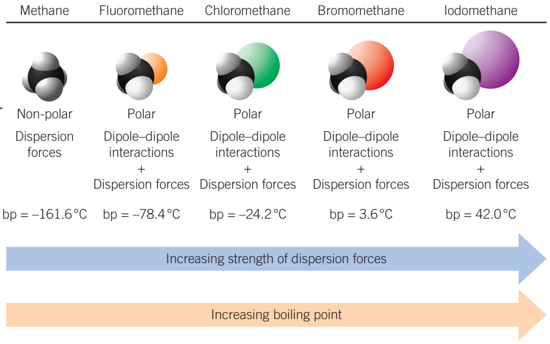
Alkanes

Alkenes



Focus on specific functional groups

7C Haloalkanes

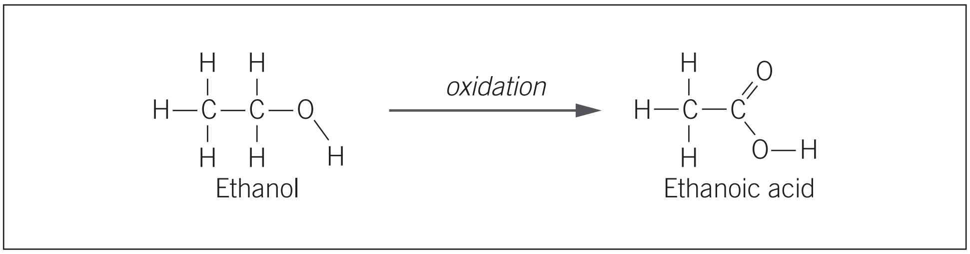


Methane	Fluoromethane	Chloromethane	Bromomethane	Iodomethane
Non-polar	Polar	Polar	Polar	Polar
Dispersion forces	Dipole–dipole interactions + Dispersion forces	Dipole–dipole interactions + Dispersion forces	Dipole–dipole interactions + Dispersion forces	Dipole–dipole interactions + Dispersion forces
bp = −161.6 °C	bp = −78.4 °C	bp = −24.2 °C	bp = 3.6 °C	bp = 42.0 °C

Increasing strength of dispersion forces

Increasing boiling point

7D Alcohols and carboxylic acids



See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

7A Organic chemistry in society

Study Design:

- 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 and artificial hearts) and an assessment of the benefits of those products for society, as well as the health and/or environmental hazards they pose

Glossary:

Biodegradable
Biomass
Carbon neutral
Crude oil
Feedstock
Fossil fuel
Fractional distillation
Greenhouse gas
Hydrocarbon
Non-renewable resource
Renewable
Sustainable

ENGAGE

Crude oil – the liquid black gold

You may not realise the extent to which crude oil has been essential to our modern civilisation. According to the Organisation of the Petroleum Exporting Countries (OPEC), more than 70 million barrels of crude oil are produced globally every day (Figure 7A–1). That equates to over 48 000 barrels, or 7.6 million litres per minute! It is also the most traded commodity in the world.

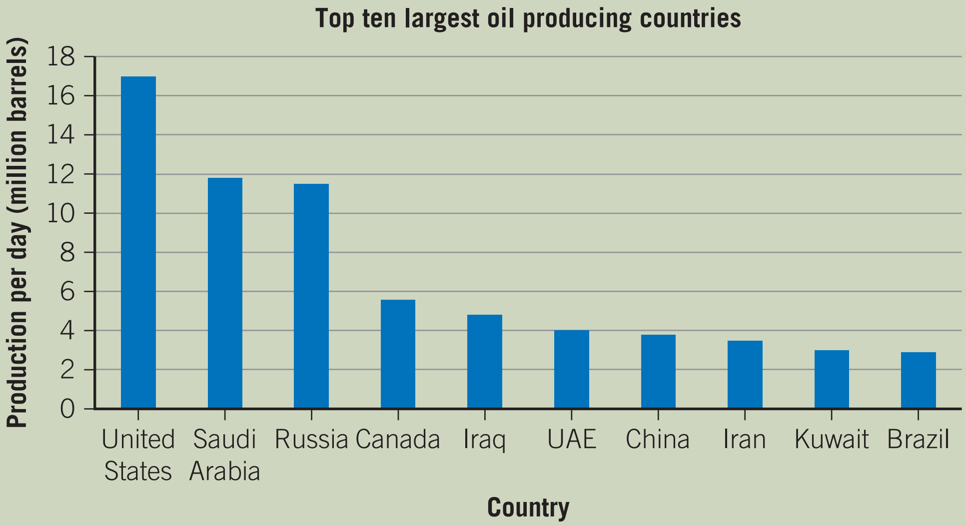


Figure 7A–1 Graph showing the top ten largest crude oil-producing countries

In many countries, access to affordable energy can mean a higher quality of life for its citizens and a faster trajectory for its economy to grow. In this regard, a country's crude oil reserves often attract investments from foreign nations. Due to the large amount of money involved in the commodity, crude oil can influence the relationship between a country's government and its citizens.

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EXPLAIN

Organic chemistry around us

Take a moment to look at the items around you. Many of these items would, at a molecular level, involve organic chemistry – the study of carbon and its compounds. Every item of food, wood, paper, plastic, medicine and all living organisms (including yourself) consists of carbon compounds. The study of their chemical composition, their properties and how they react involves organic chemistry. While you live your life surrounded by organic compounds, you may take for granted how these items are sourced and produced. This section explores and compares two major sources of organic material that we rely on to make everyday products. The first is crude oil – a finite and non-renewable resource formed over millions of years. The second is plant-based biomass – a renewable resource, which can be continuously and quickly replenished via natural processes.

Figure 7A–2 Everyday items that involve some sort of organic chemistry

Crude oil – a finite resource

Crude oil , also known as petroleum, is a naturally occurring brownish-black substance often found beneath Earth's surface. It is a **fossil fuel**, formed from the remains of prehistoric marine organisms. Chemically, crude oil is a mixture of compounds consisting mainly of carbon (85%) and hydrogen atoms (15%). These compounds are collectively known as **hydrocarbons**.

Formation of the crude oil we use today began millions of years ago. When prehistoric marine organisms, such as plants, algae and plankton died, they were buried in the sands on the sea floor. Over millions of years, sand and sediment layered on top of this organic matter, exerting heat and pressure on it. Under these conditions, the organic matter was transformed into crude oil. Certain geological formations trapped the oil in vast underground reservoirs. Today, oil rigs and platforms drill into these underground regions and use pressure to pump the oil to the surface.

Fossil fuel
a source of fuel made from decomposing plants and animals over hundreds of millions of years. It consists of mainly carbon and hydrogen atoms

Hydrocarbon
a compound made up of carbon and hydrogen atoms only

Crude oil
a naturally occurring mixture of liquid hydrocarbons that can be used as a fuel source

15B GAS STOICHIOMETRY AND THE GREENHOUSE EFFECT LINK

Greenhouse gas a gas that contributes to the greenhouse effect by trapping heat close to Earth's surface; e.g. carbon dioxide

Crude oil provides a reliable and cheap source of energy to our modern society. However, there are major disadvantages to using crude oil. The combustion of fossil fuels releases toxic gases and large amounts of carbon dioxide, a **greenhouse gas**, into the atmosphere. These gases have adverse effects on our planet's natural ability to regulate its atmospheric temperature. As shown in Figure 7A–3, over the last 120 years, greenhouse gases have exponentially accumulated (black line) in our atmosphere causing an increase in the global average temperature (coloured bars).

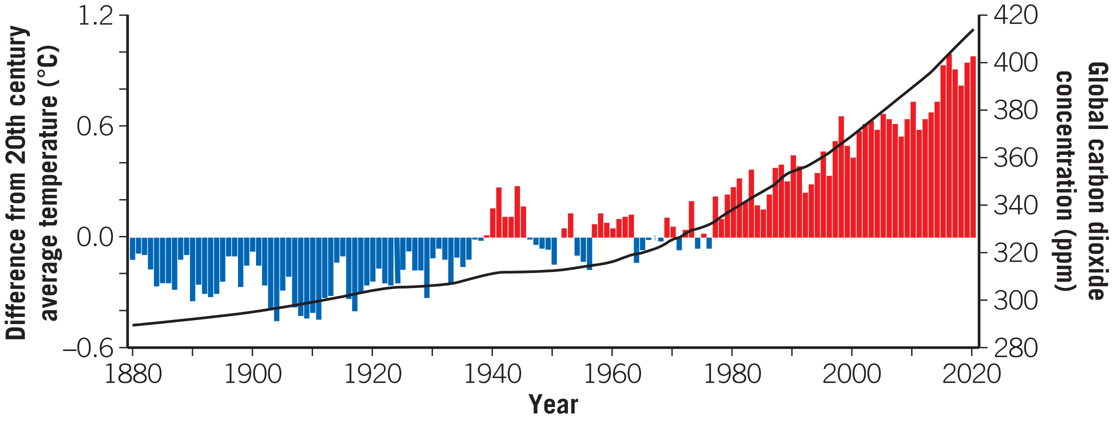

Figure 7A–3 Change in global atmospheric carbon dioxide and surface temperature from 1880–2020

Refining crude oil

Fractional distillation separation of a mixture of liquids into separate fractions based on their boiling points

Crude oil obtained from oil rigs is refined by a process called **fractional distillation**. This process involves separating the crude oil mixture based on the boiling points of its components. Large oil refineries process millions of barrels of crude oil each day and turn them into commercially useful products.

During fractional distillation, crude oil is heated in a furnace at the bottom of a fractionating column, as shown in Figure 7A–4. This supplies sufficient thermal energy to overcome the intermolecular forces between most of the compounds in the crude oil mixture. As the mixture boils, components begin to rise through the fractionating column. The temperature inside the column gradually decreases at higher elevation (further away from the furnace). Smaller hydrocarbons have lower boiling points and condense higher up in the column at cooler regions. Larger hydrocarbons have higher boiling points and condense at lower levels in the column. The products collected at each level are called fractions. Each fraction contains molecules with a certain range of boiling points. Petroleum gas or natural gas, used for cooking and heating our homes, typically consists of the smallest hydrocarbons. This fraction does not condense in the fractioning tower, with boiling points of less than 25°C. Petrol, a heavier fraction than natural gas that is typically used to power small vehicles, consists of larger molecules with higher boiling points (from 25–60°C). Diesel is an even heavier fraction consisting of hydrocarbons larger than those of petrol, with a boiling point range of 220–250°C. Diesel is suitable for powering trucks and other heavy vehicles. The components that do not boil at all are referred to as bitumen, whose main use is in the building of roads.

As shown in Figure 7A–4, crude oil refinement is a key step in the supply chain of many important goods and services in our society.

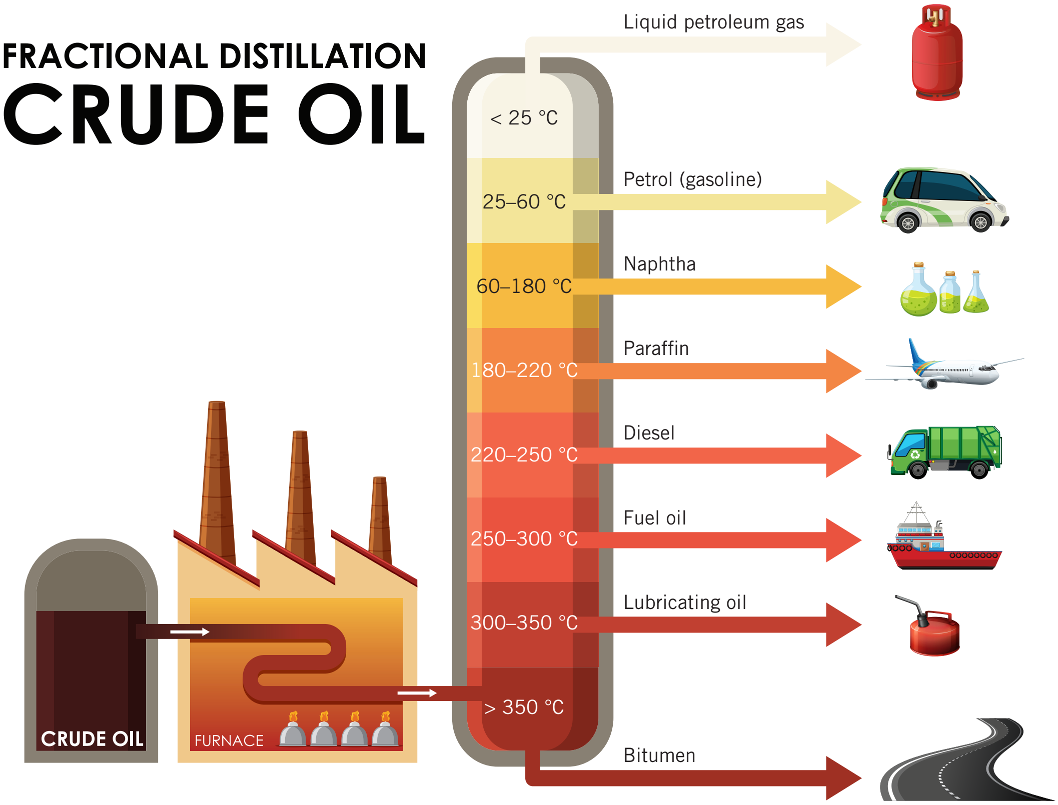

Figure 7A–4 Schematic of a fractionating tower for distillation of crude oil

Check-in questions – Set 1

1 Identify which physical property fractional distillation uses to separate mixtures.
2 Describe how hydrocarbons of different sizes are separated from crude oil using the fractional distillation technique.
3 Crude oil is usually not sold directly to the general public. It is refined by fractional distillation to obtain the products with commercial value. Referring to the composition of crude oil, explain why it is not very useful to the general public.

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Materials from crude oil

While the majority of the refined products from crude oil are used for energy, the remaining portion is used in manufacturing. The prime example is plastics (Figure 7A–5). Most plastics originate from crude oil and many of our everyday items are partially or entirely made of plastic. This is because plastic is an extremely versatile and cheap material to mass produce. Because of this, our economy and consumption behaviour have both shifted towards a use-and-discard mentality. It is often cheaper to replace a broken item than to fix it. This increased level of consumption increases our reliance on crude oil.

Figure 7A–5 Plastic resin made from materials from crude oil. This resin can be moulded into many of the plastic items we use daily.

Using crude oil-based products also generates a tremendous amount of waste that will take a long time to degrade or decompose. Plastics are mostly not **biodegradable**, meaning that microorganisms cannot break down the chemical bonds within the polymer molecules. As a result, the plastic that ends up in landfill will remain there, intact, for hundreds of years. Arguably, this constitutes one of the major long-term environmental costs of mass producing items made from plastic. It is a source of pollution that is both prolific and persistent. In Chapter 8B, you will look in more detail at the chemical structure of both fossil fuel-based plastics and bioplastics, as well as their recycling and circularity of renewability.

Biodegradable a substance or object capable of being decomposed by bacteria or other living organisms

8B POLYMER RECYCLING LINK

Figure 7A–6 Plastic is a common material found in landfills across the world.

Crude oil is also used as the raw material to produce a range of fabrics and textiles around our homes. For example, many of our clothing items, pillows and the carpet on our floors may have all originated from crude oil. Common synthetic fabrics (fabrics that are produced from materials obtained from crude oil) include polyester, nylon and spandex.

As the formation of crude oil is much slower than the rate at which we are consuming the commodity, it is a **non-renewable resource**. The world has a finite amount of crude oil that cannot be replaced once it is depleted. Prior to this point, there is an urgent need for the world to find or develop a reliable, sustainable and environmentally friendly replacement. This involves building new infrastructure to support an alternative source of energy and shifting our current supply chains toward a more circular economy using materials not derived from crude oil.

Non-renewable resource a natural resource that cannot be readily replaced by natural means at a pace quick enough to keep up with consumption

Biomass renewable organic matter that comes from plants, animals and other living organisms

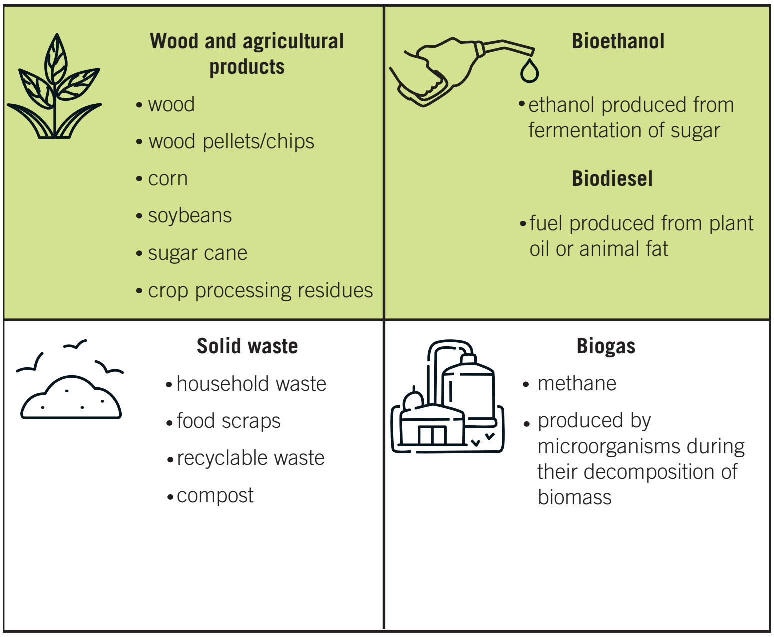
Plant-based biomass – a renewable resource

Biomass has been a resource for humans since the discovery of fire – many thousands of years ago. First Nations Australians developed many techniques for extracting, processing and preparing resources from plants such as medicines, foods, adhesives, binders and pigments, with many such techniques still in use today. *Investigation Topic 3: The chemistry of Aboriginal and Torres Strait Islander peoples' practices* in the VCE Chemistry Study Design provides the opportunity to research and discover these techniques.

The main type of biomass in use worldwide today is wood. The term **biomass** refers to organic material that comes from living organisms, such as plants, animals, and microorganisms. The raw materials obtained from biomass that are industrially relevant are called biomass **feedstock**. Figure 7A–7 shows the major types of biomass feedstocks currently used in our society. In this course, we will focus on how plant-based biomass (coloured green) can be used as an alternative to fossil fuels.

Feedstock
raw material that is used to supply an industrial process

Types of biomass feedstocks



Wood and agricultural products
- wood
- wood pellets/chips
- corn
- soybeans
- sugar cane
- crop processing residues

Bioethanol
- ethanol produced from fermentation of sugar

Biodiesel
- fuel produced from plant oil or animal fat

Solid waste
- household waste
- food scraps
- recyclable waste
- compost

Biogas
- methane
- produced by microorganisms during their decomposition of biomass

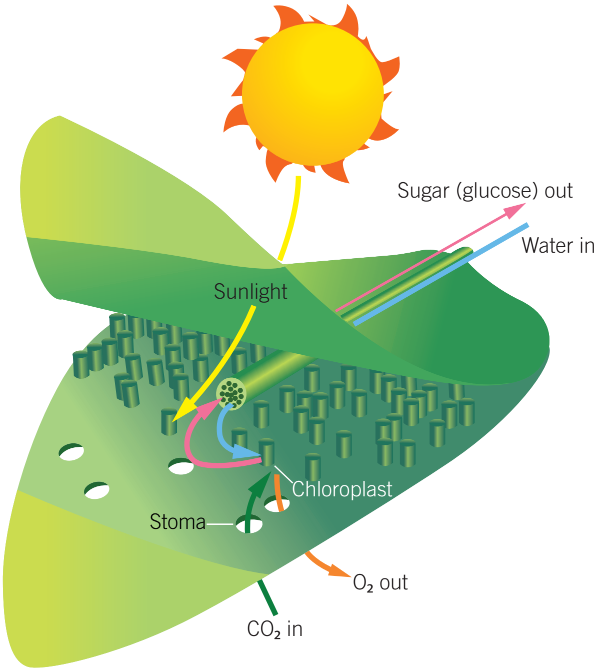
Figure 7A–7 Types of biomasses used in society

Renewable
a resource that can be replenished faster than it is consumed

Biomass feedstocks are of particular interest in today's society because they are **renewable** under proper management. This means that they can be replenished faster than they are consumed. Production of plant-based biomasses is powered by photosynthesis (Figure 7A–8), a process that uses energy from the sun while simultaneously removing CO_2 from the atmosphere. The use of biomass feedstock for our industrial processes means that we can reduce our reliance on fossil fuels.

How can plant-based biomass be used as an alternative to fossil fuels?

The power of using a renewable feedstock can be demonstrated in the example of ethanol production. Ethanol is an important organic solvent widely used in industry and scientific research. It is also a component found in alcoholic beverages. Ethanol can be produced on an industrial scale via the petrochemical route using non-renewable fossil fuels. In this case, ethene (C_2H_4) derived from crude oil is hydrated with water to give ethanol. Alternatively, the biological route involves fermentation of glucose obtained from a sugar-rich biomass such as sugarcane. The ethanol produced



$$\text{water + carbon dioxide} \xrightarrow{\text{Sunlight}} \text{glucose + oxygen}$$

$$6\,H_2O + 6\,CO_2 \xrightarrow{\text{Light energy}} C_6H_{12}O_6 + 6\,O_2$$

Figure 7A–8 The photosynthesis process to form plant-based biomass

Sustainable a process or practice that meets the needs of the current generation without compromising the needs of future generations

from the latter method is called bioethanol. The two processes are outlined in Figure 7A–9, with emphasis given to the time taken to produce the feedstock materials (ethene and glucose). Ethene, a hydrocarbon derived from distillation of crude oil, takes approximately 150 million years to form given suitable geological conditions. Production of ethanol from this method is not **sustainable** as there is a finite amount of fossil fuels. On the other hand, glucose can be extracted from sugarcane approximately 12–14 months after sowing the seed. Production of bioethanol can be sustainable as it is made from a renewable source.

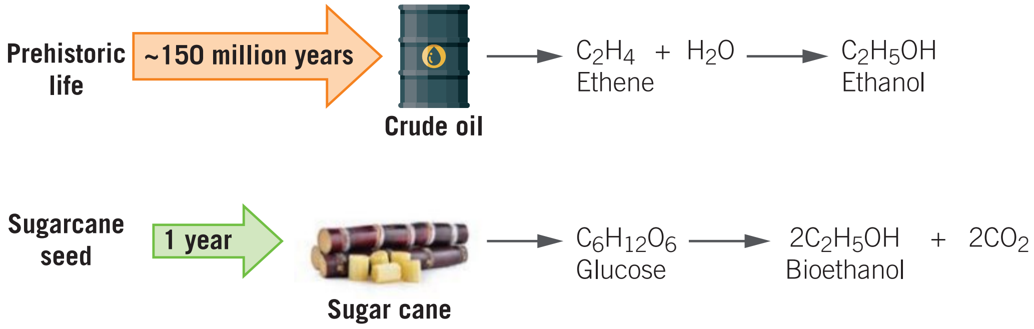


Figure 7A–9 Comparing the petrochemical and biological methods of ethanol production

Plant-based biomass can also provide numerous therapeutically important substances that are otherwise difficult to obtain from fossil fuel sources. This is because plants, through millions of years of evolution, have developed ways to produce complex molecular structures that are difficult to replicate in the laboratory. One example of this is morphine, a drug commonly used for severe-pain relief. The starting molecule for morphine synthesis was a molecule from the naphthalene family, derived from crude oil distillation (Figure 7A–10). The procedure involved 31 sequential reactions (steps) and a very low overall efficiency. To date, the most efficient laboratory synthesis of morphine still involves 14 steps. While these synthetic methods to prepare morphine are banners of great human endeavours, they are impractical from a commercial and industrial perspective. The morphine we use today is extracted directly from the plant *Papaver somniferum*, commonly known as the opium poppy. Through a series of liquid extractions, pure morphine can be obtained straight from the plant biomass without the need to perform any chemical reactions.

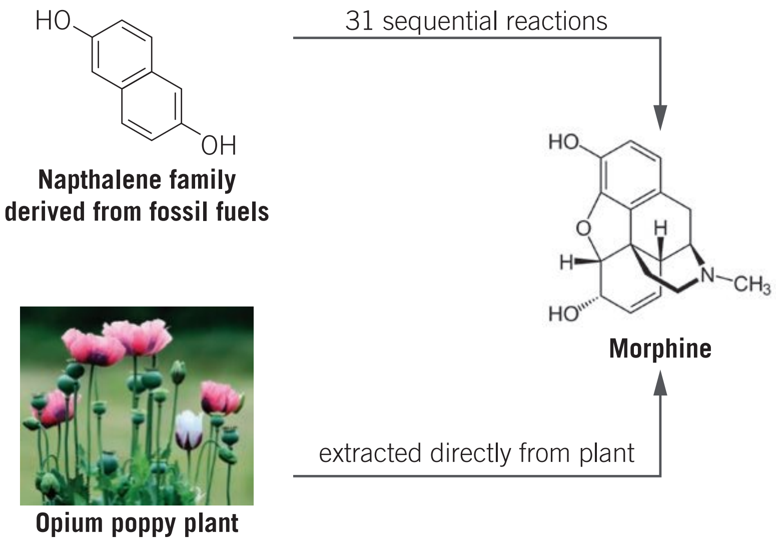


Figure 7A–10 Comparing the synthetic route and biomass route to morphine

Comparison of biomass and fossil fuels

Using plant-based biomass as an alternative to fossil fuel offers many advantages to our climate. Plant-based biomass can be a renewable resource, which is a major incentive over using fossil fuels. Growing plants allows us to capture the carbon dioxide in the atmosphere and convert it into a useful product. However, there are also several disadvantages or limitations associated with using plant-based biomass. The exact benefits and costs associated with using biomass over fossil fuels is highly dependent on the product being produced and the availability of the feedstock. This is outlined in Table 7A–1.

Table 7A–1 Advantages and disadvantages of using biomass as a fuel source

Advantages	Disadvantages
Renewability: It is possible to have continuous availability of biomass sources such as plants, manure and waste, and they are unlikely to be depleted. Plant-based biomass, such as trees, can be replanted once they are fully grown, making them renewable.	**Deforestation**: Growing plant-based biomass requires land, which often comes from clearing existing forests. While the trees from the forest are replaced by purposely grown plants, the diversity of these plants would not match that of the original forest. This would have further impact on the ecology of the land.
Reduced dependence on fossil fuels: Biomass can either be used as a fuel or feedstock for further production of goods. Both these avenues allow biomass to partially replace fossil fuel demand.	**Land management and water use**: Water is required to irrigate the biomass crops – water and land that may be required for agriculture and domestic use.
Carbon neutral: When growing plant-based biomass, the photosynthesis process captures and removes CO_2 from the atmosphere. This CO_2 may then be released into the atmosphere if the plant-based biomass is combusted or used for other purposes. Overall, the photosynthesis process balances the combustion process, making plant-based biomass a carbon-neutral energy source.	**Transportation**: Unprocessed plant matter contains a significant amount of water. This adds to the cost of transporting the material. Furthermore, when we account for the energy required for transporting biomass (for example, fuel for trucks and trains), it no longer is a truly carbon-neutral energy source.
Waste reduction: Agricultural waste, such as parts of crops that we do not consume, can be converted to biomass feedstock. This repurposes the waste we generate and avoids the cost of disposal and impact on the environment. Goods made from plants are also easier to decompose compared to goods made from plastic. This further reduces the build-up of waste at the end of a product's life.	**Cost**: Creating biomass feedstock requires specialised plants, machinery and storage facilities. The costs associated with setting up and maintaining the infrastructure may outweigh the benefit of using biomass feedstock.
Repurposing nature: Many of the materials and substances made by nature already have the molecular complexity to be repurposed into commercial products without significant processing. This includes some therapeutics, adhesives and coloured dyes.	**Seasonality**: Plants are grown and harvested according to specific seasons. This means that the availability of biomass feedstock is not consistent throughout the year. This adds an extra challenge of relying on plant-based biomass.

Carbon neutral a replenishing process where there is an overall balance between releasing and absorbing CO_2 from the atmosphere

WORKSHEET 7A–1 RENEWABLE AND NON-RENEWABLE ENERGY

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VIDEO 7A–1
SKILLS: ANSWERING COMPARATIVE QUESTIONS

7A SKILLS

Answering comparative questions

In VCE Chemistry, you will often encounter questions that ask you to compare and contrast different concepts.

Using this chapter as an example, you might be asked to compare the different types of biomass feedstocks, the petrochemical and biological methods of ethanol production or the advantages and disadvantages of using biomass as a fuel source.

Although comparison questions may seem straightforward, students often don't receive full marks for their answer. The main reason for this is that students often interpret the question as described below.

Question: Compare the petrochemical and biological methods of ethanol production.

Interpretation: List the differences between these ethanol production methods.

A comparison question requires you to answer it in two parts. First, you have to discuss the aspects that are similar between the two items, then you have to analyse the aspects that are different. This is the best way to ensure that you get maximum marks for a comparison-style question. Keep your comparisons succinct and do not repeat the question.

Let's use the example above to look at ways to approach this type of question.

Question:

Compare the petrochemical and biological methods of ethanol production.

Answer – Strategy 1: Group your responses by similarities and differences

Similarities:
- Both result in the production of ethanol.

Differences (make it clear that there is a comparison by using terms such as 'whereas', 'although', 'in contrast' and 'unlike'):
- In the petrochemical method, ethene (derived from crude oil) and water are reactants, whereas in the biological method, glucose is the reactant.
- In the petrochemical method, crude oil can take approximately 150 million years to produce the feedstock material, whereas in the biological method, glucose can be extracted from plants approximately one year after the seeds have been sowed.
- The biological method results in carbon dioxide production, whereas the petrochemical method does not.

Answer – Strategy 2: Draw a table listing key features and the responses for each item

	Petrochemical	Biological
Similarities	Both result in the production of ethanol.	
Differences	Ethene (derived from crude oil) and water are reactants	Glucose, produced by plants via photosynthesis is the reactant
	Crude oil can take approximately 150 million years to produce the feedstock material	Glucose can be extracted from plants approximately one year after the seeds have been sowed
	No carbon dioxide produced	Carbon dioxide produced

As further practice, it would be a good idea to attempt questions such as these using the strategies outlined above.

Section 7A questions

1 Explain why small molecules rise to the top of the fractionating tower during fractional distillation of crude oil.

2 Complete the following table by listing the uses of different fractions obtained from refining crude oil. Use research to find your answer.

Fraction	Used for
Petroleum (natural) gas	
Naphtha	
Petrol	
Diesel	
Fuel oil	
Lubricating oil	
Paraffin	
Bitumen	

3 Plastic is an important material obtained from fossil fuels. The table below lists several different types of common plastics. Use research to find two products made from these plastics and suggest alternative materials (for example, wood, metal, paper, fabric, stone) that can be used instead of plastic.

Plastic	Products made from the plastic	Possible alternative materials
Polyethene terephthalate (PET)		
High-density polyethene (HDPE)		
Polypropylene (PP)		

4 Household items can be made from different materials. For example, a chair can be made from plastic (derived from fossil fuels) or wood (from a biomass source).

a Explain which material is more sustainable over the long term for constructing chairs.

b Use your understanding of chemistry to state an advantage and a disadvantage of using plastic to construct household chairs.

c Plastic takes more than 450 years to decompose, while plant-based biomass takes less than 100 years. Explain the significance of this information in the context of disposal of plastic and wooden products.

5 Combustion of ethanol made from fossil fuels and combustion of bioethanol both release CO_2 into the atmosphere. Explain how these two sources differ in their environmental impact.

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7B Hydrocarbons

Study Design:

- 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)

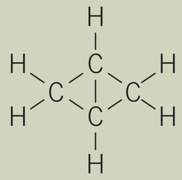
Glossary:

Addition reaction
Alkyl group
Branched alkane
Combustion reaction
Functional group
Homologous series
Parent hydrocarbon
Polymerisation
Saturated
Semi-structural formula
Straight-chain alkane
Structural formula
Structural isomer
Substitution reaction
Systematic nomenclature
Unsaturated

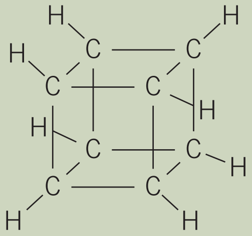
ENGAGE

Molecules under high strain

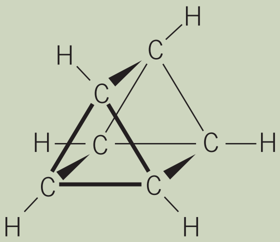
Explosive substances that release a large amount of energy have always been an area of interest to scientists. In pursuing this interest, organic chemists have synthesised the alkane molecules shown below with extremely high ring strain. An analogy for these molecules would be to imagine trying to bend a tree branch into a small loop so that it's under elevated strain to either spring open or snap. The strain placed on the bonds in these structures is likely to cause them to easily break – releasing energy in the overall process. It was hypothesised that these molecules may find uses as rocket fuels or powerful explosives. However, much to scientists' dismay, the energy-releasing powers of these molecules were rather underwhelming.



Bicyclo[1.1.0]butane



Cubane



Prismane

EXPLAIN

Alkanes

Carbon is an element that can make four covalent bonds with other atoms through the sharing of electrons. Due to this, carbon is contained within a large and diverse range of molecules. One family of these molecules are called alkanes, which are hydrocarbons with only single bonds between each carbon atom. Most alkanes are obtained directly from distillation of crude oil. Organic molecules that contain only C–C single bonds are described as **saturated**. The names and structures of the first five members of the alkane family are given in Figure 7B–1. It is important that the names of these compounds are memorised, as they will form the basis of naming all other organic molecules.

LINK **7A** ORGANIC CHEMISTRY IN SOCIETY

VIDEO 7B–1 WHAT ARE ALKANES?

Saturated
an organic molecule where all C–C bonds are single bonds

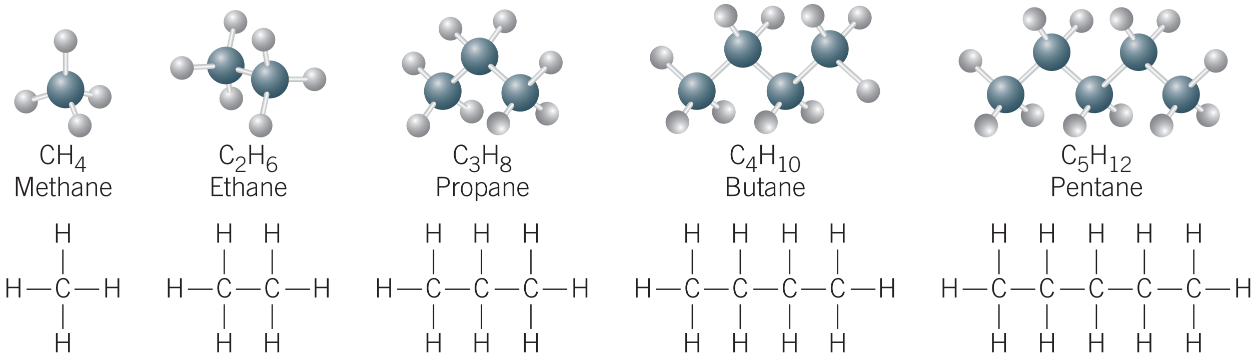


Figure 7B–1 The names and structures of the first five alkanes

The sequence of alkanes shown in Figure 7B–1 is an example of a **homologous series**. In a homologous series, each member differs from the previous one by a CH_2 group (carbon atom bonded to two hydrogen atoms). For example, adding a CH_2 group to ethane (C_2H_6) gives the molecular formula of propane (C_3H_8).

The relative number of carbon and hydrogen atoms in each alkane molecule follows a pattern. The general molecular formula for non-cyclic alkanes is C_nH_{2n+2} where n is any integer except 0. This general molecular formula can be used to determine the molecular formula of any non-cyclic alkane. If an alkane has four carbon atoms ($n = 4$), it must have 10 hydrogen atoms ($2n + 2 = 10$) and therefore it would have the molecular formula C_4H_{10}. If it has 15 carbon atoms, it must have 32 hydrogen atoms ($C_{15}H_{32}$).

Homologous series
a sequence of compounds that have identical functional groups and similar chemical properties but differ by a CH_2 group in their structure

Structural formula
a representation of how the individual atoms of the compound may be arranged in three-dimensional space

Semi-structural formula
a representation of organic molecules that shows the order in which atoms or groups of atoms are connected

Representations of alkanes

Organic molecules, such as alkanes, can be represented in one of many ways as outlined in Table 7B–1. Each representation offers different advantages and disadvantages. The choice of which representation to use often comes down to the context of the situation, the emphasis of the message and the ease of communication. A **structural formula** shows all the atoms and the bonds within the molecule. Lines are used to represent the bonds between atoms. It is worth noting that the bonds in the structural formula do not correspond to the actual angles between bonds or the shape of the molecule. A **semi-structural formula** (also called a condensed formula) simplifies the structural formula by omitting the covalent bonds within the molecule and listing the sequence of atoms bonded to each carbon. Semi-structural formulas with many repeating units can be simplified by using brackets around a repeating unit followed by a subscript number corresponding to the number of repeats. This is used in the examples of alkanes with four or more carbon atoms in Table 7B–1.

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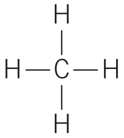
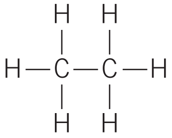
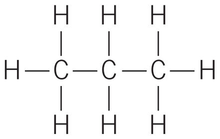
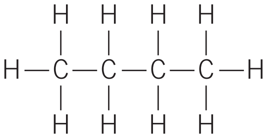
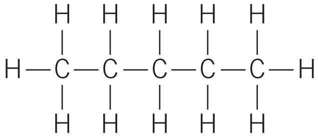
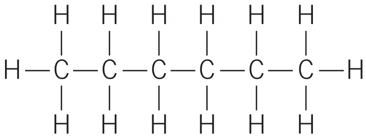
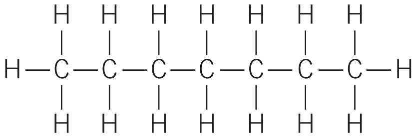
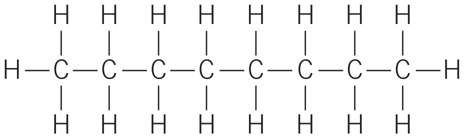
Table 7B–1 Comparison of different representations of alkanes

Number of carbon atoms	Name	Molecular formula	Structural formula	Semi-structural formula
1	Methane	CH_4	 $H-\underset{H}{\overset{H}{C}}-H$	CH_4
2	Ethane	C_2H_6	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	CH_3CH_3
3	Propane	C_3H_8	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	$CH_3CH_2CH_3$
4	Butane	C_4H_{10}	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	$CH_3CH_2CH_2CH_3$ or $CH_3(CH_2)_2CH_3$
5	Pentane	C_5H_{12}	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	$CH_3CH_2CH_2CH_2CH_3$ or $CH_3(CH_2)_3CH_3$
6	Hexane	C_6H_{14}	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	$CH_3CH_2CH_2CH_2CH_2CH_3$ or $CH_3(CH_2)_4CH_3$
7	Heptane	C_7H_{16}	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	$CH_3(CH_2)_5CH_3$
8	Octane	C_8H_{18}	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	$CH_3(CH_2)_6CH_3$

Worked example 7B–1: Writing semi-structural formulas

VIDEO
WORKED EXAMPLE 7B–1

Write the semi-structural formula of the alkane that has 12 carbons.

Solution

Step 1	Recall the general formula of alkanes.	C_nH_{2n+2}
Step 2	Calculate the number of hydrogen atoms in the molecule given that there are 12 carbon atoms ($n = 12$).	$n = 12$ $2n + 2 = 2 \times 12 + 2$ $= 26$ There must be 26 hydrogen atoms.
Step 3	Write the molecular formula.	$C_{12}H_{26}$
Step 4	Write the semi-structural formula.	$CH_3(CH_2)_{10}CH_3$
Step 5	Check that your semi-structural formula matches the molecular formula.	$CH_3(CH_2)_{10}CH_3$ has: 1 + 10 + 1 = 12 carbon atoms 3 + (2 × 10) + 3 = 26 hydrogen atoms This matches the formula $C_{12}H_{26}$

Check-in questions – Set 1

1 Determine the number of hydrogen atoms in an alkane with eight carbon atoms.

2 Give the semi-structural formula of an alkane with:

  a 10 hydrogen atoms.

  b six carbon atoms.

3 Draw the structural formula of an alkane with seven carbon atoms.

Nomenclature of straight chain alkanes

The complexity of organic molecules and the number of possible isomers exponentially increase as the number of carbon atoms increases. Rather than memorise the name of each structural isomer, chemists use a system of rules for naming structural isomers based on their structure. Since the name of the molecule describes its structure, you should also be able to deduce the structure of a compound from its name.

Straight-chain alkanes are those that feature all of their carbon atoms in a single continuous chain. The names of the first eight straight-chain alkanes are provided in Table 7B–2. These names consist of two parts: the first part of the name is the parent name and refers to the number of carbons in the **parent hydrocarbon**. The second part of the name, or the suffix, indicates the type of bonding within the parent hydrocarbon. The suffix 'ane' suggests that it is an alk*ane* and therefore consists of only C–C single bonds. Later in this section, we will explore alkenes where the suffix to the name will end in 'ene' rather than 'ane'.

Straight-chain alkane
an alkane consisting of a continuous chain of carbon atoms

Parent hydrocarbon
the longest continuous carbon chain in a molecule

Table 7B–2 Names of the first eight alkanes

Number of carbons	Parent name	Suffix	Name of alkane
1	Meth-	-ane	Methane
2	Eth-	-ane	Ethane
3	Prop-	-ane	Propane
4	But-	-ane	Butane
5	Pent-	-ane	Pentane
6	Hex-	-ane	Hexane
7	Hept-	-ane	Heptane
8	Oct-	-ane	Octane

WORKSHEET 7B–1
REPRESENTING AND NAMING ALKANES

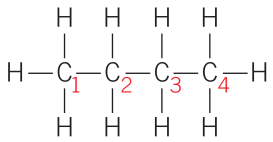
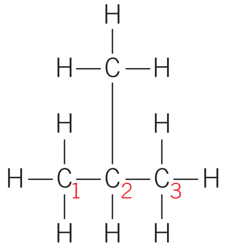
Branched alkane an alkane where the carbon atoms form a branched structure

Structural isomer a molecule with the same molecular formula but different connectivity of atoms

Structural isomers

As the size of the alkane increases beyond three carbon atoms, so too does the number of possible ways that the atoms in the molecule can bond with each other. For example, two possible structures with the molecular formula C_4H_{10} are shown in Table 7B–3. The first structure is a straight-chain alkane. The second structure has a branching CH_3 group, also called a methyl group, stemming from a continuous carbon chain. This structure is described as a **branched alkane**. These two molecules share the same molecular formula but have different atom connectivity and therefore structure. They are **structural isomers**.

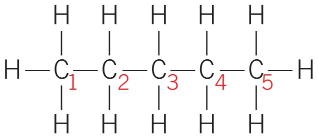
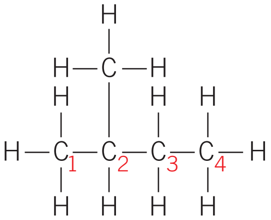
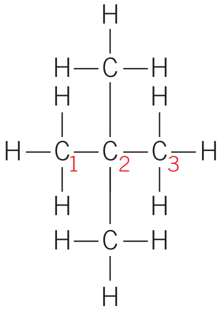
Table 7B–3 Structural isomers of butane (C_4H_{10})

Molecular formula	Structural formula	Semi-structural formula	Name	Carbon arrangement
C_4H_{10}		$CH_3CH_2CH_2CH_3$, or $CH_3(CH_2)_2CH_3$	Butane	Linear
C_4H_{10}		$CH_3CH(CH_3)CH_3$	2-methylpropane	Branched

VIDEO 7B–2 STRUCTURAL ISOMERS

As the number of carbon atoms in alkanes increases, more structural isomers become possible. For example, there are three structural isomers with the molecular formula C_5H_{12} (Table 7B–4). The straight-chain isomer is pentane, and the branched isomers are 2-methylbutane and 2,2-dimethylpropane. It is convention in the semi-structural formula to use brackets to encompass the branches.

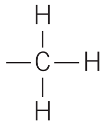
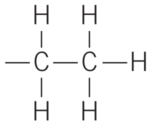
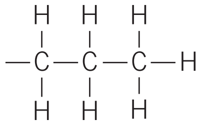
Table 7B–4 Structural isomers of pentane (C_5H_{12})

Molecular formula	Structural formula	Semi-structural formula	Name	Carbon arrangement
C_5H_{12}		$CH_3CH_2CH_2CH_2CH_3$, or $CH_3(CH_2)_3CH_3$	Pentane	Linear
C_5H_{12}		$CH_3CH(CH_3)CH_2CH_3$	2-methylbutane	Branched
C_5H_{12}		$CH_3C(CH_3)_2CH_3$	2,2-dimethylpropane	Branched

In cases where the alkane is branched, the type and position of the branch needs to be included in the name of the molecule. The branches, also called **alkyl groups**, are named according to the number of carbon atoms within them and follow the same prefix naming as shown in Table 7B–2. The names of alkyl groups end in '-yl' instead of '-ane'. Some common alkyl groups and their names are shown in Table 7B–5.

Alkyl group a group that is part of a larger molecule and is represented as an alkane minus a hydrogen atom, e.g. methyl (–CH_3)

Table 7B–5 Names and structures of alkyl side chain groups

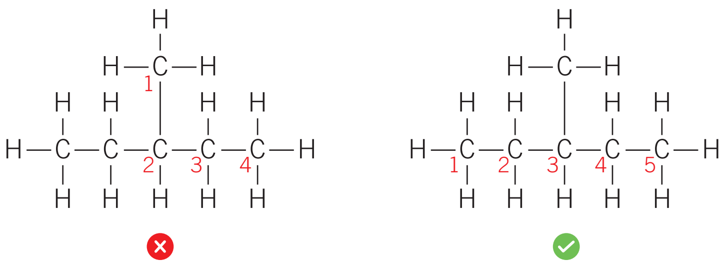
Number of carbons	Parent name	Suffix	Name of alkyl group	Structure
1	Meth-	-yl	methyl	 (CH_3)
2	Eth-	-yl	ethyl	 (C_2H_5)
3	Prop-	-yl	propyl	 (C_3H_7)

Nomenclature of branched alkanes

To put it all together, the systematic name of an alkane can be obtained by using the following rules.

Rule 1: The longest continuous carbon chain is the parent hydrocarbon of the molecule.

You should always apply this rule first by identifying the parent hydrocarbon. Note that the parent hydrocarbon does not need to be drawn in a straight line but can bend or turn within the structure. The name of the parent hydrocarbon will become the 'last name' for the molecule and appear at the end of the molecule's name. You will find some examples to help you in Figure 7B–2 below.



Incorrect
The parent hydrocarbon identified in the molecule above is not the longest continuous carbon chain.

Correct
The parent hydrocarbon is the longest continuous carbon chain. The parent chain is based on **pentane**.

Figure 7B–2 Demonstration of Rule 1

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Rule 2: The parent hydrocarbon is numbered so that alkyl groups are given the lowest possible numbers.

The carbon atoms in the parent hydrocarbon are numbered sequentially. It is possible to number the parent hydrocarbon from either end of the molecule, but the alkyl group **must** branch off the lowest possible numbered carbon atom in the parent hydrocarbon. This number will appear in front of the alkyl group, separated by a hyphen. The alkyl groups appear first in the name of the molecule. Figure 7B–3 will help you compare correct and incorrect examples of numbering the parent hydrocarbon.

3-methylbutane

✗

Incorrect

The alkyl group does not branch off the lowest numbered carbon in the parent hydrocarbon.

2-methylbutane

✓

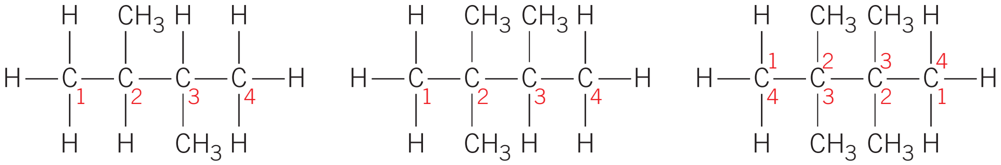
Correct

The alkyl group branches off the lowest possible numbered carbon atom in the parent hydrocarbon. (Note, numbering here is from right to left in the structural formula)

Figure 7B–3 Demonstration of Rule 2

Rule 3: Multiple identical alkyl groups are given the prefixes 'di', 'tri' and 'tetra' if they appear two, three or four times in the molecule, respectively.

The numbers of each alkyl group must appear individually even if the alkyl groups are connected to the same carbon atom on the parent hydrocarbon (for example, 2,2-dimethylpropane). Several examples of this are shown in Figure 7B–4.

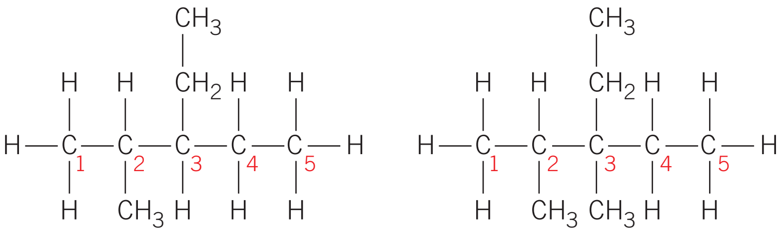


2,3-dimethylbutane 2,2,3-trimethylbutane 2,2,3,3-tetramethylbutane

Figure 7B–4 Demonstration of Rule 3

Rule 4: Multiple different alkyl groups are listed in alphabetical order, not numerical order.

If more than one type of alkyl group is attached to the parent hydrocarbon, they appear alphabetically in the name. The prefixes 'di-', 'tri-' and 'tetra-' are ignored for the purpose of alphabetising. Figure 7B–5 below shows some examples to illustrate this rule.



3-ethyl-2-methylpentane 3-ethyl-2,3-dimethylpentane

Figure 7B–5 Demonstration of Rule 4

The names obtained from these rules are called **systematic nomenclature** or IUPAC nomenclature (pronounced ‘I-U-pack’). IUPAC stands for the International Union of Pure and Applied Chemistry. A commission of scientists from IUPAC was responsible for designing these systematic nomenclature rules in Geneva, Switzerland, in 1892 and they are frequently updated. More IUPAC nomenclature rules will be introduced in subsequent sections where we explore organic molecules with more diverse **functional groups**.

LINK **7C** HALOALKANES

LINK **7D** ALCOHOLS AND CARBOXYLIC ACIDS

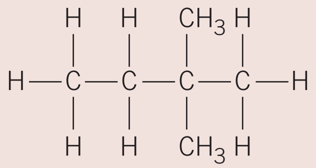
Systematic nomenclature the name given to an organic molecule that follows a set of systematic rules based on the structure of the molecule

Functional group an atom or a group of atoms in an organic molecule that largely determines the molecule’s properties and reactions

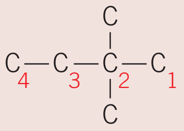
Worked example 7B–2: Writing systematic nomenclature

VIDEO WORKED EXAMPLE 7B–2

Write the systematic name for the following molecule.



Solution

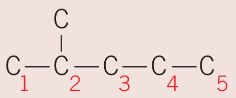
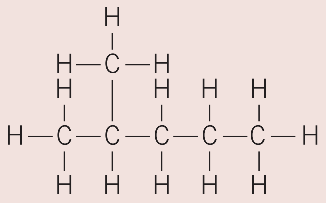
Step 1	Identify the number of carbon atoms in the longest continuous carbon chain (parent hydrocarbon).	The longest chain has four carbon atoms. Therefore, the parent hydrocarbon is a **butane**.
Step 2	Consider the two different ways of numbering the parent hydrocarbon. Number from the direction that gives the smallest number to the side chains.	
Step 3	Determine the type of side chain present.	There are two $-CH_3$ side chains, also known as a **methyl** side chains.
Step 4	Construct the name of the molecule.	Put it all together: 2,2-dimethylbutane

Worked example 7B–3: Drawing structural formulas of hydrocarbons

VIDEO WORKED EXAMPLE 7B–3

Draw the structural formula of 2-methylpentane.

Solution

Step 1	Determine and draw the parent hydrocarbon.	The parent hydrocarbon is pentane, which has five carbon atoms. C—C—C—C—C
Step 2	Number the parent hydrocarbon. Since you are drawing, you can start from either direction.	$\underset{1}{C}-\underset{2}{C}-\underset{3}{C}-\underset{4}{C}-\underset{5}{C}$
Step 3	Draw the side chain in its corresponding position.	There is a methyl side chain in position 2. 
Step 4	Complete the structure by adding hydrogen atoms.	

WORKSHEET 7B–2 STRUCTURAL ISOMERS

Combustion reaction a reaction involving oxygen that produces carbon dioxide (or carbon monoxide or carbon), water and heat

Check-in questions – Set 2

1 Draw the structural formula of 3-ethylpentane.

2 Give the IUPAC systematic names of the following molecules.

  a $CH_3CH(CH_3)CH_3$

  b $CH_3CH(CH_3)CH(CH_3)CH_3$

Chemical properties and uses of alkanes

7C HALOALKANES **7D** ALCOHOLS AND CARBOXYLIC ACIDS LINK

Alkanes are relatively unreactive compared to other families of organic molecules discussed later in this chapter. Being in the same homologous series, alkanes share very similar chemical properties, which means they all participate in a similar manner in chemical reactions.

Due to their ability to undergo **combustion reactions**, alkanes are generally used as fuels. Their different physical properties, such as melting and boiling points, lead to different alkanes being used as fuels under different circumstances. Table 7B–6 outlines applications of some members of the alkane family.

Table 7B–6 Uses of alkanes

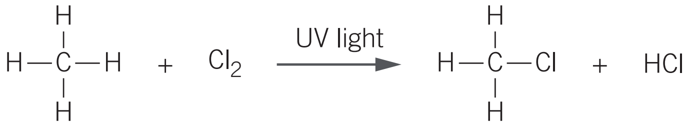
Alkane	Methane Ethane	Propane	Butane	Octane	Hentriacontane ($C_{31}H_{64}$)
Uses or occurrence					
Comments	Methane is the major component of natural gas used in gas stoves. Ethane is the second most abundant component of natural gas.	Propane is commonly stored in cylinders and used as fuel for barbeques in the form of liquified petroleum gas (LPG).	Butane is used as the fuel for portable stoves. It is also the fuel found in many lighters.	Petrol is primarily a mixture of alkanes ranging from 5–12 carbons.	Hentriacontane is a typical component of paraffin wax candles.

7C HALOALKANES LINK

UNIT 4 LINK

Alkanes can also undergo **substitution reactions**, which involve one of the hydrogen atoms being substituted for a different atom. A typical example of this involves the alkane reacting with Cl_2 gas in the presence of UV light, as shown in Figure 7B–6. The UV light is required for the reaction to occur. The products of the reaction are hydrogen chloride and a haloalkane, the latter of which will be explored in detail in Chapter 7C.

Substitution reaction a reaction where an atom or a group of atoms on a molecule is replaced with another atom or group of atoms



$$\begin{array}{c}\text{H}\\|\\\text{H}-\text{C}-\text{H}\\|\\\text{H}\end{array} + \text{Cl}_2 \xrightarrow{\text{UV light}} \begin{array}{c}\text{H}\\|\\\text{H}-\text{C}-\text{Cl}\\|\\\text{H}\end{array} + \text{HCl}$$

Figure 7B–6 Methane reacts with chlorine gas in the presence of UV light to form a haloalkane (and hydrochloric acid)

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Alkenes

Alkenes are another family of hydrocarbons, characterised by having one or more C=C double bonds. Molecules that contain C=C double bonds are described as **unsaturated**. Given that a C=C double bond must be in the structure, ethene (C_2H_4) is the smallest alkene and thus is the first member of the alkene homologous series. The structural formula of the first four members of the alkene family are shown in Figure 7B–7.

Unsaturated a molecule that has one or more C=C double bonds or C≡C triple bonds

VIDEO 7B–3 WHAT ARE ALKENES?

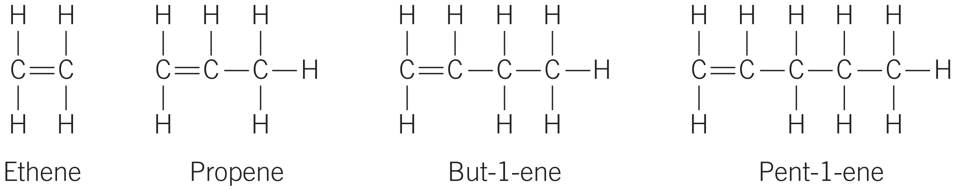


Figure 7B–7 The first four members of the alkene homologous series

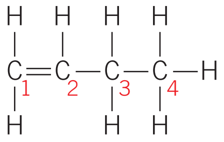
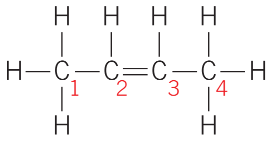
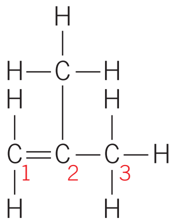
As we saw for the alkane series, the ratio of carbon to hydrogen atoms in alkenes also follows a relationship. The general molecular formula of non-cyclic alkenes with only one C=C bond is C_nH_{2n}, where n is an integer that is greater than zero. For example, an alkene with three carbon atoms ($n = 3$) must have six hydrogen atoms ($2n = 6$). Therefore, its molecular formula is C_3H_6.

The C=C double bond in alkenes is the functional group of the molecule. A functional group is an atom or a group of atoms that gives the molecule its physical and chemical properties.

Representations of alkenes

Alkenes are commonly represented by their molecular, structural or semi-structural formulas as shown in Table 7B–7. The double bond is represented by two solid lines in the structural formula. In the semi-structural formula, the double bond may or may not be explicitly shown. You should become familiar with both ways of representing alkenes. Butene is the smallest alkene to have structural isomers. The three non-cyclic isomers of butene shown in Table 7B–7 differ by the position of their double bond (but-1-ene versus but-2-ene) and by the arrangement of the carbon atoms (but-1-ene versus 2-methylpropene).

Table 7B–7 Representations of all non-cyclic isomers with the molecular formula C_4H_8

Name	Molecular formula	Structural formula	Semi-structural formula
But-1-ene	C_4H_8		$CH_2CHCH_2CH_3$ or $CH_2{=}CHCH_2CH_3$
But-2-ene	C_4H_8		$CH_3CHCHCH_3$ or $CH_3CH{=}CHCH_3$
2-methylpropene	C_4H_8		$CH_2C(CH_3)CH_3$ or $CH_2{=}C(CH_3)CH_3$

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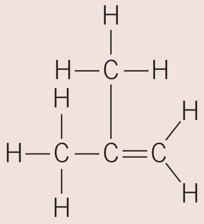
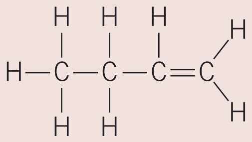
Worked example 7B–4: Drawing structural isomers

VIDEO WORKED EXAMPLE 7B–4

Draw a structural isomer of $CH_3C(CH_3)CH_2$.

Solution

Step 1	Assess and visualise the molecule. Obtain key information about the molecule.	Information that can be obtained: structural formula:  molecular formula: C_4H_8 branched isomer
Step 2	Consider how the original molecule can be rearranged using the same atoms. Make sure you don't just flip or rotate the original molecule.	Draw a linear isomer.
Step 3	Draw the isomer.	

Nomenclature

Naming alkenes relies on many of the naming rules previously mentioned, so it is important to recall what you already know about these rules. As you can see from Table 7B–8, the parent naming system remains the same as it is for alkanes. For example, eth-, prop-, and but- are still associated with two-, three- and four-carbon parent chains respectively. The names of alkenes always have the suffix '-ene'. So, the name of a three-carbon alkene would be propene.

Table 7B–8 Names of alkene parent hydrocarbons

Number of carbons	Parent name	Suffix	Name of alkene
2	Eth-	-ene	Ethene
3	Prop-	-ene	Propene
4	But-	-ene	Butene
5	Pent-	-ene	Pentene

In addition to all the nomenclature rules that apply to alkanes, the following rule applies specifically to naming alkenes:

Rule 1: The C=C double bond must be part of the parent hydrocarbon and be numbered to have the lowest possible number.

If the parent hydrocarbon is longer than three carbon atoms, the location of the double bond must be stated in the name. This is done by inserting the number corresponding to the first carbon atom of the double bond prior to the '-ene' suffix in the name but after the prefix representing the number of carbon atoms in the longest chain. The parent hydrocarbon is numbered from the side that allows the first unsaturated carbon to have the lowest possible number, before alkyl groups are considered. There is no need to specify the location of the C═C double bond for propene or ethene because it is, by default, always on the first carbon. Figure 7B–8 demonstrates the different parts of an alkene's name.

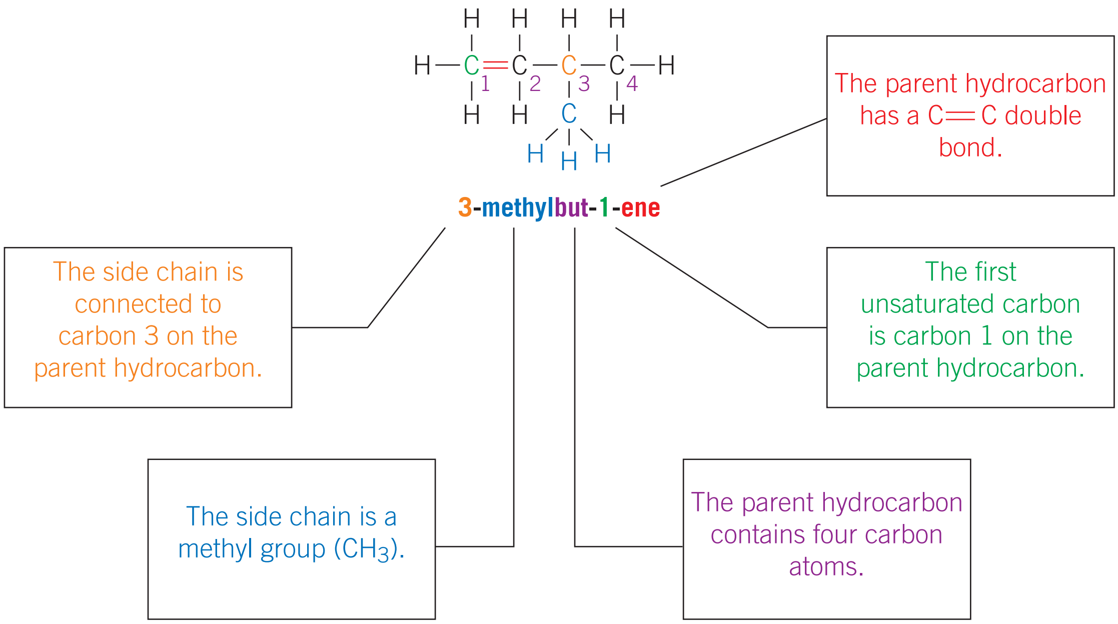


Figure 7B–8 The information given by the various parts of an alkene's name

Three examples of correctly and incorrectly named alkenes are shown in Figure 7B–9. See if you can link the incorrect names to breaking the rule stated above.

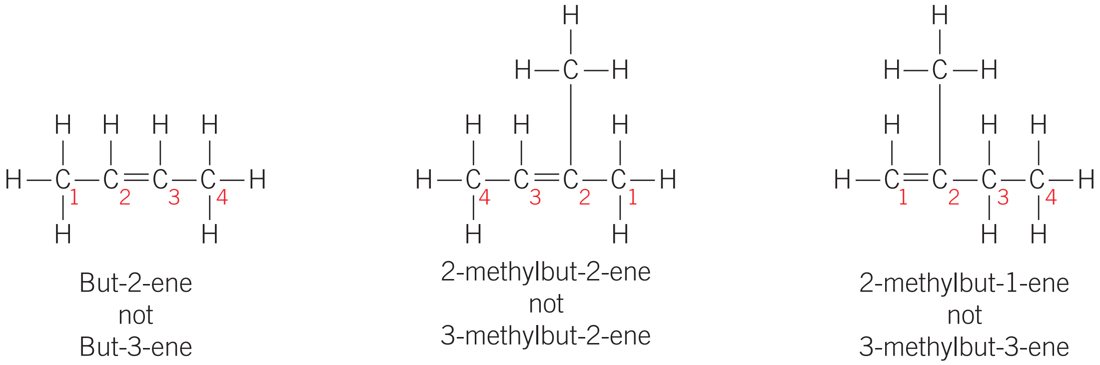


Figure 7B–9 Demonstration of correct and incorrect names of alkenes

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Worked example 7B–5: Naming organic compounds

VIDEO
WORKED
EXAMPLE 7B–5

Give the IUPAC name of $CH_3CH{=}C(CH_3)CH_3$.

Solution

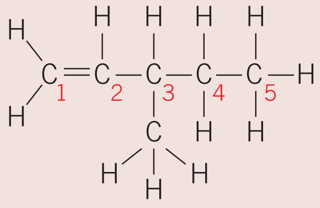
Step 1	Determine the longest continuous carbon chain including the $C{=}C$ double bond (parent hydrocarbon).	$CH_3CH{=}C(CH_3)CH_3$ The longest chain has four carbon atoms. (carbon atoms in brackets not included, as this is the branched part of chain) Therefore, the parent hydrocarbon is a **butene**.
Step 2	Determine how to give the $C{=}C$ double bond the lowest numbering.	$\underset{4}{CH_3}\underset{3}{CH}{=}\underset{2}{C}(CH_3)\underset{1}{CH_3}$ $\underset{1}{CH_3}\underset{2}{CH}{=}\underset{3}{C}(CH_3)\underset{4}{CH_3}$ Both directions position the $C{=}C$ double bond starting at carbon 2. Therefore, the parent hydrocarbon is a **but-2-ene**.
Step 3	Determine how to give any side chains the lowest possible numbers.	$\underset{4}{CH_3}\underset{3}{CH}{=}\underset{\textcircled{2}\checkmark}{C}(CH_3)\underset{1}{CH_3}$ $\underset{1}{CH_3}\underset{2}{CH}{=}\underset{\textcircled{3}\times}{C}(CH_3)\underset{4}{CH_3}$ The methyl group should be positioned on carbon 2.
Step 4	Construct the name of the molecule.	Put it all together: **2-methylbut-2-ene**

Worked example 7B–6: Drawing structural formulas of hydrocarbons

VIDEO
WORKED
EXAMPLE 7B–6

Draw the structural formula of 3-methylpent-1-ene.

Solution

Step 1	Identify and draw the carbon atoms of the parent hydrocarbon.	The parent hydrocarbon is pent-1-ene. $C{=}C-C-C-C$
Step 2	Number the parent hydrocarbon to give the $C{=}C$ double bond the lowest numbers.	$\underset{1}{C}{=}\underset{2}{C}-\underset{3}{C}-\underset{4}{C}-\underset{5}{C}$
Step 3	Identify the side chains and draw them in their corresponding positions. Also include all hydrogen atoms attached to carbon atoms.	Methyl (CH_3) group on carbon 3. 

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Check-in questions – Set 3

1 Complete the following table.

Molecular formula	Structural formula	Semi-structural formula
C_3H_6		
		$CH_3CH_2CHCHCH_2CH_3$
		$CH_2C(CH_3)CH_3$

2 Draw and name the structural formulas of all three isomers with the molecular formula C_4H_8.

3 Name the following alkenes.
   a CH_3CHCH_2
   b $CH_2CHCH_2CH_3$
   c $CH_3CH_2CHCHCH_3$
   d $CH_3CH(CH_3)CHCH_2$
   e $CH_3C(CH_3)CHCH_3$

WORKSHEET 7B–3 NAMING AND DRAWING HYDROCARBONS (ALKANES AND ALKENES)

Chemical properties and uses of alkenes

Like alkanes, alkenes can also be used as a fuel source and undergo combustion reactions.

However, unlike alkanes, the C=C double bond in alkenes allows them to participate in **addition reactions**. These reactions involve the C=C double bond breaking one of its bonds, with each carbon in that bond then forming a new bond to another element or molecule. For example, the equation below shows the addition reaction between ethene and chlorine (Cl_2). The product is a haloalkane, where the chlorine atoms are covalently bonded to the carbons where the double bond used to exist.

Addition reaction a reaction where two or more reactants combine to form a larger product. Addition reactions require an unsaturated reactant

LINK **7C** HALOALKANES

Figure 7B–10 Addition reaction between ethene and chlorine gas, resulting in the production of a haloalkane.

The ability of alkenes to participate in addition reactions gives them many interesting applications. A particular type of addition reaction called addition **polymerisation** makes use of various alkenes to produce important plastics that we commonly use today. Addition polymerisation will be explored in detail in Chapter 8A. This process allows us to produce millions of tonnes of plastics each year from cheap and readily available alkene molecules (Table 7B–9).

Polymerisation a reaction that covalently links repeated units of smaller molecules to make a larger molecule

LINK **8A** POLYMERISATION REACTIONS

Table 7B–9 Application of alkenes and products from reactions with alkenes

Alkene	Application	Explanation
Ethene		Ethene is polymerised to produce polypropene, the material used for milk bottles, plastic cling wrap, plastic bags, plastic bottles and many other everyday objects.
Propene		Propene is polymerised to produce polypropene, the material used to make reusable grocery bags, single-use masks and nappies.
Chloroethene		Chloroethene is polymerised to produce (polyvinyl chloride) (PVC). This material is used to make plastic pipes.

Physical properties of alkanes and alkenes

2B MOLECULAR SHAPES AND POLARITY LINK

Hydrocarbons are non-polar molecules. This is because carbon and hydrogen have very similar electronegativities, making the C—H bond relatively non-polar. The bond geometry around each carbon atom cancels out the small dipoles from the C—H bonds. Many of the physical properties we observe for alkanes can be understood through the lens of these two key features:

- Alkanes are overall non-polar molecules.
- Only dispersion forces exist between molecules of alkanes.

Boiling point

The small hydrocarbons we focus on in this course (C_1–C_8) are generally liquids or gases at room temperature. The dispersion forces between hydrocarbon molecules are easily overcome by a small amount of thermal energy. As shown in Table 7B–10, all straight-chain alkanes between methane and butane have boiling points below room temperature (25°C) and exist as gases at room temperature. Pentane is the smallest alkane that exists as a liquid at room temperature. One observable trend from the table is that as the size of the alkane increases in the homologous series, so too do its boiling point and melting point.

Table 7B–10 Boiling and melting points of straight-chain alkanes

Number of carbons	Name	Semi-structural formula	Boiling point (°C)	Melting point (°C)
1	Methane	CH_4	–167.7	–182.5
2	Ethane	CH_3CH_3	–88.6	–183.3
3	Propane	$CH_3CH_2CH_3$	–42.1	–187.7
4	Butane	$CH_3CH_2CH_2CH_3$	–0.5	–138.3
5	Pentane	$CH_3CH_2CH_2CH_2CH_3$	36.1	–129.8

The various boiling points of alkanes influence how we utilise them in our everyday lives. For example, natural gas (mainly consisting of methane) is easy to transport and distribute to our homes for heating and cooking. Petrol (mainly consisting of hydrocarbons between pentane and octane) is a liquid at room temperature, so it is easily stored in the fuel tanks of our cars and used by internal combustion engines.

When comparing the boiling points of structural isomers, the structure that is linear would have the highest boiling point (Figure 7B–11). This is because there is a high surface area of contact between linear molecules. Greater surface area of contact translates to stronger dispersion forces between the molecules, resulting in more thermal energy required to overcome the intermolecular forces. In comparison, branched isomers have less surface area of contact and therefore weaker dispersion forces between two molecules and lower boiling points.

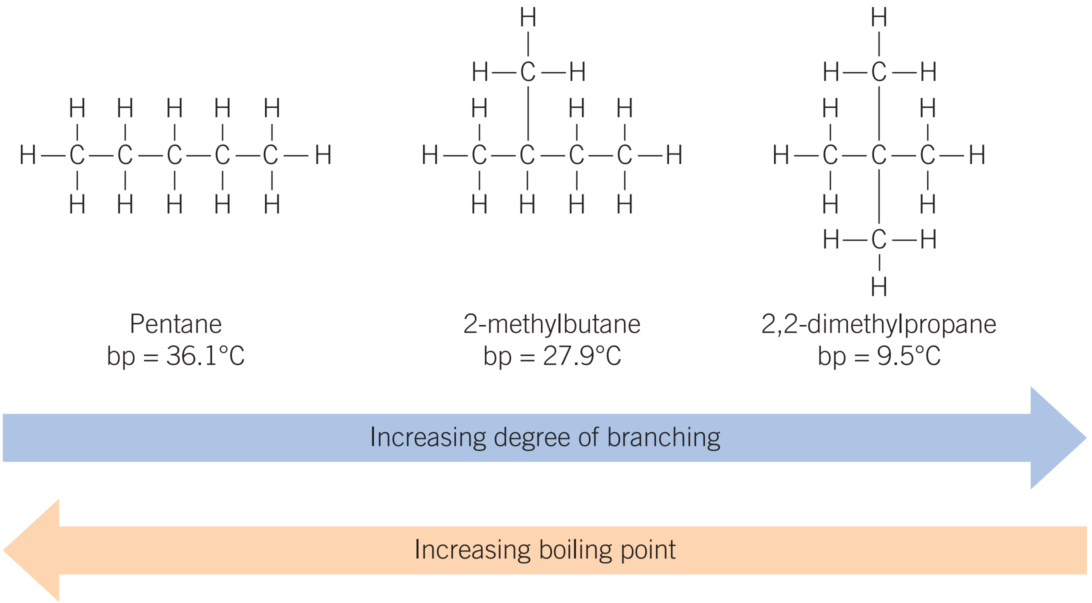

Figure 7B–11 Boiling points of all three isomers of C_5H_{12}

Similarly, larger alkenes have greater surface area of contact between molecules and therefore stronger dispersion forces than smaller alkenes. As a result, greater thermal energy is required to overcome the intermolecular forces between larger alkenes. The boiling points of selected alkenes are shown in Table 7B–11. From the table, these three key ideas are worth noting:

1. Boiling point increases as molar mass increases (due to an increase in the number of dispersion forces).
2. The boiling points of straight-chain isomers are higher than branched isomers (compare pent-1-ene to 3-methylbut-1-ene).
3. The position of the C═C double bond has relatively little influence on the boiling point (compare pent-1-ene to pent-2-ene).

Table 7B–11 Boiling points of alkenes

Name	Semi-structural formula	Boiling point (°C)
Ethene	CH_2CH_2	–103.7
Propene	CH_2CHCH_3	–47.6
But-1-ene	$CH_2CHCH_2CH_3$	–6.3
Pent-1-ene	$CH_2CHCH_2CH_2CH_3$	30
Pent-2-ene	$CH_3CHCHCH_2CH_3$	30
3-methylbut-1-ene	$CH_2CHCH(CH_3)CH_3$	20

Melting point

Substances melt when there is enough thermal energy to disrupt the intermolecular forces holding the individual molecules firmly in place. The melting point of a substance is the temperature at which the transition from a solid state to a liquid state occurs. For alkanes, this temperature is low because the intermolecular dispersion forces between the molecules are relatively weak. In general, the melting points of alkanes follow a similar trend to that of their boiling points. That is, larger alkanes have higher melting points due to stronger dispersion forces holding the alkane molecules together.

Solubility

5A POLARITY LINK

13B PREDICTING SOLUBILITY LINK

Solubility is the ability of a substance to dissolve in a solvent to form a solution, as you learned in Chapter 5A. The general rule that governs solubility is 'like dissolves like'. To elaborate, non-polar hydrocarbons will not be soluble in a polar solvent, such as water. When tragedies such as oil spills occur, the alkane molecules float on top of the ocean instead of mixing with the water. Figure 7B–12 shows the effects of the Deepwater Horizon oil spill in 2010.

Figure 7B–12 A boat attempting to combust the floating crude oil spill in the Gulf of Mexico in 2010

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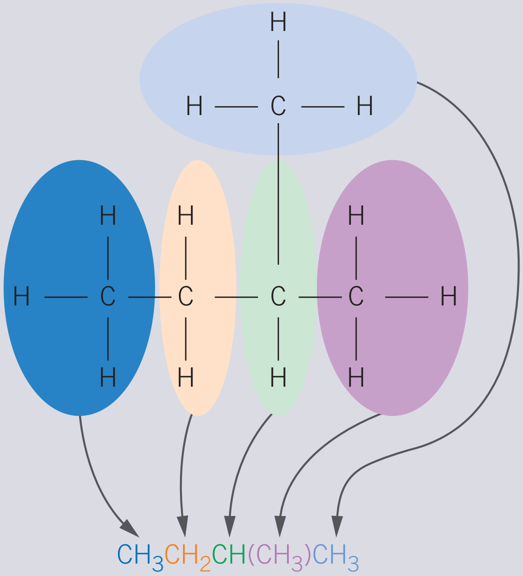
7B SKILLS

Writing semi-structural formulas

Writing semi-structural formulas from the names of molecules can often be challenging. This is because our brains usually need to visualise the molecule in its structural formula before being able to condense it into the semi-structural form. One way to navigate these problems is to draw the full structure first. While this technically may take longer, you are less likely to make careless mistakes in your answer. Converting a full structure into a semi-structure is a simpler task as outlined in the diagram below.

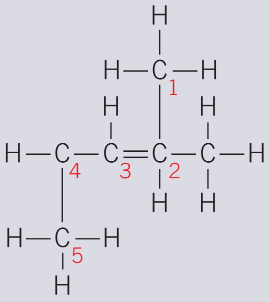
Drawing structural formulas

Drawing structural formulas of organic molecules from their IUPAC name is a typical type of question you'll encounter and therefore an important skill to master. One of the most common mistakes students make when drawing structural formulas is to include too many or too few hydrogen atoms in the molecule. Filling in the hydrogen atoms is usually the last step in drawing structural formulas. For large organic molecules, this step becomes repetitive and rather tedious. This is where students tend to become careless.



A simple method to avoid making this mistake is to perform a check using the rule: all carbon atoms **must** always have four bonds. This check should be performed after drawing a full structural formula. To do this, you simply need to scan every carbon in your structure and make sure there are exactly four bonds (lines) connecting to every one of them. Double bonds count as two bonds.

Question: Can you identify two different errors in the structure on the right?



Solution: The second carbon in the parent hydrocarbon has five bonds connected to it and the fourth carbon in the parent hydrocarbon only has three bonds connected to it.

Section 7B questions

1 Ethane is a member of the alkane homologous series.
  - a Give the molecular formula of ethane.
  - b Why is ethane considered a hydrocarbon?
  - c Why is ethane considered saturated?
  - d Draw the semi-structural formula of ethane.

2 C_4H_{10} is a hydrocarbon.
  - a Give the systematic names of two organic molecules with the above molecular formula.
  - b What is the relationship between the two molecules named in part **a**?
  - c Which molecule in part **a** has the higher boiling point? Explain your answer.

3 Draw the structural formula of the following hydrocarbons from their semi-structural formulas.

	Semi-structural formula	Structural formula
a	CH_4	
b	CH_3CH_3	
c	$CH_3CH_2CH_3$	
d	$CH_3CH_2CH_2CH_3$	

4 Write the condensed formula for the following straight-chain alkanes.

	Molecular formula	Condensed formula
a	C_2H_6	
b	C_5H_{12}	
c	C_8H_{18}	

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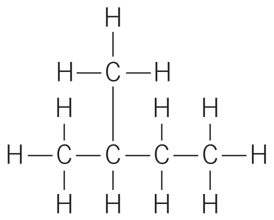
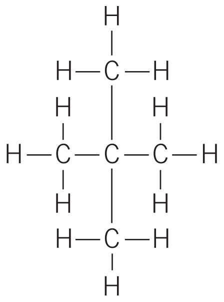
5 Give the IUPAC names of the following molecules.

a



b



6 The following molecules have been named incorrectly according to IUPAC rules. After drawing each molecule, give its correct systematic name.

a 3-methylpropane
b 2-ethylpropane
c 4-methylpentane
d 3,3-dimethylbutane

7 Recall the difference between the alkane and alkene families of compounds.

8 Categorise the straight-chain hydrocarbons below as alkanes, alkenes or neither:

$$CH_4, C_2H_6, C_3H_6, C_4H_8, C_5H_8, C_6H_{14}, C_7H_{14}, C_7H_{12}$$

9 Draw the structural formulas of the following hydrocarbons from their semi-structural formulas.

	Semi-structural formula	Structural formula
a	CH_2CH_2	
b	$CH_3CHCHCH_3$	
c	$CH_3CH_2C(CH_2CH_3)CHCH_3$	

10 Give the semi-structural formula for each of the following molecules.

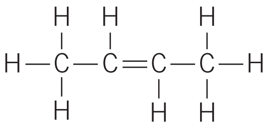
a but-1-ene
b 3-methylbut-1-ene
c 2-methylbut-1-ene
d 3-methylpent-2-ene

11 Write an equation for the following addition reactions of alkenes. Show organic molecules in structural formula.

a $CH_2CH_2 + Br_2(aq)$
b $CH_2CHCH_3 + Cl_2(aq)$
c $CH_3CHCHCH_3 + I_2(aq)$

12 Ethene, CH_2CH_2, is mixed with HCl(aq). Write a balanced equation showing all structural formulas.

13 But-2-ene is a member of the alkene homologous series. It is used as a starting material in the production of a wide range of materials that go towards making day-to-day items, such as film, bottle caps and other packaging. The structure of but-2-ene is given below.



Define 'homologous series', and give an example of another chemical in the same homologous series.

14 Methane, CH_4, and propane, C_3H_8, are non-polar molecules, which only interact by dispersion forces. Explain why propane has a higher boiling point than methane.

7C Haloalkanes

Study Design:

- 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)

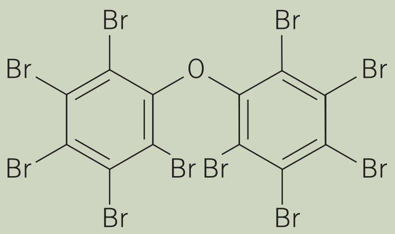
Glossary:

Bromo-
Chloro-
Fluoro-
Haloalkane
Iodo-

ENGAGE

The fireproof properties of haloalkanes

Recall that alkanes and alkenes are organic compounds that readily undergo combustion reactions. Haloalkanes, the family of molecules discussed in this section, display the opposite property: they are flame retardants, or substances that prevent the spread of fire. Molecules containing a high number of halogen atoms (group 17 elements) can be incorporated into fabrics or building materials to give them fireproof properties. One such molecule is called decabromodiphenyl ether (decaBDE). It has the molecular formula of $C_{12}OBr_{10}$ where bromine accounts for over 83% of its molecular weight.



Decabromodiphenyl ether

Haloalkane
an alkane in which one or more halogen atoms has replaced a hydrogen atom or atoms

Fluoro-
a functional group characterised by a fluorine atom bonded to a carbon

Chloro-
a functional group characterised by a chlorine atom bonded to a carbon

EXPLAIN

Haloalkanes

Haloalkanes are another family of organic molecules. The *halo* portion of the name refers to halogens or group 17 elements. Therefore, haloalkanes are alkanes where one or more halogen atom has substituted a hydrogen atom or atoms. The smallest haloalkane is fluoromethane, which has the chemical formula CH_3F. Haloalkanes can vary in complexity through the nature of the halogen atom, the number of halogen atoms in the molecule and the size of the carbon chain. It is also possible to have a mixture of different halogens on the same molecule, as shown in Table 7C–1. The halogen atom in the haloalkane molecule is considered the functional group of the molecule. For these functional groups, the '-ine' from the element name is dropped and replaced with 'o' (for example, **fluoro-**, **chloro-**, **bromo-** and **iodo-**).

VIDEO 7C–1
HALOALKANES

Bromo-
a functional group characterised by a bromine atom bonded to a carbon

Iodo-
a functional group characterised by an iodine atom bonded to a carbon

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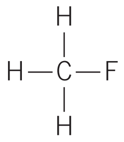
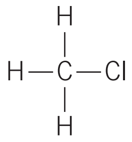
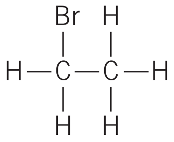
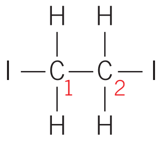
Table 7C–1 Structures and names of selected haloalkanes

Name	Structural formula	Semi-structural formula	Functional group present
Fluoromethane		CH_3F	Fluoro-
Chloromethane		CH_3Cl	Chloro-
Bromoethane		CH_2BrCH_3	Bromo-
1,2-diiodoethane		CH_2ICH_2I	Iodo-

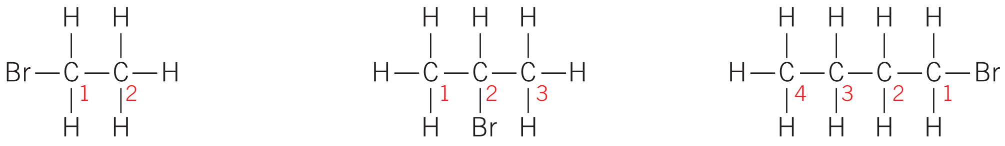
Nomenclature

In the IUPAC system, halogens appear in the name as prefixes. The systematic name also includes the location of the halogen atom on the molecule using a number separated by a hyphen. The following rules are intended as a continuation of the nomenclature rules introduced in Chapter 7B.

LINK **7B** HYDROCARBONS

Rule 1: The halogen atom must be part of the parent hydrocarbon.

The parent hydrocarbon is numbered so the halogen atom is present on the lowest numbered carbon atom.



bromoethane

There no need to include a number in this case as the bromo group is always bonded to carbon 1.

2-bromopropane

The position of the bromo group needs to be included in the name for molecules with three or more carbons.

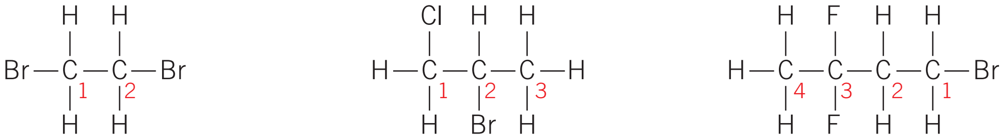
1-bromobutane

An incorrect name for this molecule is 4-bromobutane because it does not give the bromo group the lowest possible number.

Figure 7C–1 Demonstration of Rule 1

Rule 2: If there is more than one halogen atom, their names appear sequentially in alphabetical order.

This rule is similar to how alkyl side chains are also ordered alphabetically in the name. Again, prefixes, such as 'di-', 'tri-' and 'tetra-', are ignored for the purposes of alphabetical ordering. The examples in Figure 7C–2 demonstrate this rule.



1,2-dibromoethane

Numbers are required for the name of this molecule to specify the positions of the bromo groups.

2-bromo-1-chloropropane

An incorrect name for this molecule is 1-chloro-2-bromopropane because the name is not arranged alphabetically.

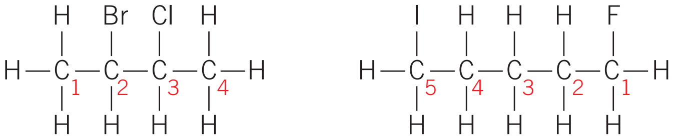
1-bromo-3,3-difluorobutane

An incorrect name for this molecule is 2,2-fluoro-4-bromobutane because it is not arranged alphabetically and does not have the lowest possible set of numbers.

Figure 7C–2 Demonstration of Rule 2

Rule 3: If the same numbers are obtained from both directions, the first group to appear in the name receives the lowest number.

As you can see in Figure 7C–3, the same set of numbers are obtained when numbering the haloalkanes from either end of the molecule (for example, 2-bromo-3-chlorobutane versus 3-bromo-2-chlorobutane). In this situation, give the lower number to the halogen that appears first in the name alphabetically – making 2-bromo-3-chlorobutane the correct name.



2-bromo-3-chlorobutane
(not 3-bromo-2-chlorobutane)

1-fluoro-5-iodopentane
(not 5-fluoro-1-iodopentane)

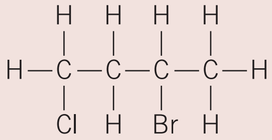
Figure 7C–3 Demonstration of Rule 3

Notice from Table 7C–1 that names such as fluoromethane and chloromethane do not require numbers to specify the location of the halogen atom. This is because numbers are not used if they are redundant in the name. However, for more complex haloalkanes, the name becomes ambiguous if numbers are not included. For example, the name dibromoethane is ambiguous because it could represent 1,1-dibromoethane or 1,2-dibromoethane – two different molecules. A good rule of thumb is that a systematic name must only specify one molecule. If you can draw more than one structure from a systematic name, then that name is either incomplete or incorrect.

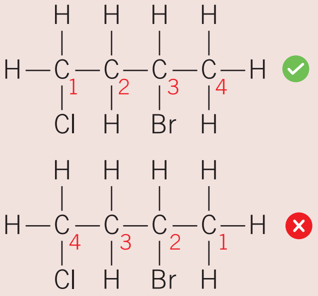
Worked example 7C–1: Naming haloalkanes

VIDEO
WORKED
EXAMPLE 7C–1

Give the systematic name of the following molecule.



Solution

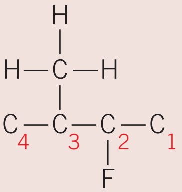
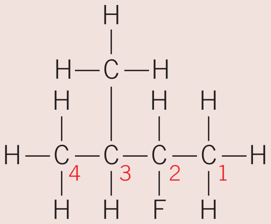
Step 1	Identify the longest continuous carbon chain (parent hydrocarbon).	The longest chain is four carbons, and there are only single bonds between the carbon atoms. Therefore, the parent hydrocarbon is **butane**.
Step 2	Identify the functional groups.	The molecule has a bromo and a chloro group.
Step 3	Consider both ways of numbering the parent hydrocarbon. Decide on the direction that gives the lowest set of numbers.	 1,3 < 2,4 so the first way of numbering is correct.
Step 4	Put it all together. Recall that the prefixes need to be ordered alphabetically.	3-bromo-1-chlorobutane

Worked example 7C–2: Drawing structural formulas of haloalkanes

VIDEO
WORKED
EXAMPLE 7C–2

Draw the structural formula of 2-fluoro-3-methylbutane.

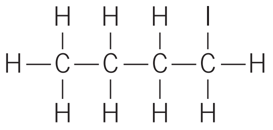
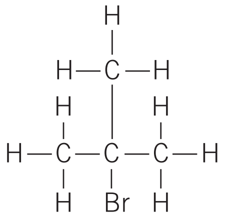
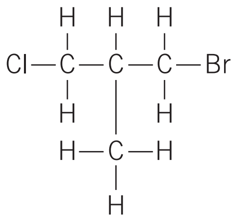
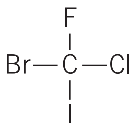
Solution

Step 1	Identify the parent hydrocarbon.	The parent hydrocarbon is butane. $C-C-C-C$
Step 2	Number the parent hydrocarbon.	$C_4-C_3-C_2-C_1$
Step 3	Draw the side chains and functional groups in their corresponding positions.	
Step 4	Complete the structure by adding the rest of the hydrogen atoms.	

Check-in questions – Set 1

WORKSHEET 7C–1 NAMING AND DRAWING HALOALKANES

1 Give the systematic name of the following molecules.

a 	**b** 
c 	**d** 

2 Give the semi-structural formula of the following molecules.

- a 2-bromobutane
- b 1,3-dichloropropane
- c 2-bromo-1-chloro-4-fluorobutane

3 The following names are incorrect. Draw the structure they represent and provide the correct IUPAC name.

- a 3-fluorobutane
- b 3-chloro-methylpropane

Physical properties of haloalkanes

Haloalkanes are usually polar molecules. This is because halogens such as F, Cl and Br are more electronegative than C, making the carbon–halogen (C–X) bond polar. Thus, the C–X bond in a haloalkane would, in most cases, break the symmetrical geometry of the bonds, creating a permanent dipole or a polar molecule. The exception is molecules like CF_4, where symmetry is restored by substituting all hydrogens for the same halogen.

Both dispersion forces and dipole–dipole forces should be considered when analysing the interactions between haloalkane molecules. Dispersion forces exist, without exception, between all molecules (including haloalkanes). As discussed in Chapter 2, the strength of dispersion forces increases as the number of electrons within the molecule increases. Polar haloalkane molecules also experience dipole–dipole forces (a stronger force than dispersion forces) between their molecules. In general, due to its stronger intermolecular forces, a haloalkane has a higher boiling and melting point than the corresponding alkane of the same carbon-chain length.

Boiling point

Figure 7C–4 shows the boiling point of methane compared to various halomethane molecules. The presence of a halogen atom increases the boiling point as the molecules become polar. The boiling point is also seen to increase as the size of the halogen atom increases. This is due to larger halogen atoms having more electrons, resulting in stronger dispersion forces. As you will recall from Chapter 2C, dispersion forces are always present between all covalent molecules.

2C PROPERTIES OF COVALENT COMPOUNDS DETERMINED BY INTERMOLECULAR FORCES LINK

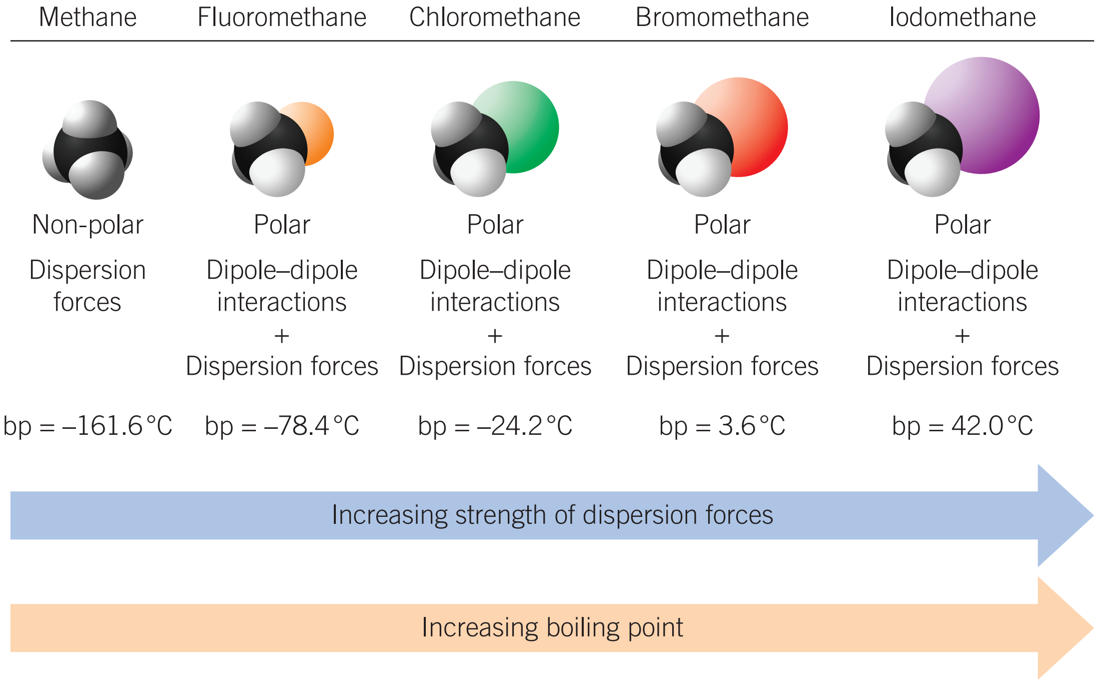
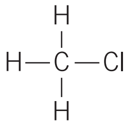
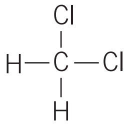
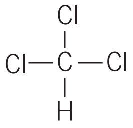
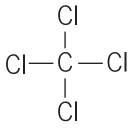


Figure 7C–4 The size of the halogen atom affects the boiling point of the molecule.

Table 7C–2 compares the boiling points of mono-, di-, tri- and tetrachloromethane. The boiling point increases as more halogen atoms are introduced into the molecule. This is again due to increasing dispersion forces as each chlorine atom contains a relatively large number of electrons, compared to hydrogen, therefore increasing the strength of the overall dispersion forces between the molecules. It is noteworthy that tetrachloromethane is non-polar and therefore dipole–dipole interactions do not exist between its molecules. However, the extra electrons from the four chlorine atoms compensate for this, causing an increase in the strength of the molecule's dispersion forces.

Table 7C–2 Comparing the polarity and boiling point of chlorinated methane molecules

Molecule and name	Structural formula	Polarity	Boiling point (°C)
CH_3Cl Chloromethane	 $\mathrm{H-\overset{H}{\underset{H}{C}}-Cl}$	Polar	–24.2
CH_2Cl_2 Dichloromethane	 $\mathrm{H-\overset{Cl}{\underset{H}{C}}-Cl}$	Polar	39.6
$CHCl_3$ Trichloromethane	 $\mathrm{Cl-\overset{Cl}{\underset{H}{C}}-Cl}$	Polar	61.2
CCl_4 Tetrachloromethane	 $\mathrm{Cl-\overset{Cl}{\underset{Cl}{C}}-Cl}$	Non-polar	76.7

Solubility

2C PROPERTIES OF COVALENT COMPOUNDS DETERMINED BY INTERMOLECULAR FORCES LINK

Haloalkanes have polar characteristics that allow some to be soluble in water. Fluoroalkanes are the most soluble in water. Solubility decreases as the size of the haloalkane increases (Figure 7C–5). This is because as the halogen atom gets larger, it becomes less electronegative. This means that the overall molecule becomes less polar and thus less soluble in water. Solubility also decreases as the size of the carbon chain increases. This is because the carbon chain is non-polar, and as it gets larger, the molecule becomes more non-polar overall.

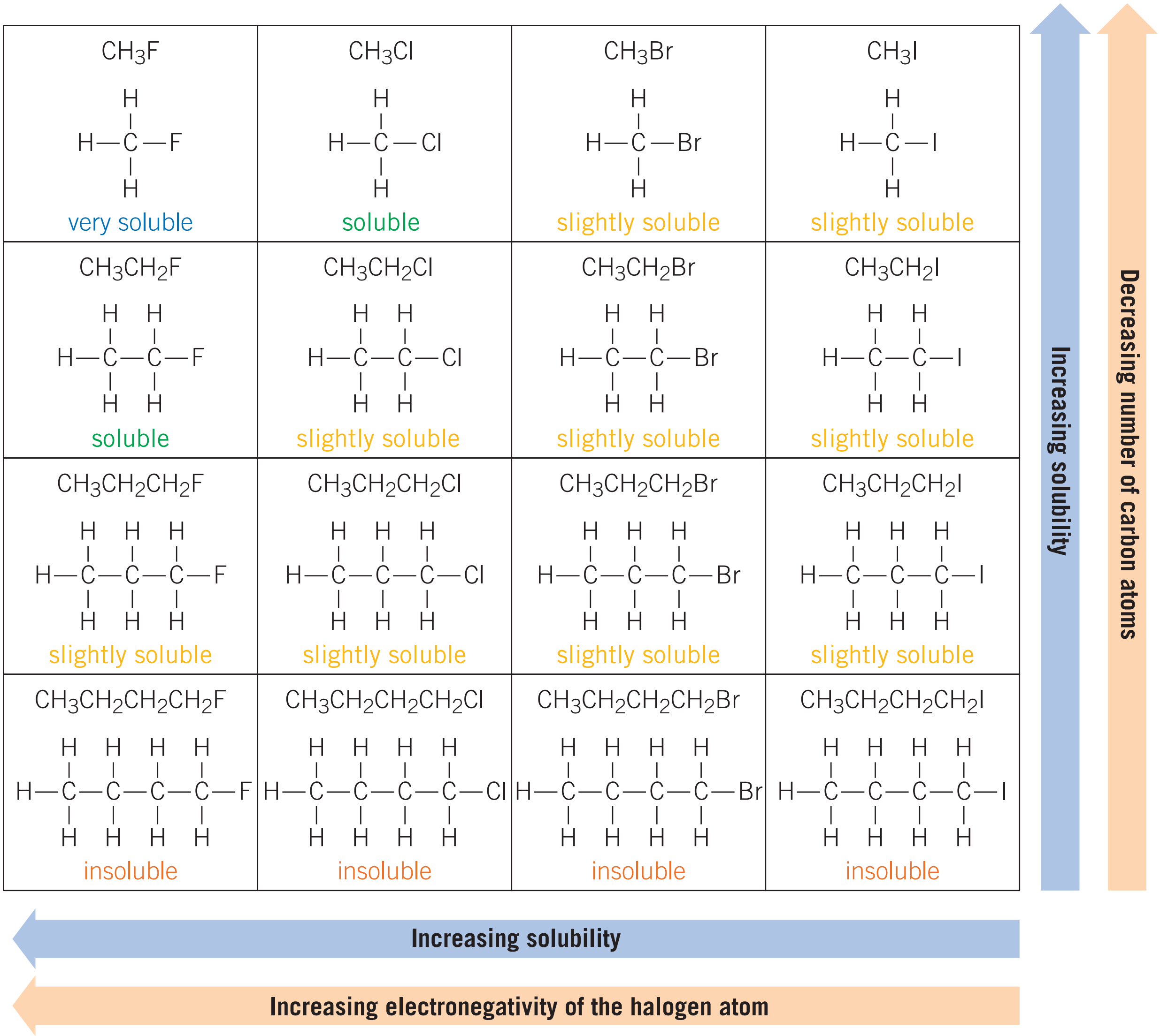


Figure 7C–5 Solubility of haloalkanes in water

Check-in questions – Set 2

1 Organise the following list by ascending boiling point: 1,2-dibromopropane, bromomethane, 1,2-dibromoethane, dibromomethane, fluoromethane

2 Explain why iodoethane has a lower boiling point than 1-iodopropane.

Chemical properties and uses of haloalkanes

One of the reactions that haloalkanes can participate in is called a substitution reaction. During this reaction, the halogen atom is substituted by another atom or group of atoms. Of particular interest in this course is the reaction between haloalkanes and aqueous sodium hydroxide or potassium hydroxide. During this reaction, the halogen atom on the organic molecule is substituted for a hydroxyl group (–OH). The product is an organic compound that has a hydroxyl functional group. This family of organic compounds, the alcohols, will be discussed in detail in Chapter 7D. The equations in Figure 7C–6 illustrate two examples of substitution reactions involving haloalkanes and metal hydroxides. The relevant parts of the molecules, where the substitution occurs, are outlined in red.

LINK **7D** ALCOHOLS AND CARBOXYLIC ACIDS

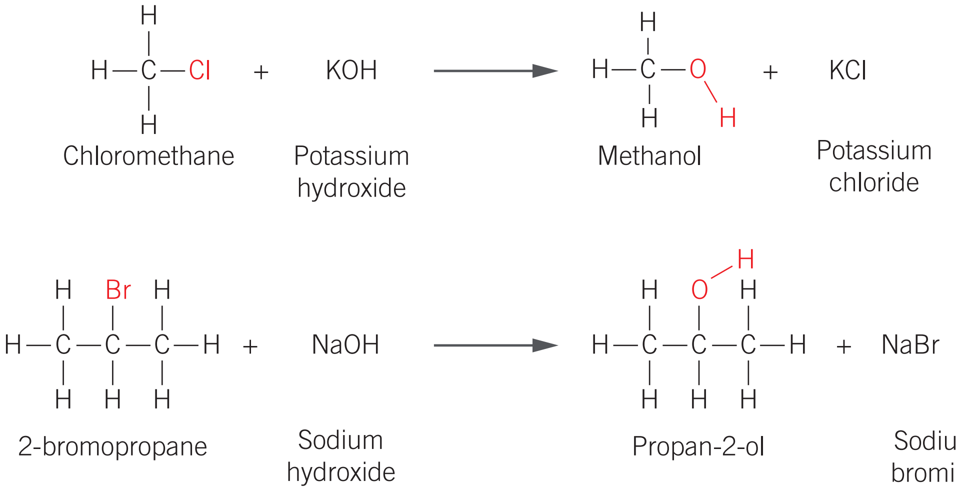


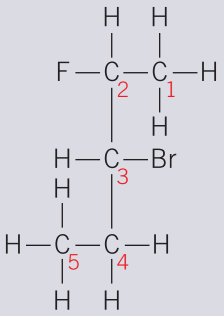
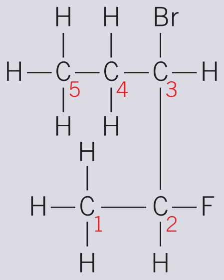
Figure 7C–6 Examples of substitution reactions involving haloalkanes

In industry, haloalkanes are important solvents and raw materials. For example, dichloromethane (CH_2Cl_2) is commonly used as an organic solvent, degreasing agent or paint remover, because it has the ability to dissolve a wide range of polar and non-polar substances. Trichloromethane ($CHCl_3$), or chloroform, is also a widely used solvent as well as being a powerful anaesthetic and sedative. Chlorofluorocarbons (CFCs) are small chlorine and fluorine-containing haloalkanes that were traditionally used as propellants for spray cans and refrigerants in appliances such as air conditioners and refrigerators. CFCs were subsequently found to be harmful to the ozone layer. Modern refrigerants are haloalkanes that only contain fluorine rather than both chlorine and fluorine.

7C SKILLS

Isomers or identical molecules?

The diverse ways to represent molecules means that it is often difficult to identify whether two molecules are isomers of each other or identical to each other. Consider the two haloalkanes on the right:

One way to determine if they are isomers or identical molecules is to number the carbons and name the molecules. If they have the same molecular formula and different IUPAC names, they are isomers. If they have the same name, they are identical compounds. In the above case, both molecules have the name 3-bromo-2-fluoropentane. They are identical and therefore, not isomers.

Using the correct terminology
The correct use of terminology when responding to questions regarding intermolecular forces is very important. You must be able to demonstrate your understanding of the idea *intermolecular* without any ambiguity. Fundamentally, intermolecular forces relate to the interaction(s) that exist *between* two particles. These particles could be two atoms, two molecules, two ions or any combinations thereof. Therefore, the word 'between' should feature in your responses when referring to intermolecular bonding.

An example of a common mistake is to write that 'water has stronger intermolecular forces than carbon dioxide'. The diagram below shows how to improve this statement.

Needs improvement	Better	Best
Water has stronger intermolecular forces than carbon dioxide.	The intermolecular forces between water molecules are stronger than the intermolecular forces between carbon dioxide molecules.	The hydrogen bonding between water molecules is stronger than the dispersion forces between carbon dioxide molecules.
This statement does not properly demonstrate that intermolecular forces is a property between molecules.	This adequately demonstrates the understanding that intermolecular forces act between molecules rather than something that belongs to a molecule.	This statement is an improvement since it gives the specific types of intermolecular forces present.

Section 7C questions

1 Write the systematic names of the following haloalkanes.
 - a CH_2FCH_3
 - b $CHBr_2CH_2CH_3$
 - c $CH_3CH_2CHClCH_3$
 - d $CH_3CI_2CH_2CH_2I$
 - e $CH_2ClCH_2CH_2Br$

2 Complete the following table by providing the correct structural and semi-structural formula.

Systematic name	Structural formula	Semi-structural formula
2-bromopropane		
1,2-dichlorobutane		
1-bromo-2-fluoroethane		
Tetrabromomethane		

3 Draw and name all four structural isomers of $C_3H_6Cl_2$. Provide your answers using structural formula.

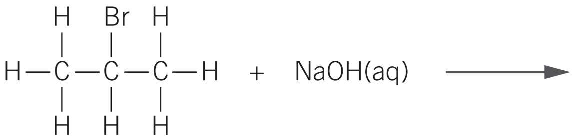
4 Explain why the addition of HCl to but-1-ene will give a mixture of two isomeric products whereas the addition of HCl to but-2-ene will only produce one product.

5 Which two of the following molecules would unlikely be separated using fractional distillation? Explain your answer.

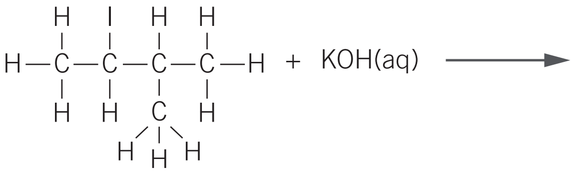
A butane
B 2-bromobutane
C 1-bromo-2-methylpropane
D 2-fluorobutane

6 Draw the structural formula of all products and by-products in the following reactions.

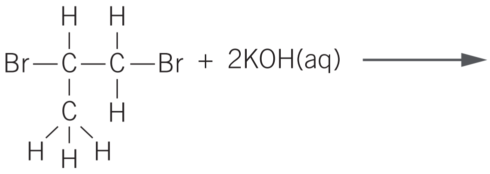
a



b



c



7D Alcohols and carboxylic acids

Study Design:

- 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)

Glossary:

Alcohol
Alkanol
Carboxyl group
Carboxylic acid
Dimer
Hydroxyl group
Molecular weight
Oxidant
Oxidation reaction
Positional isomer

ENGAGE

Wine or vinegar?

Have you ever heard someone say that the 'wine has gone bad'? This person is either pretending to be a connoisseur of wine or acting as an astute chemical detective using their senses of taste and smell to distinguish between two families of molecules introduced in this section. Most wine typically contains 6–16% ethanol, a molecule belonging to the alcohol family. When the bottle is opened and exposed to oxygen from the atmosphere, the ethanol content in wine begins to oxidise to ethanoic acid, a molecule commonly associated with vinegar. The formation of ethanoic acid in wine gives the beverage an unpleasant sour taste.

Alcohol a family of organic compounds characterised by a hydroxyl functional group

Hydroxyl group a functional group consisting of –OH

Alkanol an alkane in which a hydroxyl group (–OH) has replaced a hydrogen atom

EXPLAIN

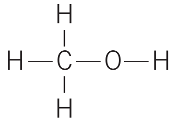
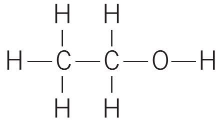
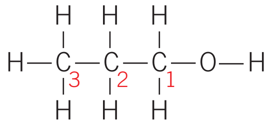
The hydroxyl functional group

VIDEO 7D–1
ALCOHOLS

Alcohols are a family of organic compounds characterised by the presence of a **hydroxyl group** (–OH). In this course we will primarily explore a specific class of alcohol known as an **alkanol**. In an alkanol, the hydroxyl group is bonded to a saturated parent alkane. Table 7D–1 shows the structural and semi-structural formula of the first three members of the alkanol homologous series.

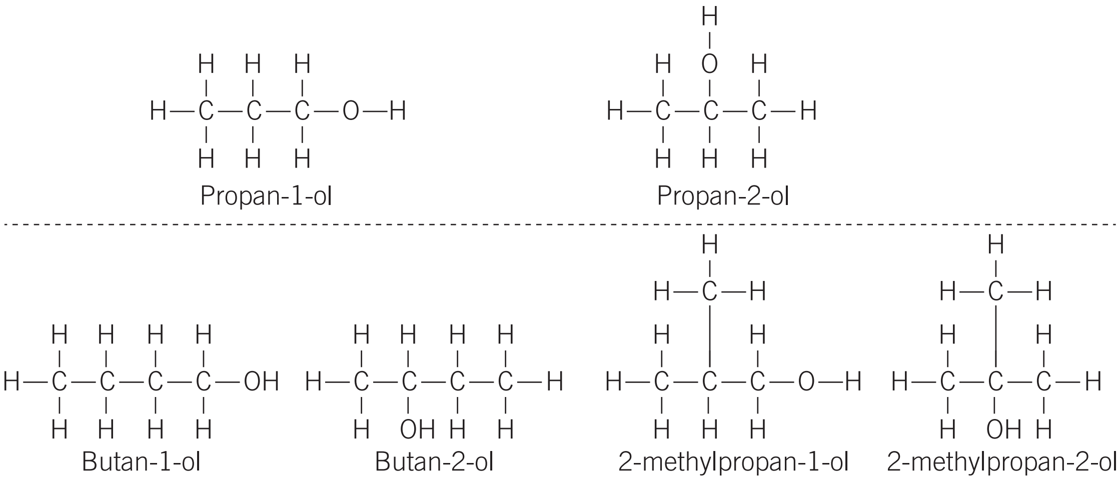
ISBN 978-1-009-22964-7 Drummond et al

Table 7D–1 Structures of the first three alkanols

Name	Molecular formula	Structural formula	Semi-structural formula
Methanol	CH_4O	 $H-\underset{H}{\overset{H}{C}}-O-H$	CH_3OH
Ethanol	C_2H_6O	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-O-H$	CH_3CH_2OH
Propan-1-ol	C_3H_8O	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-O-H$	$CH_3CH_2CH_2OH$

The first alkanol to have structural isomers is C_3H_8O. If you consider the structure of propan-1-ol in the table above, you may notice that it is also possible for the hydroxyl group to be bonded to the second carbon in the parent hydrocarbon. This produces propan-2-ol as shown in Figure 7D–1. Propan-1-ol and propan-2-ol are known as **positional isomers** because their structures differ only by the location of the functional group. Alcohols larger than propanol will have more isomers. For example, $C_4H_{10}O$ has four isomers. Butan-1-ol and butan-2-ol are positional isomers.

Positional isomer a molecule with the same molecular formula and parent hydrocarbon but different positions of functional groups



Propan-1-ol Propan-2-ol

Butan-1-ol Butan-2-ol 2-methylpropan-1-ol 2-methylpropan-2-ol

Figure 7D–1 Structural isomers of C_3H_8O (top) and $C_4H_{10}O$ (bottom)

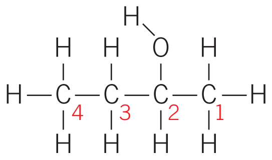
Nomenclature

The systematic names of alkanols end in '-ol' in place of '-e' at the end of the parent alkane name. All the previous rules regarding the naming of alkanes apply with the addition of the following rules:

Rule 1: For alcohols larger than ethanol, the position of the hydroxyl group on the parent hydrocarbon must be included in the name.

LINK 7B HYDROCARBONS

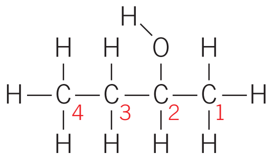
This is done by adding the number corresponding to the position of the hydroxyl (–OH) group before the '-ol'. Hyphens are used before and after the number. Figure 7D–2 contrasts an incorrect example where the position of the hydroxyl group is not included in the name (left) versus the correct version (right). The position of the hydroxyl group is not needed for methanol and ethanol because it is assumed that the functional group is on carbon 1.



butanol

✗

Incorrect
The name does not give any information on the location of the hydroxyl group.



butan-2-ol

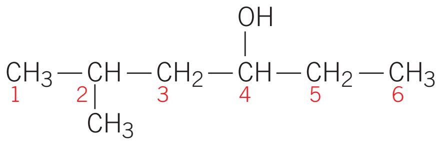
✓

Correct
The name indicates that the hydroxyl group is attached to the second carbon in the parent hydrocarbon.

Figure 7D–2 The position of the hydroxyl group should be included in the name for alcohols larger than ethanol.

Rule 2: The hydroxyl group is given the lowest possible number for the carbon it is bonded to in the parent hydrocarbon.

In other words, the molecule should be numbered starting from the end closest to the hydroxyl group. Figure 7D–3 shows two ways to name the same molecule except only the version on the right is correct. It has given the hydroxyl group the lowest number.

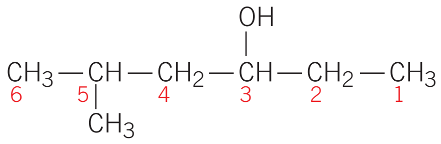


$CH_3-CH(CH_3)-CH_2-CH(OH)-CH_2-CH_3$

2-methylhexan-4-ol

✗

Incorrect
The molecule was not numbered starting from the end closest to the hydroxyl group.



$CH_3-CH(CH_3)-CH_2-CH(OH)-CH_2-CH_3$

5-methylhexan-3-ol

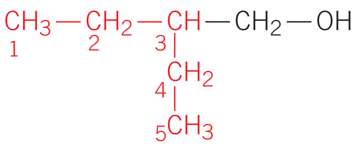
✓

Correct
The molecule was numbered starting from the end closest to the hydroxyl group giving the -OH group the lowest number.

Figure 7D–3 The molecule should be numbered from the end closest to the hydroxyl group.

Rule 3: The parent hydrocarbon is the longest carbon chain containing the hydroxyl group.

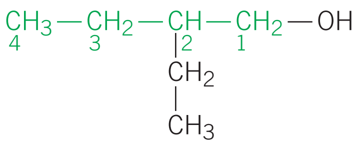
The molecule shown in Figure 7D–4 is an example where the longest carbon chain is not necessarily the parent hydrocarbon. Since the parent hydrocarbon **must** include the hydroxyl group, the parent hydrocarbon is four carbon atoms instead of five carbon atoms.



$CH_3-CH_2-CH(CH_2CH_3)-CH_2-OH$

✗

Incorrect
The parent hydrocarbon does not contain the hydroxyl group.



$CH_3-CH_2-CH(CH_2CH_3)-CH_2-OH$

2-ethylbutan-1-ol

✓

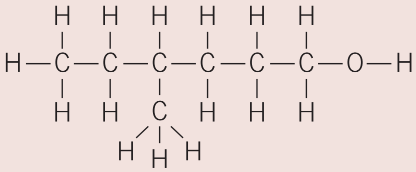
Correct
The parent hydrocarbon does contain the hydroxyl group.

Figure 7D–4 The parent hydrocarbon must include the hydroxyl group.

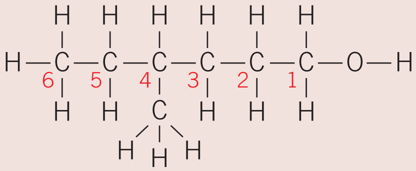
Worked example 7D–1: Naming alcohols

VIDEO
WORKED
EXAMPLE 7D–1

Name the following molecule.



Solution

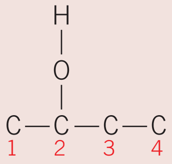
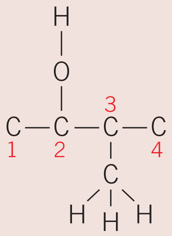
Step 1	Identify the longest carbon chain in the molecule.	There are six carbons in the longest chain. Therefore, the parent hydrocarbon is based on hexane.
Step 2	Number the parent hydrocarbon so that the hydroxyl group has the lowest possible number.	The molecule should be numbered from right to left. 
Step 3	Identify the location of the hydroxyl group.	The hydroxyl group is bonded to carbon 1. This means the name will end in '-1-ol'.
Step 4	Identify any side chains in the molecule.	There is a methyl (CH_3) group bonded to carbon 4. This means the name will start with '4-methyl'.
Step 5	Put it all together.	Name of the molecule: 4-methylhexan-1-ol

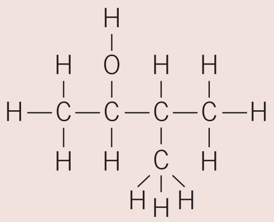
Worked example 7D–2: Drawing structural formulas of alcohols

VIDEO
WORKED
EXAMPLE 7D–2

Draw the structural formula of 3-methylbutan-2-ol.

Solution

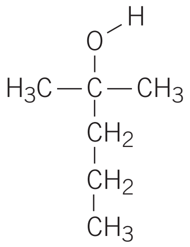
Step 1	Determine the parent hydrocarbon.	The parent hydrocarbon is a 'butan', meaning that it consists of four carbon atoms. C—C—C—C
Step 2	Number the parent hydrocarbon and draw the hydroxyl functional group.	
Step 3	Draw the alkyl group(s).	

Step 4	Draw the rest of the hydrogens.	
Step 5	Check the structure.	When the drawing is complete, check that each carbon has exactly four bonds surrounding it.

Check-in questions – Set 1

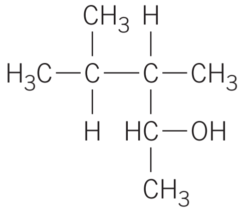
1 Write the systematic names of the following alkyl alcohols.

a



b $CH_3CHOHCH(CH_3)CH_3$

c



2 Draw the structural formula for each of the following alcohols.

a hexan-1-ol
b hexan-2-ol
c 3-ethylpentan-2-ol

3 Write the semi-structural formula for each of the following alcohols.

a 3-methylhexan-1-ol
b 3-methylbutan-2-ol

4 The name 2-ethylbutan-3-ol has broken more than one IUPAC systematic naming rule. Draw the structure it is attempting to represent and give the correct name of the molecule.

Reactions involving alcohols

Combustion reactions

Just like alkanes and alkenes, alcohols readily undergo combustion reactions. Ethanol is a particularly important example of this because it is widely used as a fuel source. The combustion of ethanol is used to power camping stoves and our car's internal combustion engines, as depicted in Figure 7D–5. The E10 fuel sold at petrol stations contains 10% ethanol sourced mostly from renewable sources. In a complete combustion reaction, ethanol reacts with oxygen to produce CO_2 and H_2O.

7A ORGANIC CHEMISTRY IN SOCIETY **LINK**

7B HYDROCARBONS **LINK**

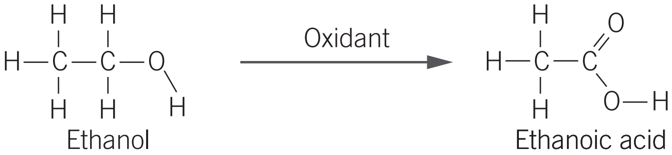
$$CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$

Figure 7D–5 Ethanol can be used as a fuel source in camp stoves and internal combustion engine cars.

Oxidation reactions

Alkan-1-ols can undergo **oxidation reactions** to give carboxylic acids. We will explore this family of organic compounds below and on the following pages. Oxidation of alkanols can occur naturally as seen in the example of wine. It can also happen in a laboratory or industry setting under a more controlled environment. Chemists can employ specific reagents called **oxidants** to allow the oxidation to occur. The oxidation process will be explored in more detail in Chapter 12 in Unit 2, and also in Unit 4 of this course. The equation shown below is an example of ethanol oxidising to ethanoic acid.

LINK **12A** REDUCTION AND OXIDATION **UNIT 4**



The carboxyl functional group

Carboxylic acids are a homologous series characterised by the **carboxyl group** (–COOH). The carboxyl group consists of an –OH group bonded to a carbonyl group (C═O). The general formula for carboxylic acids is shown in Figure 7D–6.

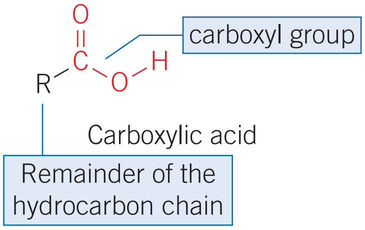


Figure 7D–6 General structure of a carboxylic acid

The structural and semi-structural formulas of the first three members of the carboxylic acid homologous series are shown in Table 7D–2. It is important to note that the carbon atom of the carbonyl group is included as part of the parent hydrocarbon. The semi-structural formula shows the carboxyl group as –COOH.

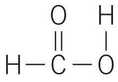
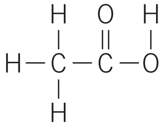
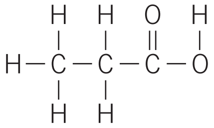
Oxidation reaction a reaction that occurs when there is an increase in the number of C–O bonds or a decrease in the number of C–H bonds

Oxidant a reagent that causes an oxidation reaction to occur

Carboxylic acid a family of organic molecules characterised by the presence of a carboxyl group

Carboxyl group a functional group consisting of –COOH

Table 7D–2 Structures of the first three carboxylic acids

Name	Structural formula	Semi-structural formula
Methanoic acid or Formic acid (common name)		HCOOH
Ethanoic acid or Acetic acid (common name)		CH_3COOH
Propanoic acid		CH_3CH_2COOH

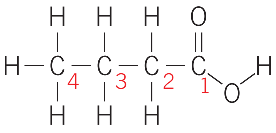
Nomenclature

VIDEO 7D–2
NOMENCLATURE OF ORGANIC COMPOUNDS

Names of carboxylic acids end in '-oic acid' instead of '-e' in the alkane. For example, a carboxylic acid with a parent hydrocarbon of three carbon atoms is named propan**oic acid**. The naming rules that apply to alkanes also apply to carboxylic acids. In addition, the following rules apply specifically to carboxylic acids.

Rule 1: The carbon in the carboxyl group must be numbered 1 in the parent hydrocarbon.

Because the carboxyl group is always at the start of the parent hydrocarbon, there is no need to include a '-1-' in the name.

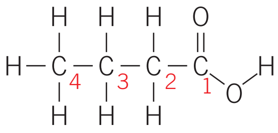


Butanoic acid

✓

Correct

The molecule is correctly numbered and named.

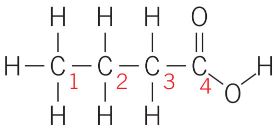


Butan-1-oic acid

✗

Incorrect

The name should not include '–1–' as it is redundant.



Butan-4-oic acid

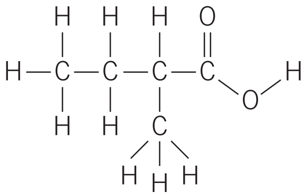
✗

Incorrect

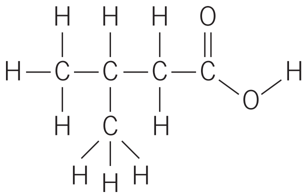
The carboxyl group is always the first carbon.

Rule 2: Branched isomers of carboxylic acids can also exist.

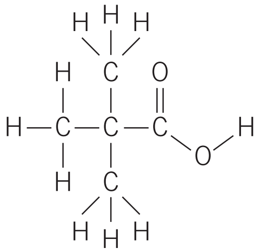
When naming these molecules, the position and type of branch is included before the parent hydrocarbon. The parent hydrocarbon is always numbered starting from the carbon atom in the carboxyl group. If more than one alkyl branch is present, they are ordered alphabetically. The following examples show the correct names of three branched carboxylic acids.



2-methylbutanoic acid



3-methylbutanoic acid

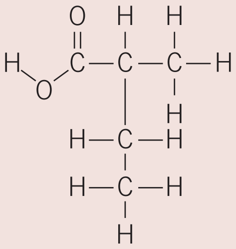


2,2-dimethylpropanoic acid

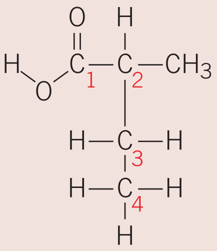
Worked example 7D–3: Naming carboxylic acids from structural formulas

VIDEO WORKED EXAMPLE 7D–3

Write the systematic name for the following molecule:



Solution

Step 1	Identify the longest carbon chain in the molecule.	The longest carbon chain is four carbons long. Therefore, the molecule has the parent name ‘butanoic acid’.
Step 2	Number the parent hydrocarbon starting from the carboxyl functional group.	
Step 3	Identify any side chains in the molecule.	There is a methyl (CH_3) group bonded to carbon 2. This means that the name will begin with ‘2-methyl’.
Step 4	Put it all together.	Name of the molecule: 2-methylbutanoic acid

Worked example 7D–4: Naming carboxylic acids from semi-structural formulas

VIDEO WORKED EXAMPLE 7D–4

Write the systematic name for the following molecule:

$$CH_3CH_2CH(CH_3)CH_2COOH$$

Solution

Step 1	Identify the longest carbon chain in the molecule.	The longest chain is five carbons. This means that the name of the parent hydrocarbon is ‘pentanoic acid’.
Step 2	Number the parent hydrocarbon starting at the carboxyl group. Recall that alkyl groups in brackets represent side chains.	$\underset{5}{CH_3}\underset{4}{CH_2}\underset{3}{CH}(CH_3)\underset{2}{CH_2}\underset{1}{COOH}$
Step 3	Identify any side groups.	There is a methyl group bonded to carbon 3 (as indicated by the CH_3 in brackets). Therefore, the name of the molecule will begin with ‘3-methyl’.
Step 4	Put it all together.	Name of the molecule: 3-methylpentanoic acid

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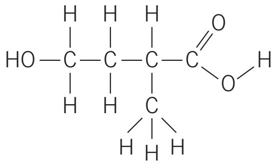
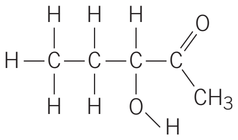
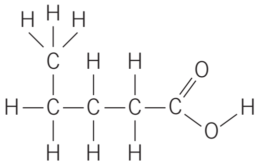
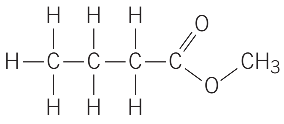
Check-in questions – Set 2

1 Refer to the following molecules to answer the following questions.


Molecule A


Molecule B

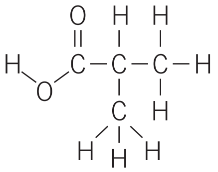

Molecule C


Molecule D

- a Which one or more of the molecules shown above is a carboxylic acid?
- b State the molecular formulas of **A–D** above and hence determine the relationship between the structures.

2 Name the following carboxylic acids.

- a $CH_3CH_2CH_2CH_2COOH$
- b



- c $HOOCCH_2CH(CH_3)CH_2CH(CH_3)CH_3$

3 Give the semi-structural formula of the following molecules.

- a Pentanoic acid
- b 3-ethylpentanoic acid
- c 2,3-dimethylbutanoic acid
- d 2-propylpentanoic acid

WORKSHEET 7D–1 NAMING AND DRAWING ALCOHOLS AND CARBOXYLIC ACIDS

4 Draw and name two isomers that have the molecular formula $C_4H_8O_2$ and belong to the carboxylic acid family.

Chemical properties and uses of carboxylic acids

CHAPTER 11 LINK

Carboxylic acids are weak acids. The concept of acids will be discussed in more detail in Chapter 11. Many carboxylic acids occur in nature, while others are produced on an industrial scale. A notable example of a common carboxylic acid is methanoic acid, which can be found in species of ants. When threatened, the ants spray the acid at their foes as a defence mechanism. Contact with methanoic acid may cause irritation or mild chemical burns. Another example of a common carboxylic acid is ethanoic acid. This acid is found in vinegar and is responsible for the sour taste of the condiment. As demonstrated earlier, ethanoic acid can be produced by the oxidation of ethanol. The vinegar that we use is also produced by this route, as shown in Figure 7D–7. Various fruits or grains are fermented by yeast to give ethanol, which is subsequently oxidised by bacteria to obtain ethanoic acid. Most vinegars are 5–8% ethanoic acid by volume.

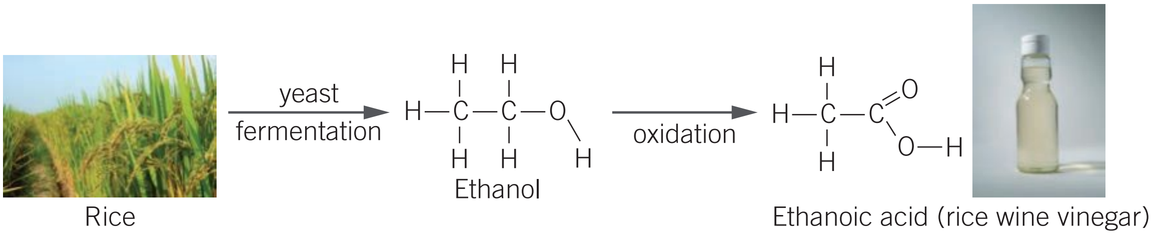


Figure 7D–7 Production of rice wine vinegar from rice

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Physical properties of alcohols and carboxylic acids

As mentioned earlier, a similarity between the carboxyl group and the hydroxyl group is that they both contain an O–H bond. Since oxygen atoms are significantly more electronegative than hydrogen atoms, the polar O–H bond allows both families of organic compounds to form hydrogen bonding between their molecules. This means that alcohols and carboxylic acids have strong intermolecular forces compared to some of the other families of organic compounds explored in this chapter. Figure 7D–8 demonstrates the hydrogen bonding between two molecules of ethanoic acid and between two molecules of ethanol.

LINK **1C** THE PERIODIC TABLE

LINK **7B** HYDROCARBONS

LINK **7C** HALOALKANES

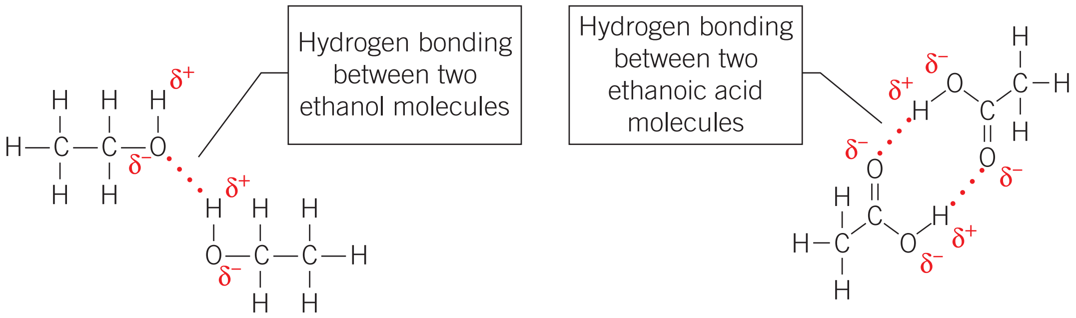


Figure 7D–8 Hydrogen bonding between alcohols (left) and carboxylic acids (right)

You may notice from Figure 7D–8 that two hydrogen bonds exist between the carboxylic acid molecules, whereas only one is shown between the alcohols. This is because the C═O bond in the carboxyl functional group is also polar and therefore is a second site that can participate in hydrogen bonding. This arrangement allows two carboxylic acids to form **dimers**. As a result, the intermolecular forces between carboxylic acids are stronger than alcohols of comparable **molecular weight**.

Dimer a complex where two molecules are linked together

Molecular weight the sum of the atomic masses of the atoms as given in the molecular formula

The previously discussed relationship between a compound's boiling point and the strength of its intermolecular forces applies here as well. More specifically, the greater the strength of a compound's intermolecular forces, the larger the thermal energy required to disrupt these bonds and therefore the higher the compound's boiling point. In this regard, when comparing compounds of similar molar masses, carboxylic acids would have a higher boiling point than alcohols. This information is illustrated in Figure 7D–9, which compares the boiling points of these two families of compounds.

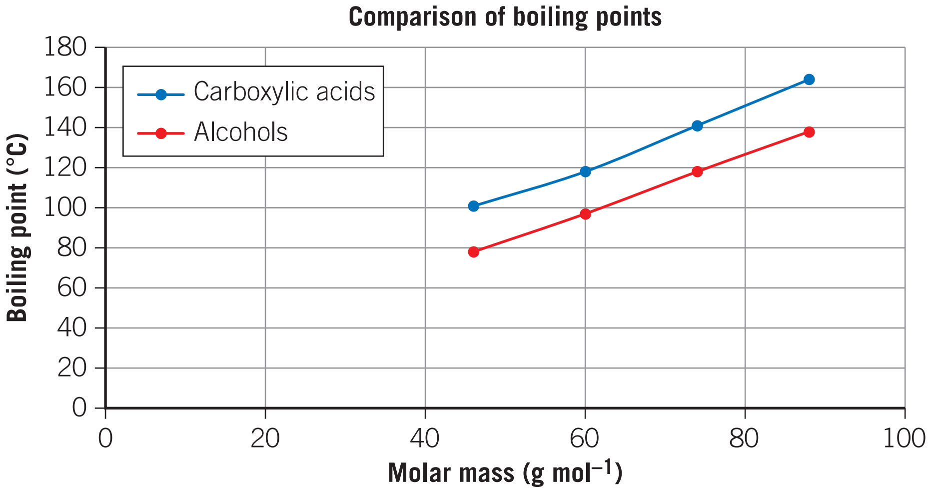
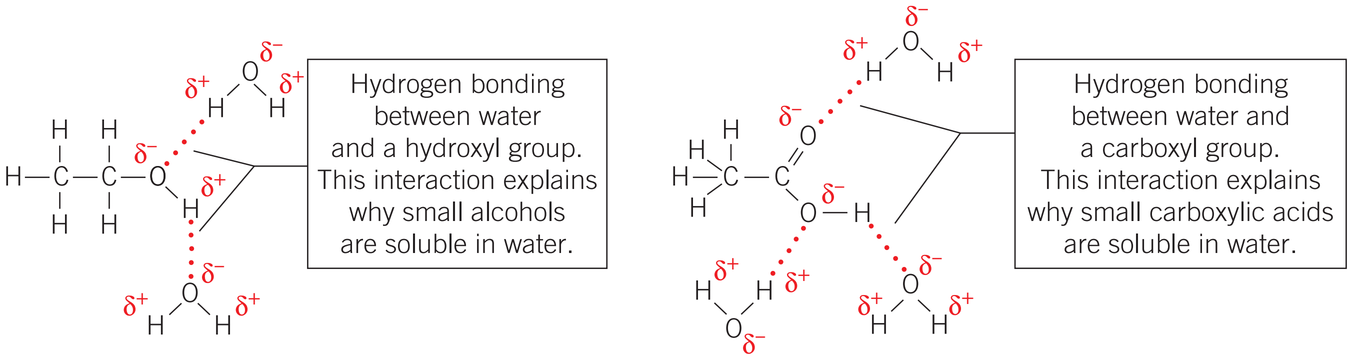


Figure 7D–9 Graphical data comparing the boiling points of alcohols to those of carboxylic acids

The hydroxyl group (–OH) and the carboxyl group (–COOH) can both interact with water via hydrogen bonding, as shown in Figure 7D–10. This means that small alcohols and carboxylic acids are soluble in water. Vinegar is an example of ethanoic acid dissolved in water, and alcoholic beverages are examples of ethanol dissolved in water. For both alcohols and carboxylic acids, as the size of the alkyl chain increases, the molecule becomes overall increasingly non-polar. This is because the alkyl chain itself is non-polar. As a result, the molecule becomes less soluble in water.

WORKSHEET 7D–2
COMPARING PROPERTIES OF DIFFERENT HYDROCARBONS



Hydrogen bonding between water and a hydroxyl group. This interaction explains why small alcohols are soluble in water.

Hydrogen bonding between water and a carboxyl group. This interaction explains why small carboxylic acids are soluble in water.

Figure 7D–10 Carboxylic acids and alcohols can form hydrogen bonding with water. This interaction allows small carboxylic acids and alcohols to be soluble in water.

VIDEO 7D–3
SKILLS: COMPARING THE PROPERTIES OF HYDROCARBONS WITH DIFFERENT FUNCTIONAL GROUPS

7D SKILLS

Comparing the properties of hydrocarbons with differing functional groups

Having reached the completion of organic chemistry in this chapter, it is important that you can recall the different types of functional groups and the types of intermolecular bonding that exists between these organic molecules. Let's first begin with a summary of the different families of molecules covered:

Type of organic compound	Polar or non-polar	Intermolecular force(s) present between molecules	Soluble in water
Alkanes	Non-polar	Dispersion forces	No
Alkenes	Non-polar	Dispersion forces	No
Haloalkanes	Polar (mostly)	Dipole–dipole interactions Dispersion forces	Partially
Alcohols	Polar	Hydrogen bonding Dispersion forces	Yes
Carboxylic acids	Polar	Hydrogen bonding Dispersion forces	Yes

You will notice from the table that in general, when a molecule is non-polar, it is not soluble in water, and when a molecule is polar, it is soluble in water. This is because, as you will recall from Chapters 2, 4 and 5, water is a bent polar molecule. Another way to remember this is early in this chapter, we referred to the 'like dissolves like' rule.

Another property not presented in the table above is boiling point. Usually when comparing the boiling points of different hydrocarbons with varying functional groups we focus on hydrocarbons with a comparable number of carbon atoms. In order to then explain the differences in boiling point between different organic molecules, we need to discuss differences in their intermolecular bonding between molecules. An example question and solution are provided below. Note the level of detail in the answer provided. You should never assume that your teacher or assessor knows what you are talking about if you cannot go into detail. It is clearly outlined here what each type of molecule is, whether they are polar or non-polar and the types of intermolecular forces present.

Question:

Four molecules, all containing the same number of carbon atoms, are listed below:

- Propane
- 2-chloropropane
- Propan-2-ol
- Propanoic acid

Order these molecules from highest to lowest boiling point. Explain the reasoning behind your ordering.

Solution:

Propanoic acid > propan-2-ol > chloropropane > propane

Propane is a member of the alkane family of hydrocarbons containing only C—C single bonds. It has the lowest boiling point as it is a non-polar hydrocarbon. As such, it only has dispersion forces present between its molecules.

2-chloropropane is a haloalkane due to the presence of a chlorine atom attached to the middle carbon in the parent hydrocarbon. This makes the molecule polar and as such, in addition to dispersion forces, there are also dipole–dipole interactions between these molecules. Dipole–dipole interactions are a stronger intermolecular force compared to dispersion forces.

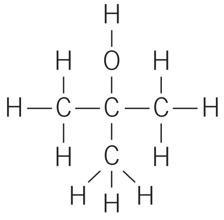
Both propan-2-ol (an alcohol) and propanoic acid (a carboxylic acid) are polar and while also having dispersion forces present between molecules, importantly, there is also hydrogen bonding. This is due to both molecules containing a hydroxyl (–OH) group. However, propanoic acid has a higher boiling point due to its ability to form a dimer. This is when two propanoic acid molecules are connected due to the electrostatic force of attraction of two hydrogen bonds.

Hydrogen bonding is a stronger intermolecular force than both dipole–dipole interactions and dispersion forces.

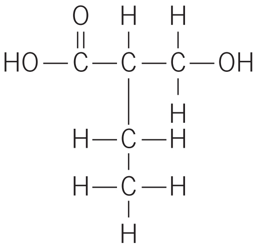
Section 7D questions

1 Circle and label the functional groups present in the following molecules.

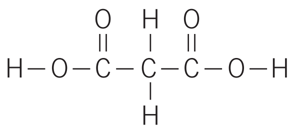
a



b

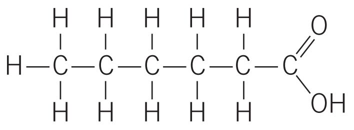


2 Write the molecular formula of the following compounds.
- a $CH_3CH_2CH_2COOH$
- b propan-2-ol
- c



$$\mathrm{H{-}O{-}\overset{\displaystyle O}{\overset{\|}{C}}{-}\underset{\underset{\displaystyle H}{|}}{\overset{\overset{\displaystyle H}{|}}{C}}{-}\overset{\displaystyle O}{\overset{\|}{C}}{-}O{-}H}$$

3 Complete the following table of alcohols and carboxylic acids.

Name	Molecular formula	Structural formula	Semi-structural formula
Butan-1-ol			
Methylpropan-2-ol			
			$CH_3CH(OH)CH_3$
			$CH_3CH_2CH_2CH_2CH_2$ $CH(OH)CH_2CH_3$
Propanoic acid			
			

4 Name and draw the structural formula of all four isomers with the molecular formula $C_4H_{10}O$.

5 Haloalcohols, or halohydrins, are a class of compounds containing both halogen and alcohol functional groups. Draw the structural formula for the haloalcohol, 1,3-dichloropropan-2-ol.

6 Write the IUPAC names of the following compounds.
- a $CH_3CHOHCH_2CH_2CH_3$
- b $CH_3CH_2CH(CH_2CH_3)CH_2COOH$
- c $HOOCCH_2CH_2CH_2CH(CH_3)CH(CH_3)CH_2CH_3$
- d $CH_3CH_2CHOHCH_2CH(CH_3)CH_2CH_3$

7 Draw a diagram showing the hydrogen bonding between two molecules of propanoic acid.

8 The table below shows the boiling points of three different organic molecules.

Molecule	Boiling point (°C)	Molar mass (g mol^{-1})
Pentan-1-ol	138	88
Hexan-1-ol	157	102
Pentanoic acid	186	102

- a Explain why hexan-1-ol has a higher boiling point than pentan-1-ol.
- b Explain why pentanoic acid has a higher boiling point than hexanol.

Chapter 7 review

Summary

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

Success criteria – I am now able to:		Linked question
7A.1	Recall that fossil fuels provide a non-renewable source of chemicals that can be used for energy generation and production of goods	8☐
7A.2	Explain how plant-based biomass can serve as an alternative to fossil fuels in the production of materials and goods	10☐, 24☐
7A.3	Explain why plant-based biomass is renewable, whereas fossil fuels are non-renewable	24☐
7A.4	Analyse the impact of materials and products made from organic chemistry on society and the environment	9☐, 24☐
7B.1	Define the characteristics of alkanes and alkenes	3☐, 11☐, 12☐
7B.2	Determine the molecular formula of alkanes and alkenes using their general formulas	1☐
7B.3	Represent linear and branched alkanes and alkenes using structural and semi-structural formulas	2☐, 16☐
7B.4	Write systematic names for linear and branched alkanes and alkenes (including isomers)	5☐, 20☐
7B.5	Connect physical properties to the intermolecular forces between molecules of hydrocarbons	6☐
7B.6	Describe the chemical properties and uses of alkanes and alkenes	4☐
7C.1	Define the characteristics of haloalkanes	14☐
7C.2	Represent haloalkanes using structural and semi-structural formulas	17☐
7C.3	Write systematic names for haloalkanes (including isomers)	7☐, 20☐
7C.4	Connect physical properties to the intermolecular forces between molecules of haloalkanes	6☐
7C.5	Describe the chemical properties and uses of haloalkanes	6☐
7D.1	Define the characteristics of alcohols and carboxylic acids	22☐
7D.2	Represent alcohols and carboxylic acids using structural and semi-structural formulas	17☐
7D.3	Write systematic names for alcohols and carboxylic acids (including isomers)	21☐, 22☐

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Success criteria – I am now able to:		Linked question
7D.4	Connect physical properties to the intermolecular forces between molecules of alcohols and carboxylic acids	18☐
7D.5	Describe the chemical properties and uses of alcohols and carboxylic acids	23☐

Multiple-choice questions

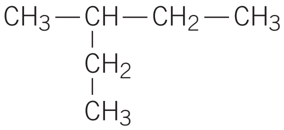
1 A straight-chain molecule with five carbon atoms and ten hydrogen atoms would be described as
- **A** an alkane with the prefix 'prop-'.
- **B** an unsaturated hydrocarbon with the prefix 'prop-'.
- **C** a saturated hydrocarbon with the prefix 'pent-'.
- **D** an alkene with the prefix 'pent-'.

2 The correct semi-structural formula of but-2-ene is
- **A** $CH_3CHCHCH_3$
- **B** $CH_2CHCH_2CH_3$
- **C** $CH_3CH_2CHCH_3$
- **D** $CH_3CH_2CH_2CH_3$

3 Which of the following lists contains compounds that all belong to the to the same homologous series?
- **A** $CH_3CH_2CH_3$, C_4H_8, C_2H_6
- **B** C_6H_{14}, C_6H_{12}, C_6H_6
- **C** C_4H_8, C_7H_{14}, CH_2CHCH_3
- **D** C_5H_{12}, $CH_2CHCH_2CH_2CH_2CH_3$, C_5H_{10}

4 Which of the following is likely to form during a substitution reaction involving butane?
- **A** $CH_3CH_2CH_2CH_3$
- **B** $CH_3CH_2CHBrCH_3$
- **C** $CH_3CH_2CHBrCH_2Br$
- **D** $CH_3CH{=}CHCH_3$

5 What is the systematic name for the molecule shown below?



$$\begin{array}{l} CH_3-CH-CH_2-CH_3 \\ \qquad\quad | \\ \qquad\;\; CH_2 \\ \qquad\quad | \\ \qquad\;\; CH_3 \end{array}$$

- **A** 3-methylpentane
- **B** 2-ethylbutane
- **C** 3-ethylbutane
- **D** 3-methylhexane

6 Which of the following compounds has the highest boiling point?
- **A** butane
- **B** 1,1-diiodopropane
- **C** chloroethane
- **D** 1-bromopropane

7 Which one of the following obeys the IUPAC system of naming organic compounds?
- **A** 1-methylpropane
- **B** 2,2-diethylbutane
- **C** 1-chloroethane
- **D** 2,3-dibromobut-2-ene

8 Which of the following is most likely made from a petroleum source?
- **A** cardboard box
- **B** stainless steel fork
- **C** ceramic plate
- **D** plastic milk bottle

9 Which of the following statements regarding plant-based biomass is **not** true?
- **A** Plant-based biomass is a renewable resource.
- **B** Plant-based biomass is only obtained from farming.
- **C** Plant-based biomass can be carbon neutral.
- **D** Plant-based biomass can reduce our reliance on fossil fuels.

10 Which of the following is a true statement when comparing fossil fuels to plant-based biomass?
- **A** Formation of biomass involves living organisms, whereas formation of fossil fuels never involved any form of life.
- **B** Combustion of fossil fuels releases CO_2, whereas combustion of plant-based biomass does not release CO_2.
- **C** Biomass can be replenished but there are no fossil fuels being formed anymore.
- **D** Plant-based biomass and fossil fuels are both used as sources of energy and feedstock for industrial processes.

Short-answer questions

11 Define 'structural isomer'. (1 mark)

12 State the shortest saturated hydrocarbon to have isomers. (1 mark)

13 Write an isomer of $CH_3CH_2CH{=}CH_2$ using a semi-structural formula. (1 mark)

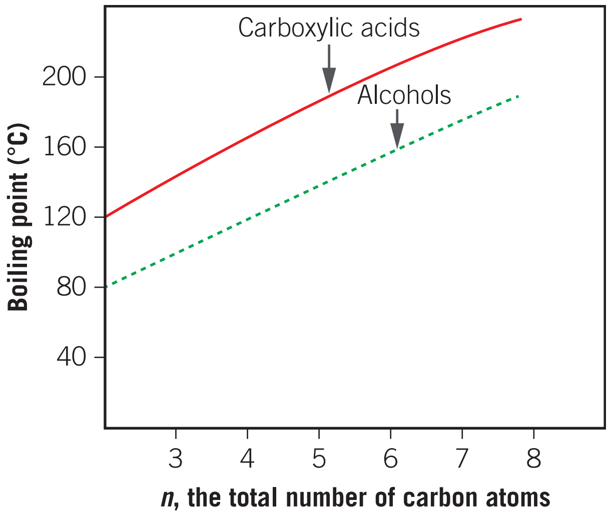
14 State the molecular formula of 2,2-dichloropentane. (1 mark)

15 Write a balanced chemical equation, using semi-structural formulas, for each of the following reactions.
- **a** Gaseous ethene is bubbled through a solution of bromine. (2 marks)
- **b** Gaseous propene is reacted with gaseous HCl. (2 marks)
- **c** A mixture of butane and Cl_2 gas is irradiated with UV light. (2 marks)

16 Butane has two structural isomers, whereas butene has three.
- **a** Draw the full structural formulas, showing all bonds, of the branched isomers of butene. (3 marks)
- **b** Give the systematic name of the three compounds drawn in part **a**. (3 marks)

17 The following IUPAC names are written incorrectly. Draw the structural formula they are attempting to represent then give the correct systematic name.
- **a** 2,2-dimethylethanol (2 marks)
- **b** 2-propylbut-3-ene (2 marks)
- **c** 1-chloro-1,1,2,2-tetramethylpropane (2 marks)

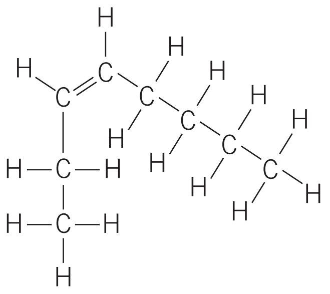
18 **a** According to the graph below, the boiling points of carboxylic acids are consistently higher than the boiling points of alcohols for molecules of similar size. Provide a detailed explanation for this observation. You may wish to include relevant diagrams in your response. (3 marks)



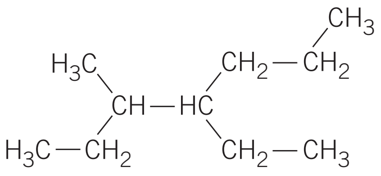
b Explain why the boiling points of alcohols and carboxylic acids increase as the number of carbon atoms increases. (2 marks)

19 Name the following hydrocarbons.

a (1 mark)



b (1 mark)

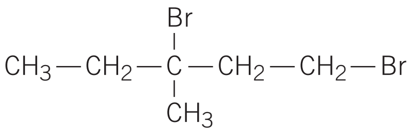


c $CH_3CH_2CHCHCH_2CH(CH_3)CH_3$ (1 mark)

20 Name the following haloalkanes.

a $CH_2ClCHClCH_3$ (1 mark)

b (1 mark)



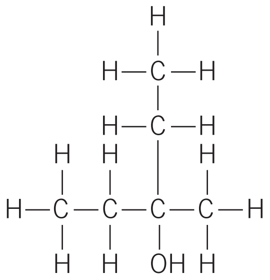
$$CH_3-CH_2-\underset{CH_3}{\overset{Br}{\underset{|}{\overset{|}{C}}}}-CH_2-CH_2-Br$$

c C_2H_5F (1 mark)

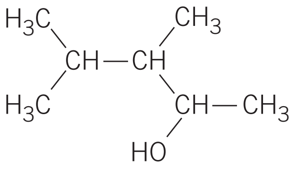
21 Name the following alcohols.

a $CH_3CH_2CH_2OH$ (1 mark)

b (1 mark)

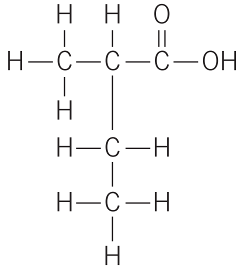


c (1 mark)

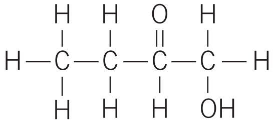


22 Circle and name only those that are carboxylic acids from the four molecules shown below.

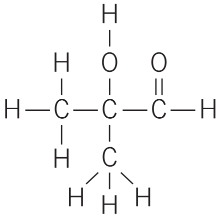
A



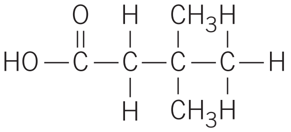
B



C



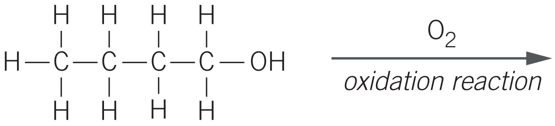
D



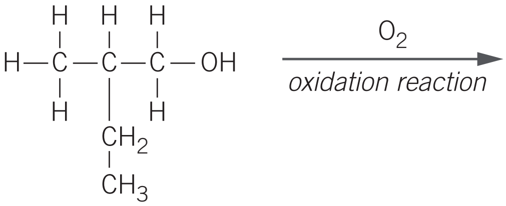
(2 marks)

23 Draw the products of the following oxidation reactions.

a (2 marks)



b (2 marks)



24 Traditional oil-based paints were made from volatile organic compounds derived from fossil fuels. The evaporation of the volatile components of paint allowed it to cure and dry. Scientists have developed a method to replace the volatile molecules in oil-based paints with a material derived from soybean oil and sugar. This new formulation has cut the number of volatile substances in paints by 50%.

a Discuss the potential hazard to the environment of using paint. (2 marks)
b Identify two advantages of using soybean oil as an alternative to fossil fuels. (2 marks)
c Identify two potential disadvantages of relying on soybean oil as a feedstock for paint. (2 marks)

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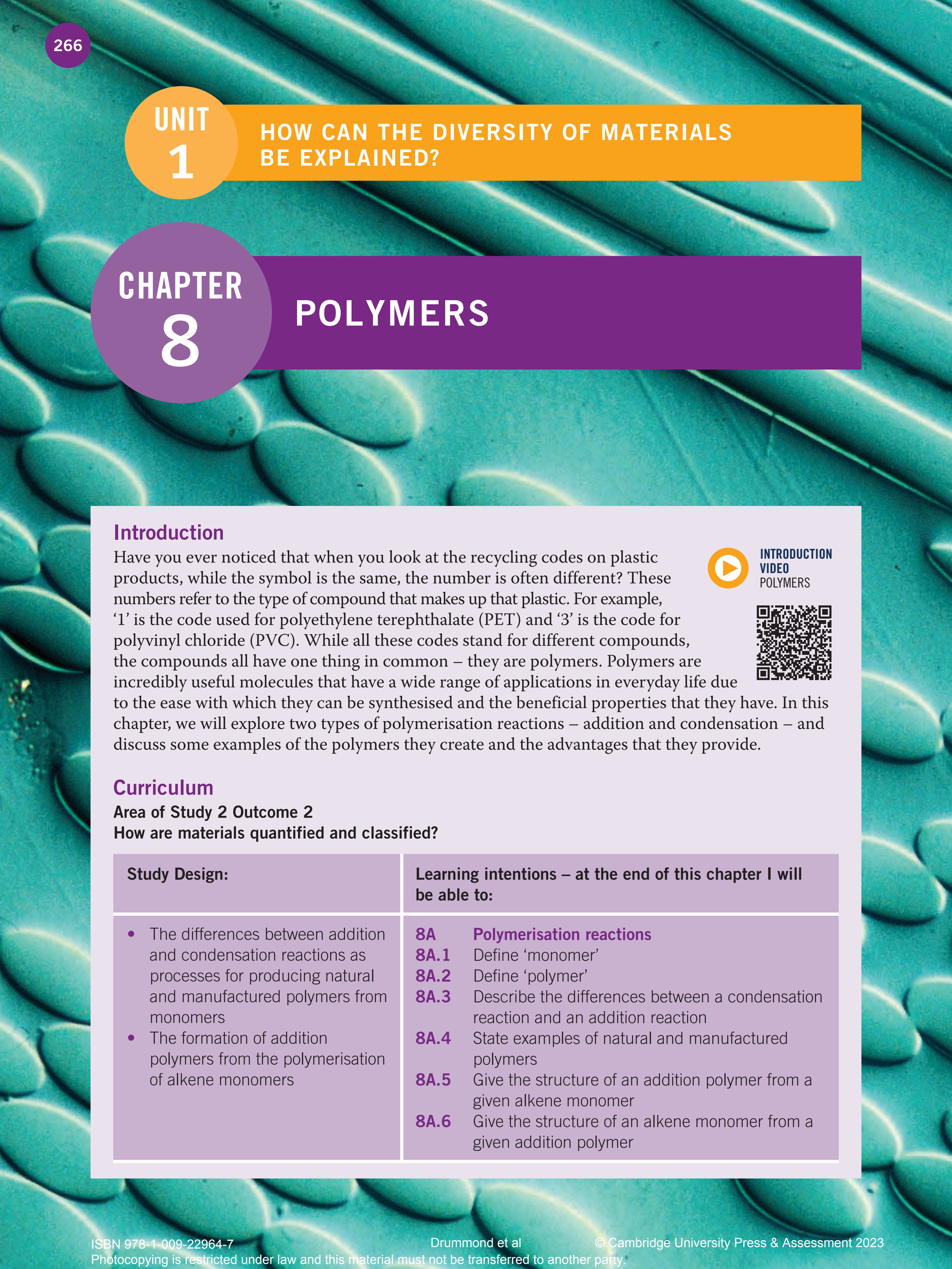


UNIT 1 HOW CAN THE DIVERSITY OF MATERIALS BE EXPLAINED?

CHAPTER 8 POLYMERS

Introduction

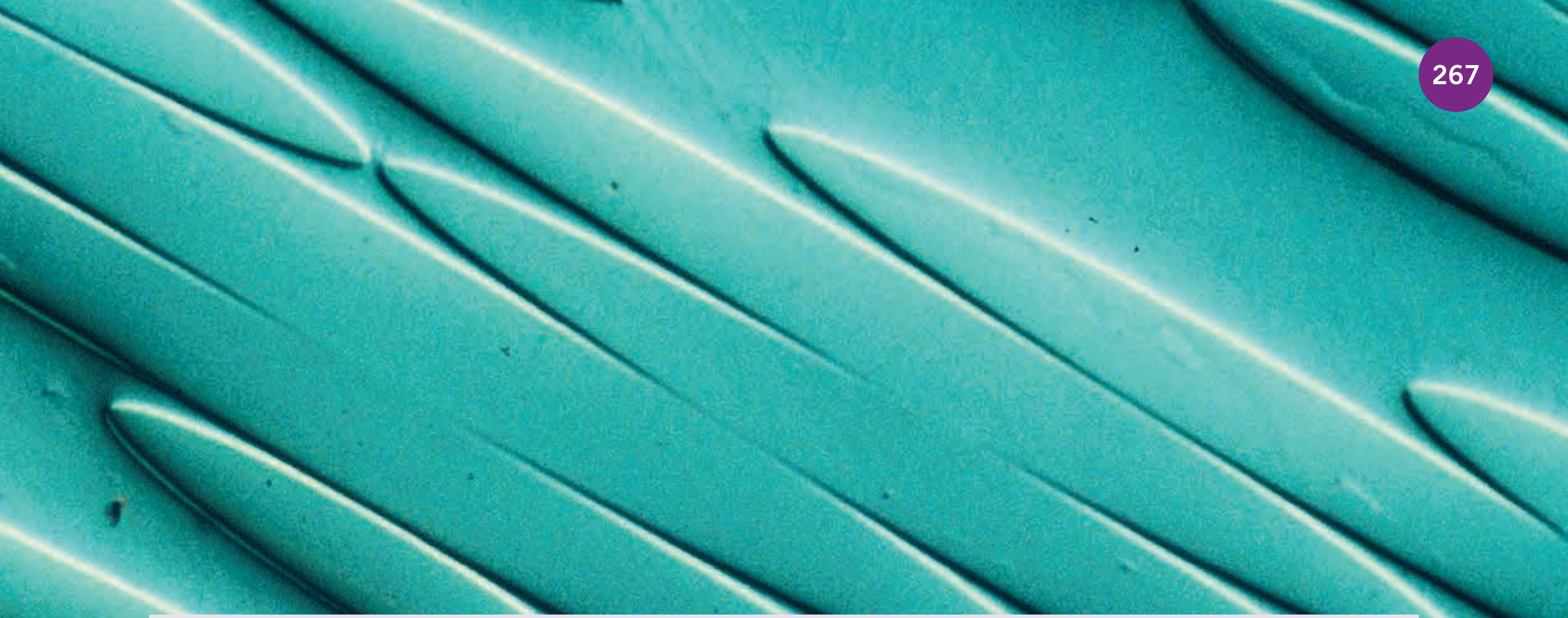
INTRODUCTION VIDEO
POLYMERS

Have you ever noticed that when you look at the recycling codes on plastic products, while the symbol is the same, the number is often different? These numbers refer to the type of compound that makes up that plastic. For example, ‘1’ is the code used for polyethylene terephthalate (PET) and ‘3’ is the code for polyvinyl chloride (PVC). While all these codes stand for different compounds, the compounds all have one thing in common – they are polymers. Polymers are incredibly useful molecules that have a wide range of applications in everyday life due to the ease with which they can be synthesised and the beneficial properties that they have. In this chapter, we will explore two types of polymerisation reactions – addition and condensation – and discuss some examples of the polymers they create and the advantages that they provide.

Curriculum

Area of Study 2 Outcome 2
How are materials quantified and classified?

Study Design:	Learning intentions – at the end of this chapter I will be able to:
• The differences between addition and condensation reactions as processes for producing natural and manufactured polymers from monomers • The formation of addition polymers from the polymerisation of alkene monomers	**8A Polymerisation reactions** **8A.1** Define ‘monomer’ **8A.2** Define ‘polymer’ **8A.3** Describe the differences between a condensation reaction and an addition reaction **8A.4** State examples of natural and manufactured polymers **8A.5** Give the structure of an addition polymer from a given alkene monomer **8A.6** Give the structure of an alkene monomer from a given addition polymer



Study Design:	Learning intentions – at the end of this chapter I will be able to:
- 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	**8A.7** Compare the properties of thermoplastic and thermosetting polymers **8A.8** Compare the structures of thermoplastic and thermosetting polymers **8A.9** Use the structure of thermoplastic and thermosetting polymers to explain their properties **8A.10** Describe how the structure and properties of a monomer determine the properties of its polymer **8A.11** Describe how chain length affects the properties of a polymer **8A.12** Describe how degree of branching affects the properties of a polymer
- The categorisations 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	**8B** **Polymer recycling** **8B.1** Define 'fossil fuel-based plastics' and 'bioplastics' **8B.2** Categorise a polymer as a fossil fuel-based plastic or a bioplastic, based on the origin of the monomer **8B.3** Compare mechanical, chemical and organic recycling of plastics **8B.4** List the requirements for a plastic to be considered compostable **8B.5** Describe how hydrolysis reactions are used in the recycling of polymers **8B.6** Compare the linear and circular economy models for polymers

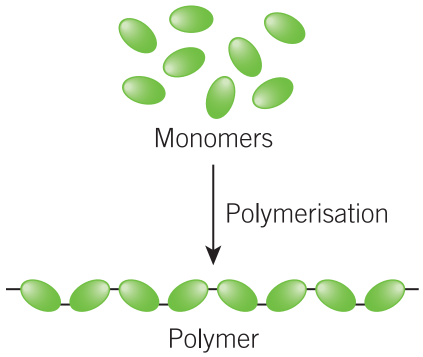
Glossary

Bioplastic	Monomer
Fossil fuel-based plastic	Polymer

Concept map

Formation of polymers from monomers by addition and condensation reactions, and their properties

8A Polymerisation reactions



Recycling and sustainability of plastics, including innovation in manufacturing and breakdown by hydrolysis

8B Polymer recycling

See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

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8A Polymerisation reactions

Study Design:

- The differences between addition and condensation reactions as processes for producing natural and manufactured polymers from monomers
- The formation of addition polymers from 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

Glossary:
Monomer
Polymer

ENGAGE

Polymers are all around and within us!

Did you know that most of the biomacromolecules in your body are all just made from joining together lots of smaller, repeating building blocks called monomers? DNA is a series of nucleotide monomers, proteins are strings of amino acids monomers and even carbohydrates are just created by linking together sugars. These are examples of biological polymers. In this section, you will learn more about the production of synthetic polymers and how they are used to produce the vast majority of plastic products that we use in our daily lives.

Figure 8A–1 DNA is an example of a naturally occurring polymer, but in this chapter, we are going to focus on synthetic polymers and plastics.

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EXPLAIN

Monomers and polymers

Polymer
a long-chain molecule formed by large numbers of repeating units (monomers)

Monomer
a small molecule that reacts with other monomer molecules to form long chains (polymers)

A **polymer** is any larger molecule that is made up of a series of smaller, repeating units. These units are known as **monomers**. We can visualise this by looking at Figure 8A–2.

There are two main types of reactions that can cause this polymerisation: addition polymerisation reactions and condensation polymerisation reactions. Perhaps not surprisingly, the products of these reactions are known as addition polymers or condensation polymers, respectively.

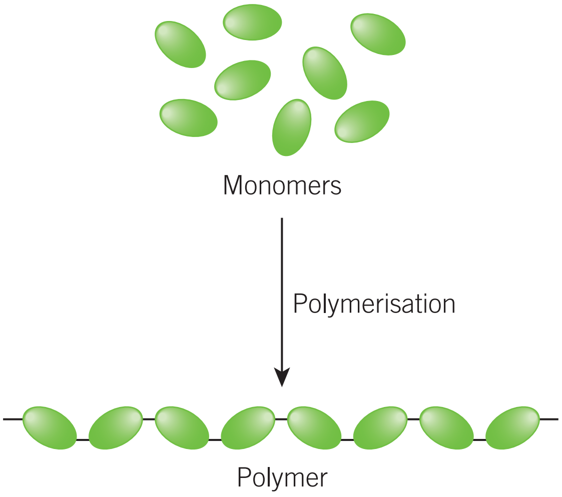


Figure 8A–2 A polymer is a series of repeating units, called monomers, linked together.

Addition polymerisation

7B HYDROCARBONS **LINK**

In an addition polymerisation reaction, the monomers are unsaturated, which means that they contain at least one carbon–carbon double or triple bond. In these reactions, a single product is produced – the polymer. An example of an addition polymerisation reaction is shown in Figure 8A–3, where many ethene monomers (three of which are shown in the figure) react to give polyethene (the polymer). Polyethene is also sometimes referred to as polyethylene, and we will discuss it in more detail later in this chapter.

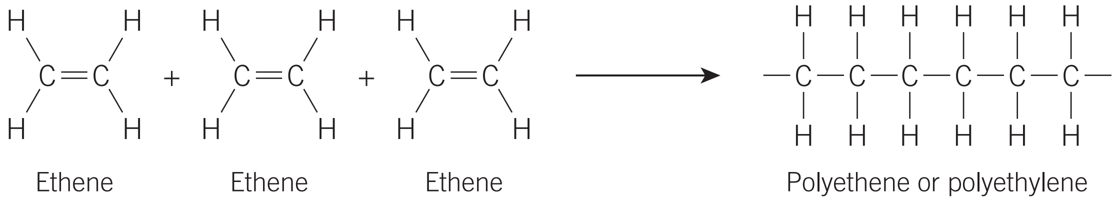


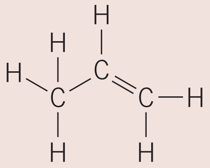
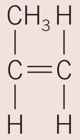
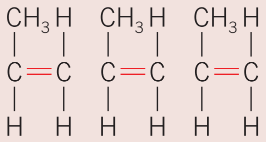
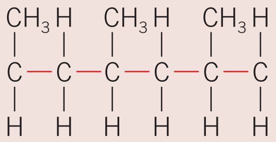
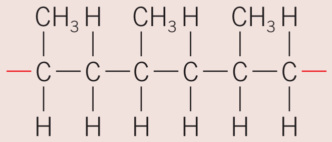
Figure 8A–3 The formation of the polymer polyethene from the monomer ethene is an example of addition polymerisation.

Worked example 8A–1: Drawing polymers

VIDEO
WORKED
EXAMPLE 8A–1

Draw a section of polypropene showing at least three units of the monomer.

Solution

Step 1	Determine the monomer. Polypropene is a polymer made up of repeating units of propene.	The structure of propene is 
Step 2	It is going to be easier to draw the polymer if we draw the monomer with the carbon–carbon double bond in a central position.	 $\begin{array}{ccc} CH_3 & & H \\ \vert & & \vert \\ C & = & C \\ \vert & & \vert \\ H & & H \end{array}$
Step 3	Draw three of the propene monomers in a row with the carbon–carbon double bonds lined up.	The carbon–carbon double bond will be erased in the next step, so it is suggested you draw it in pencil.  $\begin{array}{ccccccccc} CH_3 & & H & CH_3 & & H & CH_3 & & H \\ \vert & & \vert & \vert & & \vert & \vert & & \vert \\ C & = & C & C & = & C & C & = & C \\ \vert & & \vert & \vert & & \vert & \vert & & \vert \\ H & & H & H & & H & H & & H \end{array}$
Step 4	Erase the carbon–carbon double bonds and replace with a carbon–carbon single bond. Add a single bond from the carbon atom in one monomer to the carbon atom in the neighbouring monomer. The carbon atoms involved here are the carbon atoms that had the double bond in the monomer.	 $\begin{array}{ccccccccccc} CH_3 & & H & & CH_3 & & H & & CH_3 & & H \\ \vert & & \vert & & \vert & & \vert & & \vert & & \vert \\ C & - & C & - & C & - & C & - & C & - & C \\ \vert & & \vert & & \vert & & \vert & & \vert & & \vert \\ H & & H & & H & & H & & H & & H \end{array}$
Step 5	Complete the structure by adding open bonds to each end to show that the structure continues in both directions.	 $\begin{array}{ccccccccccccc} & CH_3 & & H & & CH_3 & & H & & CH_3 & & H & \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ - & C & - & C & - & C & - & C & - & C & - & C & - \\ & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & \\ & H & & H & & H & & H & & H & & H & \end{array}$

Condensation polymerisation

Condensation polymerisation does not require the monomers to have a carbon–carbon double or triple bond. However, it does require the monomers to have two functional groups. These two functional groups can be the same or they can be different. Condensation polymerisation reactions produce two products – the polymer and a second small molecule. This by-product is often water, but it is important to realise that other products are also possible, such as hydrogen chloride (HCl), methanol (CH_3OH) or acetic acid (CH_3COOH). An example of a condensation reaction is shown in Figure 8A–4.

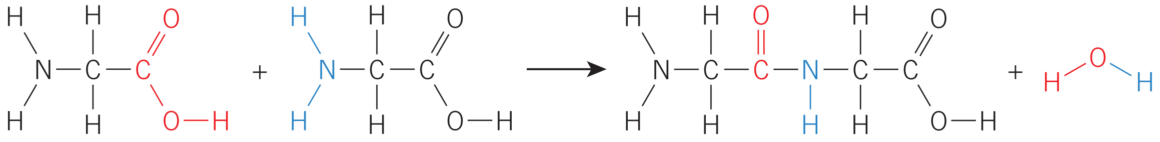


Figure 8A–4 The reaction of two glycine amino acids to form a dipeptide is an example of a condensation reaction. It can be seen that water is produced as a second product (by-product) in this reaction.

Check-in questions – Set 1

1 What is the structural feature that a molecule must have to be able to undergo addition polymerisation?

2 How many different products are produced in an addition polymerisation reaction? How many are produced in a condensation polymerisation reaction?

3 What is the general name given to the subunits that make up polymers?

7A ORGANIC CHEMISTRY IN SOCIETY LINK

WORKSHEET 8A–1 MONOMERS AND POLYMERS

Polymers can either be described as natural or manufactured. Natural polymers, as the name suggests, are naturally occurring polymers. Some examples of natural polymers are wool, silk, DNA, proteins, cellulose and cotton. On the other hand, manufactured polymers are polymers synthesised by humans, often from fossil fuel-derived monomers. These include nylons, polyesters, polyethene and polystyrene. Natural and manufactured polymers can be formed by both addition and condensation polymerisation reactions.

Thermoplastic and thermosetting polymers

VIDEO 8A–1 THERMOSETTING VERSUS THERMOPLASTIC

As well as natural and manufactured, polymers can also be classified as being either thermoplastic (also called thermosoftening) or thermosetting. The main difference between these two categories is what occurs when they are heated.

Thermoplastic polymers, such as polyethene and polyvinyl chloride (PVC), become soft when heated. This allows these polymers to be reshaped after heating. Thermosetting polymers, however, are irreversibly rigid and become hardened after heating. The reason for this is due to their chemical structure, and in particular, the forces that hold together the different polymer molecules.

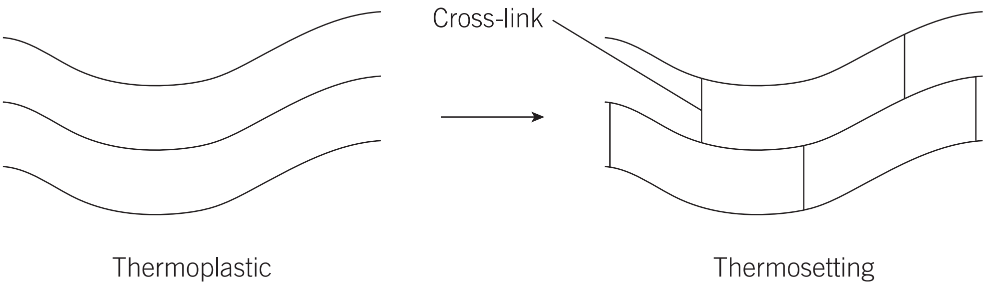


Figure 8A–5 Thermoplastic polymers (left) do not have any bonds between polymer chains, whereas thermosetting polymers (right) have several strong covalent bonds (cross-links) that act to link the different polymer chains together.

As you can see in Figure 8A–5, thermoplastics do not have any bonds that link together the different polymer molecules. Therefore, the structure of a thermoplastic is held together only by weak dispersion forces between the polymers. When heated, the polymer molecules will have increased kinetic energy, which allows them to break these dispersion forces and enable them to easily slide over each other without breaking the covalent bonds within the polymer itself. This explains the low melting point of most thermoplastics, and why they are soft, flexible and can be remoulded after heating.

Thermosetting polymers have extensive covalent bonds between polymer chains. These bonds are known as cross-links and are shown in Figure 8A–5. These cross-links strengthen the overall structure. The greater the level of cross-linking within the polymer, the harder it is. For this reason, thermosetting polymers are well-suited to high-temperature applications, as heating won't cause them to soften. In saying this, if thermosetting polymers are heated beyond a certain temperature, they will decompose, rather than be able to be remoulded. This is because at these very high temperatures the covalent bonds holding together the polymer will break. It is important to note that this will not only be the bonds that form the cross-links but also likely the bonds that are linking the monomers in the individual polymer chains.

Properties and uses of addition polymers

The degree of cross-linking is just one factor that can influence the properties of polymers. Other factors that also influence the properties of polymers include the:

- selection of monomer
- polymer chain length
- degree of branching.

Understanding the influence of these factors can help in selecting the appropriate polymer for a particular application or product.

Selection of monomer

Unsurprisingly, the monomer that makes up the polymer has a major influence on the properties of the polymer and the applications it can be used for. The structure of the monomer can also influence the chain length and the degree of branching that can occur, which is discussed in more detail below.

Chain length and degree of branching

The chain length is another factor that can affect the properties of a polymer. As a general rule, a polymer with a longer chain length will have a higher melting point and be stronger compared to one with a shorter chain length. As well as chain length, another factor that can affect the properties of a polymer is the degree of branching (Figure 8A–6).

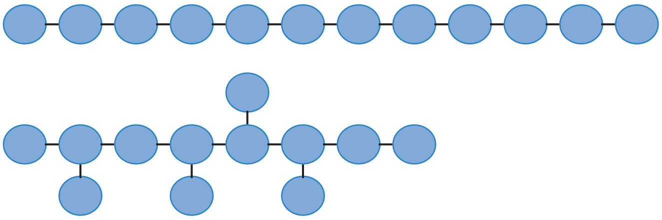


Figure 8A–6 Polymer chains can have no branching (top) or high levels of branching (bottom), which affects their properties.

The effect that the degree of branching has on the properties of the polymer is evident when comparing high density polyethene (HDPE) and low density polyethene (LDPE). As the name suggests, both of these polymers contain the same monomeric unit: ethene. However, as shown in Figure 8A–7, they are very different in terms of the amount of branching that they contain. HDPE contains very little branching, whereas LDPE contains a high degree of branching.

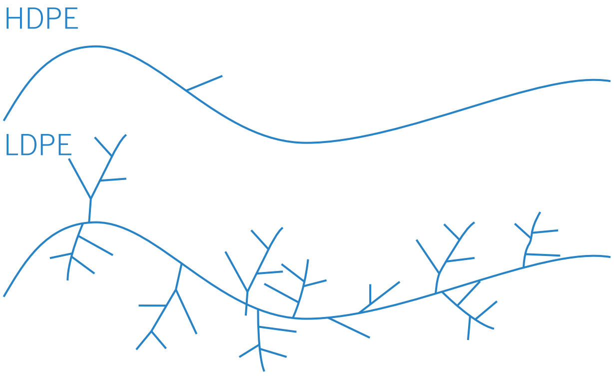


Figure 8A–7 HDPE contains a very low level of branching (top), whereas LDPE has a high degree of branching (bottom).

This branching is important because it impacts how closely the polymers can pack together. With not much branching, the polyethene molecules within HDPE are able to arrange themselves in close proximity to each other. The opposite is true for LDPE, where the high degree of branching prevents them from packing closely and instead there is separation of the polyethene polymers. When polymers are able to interact more closely with each other, like in HDPE, it leads to stronger dispersion forces. This helps explain why HDPE is much stronger and less flexible than LDPE. This variation in degree of branching, which leads to these different properties, is also reflected in the products that are manufactured from them. HDPE is used for making pipes, bins and plastic tables and chairs. LDPE, on the other hand, is used to make cling film and plastic bags (Figure 8A-8).

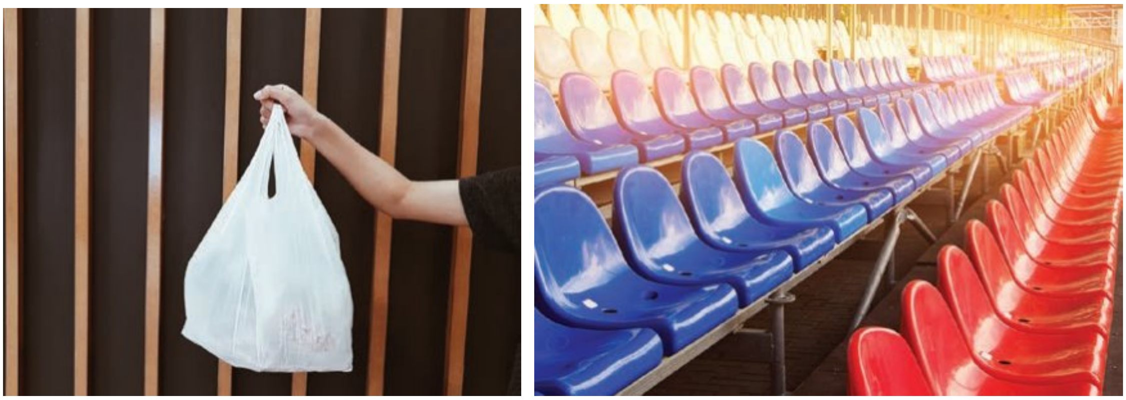


Figure 8A–8 Despite having the same chemical formula, LDPE and HDPE have very different properties and applications. LDPE is soft and flexible and is used for items such as plastic bags (left), whereas HDPE is strong and rigid and can be manufactured into products like plastic seats (right).

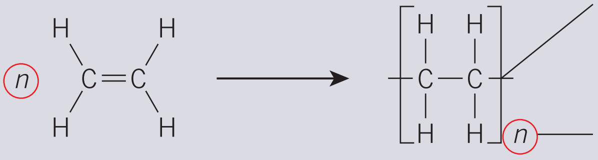
8A SKILLS

Drawing polymer structures

It is important to note that there are a couple of different ways in which you can draw polymer structures. In this skills section, we will go through the two different options.

Option 1 – With brackets

In this option, a single monomer unit is represented within brackets. There are two important things to remember if you are using this method. Firstly, the bonds at either end of the monomer need to extend outside the brackets. This indicates that the monomer will be bonded to other repeating monomer units to form the polymer. Secondly, the letter ‘n’ is written outside the brackets. The ‘n’ represents a very large positive integer.

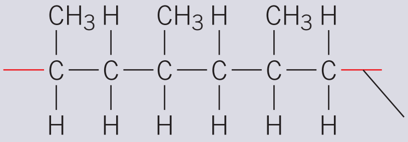


Bonds extend outside the brackets to indicate bonding to another monomer.

n represents a large number of monomers repeating in the chain.

Option 2 – Without brackets

In this option, you can draw one or more monomer units. Note that this is different from option 1, where you only ever show a single monomer. However, similarly to option 1, it is important that the bonds at each end are shown to extend, indicating that they would be able to bond to more monomer units to continue to grow the polymer. Despite this, the letter ‘n’ is not included in these representations.



Bonds at ends indicate bonding to another monomer. Note that *n* is not used.

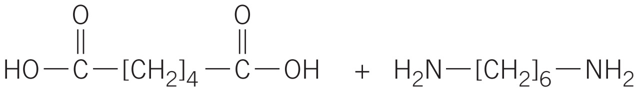
Section 8A questions

1 Define ‘natural polymer’ and ‘synthetic polymer’ and provide an example of each.

2 Define ‘monomer’ and ‘polymer’, using propene as an example.

3 Describe two differences between addition and condensation polymerisation.

4 Nylon is a polymer used in modern carpets due to its strength and abrasion resistance. It is synthesised using the following monomers:



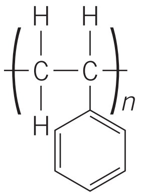
$$HO-\overset{\overset{O}{\|}}{C}-[CH_2]_4-\overset{\overset{O}{\|}}{C}-OH \quad + \quad H_2N-[CH_2]_6-NH_2$$

a Write an equation to show the synthesis of the nylon polymer from these monomers. (Note: Balancing is not required.)

b The table below shows polystyrene and some other common classes of polymers and their physical properties. Using only the information in the table, categorise which of the polymers would not be suitable to be used to produce carpet.

Polymer	Physical properties
Acrylic	Chemical- and temperature-resistant
Polystyrene	Moderately strong and brittle
Olefin	Lightweight, stain- and water-resistant
Polyester	Strong, durable, moisture-resistant

5 The polymer polystyrene is illustrated below:



a Write a balanced equation to show the synthesis of polystyrene.

b Describe the intermolecular forces that arise between polymer chains of polystyrene.

c Explain whether polystyrene would be considered a thermoplastic or a thermosetting polymer.

6 Super glues are commonly polymers.

a Describe the properties of a polymer that would be useful to this application.

b State whether this polymer would likely be a thermoplastic or a thermosetting polymer.

7 High-density polyethene (HDPE) and low-density polyethene (LDPE) are polymers made from an ethene monomer. The difference between these polymers is that LDPE has a high degree of branching in the polymer chains compared to HDPE.

a Explain why HDPE is often described as a rigid polymer, while LDPE is described as more flexible.

b Cross-linked polyethene (XLPE) is considered a thermosetting polymer. Consider to what extent XLPE would be considered rigid or flexible in comparison with HDPE and LDPE.

8B Polymer recycling

Study Design

- The categorisations of different plastics as fossil fuel-based (HDPE, PVC, LDPE, PP and PS) and as bioplastics (PLA, Bio-PE and Bio-PP); plastic recycling (mechanical, chemical and organic), compostability, circularity and renewability of raw ingredients
- Innovations in polymer manufacture using condensation reactions and the break down of polymers using hydrolysis reactions, contributing to the transition from a linear economy towards a circular economy.

Glossary

Bioplastic
Fossil fuel-based plastic

ENGAGE

Your plastic bottle could still be around in the year 2450!

Plastics and polymers are useful for so many products in our everyday lives. However, they can also cause many problems. Global statistics from the Earth Day organisation show that humans use 1.2 million plastic bottles per minute! Unfortunately, they also highlight that 91% of plastic doesn't get recycled. This is particularly worrying when you take into account that it has been estimated that a single PET plastic bottle will take around 450 years to break down into the ground. Significantly, a lot of this plastic waste does not even end up in the right location, with much of it being dumped into the ocean. It is suggested that plastic waste in the ocean is responsible for the death of over one million marine creatures each year. Considering these concerning statistics, it is not surprising that better alternatives and processes are being invented and implemented. In this section, we will explore the development of bioplastics and the move towards a circular economy.

Figure 8B–1 Plastic waste causes many problems, including harming marine life when it ends up in the ocean.

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EXPLAIN

Fossil fuel-based plastics versus bioplastics

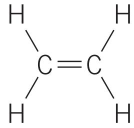
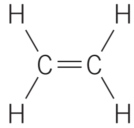
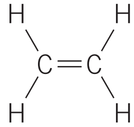
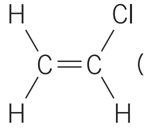
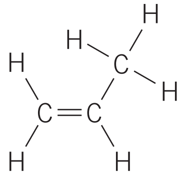
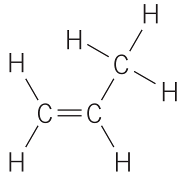
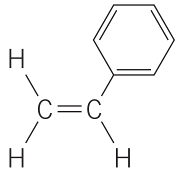
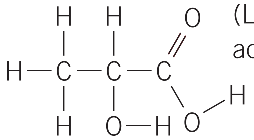
As mentioned at the start of the chapter, you may have noticed that plastic products often have different recycling codes that relate to the polymer that has been used to create them. More generally, we can separate polymers based on the source of their monomers. There are two broad categories: fossil fuel-based plastics and bioplastics. **Fossil fuel-based plastics** have monomers derived from either petroleum or natural gas. **Bioplastics**, on the other hand, are produced from renewable sources, such as corn starch, recycled food waste, sugar cane, straw and vegetable fats and oils. A summary of some of the major polymers, including their classification as either a fossil fuel-based plastic or bioplastic, can be found in Table 8B–1.

LINK 8A POLYMERISATION REACTIONS

Fossil fuel-based plastic plastic in which the monomers are derived from either petroleum or natural gas

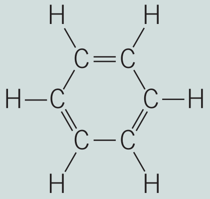
Bioplastic plastic produced from renewable plant sources

Table 8B–1 Common polymers, their monomers, the reaction that produces them and their classification as a fossil fuel-based plastic or bioplastic

Name	Monomer	Addition or condensation polymer?	Fossil fuel-based or a bioplastic?
HDPE High-density polyethene	 (Ethene)	Addition	Fossil fuel
LDPE Low-density polyethene	 (Ethene)	Addition	Fossil fuel
Bio-PE Biopolyethene	 (Ethene)	Addition	Bioplastic
PVC Polyvinyl chloride	 (Chloroethene)	Addition	Fossil fuel
PP Polypropene (also known as polypropylene)	 (Propene)	Addition	Fossil fuel
Bio-PP Biopolypropene	 (Propene)	Addition	Bioplastic
PS Polystyrene	 (Styrene)	Addition	Fossil fuel
PLA Polylactic acid	 (Lactic acid)	Condensation	Bioplastic

NOTE

is a symbol that represents



Note that Table 8B–1 shows a couple of instances where the same polymer can be made using both natural and fossil fuel-based sources. Polyethene (PE) has been traditionally manufactured through the use of feedstock obtained from natural gas or oil refining. In these instances, PE would be described as a fossil fuel-based polymer. However, recent advances have seen the development of Bio-PE. In this process, sugar cane is fermented to make ethanol. This can then be converted into 'ethene', which is then polymerised into Bio-PE. As sugar cane can be grown, it is seen as a renewable resource. While this seems advantageous over the use of non-renewable fossil fuels, there is still debate over whether this is a good alternative. In particular, many people are unsure whether it is appropriate to use a potential food source to make polymers. Bio-PE is also more expensive to produce than PE and is not biodegradable so does not confer that advantage either.

Biopolypropene (Bio-PP) is another polymer that can be manufactured from natural materials. In this instance, a variety of sources can be used, including vegetable oils, corn and sugar cane. Similar to Bio-PE, the production of Bio-PP could be considered as a better alternative to regular polypropene manufacturing, as the monomer is derived from renewable sources. However, the same potential cost and environmental concerns are also an issue. The environmental impacts of polymer manufacturing and disposal are becoming a key topic of discussion as the use of plastics continues to increase throughout the world.

WORKSHEET 8B–1 FOSSIL FUEL-BASED OR BIOPLASTIC?

Sustainability

The linear economy for polymers is based on the 'take, make, use, dispose' system and results in around 95% of polymer products having lost their value after a single use. Conversely, the circular economy model has a philosophy of 'materials designed to be used but not used up'. The general principles behind linear and circular economies were discussed in Chapter 3.

3C METAL RECYCLING LINK

VIDEO 8B–1 LINEAR VERSUS CIRCULAR ECONOMY OF POLYMERS

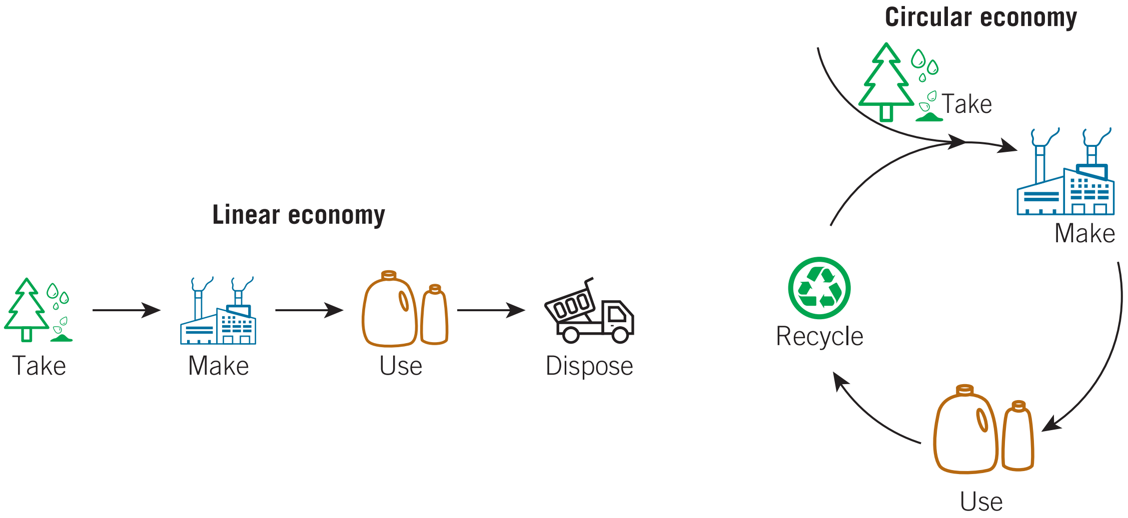


Figure 8B–2 The shift from a linear economy model of plastic use (left) to a circular economy model (right) relies on effective recycling techniques.

As you can see from Figure 8B–2, the key difference between a linear economy and a circular economy for the manufacturing and use of plastics is the addition of a recycling step. Recent Australian reports have shown that while around 3.5 million tonnes of plastic is consumed in Australia annually, only 10% of this is recycled. Therefore, it is important to understand how plastics can be recycled and discover innovative new approaches that can help maximise the recovery of resources from the use of polymers. Generally, there are three current techniques that can be used for the recycling of plastics: mechanical, chemical and organic.

Check-in questions – Set 1

1 What is the difference between a fossil fuel-based plastic and a bioplastic?
2 What are two types of reactions that can form polymers?
3 For an addition polymerisation reaction to occur, what must the monomer contain?
4 What is the main difference between HDPE and LDPE? What effect does this have on the properties of these polymers?
5 What is the source of the monomer for polyethene versus biopolyethene?
6 What is the difference between what happens to polyethene in a linear economy and a circular economy?

Mechanical recycling

Mechanical recycling involves processing the waste that is generated into raw components that can be re-used without impacting on the original chemical structure of the material. Generally, mechanical recycling can be described through the series of steps shown in Figure 8B–3. In this process, the collected plastics must first be sorted, as mechanical recycling is only suitable for certain types of polymers. This sorted material is then ground down into tiny fragments, which are washed to remove impurities and then often melted and formed into pellets. These pellets are then available to be used to create new products. As the mechanical recycling procedure leads to a loss in the quality of the polymer, the material at the end is usually not available to be used to create the same products; it can only be used when a lower quality is required. This is often called ‘downcycling’. For this reason, there is a limit to the number of times that products can undergo mechanical recycling. Most estimates suggest that after three to six cycles of mechanical recycling, the polymer will no longer be able to be re-used.

Figure 8B–3 Mechanical recycling of plastics involves collecting the plastic waste, sorting it and then grinding it into small pieces, which are then washed and available to be re-used.

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Chemical recycling

Chemical recycling is not something that is done on an industrial scale yet, though many trials are being performed to assess the best possible approaches. In this procedure, the recycled polymer chains are broken down into substances that can be used as a fuel or as a feedstock to make new polymers. This is a key difference to mechanical recycling, and it comes with a number of potential benefits. Firstly, it means that no sorting of the plastics is required, as the process should be able to recycle most polymers together simultaneously. Secondly, there is no quality loss through this process and so the polymers can theoretically be recycled infinitely. This also means that the products made from the recycled material can be of the same quality or even better quality than the original. However, there are also some drawbacks to the use of chemical recycling over mechanical recycling. It is thought that chemical recycling will be more energy intensive than mechanical recycling, while also requiring solvents and acids that could be potentially environmentally hazardous.

Organic recycling

UNIT 4 LINK

The third type of recycling is organic recycling. Organic recycling refers to the use of living organisms to break down the polymers. In particular, the use of bacterial enzymes is becoming a viable solution for the recycling of plastics. These enzymes have likely developed within these organisms to help decompose and utilise plastic waste that has built up in various ecosystems. Enzymes are biological catalysts that speed up the rates of chemical reactions. You will learn more about them in Unit 4; however, their basic method of action is shown in Figure 8B–4. In these instances, the polymer binds to a region on the enzyme known as the active site. The enzyme then helps break the bonds that hold the polymer together, releasing the monomer subunits as products of the reaction.

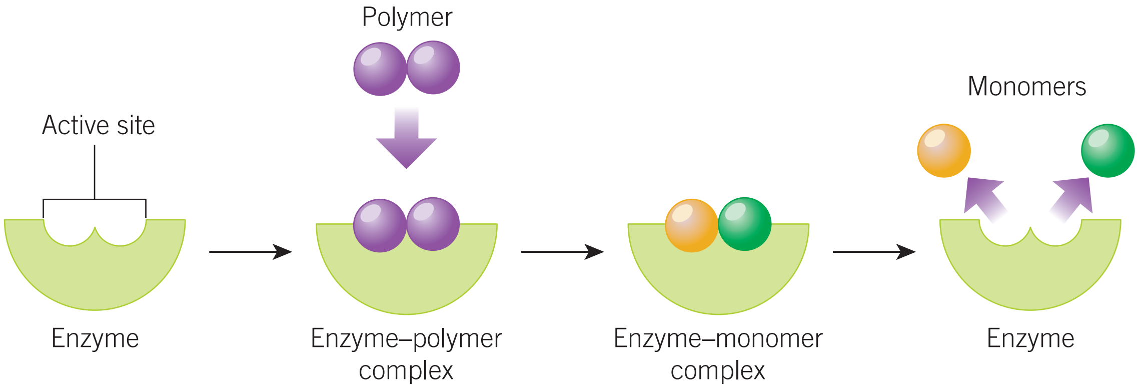

Figure 8B–4 Enzymes discovered in a range of bacteria are being investigated as a possible organic approach to polymer recycling.

While these enzymes represent a possible ecofriendly approach to the recycling of polymers, there is still considerable work that needs to be done to make them suitable for use on an industrial scale. Improvements that are currently being investigated include stabilising the enzymes at higher temperatures to speed up the rate of the reaction and enhancing the binding of the polymer to the active site of the enzyme. In April 2022, a research team from Texas published a paper in the journal *Nature*, highlighting their development of a modified enzyme called FAST-PETase. This enzyme is capable of breaking down polyethene terephthalate (PET), a common polymer used in packaging, in a matter of days. Usually, this plastic would take centuries to degrade.

Compostability

Composting is very similar to organic recycling. However, the end goal of composting is slightly different. Unlike in organic recycling, where the polymer is broken down into its monomers for re-use, the goal of composting is to provide a more environmentally friendly form of disposal. There are strict guidelines that must be met for a polymer to be considered compostable, which include the following:

- 90% of the material must have undergone biodegradation (that is, it must be biodegradable) within 180 days.
- Within 12 weeks, 90% of the material should have disintegrated into pieces that are no greater than 2 mm in size.
- The degraded material in the compost cannot be toxic to plants or earthworms.

An example of a polymer that is compostable is polylactic acid (PLA). PLA became well known during the 2012 London Olympics, after the organisers announced it would be the first 'zero-waste' Olympics. After the completion of the event, around 8000 tonnes of plastic packaging waste was composted over a nine-week period and then used to fertilise crops in the local area.

Chemical and organic recycling, including composting, involve hydrolysis reactions. Hydrolysis reactions involve the addition of water to a compound to produce two or more new compounds, through the breaking of bonds. In this way, they can be thought of as the reverse of condensation reactions. In Figure 8B–5, you can see the hydrolysis of a dipeptide made of two glycine residues. You may notice that this is the reverse of what was shown in Figure 8A–4, where a condensation reaction occurred between two glycine amino acids to form a dipeptide.

LINK **8A** POLYMERISATION REACTIONS

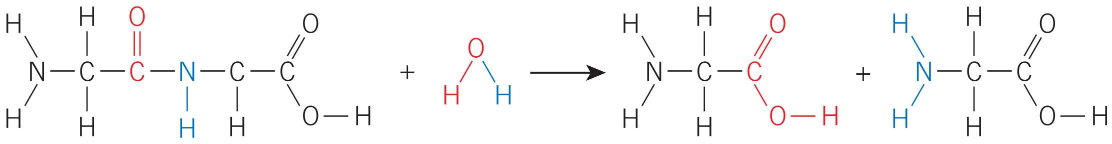

Figure 8B–5 A hydrolysis reaction where a dipeptide is separated into the two monomeric glycine subunits

Hydrolysis reactions are a cornerstone of a circular economy, as they allow for condensation polymers to be broken down into their monomers to be re-used as the same or different products. The microbial enzymes, which were discussed earlier in this section, use hydrolysis reactions to break down polymers back into their monomers for re-use.

8B SKILLS

Answering comparative questions

In VCE Chemistry, you may be asked questions that require you to compare and contrast different molecules or concepts. As an example, you might be asked to compare the different types of polymers (natural versus synthetic) or the different types of polypropene (PP versus Bio-PP). Although comparison questions may seem straightforward, students often don't receive full marks for their answer. The main reason for this is that students often interpret the question as follows:

Question:

Compare HDPE, LDPE and Bio-PE.

Interpretation:

List the differences between HDPE, LDPE and Bio-PE.

A comparison question requires you to answer it in two parts. First, you have to discuss the aspects that are similar between the two items, then you have to analyse the aspects that are different. This is the best way to ensure that you get maximum marks for a comparison-style question. Keep your comparisons succinct and do not repeat the question. Let's use the example above to look at ways to approach this type of question.

Answer – strategy 1:

Group your responses by similarities and differences.

Similarities:

- All are composed of ethene monomers.
- All are formed from addition polymerisation.

Differences (make it clear that there is a comparison by using terms such as 'whereas', 'although', 'in contrast' or 'unlike'):

- HDPE and LDPE are fossil fuel-based polymers, whereas as Bio-PE is a bioplastic.
- HDPE and LDPE are relatively cheap to produce, whereas Bio-PE is expensive.

Answer – strategy 2:

- Draw a table listing key features and the responses for each item.

	HDPE	LDPE	Bio-PE
Monomer	Ethene	Ethene	Ethene
Monomer source	Petroleum	Petroleum	Sugar cane
Fossil fuel-based or bioplastic?	Fossil fuel-based	Fossil fuel-based	Bioplastic
Production cost	Cheap	Cheap	Expensive

As you can see, using either strategy helps ensure that your answer is complete and thorough. Also, setting your answer out in either of these two ways helps the examiner to see clearly that you have identified all the key elements required for full marks.

Section 8B questions

1 Using the table format below, categorise the following polymers as fossil fuel-based plastics or bioplastics.

Biopolyethene
Biopolypropene
High-density polyethene
Low-density polyethene
Polyethylene
Polylactic acid
Polypropene
Polystyrene
Polyvinyl chloride

Fossil fuel-based plastics	Bioplastics

2 Describe the processes of mechanical and chemical recycling of plastics.

3 Compare a linear and circular plastic economy, referring to polyethene as an example.

4 Explain whether polypropene and polylactic acid would undergo hydrolysis in the presence of water.

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Chapter 8 review

Summary

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

Success criteria – I am now able to:		Linked question
8A.1	Define 'monomer'	13☐
8A.2	Define 'polymer'	13☐
8A.3	Describe the differences between a condensation reaction and an addition reaction	3☐
8A.4	State examples of natural and manufactured polymers	8☐
8A.5	Give the structure of an addition polymer from a given alkene monomer	11☐
8A.6	Give the structure of an alkene monomer from a given addition polymer	13☐
8A.7	Compare the properties of thermoplastic and thermosetting polymers	9☐
8A.8	Compare the structures of thermoplastic and thermosetting polymers	12☐
8A.9	Use the structure of thermoplastic and thermosetting polymers to explain their properties	7☐, 12☐
8A.10	Describe how the structure and properties of a monomer determine the properties of its polymer	2☐
8A.11	Describe how chain length affects the properties of a polymer	1☐, 2☐
8A.12	Describe how degree of branching affects the properties of a polymer	1☐, 2☐
8B.1	Define 'fossil fuel-based plastics' and 'bioplastics'	6☐
8B.2	Categorise a polymer as a fossil fuel-based plastic or a bioplastic, based on the origin of the monomer	6☐
8B.3	Compare mechanical, chemical and organic recycling of plastics	5☐
8B.4	List the requirements for a plastic to be considered compostable	10☐
8B.5	Describe how hydrolysis reactions are used in the recycling of polymers	13☐
8B.6	Compare the linear and circular economy models for polymers	4☐

Multiple-choice questions

1 Which of the following would most likely lead to a polymer with high flexibility?
  - **A** long polymer chain length
  - **B** high degree of branching
  - **C** low degree of branching
  - **D** high amount of cross-linking

2 Which of the following does **not** affect the properties of polyethene?
  - **A** source of the monomer
  - **B** structure of the monomer
  - **C** degree of branching
  - **D** chain length

3 What is one of the main differences between a condensation polymerisation reaction and an addition polymerisation reaction?
  - **A** Condensation polymerisation reactions involve carbon–carbon double bonds, whereas addition polymerisation reactions do not.
  - **B** Condensation polymerisation reactions only occur with natural monomers, whereas addition polymerisation reactions only occur with manufactured monomers.
  - **C** Condensation polymerisation reactions always produce polymers with high flexibility, whereas addition polymerisation reactions always produce rigid polymers.
  - **D** Condensation polymerisation reactions result in two different products being formed, whereas addition polymerisation reactions only create one product.

4 Which of the following steps is **not** part of the linear economy model of polymer use?
  - **A** make
  - **B** recycle
  - **C** use
  - **D** waste

5 Which of the following processes does **not** change the chemical structure of the polymer?
  - **A** chemical recycling
  - **B** composting
  - **C** mechanical recycling
  - **D** organic recycling

6 Knowing that the monomer for a polymer is derived from sugar cane tells you that it is
  - **A** thermosetting.
  - **B** cross-linked.
  - **C** a fossil fuel-based plastic.
  - **D** a bioplastic.

7 Which of the following best describes a thermosetting polymer?
  - **A** no cross-linking, softens after heating
  - **B** no cross-linking, remains rigid after heating
  - **C** cross-linking, softens after heating
  - **D** cross-linking, remains rigid after heating

8 Which of the following is a natural polymer?
  - **A** polylactic acid
  - **B** cellulose
  - **C** polypropene
  - **D** polystyrene

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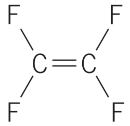
9 A thermoplastic is
- **A** able to be remoulded after heating.
- **B** highly heat-resistant.
- **C** always cross-linked.
- **D** not able to be recycled.

10 Which of the following is **not** a requirement for a polymer to be considered compostable?
- **A** The degraded material in the compost cannot be toxic to plants or earthworms.
- **B** Within 12 weeks, 90% of the material should have disintegrated into pieces that are no greater than 2 mm in size.
- **C** The monomer must be derived from a fossil fuel.
- **D** 90% of the material must have undergone biodegradation within 180 days.

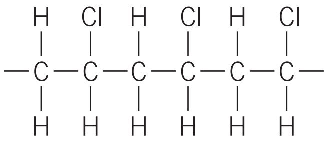
Short-answer questions

11 Draw the structural formula of three repeating units of the polymer made from the tetrafluoroethene monomer below. (2 marks)



12 Polystyrene is an example of a thermoplastic, whereas epoxy resin is classified as being a thermosetting polymer.
- **a** Using only the information given above, describe how the structures of polystyrene and epoxy resin would differ. (3 marks)
- **b** Would you expect polystyrene or epoxy resin to be more flexible? Explain your reasoning. (2 marks)

13 Polyvinyl chloride (PVC) is an addition polymer that can be used to make a variety of things, from shoes to drainpipes. The structure of PVC is shown below.



$$\begin{array}{ccccccccccccc} & H & & Cl & & H & & Cl & & H & & Cl & \\ & | & & | & & | & & | & & | & & | & \\ - & C & - & C & - & C & - & C & - & C & - & C & - \\ & | & & | & & | & & | & & | & & | & \\ & H & & H & & H & & H & & H & & H & \end{array}$$

- **a** Define ‘monomer’ and ‘polymer’. (2 marks)
- **b** What is the structure of the monomer that polymerises to produce PVC? (1 mark)
- **c** Condensation reactions are also able to produce polymers. Describe how hydrolysis reactions can be used to recycle polymers made through these reactions. (3 marks)

UNIT 1 HOW CAN THE DIVERSITY OF MATERIALS BE EXPLAINED?

CHAPTER 9 RESEARCH INVESTIGATIONS

Introduction

All scientists have the key role of researching practices that have previously been conducted. Outcome 3 for Unit 1 requires you to investigate and communicate a response to a recent discovery, innovation, advance, case study, issue or challenge linked to the knowledge and skills developed throughout Unit 1 Chemistry. It may include research into green chemistry principles, sustainable development and/or the transition to a circular economy. To support you with your research investigation, this chapter will assist you in how to prepare for such an investigation. It will provide you with strategies using 'thinking organisers', enabling you to digest the wide range of views and information you will uncover, as well as how to analyse both the validity and reliability of information. As part of your research, you will need to consider different implications on your chosen investigation. The chapter concludes with a suggested process for framing your investigation, weaving together the different threads of your understanding.

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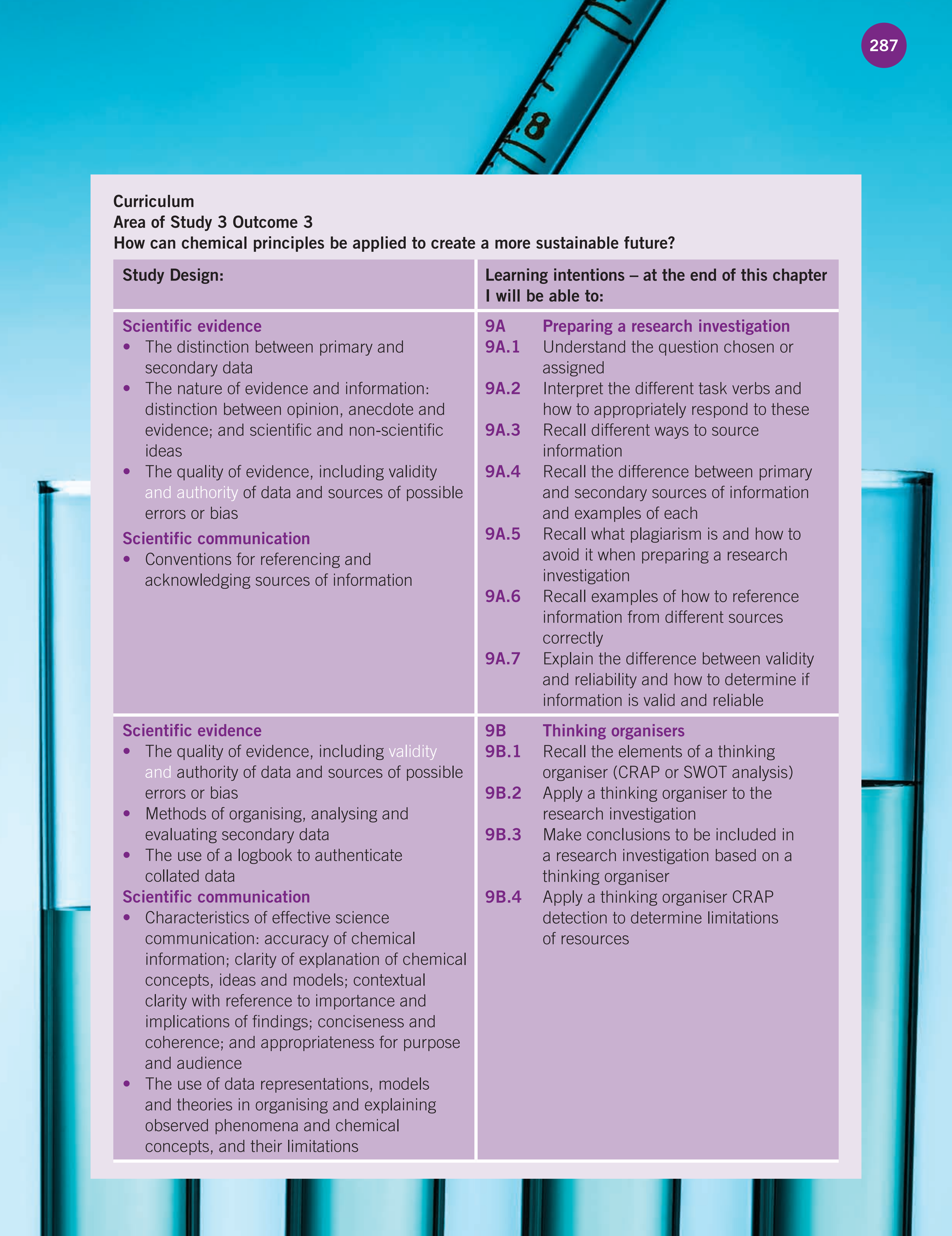

Curriculum
Area of Study 3 Outcome 3
How can chemical principles be applied to create a more sustainable future?

Study Design:	Learning intentions – at the end of this chapter I will be able to:
Scientific evidence • The distinction between primary and secondary data • 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 **Scientific communication** • Conventions for referencing and acknowledging sources of information	**9A Preparing a research investigation** **9A.1** Understand the question chosen or assigned **9A.2** Interpret the different task verbs and how to appropriately respond to these **9A.3** Recall different ways to source information **9A.4** Recall the difference between primary and secondary sources of information and examples of each **9A.5** Recall what plagiarism is and how to avoid it when preparing a research investigation **9A.6** Recall examples of how to reference information from different sources correctly **9A.7** Explain the difference between validity and reliability and how to determine if information is valid and reliable
Scientific evidence • The quality of evidence, including validity and authority of data and sources of possible errors or bias • Methods of organising, analysing and evaluating secondary data • The use of a logbook to authenticate collated data **Scientific communication** • 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 • The use of data representations, models and theories in organising and explaining observed phenomena and chemical concepts, and their limitations	**9B Thinking organisers** **9B.1** Recall the elements of a thinking organiser (CRAP or SWOT analysis) **9B.2** Apply a thinking organiser to the research investigation **9B.3** Make conclusions to be included in a research investigation based on a thinking organiser **9B.4** Apply a thinking organiser CRAP detection to determine limitations of resources

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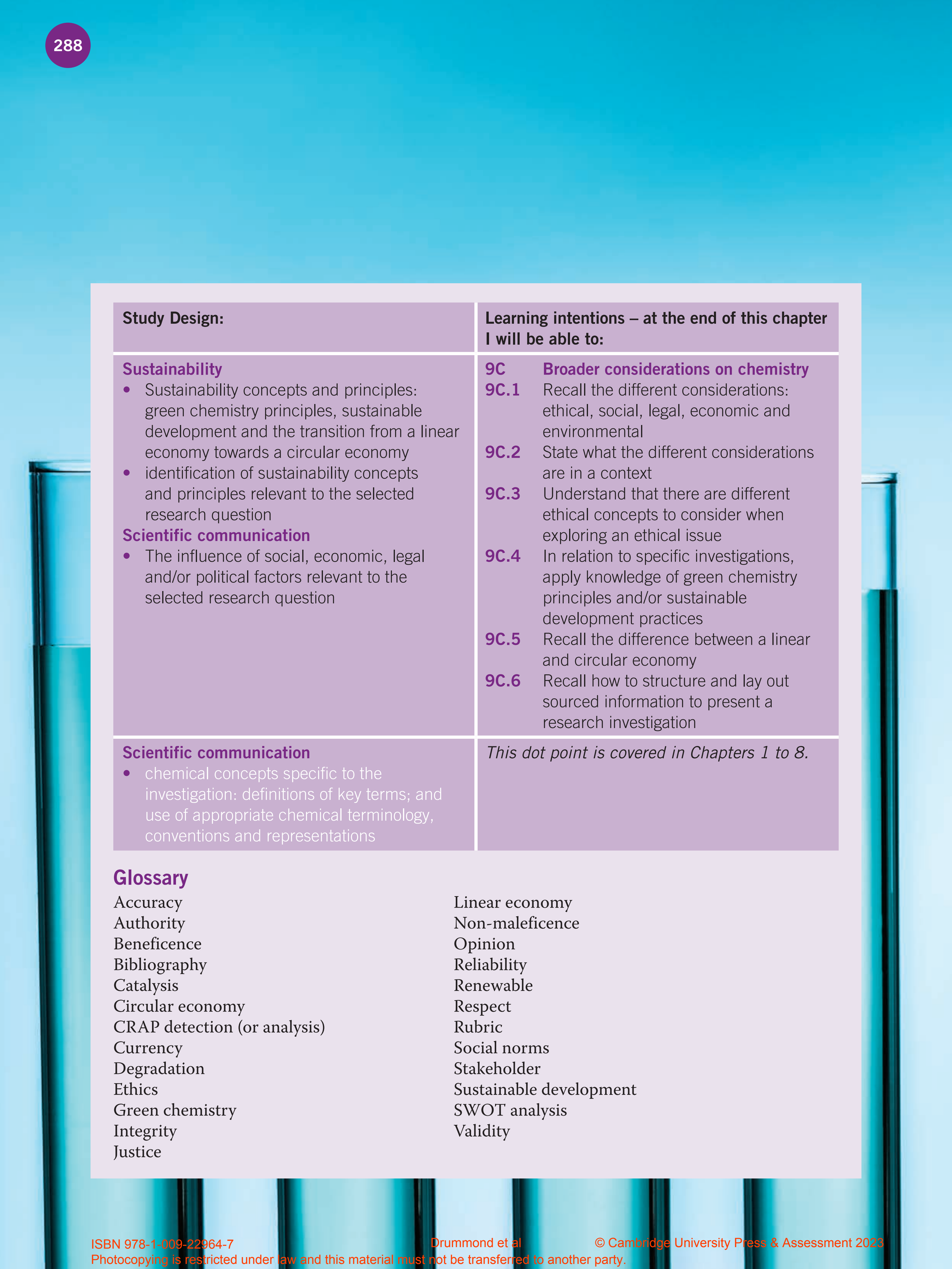

Study Design:	Learning intentions – at the end of this chapter I will be able to:
Sustainability • Sustainability concepts and principles: green chemistry principles, sustainable development and the transition from a linear economy towards a circular economy • identification of sustainability concepts and principles relevant to the selected research question **Scientific communication** • The influence of social, economic, legal and/or political factors relevant to the selected research question	**9C Broader considerations on chemistry** **9C.1** Recall the different considerations: ethical, social, legal, economic and environmental **9C.2** State what the different considerations are in a context **9C.3** Understand that there are different ethical concepts to consider when exploring an ethical issue **9C.4** In relation to specific investigations, apply knowledge of green chemistry principles and/or sustainable development practices **9C.5** Recall the difference between a linear and circular economy **9C.6** Recall how to structure and lay out sourced information to present a research investigation
Scientific communication • chemical concepts specific to the investigation: definitions of key terms; and use of appropriate chemical terminology, conventions and representations	*This dot point is covered in Chapters 1 to 8.*

Glossary

Accuracy
Authority
Beneficence
Bibliography
Catalysis
Circular economy
CRAP detection (or analysis)
Currency
Degradation
Ethics
Green chemistry
Integrity
Justice
Linear economy
Non-maleficence
Opinion
Reliability
Renewable
Respect
Rubric
Social norms
Stakeholder
Sustainable development
SWOT analysis
Validity

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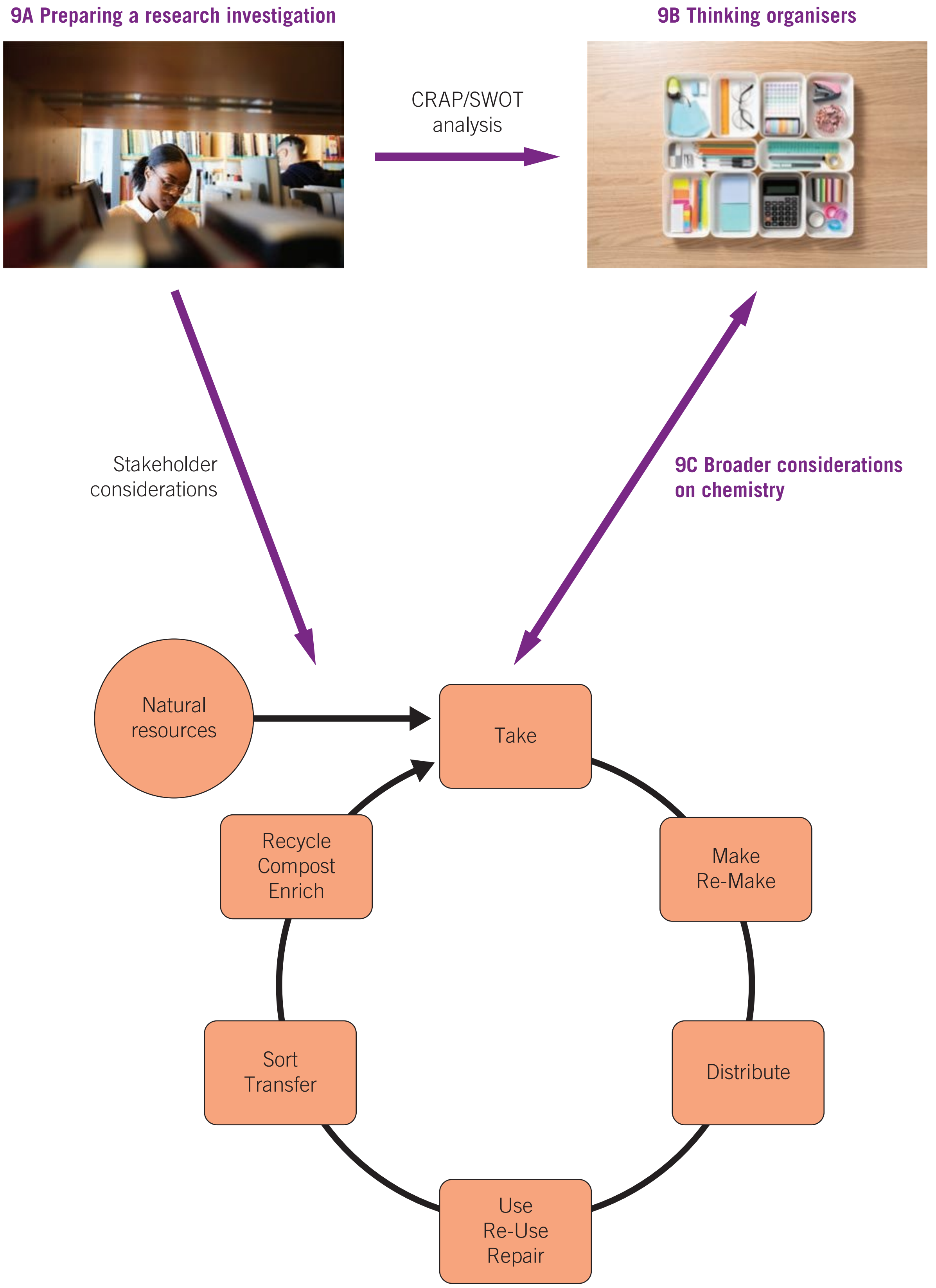
Concept map



See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

9A Preparing a research investigation

Study Design:
Scientific evidence
- The distinction between primary and secondary data
- 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

Scientific communication
- Conventions for referencing and acknowledging sources of information

Glossary:
Bibliography
Opinion
Reliability
Validity

ENGAGE

Reading critically

If you visit the 'Dihydrogen monoxide – DHMO' website at dhmo.org, you will be presented with information, claimed to be unbiased, about dihydrogen monoxide, including claims about the many health effects of this substance. If you read the information thoroughly, you may realise (or you might already know) that dihydrogen monoxide is in fact water.

This site is an example of the need to be careful when reading information online and to cast a critical eye on the details being presented.

Much of your research will be conducted online. The following letters at the end of each URL indicate the type of organisation that published the information on the website. Knowing this can help you to assess the reliability of the information.

- .com = commercial
- .edu = educational
- .gov = government
- .org = non-profit

EXPLAIN

Choosing a research question

The starting point for your Unit 1 Area of Study 3 Research Investigation is to pick a topic or question. As per the VCAA Study Design, this must relate to the evaluation of a recent discovery, innovation, advance in chemistry, case study, issue or challenge linked to the knowledge and skills developed in Unit 1 Area of Study 1 or 2. This may include consideration of sustainability concepts, such as green chemistry principles, sustainable development and the transition towards a circular economy.

Examples of investigation topics and possible research questions have been provided in the Study Design and can be found in the supporting documents in the digital resource. It is worth noting though that these are only included as a guide, and a question or topic may be provided by your teacher – or alternatively, you may be given free rein to develop your own question.

The key is not to choose something that seems easy to research. Rather, to enhance your understanding of the chemistry principles, choose a question that you are passionate about or that interests you as something you want to know more about. A good example is an understanding of chemistry and/or processes used and still in use by Aboriginal and Torres Strait Islander peoples compared to westernised ways of doing things.

Clarifying your research question

Once you have chosen your investigation question and understand what your teacher requires of you, your next task is to break down your research question more. One thing you could try is to write down the question in your own words. This can help you identify quickly if there are any ideas you don't understand.

Firstly, isolate key task words or verbs in the question, such as those shown in Table 9A–1. The definitions of such terms give an indication of what exactly you need to do. This is also included in Table 9A–1.

Opinion
a view or judgement that may be based on personal feelings, not necessarily on facts or research

Table 9A–1 Key task words/verbs with accompanying information about what each means

Task word/verb	What is required to include
Analyse	Identify components/elements and the significance of the relationship between them; draw out and relate implications; determine logic and reasonableness of information.
Anecdote	Information that is typically casual and passed along more by word of mouth as opposed to rigorously scientifically documented.
Assess	Make a judgement about, or measure, determine or estimate, the value, quality, outcomes, results, size, significance, nature or extent of something.
Compare	Recognise similarities and differences and the significance of these similarities and differences.
Describe	Provide characteristics, features and qualities of a given concept, **opinion**, situation, event, process, effect, argument, narrative, text, experiment, artwork, performance piece or other artefact in an accurate way.
Discuss	Present a clear, considered and balanced argument or prose that identifies issues and shows the strengths and weaknesses of, or points for and against, one or more arguments, concepts, factors, hypotheses, narratives and/or opinions.
Evaluate	Ascertain the value or amount of; make a judgement using the information supplied, criteria and/or own knowledge and understanding to consider a logical argument and/or supporting evidence for and against different points, arguments, concepts, processes, opinions or other information.
Examine	Consider an argument, concept, debate, data point, trend or artefact in a way that identifies assumptions, possibilities and interrelationships.
Explain	Give a detailed account of why and/or how with reference to causes, effects, continuity, change, reasons or mechanisms; make the relationships between things evident.
Infer	Derive conclusions from available information or evidence, or through reasoning, rather than through explicit statements.
Investigate	Observe, study or carry out an examination in order to establish facts and reach new conclusions.
Research	The systematic investigation, including the analysis of previous research, to establish facts and reach new conclusions.
Summarise	Retell concisely the relevant and major details of one or more arguments, texts, narratives, methodologies, processes, outcomes and/or sequences of events.

Secondly, look at topic words. These are often the nouns used to define the topic content. They give an indication of the content expected in the investigation. They are usually your ticket to speedy, effective research. In the examples below, taken from the Study Design, the topic words have been underlined.

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 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?

The WH? Plan

What is the 'WH? Plan'? It refers to these five questions:

- Who?
- What?
- Where?
- Why?
- How?

Asking these questions begins the thinking process in your research. For example, if the question was:

How do the properties of the metalloids (such as germanium, antimony, tellurium) differ from their neighbours on the periodic table, and how have these properties made them highly important for society and consequentially scarce in supply?

You would need to:

- Understand **what** metalloids are
- Understand **what** are the properties of metalloids
- Know **where** metalloids come from
- Know **where** these metalloids are located on the periodic table
- Understand **what** the properties of these neighbouring elements are
- Explain **how** the properties of metalloids make them important for society
- Understand **why** they are scarce in supply
- Know **who** the current stakeholders involved in the life cycle of these elements are and **how** they benefit.

As part of this research investigation, you would need to then apply your question to relevant chapters, chapter sections or digital resources within Unit 1.

Developing a plan for your research question

Once you have unpacked your research question and what is required of you, the next step is to begin researching information. For this step, you must remember to:

- search multiple online databases

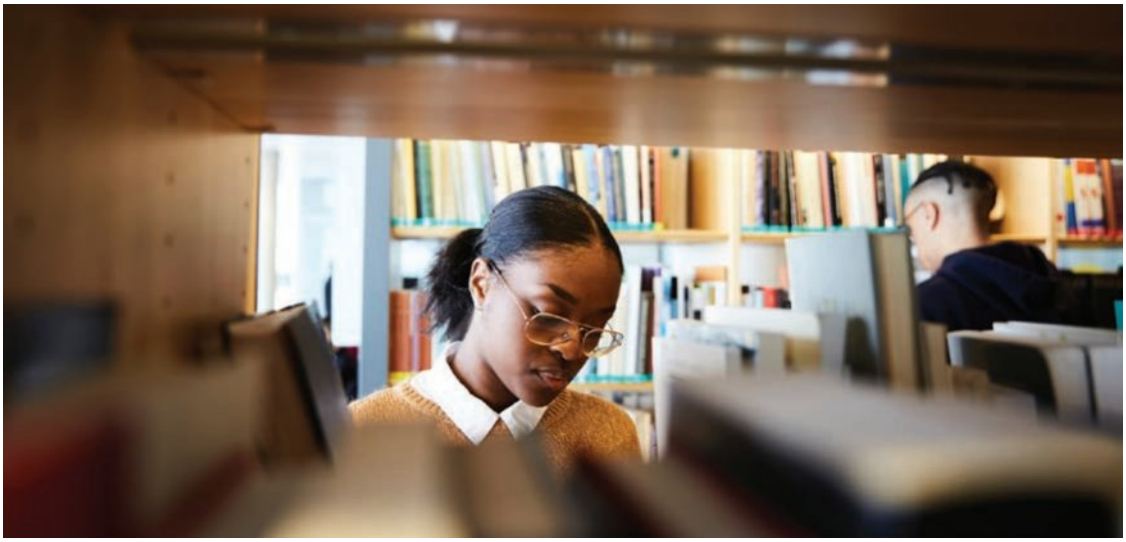
NOTE

If you are using Google Chrome, there are extensions available such as 'CiteMaker' that will generate references for you based on a website or a book's ISBN and then also collate these into a final **bibliography**.

Bibliography a list of referred to resources, listed in alphabetical order according to the author's surname or organisation, appearing at the end of a report

- utilise your school library databases and any physical resources
- look at both hard-copy and digital catalogues for information
- visit chemistry laboratories, local chemical manufacturers or industrial plants

- contact and interview experts involved in materials science or sustainability, or Aboriginal and Torres Strait Islander elders and leaders, following the appropriate protocols
- read the newspapers and follow social media groups reporting on recent research findings or experts' published points of view
- look at local, national or international governments for changes to their funding, policies or initiatives (for example, promotions of moving from a linear to a circular economy)
- watch TED talks, Clickview or other YouTube presentations.



Throughout all your research, remember to take your own notes and ensure that you do the following:

- Use a variety of sources.
- Use headings and sub-headings to group your information.
- Use abbreviations and content words.
- Select only the information relevant to your topic (as you will come across a lot of information and much irrelevant information.)
- Identify the sociocultural, economical, political, legal and ethical implications of the investigation with reference to sustainability.
- If required, form an opinion.

Primary versus secondary sources

When conducting your research, you will come across many different resources. There are two main types.

Primary sources

Primary sources, also known as primary material, is first-hand information that has been created by people who witnessed the events that they have recorded. They contain original data and evidence. Primary sources include:

- articles within scientific journals or from research presented at conferences and symposiums. They will include information directly from research discoveries and/or findings. These are often peer-reviewed (proofed) by qualified scientists before going to print
- interviews directly with scientists involved in discoveries
- photographs
- maps
- audio recordings.

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Secondary sources

Secondary sources, also known as secondary material, interpret and comment on primary material. These include any information from primary sources that has been further analysed or summarised. Secondary sources can also potentially provide a different interpretation of the significance of the primary resource. These sources include:

- textbooks
- magazines and newspaper articles
- government reports
- databases containing statistics and surveys
- laboratory work
- interviews with other professionals in the scientific or broader community who were not directly linked to the discoveries or findings.

Identifying bias

With any secondary sources in particular, it is important to know how to determine if these are credible. This is especially the case when evaluating websites, because anyone can post information online. While the content may seem believable, you still need to consider the following aspects as outlined in Table 9A–2.

Table 9A–2 What to look for when determining if a source is credible

Author	• Who is the author of the information? • Are they qualified and/or do they have an irrefutable reputation to be reporting on the information? • Why have they written about this? • Where was the information published? • When was the information published? (Particularly with the speed of advancements in science, if it is outside of the last 10 years, evaluate the relevance.) • Are their conclusions supported by data and concrete evidence? • Are they unbiased? Or have they listed any potential conflict of interest or any association with groups funding the research or information?

Websites	• Consider all the questions raised above about the author/creator of the site for this information
	• Is the domain name indicative of a reputable organisation? A few reputable domain extension include:
	  • '.gov' (government sites)
	  • '.edu' (education sites)
	  • 'org' (often non-profit organisations)
	• Is the purpose of the site to sell a product or to provide information?
	• How accessible is the site? Is it regularly maintained with current working links? Is it easy to use?
	• When was the site last updated?

Validity and reliability of information

If you've confirmed that the sources of information you are using are reputable, the next question to ask of these is 'Are they both valid and reliable?'.

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You may have come across the terms 'validity' and 'reliability' previously in the context of a practical investigation. However, in terms of a research investigation, these terms have slightly different meanings. These are outlined in Table 9A–3.

Table 9A–3 The meaning of 'validity' and 'reliability'

Term	Meaning
Validity	The extent to which evidence supporting an argument is true or accurate
Reliability	If research findings are repeated or supported by evidence more than once

Validity
the extent to which evidence supporting an argument is true or accurate

Reliability
if research findings are repeated or supported by evidence more than once

Plagiarism and referencing

There is no doubt your teachers will have introduced you to the term 'plagiarism' in previous years.

In case they haven't, plagiarism is where you directly use another author's writing or images without acknowledging them. Even if you're paraphrasing pieces of information or changing it into your own words, it is still important to include a reference as to where you sourced this information.

This is why as you source information, you must ensure that you record (reference) accurately the sources you have read and accessed for information. Referencing is necessary to avoid plagiarism. Remember, your final written investigation will be a form of secondary source, and if your reader is interested in finding out more about your question, they may want to locate the primary source from where you found your information.

Figure 9A–1 Plagarism involves passing someone else's work off as your own.

Bibliography

A bibliography, or reference list, appears at the end of your document and is started on a separate page. These terms are often used interchangeably, but a reference list only includes items that you have directly referenced in your investigation. A bibliography also includes sources used in the preparation of your investigation.

For your investigation, it should have the heading 'References', which is in bold, centred on the page and not in italics or underlined. Each reference is then indented slightly on a new line, which is **not** numbered or dot-pointed. If you are referencing a website, the reference must include the URL and the date on which the website was accessed.

NOTE

To save you time and energy later, it is useful to begin accumulating your list of references as soon as you begin doing research or writing.

Referencing

There are a few common formats used for referencing sources of information. This textbook is going to focus on the main one currently used by universities, including The University of Melbourne and Monash University.

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DOCUMENT 9A–1 APA-STYLE REFERENCING

American Psychological Association (APA) style referencing

This format for referencing utilises both in-text citations and a full list of details in the bibliography. Examples of how this would appear based on the number of authors and type of text is shown in the Skills section at the end of this chapter.

In-text citations

In-text citations is for information directly used in the text of your investigation. They include only the author's name(s) and publication year. Specific pages are required only if you have directly quoted from someone else's work.

In the Skills section below, you will see that for three or more authors, only the surname of the first author is used, followed by 'et al.' (a Latin abbreviation meaning 'and others') and the year.

If copying or using the exact words or direct quote of another author, you must enclose those in single quotation marks, then immediately follow with an in-text citation, including the author, year and page number, or section/paragraph identified, in brackets. For longer quotations, you should use a separate paragraph that is indented from both the left and right margins. Italics are also used instead of quotation marks to distinguish lengthier quotations from the main text.

Even when paraphrasing, you should still provide the page number or section identifier. The reason for this is to not only allow a reader to easily locate the primary source information, but also you as the writer, if you need to refer back to this.

VIDEO 9A–1 SKILLS: CORRECTLY CITING REFERENCES

9A SKILLS

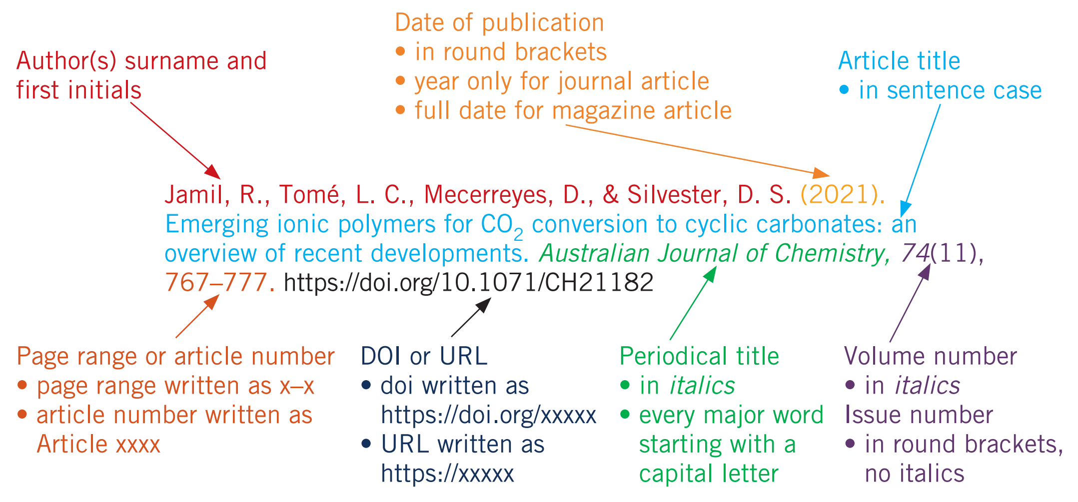
Worked examples of how to cite information obtained from different types of sources with different numbers of authors

Number of authors	Information to include	Examples
One author or one organisation	• Author's surname and publication year	• Maaser (2023) has proposed that … • … effect was shown in a recent study (Drummond, 2023).
	• Organisation name and publication year • Organisation's full name used in first citation; an abbreviation used thereafter	• In 2019, the International Union of Pure and Applied Chemistry (IUPAC) produced a document …
Two authors	• Both surnames every time, and the publication year • Use '&' only within brackets.	• (Watson & Crick, 1953) • Car and Parrinello (1985) found that …
Three or more authors	• First author's surname followed by 'et al.' and publication year	• (Shaw et al., 2019) • Adem et al. (2022) found that …

Reference lists

All sources cited in text, such as in the introduction section of a scientific poster, must be included in the reference list. References should be listed in alphabetical order by author surname or organisation name. There are very specific formatting rules for references lists, and different sources require different details. In general, four elements are included in a reference: author, date, title and source. The examples below show a reference entry for a journal article, a book and a website. For more examples and explanations of how to reference other sources such as Wikipedia, organisational reports, films, YouTube clips, TED Talks and podcasts, a referencing quick guide is provided in the Interactive Textbook.

Periodical (journal, magazine or newspaper article; print or online)



Author(s) surname and first initials

Date of publication
- in round brackets
- year only for journal article
- full date for magazine article

Article title
- in sentence case

Jamil, R., Tomé, L. C., Mecerreyes, D., & Silvester, D. S. (2021). Emerging ionic polymers for CO_2 conversion to cyclic carbonates: an overview of recent developments. *Australian Journal of Chemistry, 74*(11), 767–777. https://doi.org/10.1071/CH21182

Page range or article number
- page range written as x–x
- article number written as Article xxxx

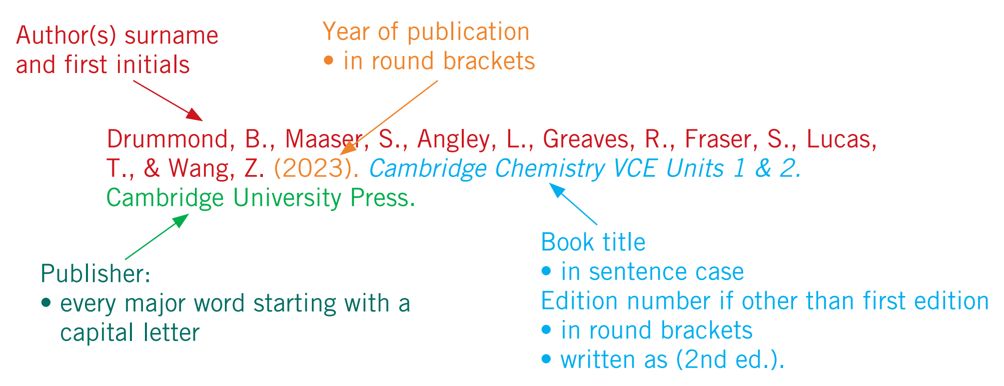
DOI or URL
- doi written as https://doi.org/xxxxx
- URL written as https://xxxxx

Periodical title
- in *italics*
- every major word starting with a capital letter

Volume number
- in *italics*

Issue number
- in round brackets, no italics

Book



Author(s) surname and first initials

Year of publication
- in round brackets

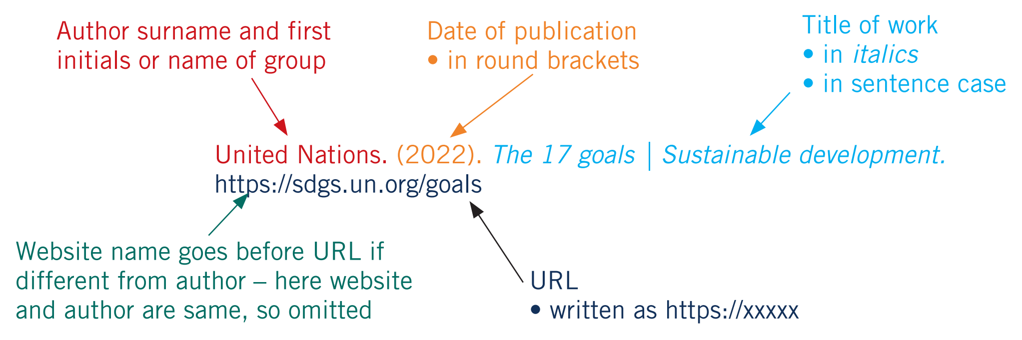
Drummond, B., Maaser, S., Angley, L., Greaves, R., Fraser, S., Lucas, T., & Wang, Z. (2023). *Cambridge Chemistry VCE Units 1 & 2.* Cambridge University Press.

Publisher:
- every major word starting with a capital letter

Book title
- in sentence case

Edition number if other than first edition
- in round brackets
- written as (2nd ed.).

Website



Author surname and first initials or name of group

Date of publication
- in round brackets

Title of work
- in *italics*
- in sentence case

United Nations. (2022). *The 17 goals | Sustainable development.* https://sdgs.un.org/goals

Website name goes before URL if different from author – here website and author are same, so omitted

URL
- written as https://xxxxx

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Activity 9A–1

Read the following article, a response by a professor of food science and technology at Ohio State University to an 11 year old's question about plastics being biodegradable.

If plastic comes from oil and gas, which come originally from plants, why isn't it biodegradable? – Neerupama, age 11, Delhi, India

To better understand why plastics don't biodegrade, let's start with how plastics are made and how biodegradation works.

Oil, also known as petroleum, is a fossil fuel. That means it's made from the remains of very old living organisms, such as algae, bacteria and plants. These organisms were buried deep underground for millions of years. There, heat and pressure turned them into fossil fuels.

Petroleum contains a lot of a chemical called propylene. To make plastic, refiners heat the propylene together with a catalyst – a substance that speeds up chemical reactions. This causes individual molecules of propylene to hook together like beads on a string.

8A POLYMERISATION REACTIONS **LINK**

The chain is called a polymer – a large molecule made of many small molecules strung together. Its name, polypropylene, literally means "many propylenes." And the bonds between these molecules are super strong.

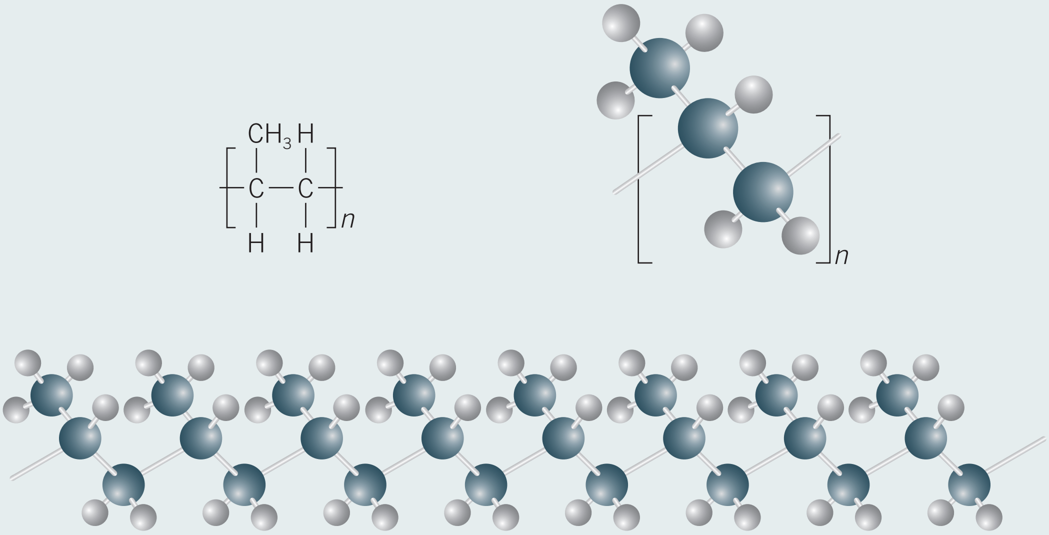

Figure 9A–2 This image shows the chemical structure of a propene monomer as found in polypropene (top left), a ball-and-stick model of the propene monomer (top right), and a chain of propene molecules linked together to make a molecule of polypropene.

When something that is biodegradable, like a cardboard box, breaks down, microorganisms that are present in nature break down and digest the polymers in it. They do this using enzymes – proteins that help speed up the breakdown of compounds such as lignin, a natural polymer found in plant tissues.

If oxygen is present, which usually means the microbes and the thing they are breaking down are exposed to air, the polymers will biodegrade completely. Eventually, all that's left will be carbon dioxide, water and other biological material.

Oxygen is essential because it helps the microorganisms that degrade the material live longer. Biodegradation is usually fastest in hot, wet environments where there are enough microorganisms – for example, damp leaves on the ground in a warm tropical forest.

But polymers like polypropylene are not abundant in nature. The enzymes in the microorganisms that break down biodegradable materials don't recognize the bonds that hold polymers together.

Eventually, the polymers in plastic waste may break down, perhaps after hundreds of thousands of years. But when it takes such a long time, the damage is already done to the environment. Plastic trash can release harmful chemicals into soil and water, or break into tiny bits that animals, fish and birds eat.

Figure 9A–3 In the outdoors, plastic trash breaks down into smaller and smaller pieces, but it doesn't biodegrade completely for thousands of years.

In my laboratory, we're developing what we hope will be the plastics of the future – materials that work like regular plastic, but don't spoil the environment because they can degrade when people are done with them.

We work with bioplastics – materials that are made by tiny living bacteria. The bacteria make these substances for uses such as storing energy or protecting themselves from their surroundings. They can do this over and over, so we have lots of bioplastic to work with.

We blend these polymers with natural rubber, an abundant resource that comes from rubber plants, and with oil removed from waste grounds that are left over from making coffee. The rubber makes our bioplastic flexible, and we chemically modify the coffee ground oil to help make the material flow in the industrial machines that we use to shape it.

Making bioplastics isn't cheap, because there is not enough of the different ingredients right now that go into making these materials, and it costs a lot of money to set up the equipment to make them. But when enough people want them, the price will come down. I hope that one day these new biodegradable materials will replace plastics made from fossil fuels.

Source:

Vodovotz, Y. (2022). 'If plastic comes from oil and gas, which come originally from plants, why isn't it biodegradable?' *The Conversation.* May 30, 2022. https://theconversation.com/if-plastic-comes-from-oil-and-gas-which-come-originally-from-plants-why-isnt-it-biodegradable-179634

1 Answer the following questions about the article from *The Conversation* on the previous pages.
   - a Are there any task verbs/words in the question at the start of the article? If so, identify these. If not, explain how the question could be changed, from the author's perspective, to include one.
   - b Explain if you can determine if the writer is credible to answer the question posed by the child.
   - c Imagine you were the writer in charge of researching information to help answer this 11 year-old's question. Develop a 'WH?' plan of questions that would help you to go about your research.
   - d Is there any bias from the author in this response? Explain.
   - e Define 'validity' and 'reliability'.
   - f Justify, using both your knowledge and evidence from this article, if this source of information is valid and/or reliable.

9B Thinking organisers

Study Design:

Scientific evidence

- The quality of evidence, including validity and authority of data and sources of possible errors or bias
- Methods of organising, analysing and evaluating secondary data
- The use of a logbook to authenticate collated data

Scientific communication

- 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
- The use of data representations, models and theories in organising and explaining observed phenomena and chemical concepts, and their limitations

Glossary:

Accuracy
Authority
CRAP detection/ analysis
Currency
SWOT analysis

ENGAGE

Why you should be organised

Research tells us that, as humans, we need some incentive to be motivated to complete even the simplest of tasks.

With anything, including undertaking our own research, having good systems in place will help us to save time. Perhaps, equally as important, implementing good routines can save brainpower and minimise stress. Stress arises from clutter or disorder. If you are not keeping track of your sources while you research, it can be difficult to go back and look for the sources from which you obtained information and reference them correctly.

This section will explore two examples of thinking organisers that will help you to be more organised when researching information for your own investigation. It will help you to reduce the overabundance of information you will encounter and make it easier to collate information into similar ideas.

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EXPLAIN

Using the CRAP and SWOT methods

Two key thinking organisers can help you to assess the information you access while conducting your research:

- **CRAP detection (or analysis)** – use this to determine whether the source of the information is reliable
- **SWOT analysis** – use this to develop arguments for and against an issue so that you are able to present a balanced research investigation.

CRAP detection (or analysis) a method of determining whether a source is reliable, based on the currency, reliability, authority and purpose of the information

SWOT analysis a method of assessing an issue by considering strengths, weaknesses, opportunities and threats associated with the issue

Currency timeliness; how up to date information is

Authority the level of expertise and relevance of an author in relation to the information they provide

CRAP detection

CRAP is an acronym for **currency**, reliability, **authority**, purpose. You may need to include a CRAP analysis as part of your investigation as evidence of being critical about the resources you have based your arguments or information on. Note that, although medical and scientific journals are reliable, you do not need to limit yourself to these as your only resource. Whatever source you use, assess and acknowledge its reliability in supporting a particular argument.

Table 9B–1 lists questions you should ask about each piece of information you find. Use these questions to assess the reliability and credibility of the information.

Table 9B–1 CRAP detection: questions to ask

C (Currency)	When was the information published? Has it been updated and when?
R (Reliability)	Is this information based on opinion or factual evidence? Is it a primary or secondary source? Does it cite, or refer to, other research or clinical studies? Have you checked those references? Is the content general or detailed, balanced or biased?
A (Authority)	Who is the author of, or organisation responsible for, the information? Do they have expertise/experience in the topic being covered?
P (Purpose)	Why and for whom was the information written? Does the source have a vested interest in the topic and the information that is included?

DOCUMENT 9B–1
CRAP DETECTION TEMPLATE

SWOT analysis

SWOT stands for strengths, weaknesses, opportunities, threats. A SWOT analysis is a way of ordering your thinking to help you build the discussion section of your research report. Do not paste your SWOT analysis as a table into your research report unless you have been instructed to do so.

DOCUMENT 9B–2
SWOT ANALYSIS TEMPLATE

> NOTE
>
> SWOT is not a framework for how to construct your discussion. It is a guide to help you elaborate, make connections and present balanced arguments or evidence in your discussion.

Table 9B–2 describes how to use SWOT to help build your discussion.

Table 9B–2 SWOT analysis

S (Strengths)	**W (Weaknesses)**
Record the basic facts about the positive aspects of the technology.	Record the basis facts about the negative aspects of the technology.
This is where you place a list of facts from your sources. It can be in dot point form. It can be cut and pasted from resources. Be sure to include the source for each point listed.	
O (Opportunities)	**T (Threats)**
Using your own words, elaborate on the facts (listed in the dot points in your Strengths list). Identify potential future applications for the technology. Discuss positive links to the stakeholders. Link the identified opportunities to the four considerations (ethical, social, economic, legal).	Using your own words, elaborate on the facts (listed in the dot points in your Weaknesses list). Identify potential setbacks for the technology in the future (or those that have occurred in the past). Discuss negative links to the stakeholders. Link the identified threats to the four considerations (ethical, social, economic, legal).
This is where you put the content into your own words and make connections.	

Science communication

You have seen how the thinking organisers, CRAP detection and SWOT analysis, can help you to source, assess and organise information to begin constructing a balanced research investigation using reliable resources. To present a high-quality research investigation, these ideas must also be communicated effectively.

There are a number of key principles that you should try and keep in mind, which are outlined in the Table 9B–3.

Table 9B–3 Key principles of science communication

Principle	Importance
Accuracy	While science communication often involves simplifying complicated concepts, it is critical that this doesn't result in a loss of accuracy. Ensuring that the correct information is provided, while making it easy to understand, is crucial to successful communication.
Clarity	It is important to ensure that your messages are presented clearly. To do this, you should always keep your key points in mind and refer to them repeatedly. Finding visual ways to display the information can help improve clarity as well.
Context	Context involves placing your information into the 'bigger picture'. This could involve highlighting what the possible future directions are of the work or showing how it links into other research that has been conducted previously. This helps your audience gain a better understanding of your work and why they should care about it.

Accuracy
how close the measurements are to the 'true' value of the quantity being measured

Table 9B–3 *Continued*

Conciseness	In order to communicate effectively, you must retain the attention of your audience. The easiest way to do this is to present your information in a thorough but brief way. Techniques such as using dot points or visual representations, such as infographics, can help ensure that you maximise the information you provide while staying concise.
Coherence	Coherence refers to the logical order that you communicate your findings. If ideas are presented haphazardly, it becomes very difficult for the audience to follow. Proofreading can be a very effective technique to help ensure that your work is presented coherently.
Appropriate for purpose	It is always important to remember what the overall aims of your communication are and ensure that you are meeting those goals. The style of presentation will differ depending on what the intention is: are you trying to provide a general overview of a concept (where you can be brief and broad) or trying to answer a scientific research question (where you will have more detail and support your argument with evidence)?
Appropriate for audience	The language used should be specific to the people that will be reading your work. As an example, the terminology that you use should be different if you are writing for people with a chemistry background compared to if you are writing for a non-scientific audience.

Data visualisation

An important consideration in science communication is how you present any data. Tables and graphs provide a great way to organise and visualise data, rather than simply listing values. Tables compile data into columns and rows so that it may be easily interpreted. As with scientific writing, it is important to structure tables so your reader can easily understand them. Consider Table 7B–10 from Chapter 7, reproduced below, which demonstrates some key features:

7B HYDROCARBONS LINK

- a clear title that describes the table content
- clear headings with relevant units, which indicate what each column represents
- rows that are listed in a logical order (increasing number of carbon atoms in this example).

Table 7B–10 Boiling and melting points of straight-chain alkanes

Number of carbons	Name	Semi-structural formula	Boiling point (°C)	Melting point (°C)
1	Methane	CH_4	–167.7	–182.5
2	Ethane	CH_3CH_3	–88.6	–183.3
3	Propane	$CH_3CH_2CH_3$	–42.1	–187.7
4	Butane	$CH_3CH_2CH_2CH_3$	–0.5	–138.3
5	Pentane	$CH_3CH_2CH_2CH_2CH_3$	36.1	–129.8

This table enables the reader to easily access the boiling and melting point of a particular alkane at a glance.

Trends and patterns in tabulated data may not be immediately evident to the reader. Due to their visual nature, graphs can instead be used to effectively show the overall shape of data, including any patterns or trends, and to make comparisons between different groups. The two types of graphs that you will most commonly encounter are line graphs and bar graphs. Line graphs plot individual data points and show how they are connected with lines. Bar graphs use bars of different heights to represent data.

LINK **7D** ALCOHOLS AND CARBOXYLIC ACIDS

Consider the line graph shown in Figure 7D–9 from Chapter 7, reproduced below, which compares the boiling points of alcohols and carboxylic acids. You should be able to easily identify the following trends from inspection of this graph.

1. The boiling points of alcohols and carboxylic acids increase with increasing molar mass.
2. The boiling points of carboxylic acids are higher than alcohols with corresponding molar masses.

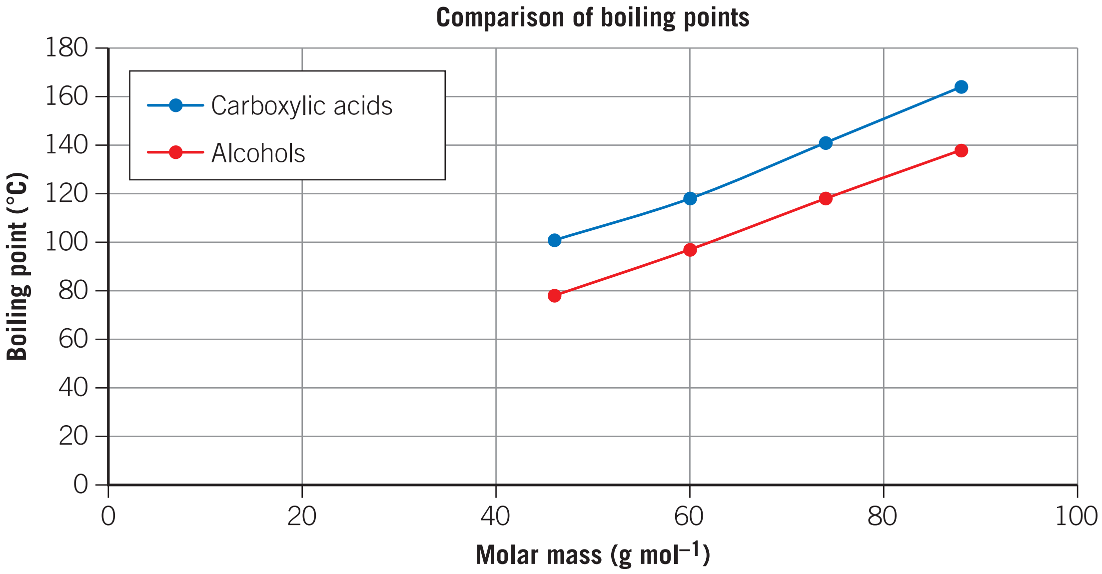


Figure 7D–9 Graphical data comparing the boiling points of alcohols to those of carboxylic acids

In your research investigation, you may wish to present tables or graphs from other sources. Ensure that you appropriately cite the source and refer to all figures and tables in the main body of your text.

Another way of being organised

The use of a logbook is another way to help with your organisation when collecting information. The purpose and use of a logbook is explored in detail from a practical investigation perspective in Chapter 16.

LINK **16A** INVESTIGATIVE PLANNING AND DESIGN

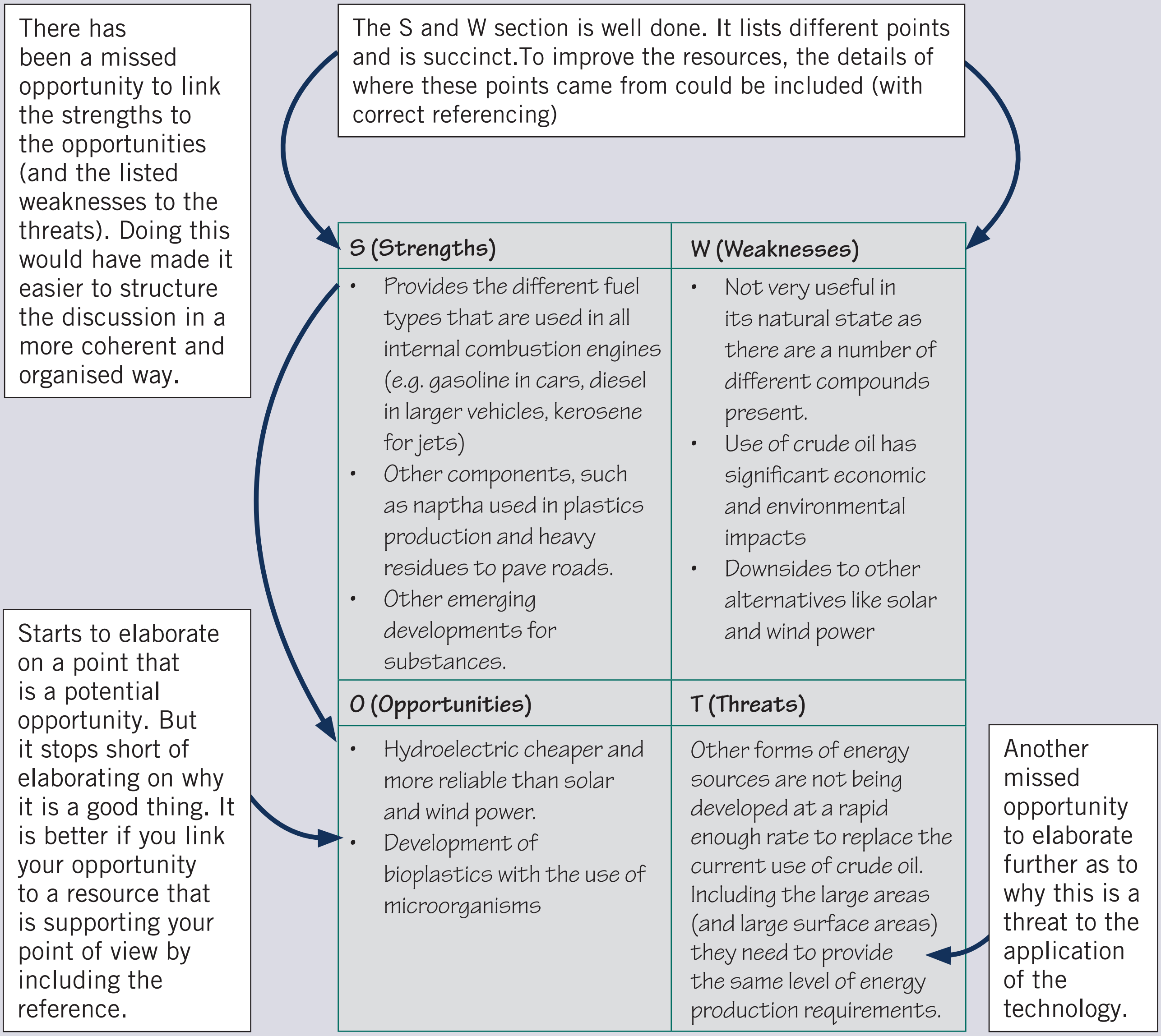
VIDEO 9B-1 SKILLS: HOW TO USE A SWOT ANALYSIS

ISBN 978-1-009-22964-7 Drummond et al

9B SKILLS

How to use a SWOT analysis: making connections between sections

Below are two examples from actual student attempts to use a SWOT analysis during the planning of their research task. Note the annotations around the examples, highlighting features done well and areas for improvement.



There has been a missed opportunity to link the strengths to the opportunities (and the listed weaknesses to the threats). Doing this would have made it easier to structure the discussion in a more coherent and organised way.

The S and W section is well done. It lists different points and is succinct.To improve the resources, the details of where these points came from could be included (with correct referencing)

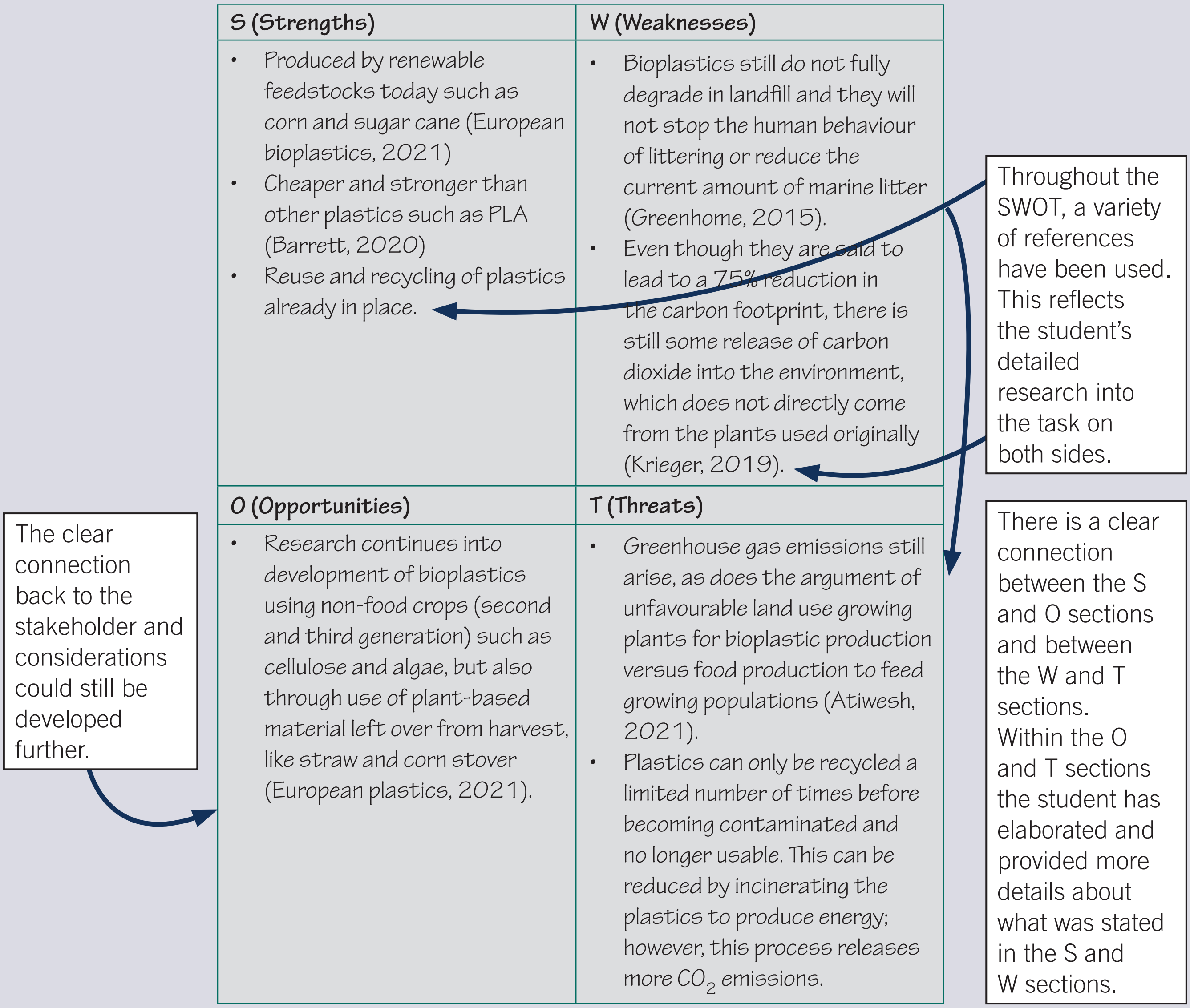
S (Strengths)	W (Weaknesses)
• Provides the different fuel types that are used in all internal combustion engines (e.g. gasoline in cars, diesel in larger vehicles, kerosene for jets) • Other components, such as naptha used in plastics production and heavy residues to pave roads. • Other emerging developments for substances.	• Not very useful in its natural state as there are a number of different compounds present. • Use of crude oil has significant economic and environmental impacts • Downsides to other alternatives like solar and wind power
O (Opportunities)	**T (Threats)**
• Hydroelectric cheaper and more reliable than solar and wind power. • Development of bioplastics with the use of microorganisms	Other forms of energy sources are not being developed at a rapid enough rate to replace the current use of crude oil. Including the large areas (and large surface areas) they need to provide the same level of energy production requirements.

Starts to elaborate on a point that is a potential opportunity. But it stops short of elaborating on why it is a good thing. It is better if you link your opportunity to a resource that is supporting your point of view by including the reference.

Another missed opportunity to elaborate further as to why this is a threat to the application of the technology.

Figure 9B–1 Student A's SWOT analysis of whether we actually need crude oil

As highlighted in Figure 9B–1, student A completed all the sections of the SWOT analysis. A successful SWOT requires connections between the S and O and between the W and T sections, which Student A did not demonstrate. In the O and T sections, elaboration of the points listed in the S and W sections is required. Not making these connections and elaborations will make completing the discussion section of your research task more difficult.



S (Strengths)	W (Weaknesses)
• Produced by renewable feedstocks today such as corn and sugar cane (European bioplastics, 2021) • Cheaper and stronger than other plastics such as PLA (Barrett, 2020) • Reuse and recycling of plastics already in place.	• Bioplastics still do not fully degrade in landfill and they will not stop the human behaviour of littering or reduce the current amount of marine litter (Greenhome, 2015). • Even though they are said to lead to a 75% reduction in the carbon footprint, there is still some release of carbon dioxide into the environment, which does not directly come from the plants used originally (Krieger, 2019).
O (Opportunities)	**T (Threats)**
• Research continues into development of bioplastics using non-food crops (second and third generation) such as cellulose and algae, but also through use of plant-based material left over from harvest, like straw and corn stover (European plastics, 2021).	• Greenhouse gas emissions still arise, as does the argument of unfavourable land use growing plants for bioplastic production versus food production to feed growing populations (Atiwesh, 2021). • Plastics can only be recycled a limited number of times before becoming contaminated and no longer usable. This can be reduced by incinerating the plastics to produce energy; however, this process releases more CO_2 emissions.

Throughout the SWOT, a variety of references have been used. This reflects the student's detailed research into the task on both sides.

There is a clear connection between the S and O sections and between the W and T sections. Within the O and T sections the student has elaborated and provided more details about what was stated in the S and W sections.

The clear connection back to the stakeholder and considerations could still be developed further.

Figure 9B–2 Student B's SWOT analysis of the advantage of bioplastics over fossil fuel-based plastics

Student B (Figure 9B–2) showed a greater level of understanding of the content and how to use the SWOT analysis framework than Student A. All sections were completed with appropriate linking to subsequent sections. Throughout the analysis, Student B included references, and this reflected the research they had done to understand the positive and negative aspects of the application of technology. An area of improvement for Student B would be in connecting the O and T sections back to the stakeholder and the four considerations required in the research task. This would have elevated the SWOT analysis even further. Constructing your discussion section in the final research report will be easier if you put effort into the SWOT analysis.

Section 9B questions

1 What do the letters in CRAP detection stand for?
2 What is the main purpose of using CRAP detection?
3 What sections of a SWOT analysis need to be linked together?
4 Outline the benefit of using a SWOT analysis when completing a research task.

9C Broader considerations on chemistry

Sustainability

- Sustainability concepts and principles: green chemistry principles, sustainable development and the transition from a linear economy towards a circular economy
- Identification of sustainability concepts and principles relevant to the selected research question

Scientific communication

- The influence of social, economic, legal and/or political factors relevant to the selected research question

Glossary:

Beneficence
Catalysis
Circular economy
Degradation
Ethics
Green chemistry
Integrity
Justice
Linear economy
Non-maleficence
Respect
Rubric
Social norms
Stakeholder
Sustainable development

ENGAGE

The unforeseen effects of DDT

New chemical discoveries can often have unintended consequences. A famous example of this is the compound dichlorodiphenyltrichloroethane, otherwise known as DDT. DDT was first synthesised in 1874, and its insecticidal qualities were discovered in 1939 by the Swiss chemist Paul Hermann Müller. DDT was first used outside the lab during World War II to prevent the spread of tropical diseases, such as malaria, with incredible success. Müller received the Nobel Prize in Physiology or Medicine in 1948 for his discovery. After the war, DDT was repurposed for agricultural use and was introduced to Australia and New Zealand in the 1950s. By the 1960s, approximately 80 million kilograms of DDT were made and used annually.

However, DDT had unexpected impacts on the environment. In 1962, American writer and marine biologist Rachel Carson wrote the book *Silent Spring*, an impassioned criticism of the usage of synthetic pesticides, such as DDT. The title of the book refers to the decrease in bird populations in areas sprayed with pesticides, leaving silence where there was once birdsong. Carson wrote on how DDT worked its way through the food chain, killing all manner of organisms but particularly birds. This is part of a process known as bioaccumulation, which you will read about further in Chapter 14. Partly in response to the book, most DDT usage was banned in the United States by 1972, and bird populations began to increase in areas previously affected. Australia banned its usage in 1987.

CHAPTER 14 LINK

14A SOURCES OF SALT IN WATER LINK

Silent Spring is considered one of the most important scientific books of all time, and it had significant impact on the modern environmental movement. However, DDT remains highly effective against malaria and is still used to this day in many places where the disease is prominent.

Figure 9C–1 A plane spraying DDT over a field

EXPLAIN

Considerations for your stakeholder

Once you have determined the technology you are going to focus on for your research task, it is important that you define who you are going to identify as your **stakeholder**. The stakeholder is anyone or anything (such as an organisation) that will be in some way affected by the technology. How broadly or specifically you define your stakeholder will dictate how broad or specific the scope of your research will be.

Ethical, sociocultural, economic, legal and/or political factors relevant to the selected research question are the five areas you will need to discuss when outlining the impact of your chosen technology on your stakeholder. In this course, ethical understanding is a key science skill and cross-study specification. Table 9C–1 outlines each of these considerations. The Skills section then demonstrates how to build connections between four types of these considerations. The difficulty of this research task lies in how well you can weave the various considerations together while being specific to green chemistry principles and sustainability practices.

VIDEO 9C–1
DIFFERENT TYPES OF IMPLICATIONS

Stakeholder
the person or organisation that will be affected by the factor under consideration

Ethics
moral principles that guide our beliefs about what is right or wrong conduct

Table 9C–1 Ethical, social, legal and economic considerations in relation to the impact of technology on stakeholders

Consideration	Definition	Relation to stakeholders
Ethical	A sense of right or wrong, based upon morals and beliefs, in producing or obtaining the technology	A range of technologies is available, and that gives an individual a choice as to which service to use, if any. The decision to engage these technologies will be based upon an individual's personal ethical considerations. To conduct scientific research involving humans, animals or plant products requires approval by an ethics committee.
Sociocultural	The influence of society on **social norms**	An individual's decision-making will be influenced by the context they are in. For example, people from different countries or cultural backgrounds may have different attitudes towards what is socially acceptable in regard to different technologies or products and the development of these. Attitudes may also differ across different communities within the same country.

Social norms
standards of what is considered socially acceptable within a community

Table 9C–1 *Continued*

Legal/political	The legislative rules and processes applicable in the jurisdiction in which the stakeholder or technology is located	The ability to use land, for example, to grow crops for the purpose of fuel or bioplastics as opposed to food, especially in developing countries. This is known as the food-versus-fuel debate. Starting up a service with a technology may require permits under relevant legislation in order for research or production to be conducted.
Economic	The availability of funds in order to obtain or produce the technology	The use of equipment, for example, for the purpose of metal recycling, can be extremely expensive. These technological services may not be equitable for all countries or specific communities to access. Also consider who is funding the research and development for these technologies. For example, a company might bias its results by emphasising positive outcomes and downplaying negative outcomes, such as greenhouse gas emissions or other waste products produced.

As mentioned earlier, how you present your information will depend on which stakeholder's point of view you use when addressing these considerations, as well as which type of investigation topic you are focusing on. You will need to conduct further research to understand how each consideration can influence your stakeholder in the context of your chosen chemical product, development or sustainability practice.

Approaching the ethical understanding

To assist you in unpacking the ethical understanding required in this course, there are further approaches to bioethics and ethical concepts (see Tables 9C–2 and 9C–3). How you approach the ethical consideration is an essential element of the research task. In light of this importance, when framing how you will explore the ethical consideration, you should be influenced by one or more of the approaches and ethical concepts outlined in tables 9C–2 and 9C–3.

Table 9C–2 Approaches to ethical considerations

Approach	What is it?
Consequence-based	Places importance on how the consequence of an action is achieved with maximum positive outcomes and minimal negative effects
Duty-based *and/or* rule-based	Focuses on the importance of people acting in a particular way so the ethical rules are followed, regardless of the consequences that could follow
Virtues-based	Places importance on the moral character of the person carrying out the action, where they have been influenced by the characteristics and behaviours a good person would seek to achieve to be able to act in the right way

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Table 9C–3 Ethical concepts

Ethical concept	What is it?
Integrity	The commitment to searching for knowledge and understanding, and the honest reporting of all sources of information and communication of results
Justice	The moral obligation to ensure that there is fair consideration of competing claims, that there is no unfair burden on a particular group from an action and that there is fair distribution and access to the benefits of an action
Beneficence	The commitment to maximising benefits and minimising the risks and harms involved in taking a particular position or course of action
Non-maleficence	Involves avoiding the causation of harm. Scientific research may involve some degree of harm, and therefore the harm resulting from any position or course of action should not be disproportionate to the benefits from that position or course of action
Respect	Involves consideration of the extent to which living things have an intrinsic value and/or instrumental value; giving due regard to the welfare, liberty and autonomy, beliefs, perceptions, customs and cultural heritage of both the individual and the collective

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In addition to these ethical concepts, special consideration should be given to the ethics of any investigation (such as Investigation topic 3) involving Aboriginal and Torres Strait Islander knowledge and culture. Refer to the *Overview* at the front of this textbook and in the online guidelines listed there for further information. Sustainable development and green chemistry principles also need to be considered.

Sustainable development can be defined as an approach that addresses the needs of the present without negatively impacting the ability of future generations to meet their own needs. As a result, the United Nations developed their 17 Sustainable Development Goals to address current global challenges. The goals from this list that are especially relevant to the four investigation topics that may be researched as part of your investigation are outlined in Table 9C–4.

Table 9C–4 United Nations Sustainable Development Goals linked to Unit 1 Investigation topics

Investigation topic(s)	Specific investigation topic	Sustainable Development Goal(s)
Topic 1: Endangered elements in the periodic table	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?	9 INDUSTRY, INNOVATION AND INFRASTRUCTURE
	Why is helium classified as a critical and endangered element, and how can it be saved given that its atmospheric recovery is almost impossible?	9 INDUSTRY, INNOVATION AND INFRASTRUCTURE 12 RESPONSIBLE CONSUMPTION AND PRODUCTION

Integrity
the commitment to searching for knowledge and understanding, and the honest reporting of sources and results

Justice
ensuring fair consideration of competing claims, with no unfair burden on any particular group, and fair distribution of benefits

Beneficence
the commitment to maximising benefits and minimising risks and harms in a particular position or action

Non-maleficence
avoiding harm; in scientific research the harm resulting from a position or course of action should not be disproportionate to the benefits

Respect
giving consideration to the extent to which living things have an intrinsic value; giving due regard to the welfare, liberty, autonomy, beliefs, perceptions, customs and cultural heritage of individual and collective

Sustainable development
an approach that meets the needs of the present without compromising the ability of future generations to meet their own needs

Table 9C–4 *Continued*

	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?	9 INDUSTRY, INNOVATION AND INFRASTRUCTURE
	How do the properties of the metalloids (such as germanium, antimony and tellurium) differ to their neighbours on the periodic table, and how have these properties made them highly important for society and consequentially scarce in supply?	12 RESPONSIBLE CONSUMPTION AND PRODUCTION
	How are precious metals from electronic waste (e-waste) recycled, and what are the environmental and economic benefits of these recovery processes?	11 SUSTAINABLE CITIES AND COMMUNITIES
Topic 2: Producing and using 'greener' polymers	What are plant-based biopolymers and what are the impacts of their production on the environment?	13 CLIMATE ACTION 15 LIFE ON LAND
	How do biodegradable and degradable polymers, compostable polymers and recyclable polymers differ in structure, production and environmental impacts?	15 LIFE ON LAND
	What is the difference between micropolymers and nanopolymers, and how are used plastic materials and litter managed and repurposed?	11 SUSTAINABLE CITIES AND COMMUNITIES 15 LIFE ON LAND

Table 9C–4 *Continued*

	Is the recycling of packaging products containing aluminium more sustainable than LDPE polymer-based packaging products?	11 SUSTAINABLE CITIES AND COMMUNITIES 15 LIFE ON LAND
	Why is the sale of plastic water bottles and single-use plastics banned in many countries?	14 LIFE BELOW WATER 11 SUSTAINABLE CITIES AND COMMUNITIES 15 LIFE ON LAND
	How do animal proteins compare with non-animal proteins for different applications, such as meat substitutes and non-animal leather?	2 ZERO HUNGER 13 CLIMATE ACTION
	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?	11 SUSTAINABLE CITIES AND COMMUNITIES 6 CLEAN WATER AND SANITATION

Table 9C–4 *Continued*

	What impact does the vulcanisation of rubber have on the environment and the communities where rubber is sourced and produced?	15 LIFE ON LAND 11 SUSTAINABLE CITIES AND COMMUNITIES
	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; single-use plastics, such as takeaway cups; containers; and electrical and electronic products, such as mobile phone cords and USB flash drives)?	15 LIFE ON LAND
	How are polymers used in forms of renewable energy, such as in components for wind turbines and solar panels?	7 AFFORDABLE AND CLEAN ENERGY
Topic 3: The chemistry of Aboriginal and Torres Strait Islander peoples' practices	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)?	12 RESPONSIBLE CONSUMPTION AND PRODUCTION
	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?	2 ZERO HUNGER
	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?	11 SUSTAINABLE CITIES AND COMMUNITIES
	How do binders and fixatives work to allow Aboriginal and Torres Islander peoples' paintings to be preserved for thousands of years?	11 SUSTAINABLE CITIES AND COMMUNITIES

Table 9C–4 *Continued*

	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?	11 SUSTAINABLE CITIES AND COMMUNITIES
	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 Nations peoples' fishing practices around the world?	14 LIFE BELOW WATER
	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'?	12 RESPONSIBLE CONSUMPTION AND PRODUCTION
Topic 4: The sustainability of a commercial product or material	What is 'green steel' and what are the implications of its production for human health and the environment?	9 INDUSTRY, INNOVATION AND INFRASTRUCTURE 13 CLIMATE ACTION
	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?	9 INDUSTRY, INNOVATION AND INFRASTRUCTURE
	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?	2 ZERO HUNGER 11 SUSTAINABLE CITIES AND COMMUNITIES

Table 9C–4 *Continued*

	Select a product whose composition has changed over time – for example, hair combs (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'?	11 SUSTAINABLE CITIES AND COMMUNITIES
	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?	12 RESPONSIBLE CONSUMPTION AND PRODUCTION
	How is the use of common pharmaceutical products, such as antibiotics and the contraceptive pill, impacting our water supply and the health of humans and animals?	6 CLEAN WATER AND SANITATION

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NOTE

The italicised questions are additional questions not present in the VCAA Study Design.

Green chemistry principles

Green chemistry is the design of new chemical products and manufacturing processes that are safer and more sustainable than traditionally used products and processes. A set of 12 principles underpin green chemistry to ensure the minimisation of the impacts of the products and processes on the environment, human health, the survival of other living organisms and ultimately on Earth's resources. Hence, it is about minimising the raw materials consumed with the aim of using renewable resources that can be produced or recycled with little energy, waste or other by-products.

Green chemistry the design of new chemical products and manufacturing processes that are safer and more sustainable than traditionally used products and processes

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Of these 12 principles, seven are relevant to VCE Chemistry. These are listed in Table 9C–5.

Table 9C–5 The seven green chemistry principles relevant to VCE Chemistry, developed initially by Paul T. Anastas and John C. Warner

Green chemistry principle	Description of green chemistry principle
Atom economy	The processes or pathways designed to maximise the incorporation of all reactants used into final products made.
Catalysis	Catalysts used generate the desired product with less energy usage and less waste.
Design for **degradation**	At the end of a product's use, it can be broken down into harmless degradation products and not remain in the environment.
Design for energy efficiency	The processes or pathways maximise energy efficiency with minimal negative environmental and economic impact.
Designing safer chemicals	Products achieve desired function while minimising toxicity.
Prevention	Waste is prevented as opposed to needing to treat or clean up waste.
Use of **renewable** feedstocks	Raw materials or feedstocks are made from renewable materials (like plants) as opposed to fossil fuels.

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Catalysis a process where a substance known as a catalyst is used during a reaction, speeding up the reaction rate

Degradation the process of a substance breaking down into a less complex state

Renewable a resource that can be replenished at a higher rate than it is consumed

Linear economy an economy that functions on a model whereby products are made and then simply discarded after use

Linear and circular economies

In previous chapters throughout Unit 1, a transition of different processes and material production or recycling from a linear to a more circular economy has been discussed.

The purpose of moving to a circular economy is to conserve resources that are at a high risk of becoming scarce in the future due to the increasing demands of growing populations. Industries are adapting their practices to set up research divisions, manufacturing processes and recycling techniques with a big focus on this. However, this can be quite costly to set up.

A **linear economy** is one that functions on a 'take, make and dispose' model where products are made and then simply discarded after use. This is shown in Figure 9C–2.

LINK **1D** CRITICAL ELEMENTS AND RECYCLING PROCESSES

LINK **3C** METAL RECYCLING

LINK **8B** POLYMER RECYCLING

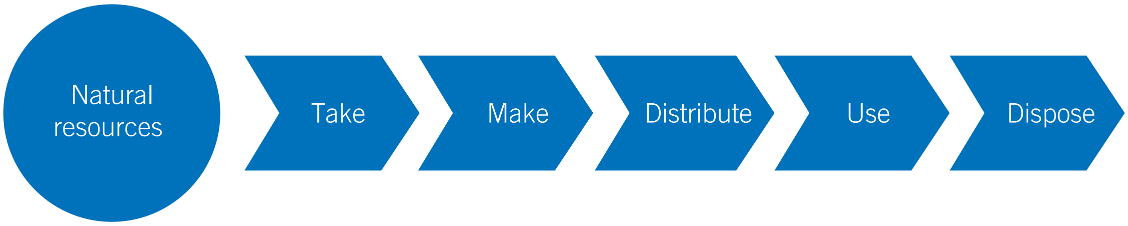


Figure 9C–2 A simple representation of a linear economy

Circular economy an economy that functions on a continuous cycle with a focus on optimising the re-use of resources and materials

A **circular economy** is one that functions on a continuous cycle with a focus on optimising the re-use of resources and materials. This is shown in Figure 9C–3.

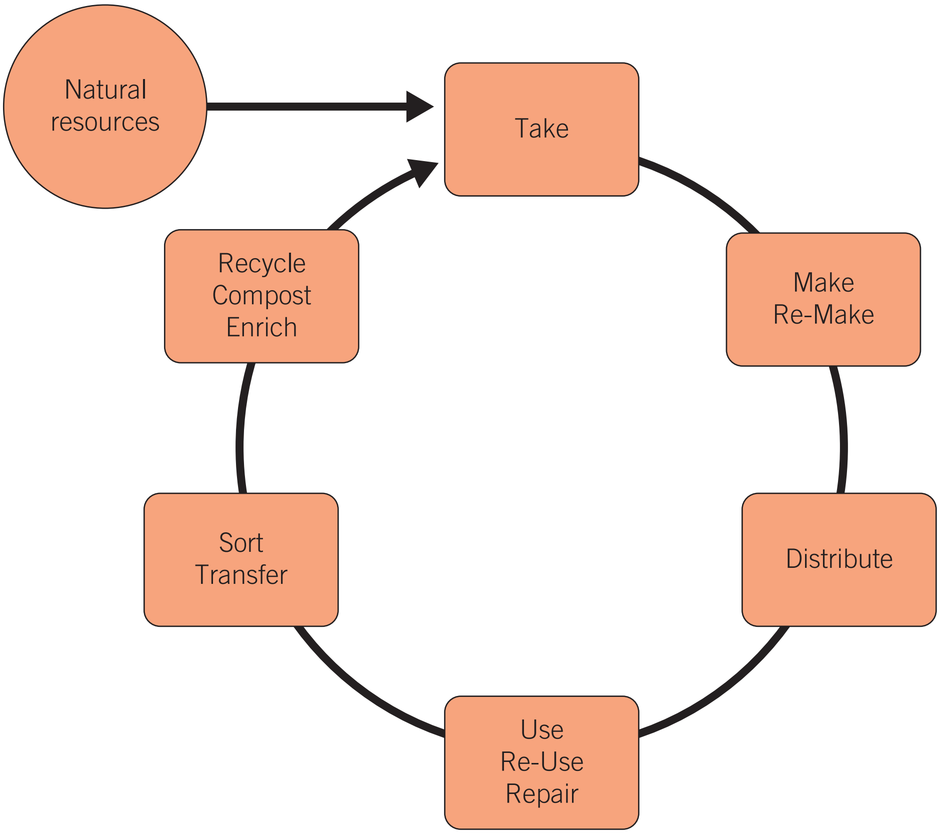


Figure 9C–3 A simple representation of a circular economy

The approach you decide to use to frame your research investigation will depend on the technology, your stakeholder and the context of the issue. For example, if you are coming from an industrial chemistry company's point of view, you could discuss the ethical concerns from a consequence-based approach, where all the outcomes need to be positive, with very few negative effects. While this company is conducting its own research, it would need to be mindful of the ethical concept of 'respect' in relation to use of resources, environmental impact and impact on the viability of humans and other organisms.

The skills section explores how you can build your own framework by starting with the ethical consideration.

VIDEO 9C–2 SKILLS: CREATING A FRAMEWORK TO MAKE CONNECTIONS

9C SKILLS

Creating a framework to make connections

This research investigation involves many ideas and considerations, all of which appear to be interwoven, and this can make it difficult to present your information in a coherent way while demonstrating a deep understanding of the technology and its impact.

The purpose of this section is to model a method of constructing your own framework, providing one possible template for presenting information. There are many ways to look at the different considerations. This is just one way that you may find useful. You could also use this process to construct a framework of your own that suits the way you want to write.

Step 1: Pick a central idea to which all factors connect.

The ethical consideration has strong links to all the remaining considerations, as well as connection to the green chemistry and sustainable development principles. The social setting, the legal and political environment and any financial burden influence what is right or wrong. As a result, a good starting point to explore the technology is from an ethical point of view in the context of your stakeholder.

Things to consider when exploring the ethical consideration include:
- Who would benefit from the use of the technology?
- Who might be harmed by the use of the technology?
- Are the interests of the community being considered and respected?

As you continue, keep referring back to who is affected and how they are affected.

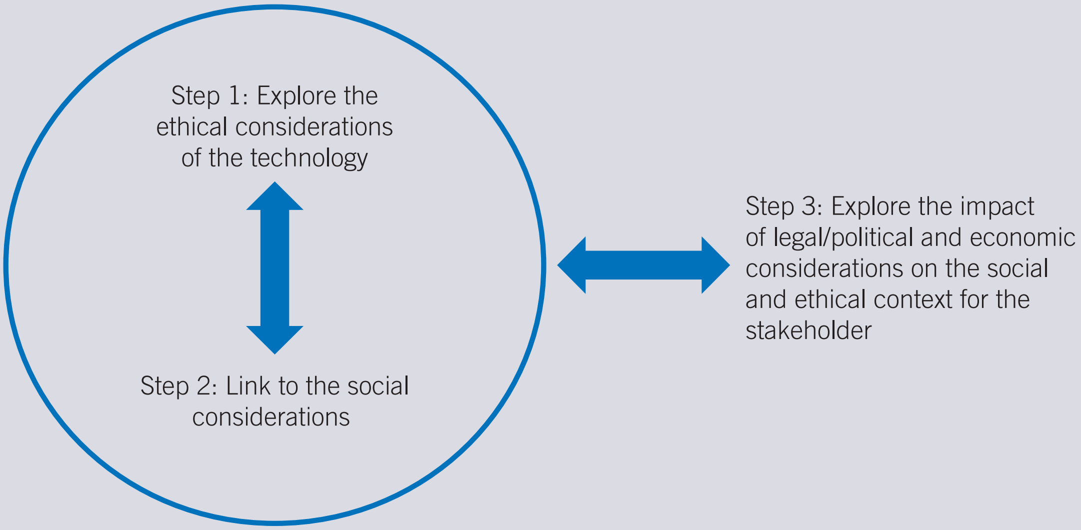
Step 2: Explore the next consideration that affects the central idea.

The sociocultural setting in which the technology is available influences how different stakeholders would behave. These two considerations appear closely linked and as a result would be explored after the ethical issues have been fleshed out. First, outline the social context of the technology and your identified stakeholder. Follow this up by demonstrating the interconnectedness between the two.

Continued on the next page

Step 3: Complete the connections.

Legal/political and economic considerations can equally and interchangeably affect and reflect social beliefs about what is right or wrong. As a result, these last two considerations work well to draw all the ideas together and complete the picture. First, define the context of the legal/political and economic climate for the stakeholder, and then explore how these relate back to the ethical and social considerations.


Step 1: Explore the ethical considerations of the technology

Step 2: Link to the social considerations

Step 3: Explore the impact of legal/political and economic considerations on the social and ethical context for the stakeholder

Writing your research investigation

Whether you structure your research investigation as a typical report saved as a Word document or PDF format, or present your final report as a PowerPoint or Prezi presentation, there is a certain structure that is good to follow.

There are three main sections in a research report: introduction, discussion and conclusion. When you are writing the report, it is important that you write in third person past tense. Any evidence or arguments presented should be based on research you have conducted (remember that Chapter 9B explains how to use CRAP detection to determine how reliable a resource is), not just your own opinion. Be wary of biased resources, as the technologies you are researching can have polarising agendas and opinions. Lastly, when you are using resources it is important to acknowledge where your information has come from by referencing your sources, as was outlined in Chapter 9A.

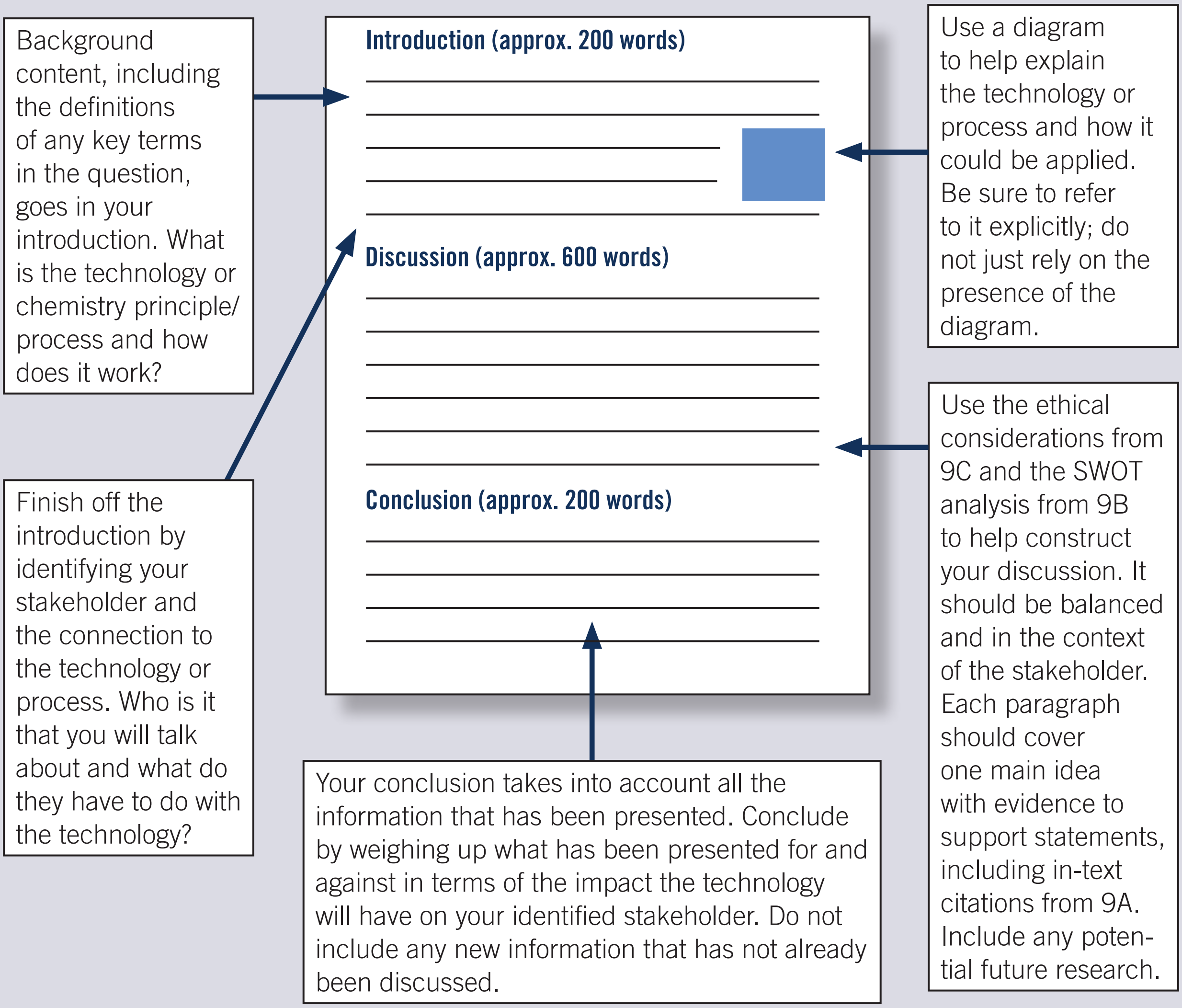
It is also a good idea to consult the assessment guide or **rubric** for how the investigation will be marked, to make sure your report complies with its requirements.

9B THINKING ORGANISERS LINK

9A PREPARING A RESEARCH INVESTIGATION LINK

Rubric
a scoring guide; a set of criteria used to evaluate students' responses

This structure is outlined in Figure 9C–4.



Background content, including the definitions of any key terms in the question, goes in your introduction. What is the technology or chemistry principle/ process and how does it work?

Finish off the introduction by identifying your stakeholder and the connection to the technology or process. Who is it that you will talk about and what do they have to do with the technology?

Use a diagram to help explain the technology or process and how it could be applied. Be sure to refer to it explicitly; do not just rely on the presence of the diagram.

Use the ethical considerations from 9C and the SWOT analysis from 9B to help construct your discussion. It should be balanced and in the context of the stakeholder. Each paragraph should cover one main idea with evidence to support statements, including in-text citations from 9A. Include any potential future research.

Your conclusion takes into account all the information that has been presented. Conclude by weighing up what has been presented for and against in terms of the impact the technology will have on your identified stakeholder. Do not include any new information that has not already been discussed.

Figure 9C–4 An example structure for preparing your research investigation

Before submitting your research investigation

Once you feel you have completed your final research investigation report, it is a good idea to use the Table 9C–6 on the following page as a checklist to ensure you have completed everything.

Continued on the next page

Table 9C–6 Checklist before submitting your research investigation

Item/process to be checked	Completed
I have answered the question with the perspective of a particular stakeholder(s).	
The report is structured correctly (as per Figure 9C–4 or the assessment guide provided by the school).	
I have checked off all parts of the report against the rubric or assessment criteria.	
All information included adds to the information in relation to the question.	
I have proofread my own work and checked for spelling mistakes or grammatical errors, and ensured that it is presented professionally.	
I have asked someone else to review my report to check that it answers the question asked and that it is interesting.	
The report is academic and demonstrates what I know compared to what others have written from research.	
I have referenced all sources of information correctly and with appropriate formatting.	
I have complied with the guidelines and ethics for studying Aboriginal and Torres Strait Islander knowledge and cultures (if included in the investigation).	

Section 9C questions

1 Aside from ethical issues, state three other examples of considerations that need to be addressed in carrying out your research task.

2 Research the following investigations topic from the VCAA Study Design in light of all the considerations discussed in this section, and answer the questions below.

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)?

  a Identify which groups could be positively affected by the use of these plants for their medicinal properties.

  b Identify which groups could be negatively affected by the use of these plants for their medicinal properties.

  c Outline whether you would harvest these plants to produce more natural medication. Give your reason(s).

3 Provide an annotated (with notes) flow chart to connect the considerations discussed in Table 9C–1.

Chapter 9 review

Summary

Create your own set of summary notes for this chapter, on paper or in a digital document. A model summary is provided in the Teacher Resources and can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have achieved the criteria.

Success criteria		Linked question
9A.1	Understand the question chosen or assigned	11☐
9A.2	Interpret the different task verbs and how to appropriately respond to these	11☐
9A.3	Recall different ways to source information	5☐
9A.4	Recall the difference between primary and secondary sources of information and examples of each	12☐
9A.5	Recall what plagiarism is and how to avoid it when preparing a research investigation	1☐
9A.6	Recall examples of how to reference information from different sources correctly	2☐
9A.7	Explain the difference between validity and reliability and how to determine if information is valid and reliable	3☐
9B.1	Recall the elements of a thinking organiser (CRAP or SWOT analysis)	4☐
9B.2	Apply a thinking organiser to the research investigation	13☐
9B.3	Make conclusions to be included in a research investigation based on a thinking organiser	13☐
9B.4	Apply a thinking organiser CRAP detection to determine limitations of resources	13☐
9C.1	Recall the different considerations: ethical, social, legal, economic and environmental	8☐
9C.2	State what the different considerations are in a context	8☐
9C.3	Understand that there are different ethical concepts to consider when exploring an ethical issue	7☐
9C.4	In relation to specific investigations, apply knowledge of green chemistry principles and/or sustainable development practices	9☐, 14☐
9C.5	Recall the difference between a linear and circular economy	6☐
9C.6	Recall how to structure and lay out sourced information to present a research investigation	10☐

Multiple-choice questions

1 You are guilty of plagiarism if you

- **A** copy sentences or paragraphs from internet sources without citing the source.
- **B** paraphrase from several different sources without citing these sources.
- **C** submit someone else's research report as your own.
- **D** all of the above

2 Which of the following is a correctly formatted journal article reference?

- **A** Carbene reactivity from alkyl and aryl aldehydes by L. Zhang in *Science*, 2022. https://doi.org/10.1126/science.abo6443
- **B** Zhang, L., Demuynck, B. M., Paneque, A. N., Rutherford, J. E., Nagib, D. A. (2022). Carbene reactivity from alkyl and aryl aldehydes. *Science*, *377*
- **C** Zhang, L., Demuynck, B. M., Paneque, A. N., Rutherford, J. E., & Nagib, D. A. (2022). Carbene reactivity from alkyl and aryl aldehydes. *Science*, *377*, Article 6606. https://doi.org/10.1126/science.abo6443
- **D** Lumin Zhang, Bethany M. Demuynck, Alyson N. Paneque, Joy E. Rutherford and David A. Nagib (2022). Carbene reactivity from alkyl and aryl aldehydes. *Science*, *377*, Article 6606. https://doi.org/10.1126/science.abo6443

3 Which of the following statements is correct?

- **A** An argument that is supported by accurate evidence is deemed to be valid.
- **B** Validity and reliability have the same meaning.
- **C** Reliability refers to the extent to which evidence supporting an argument is accurate.
- **D** Validity refers to research findings that are supported by evidence more than once.

4 Which of the following is **not** part of a CRAP analysis?

- **A** reliability
- **B** accuracy
- **C** currency
- **D** purpose

5 Which of the following is **not** likely to be a reliable source of information?

- **A** statistics on water usage from the Australian Bureau of Statistics website
- **B** a 2020 study published in the Royal Society of Chemistry's journal *Chemical Science*
- **C** a blog article titled 'Why Fake Meat is Worth the Hype'
- **D** a news article published by the ABC (Australia Broadcasting Corporation)

6 Which of the following steps is found in a linear economy but not a circular economy?

- **A** make
- **B** use
- **C** distribute
- **D** dispose

7 Which of the following ethical concepts is correctly defined?

- **A** Beneficence involves avoiding causing harm.
- **B** Justice is the moral obligation to ensure a fair consideration of competing claims.
- **C** Integrity involves consideration of the extent to which living things have value.
- **D** Respect is the honest reporting of all sources of information.

8 Establishing a technology that requires permits is an example of what sort of consideration?

- **A** ethical
- **B** sociocultural
- **C** economical
- **D** legal

9 Which of the following is **not** a green chemistry principle?

- **A** atom economy
- **B** building more research laboratories
- **C** catalysis
- **D** designing safer chemicals

10 In what section of a research investigation should you provide background information?

- **A** introduction
- **B** results
- **C** discussion
- **D** conclusion

Short-answer questions

11 Consider the following research investigation topic, taken from the VCAA Study Design: *How do animal proteins compare with non-animal proteins for different applications, such as meat substitutes and non-animal leather?*

- **a** Break down this research topic by identifying the following:
  - **i** a key task word/verb (1 mark)
  - **ii** topic words (1 mark)
- **b** Using the 'WH?' plan, suggest a relevant question that you could ask to begin planning a research investigation for each of the five questions: Who? What? Where? Why? and How? (5 marks)

12 Primary sources are different from secondary sources.

- **a** Define, with examples, what primary and secondary sources are. (2 marks)
- **b** Classify each of the following as either a primary or secondary source.
  - **i** the article used in Question 1 of Chapter 9A
  - **ii** *Cambridge Chemistry VCE Units 1 & 2* (this textbook)
  - **iii** an article on metal recycling written in the *Herald–Sun* newspaper
  - **iv** a radio interview with the lead author of a research study (4 marks)

13 CRAP analysis is a key tool that can help you assess the information you access while conducting research.

- **a** What is the purpose of CRAP analysis? (2 marks)
- **b** Visit the 'Dihydrogen monoxide – DHMO' website: dhmo.org. Perform a CRAP analysis on this website. (2 marks)
- **c** Perform a CRAP analysis based on the following information about a journal article. (2 marks)

  A study was performed at Harvard University by a team of researchers that are experts in organic chemistry. Their data, which was published in a 2018 issue of the journal *Nature Medicine*, described a novel pain medication. This work was an extension of previous research conducted by a group of scientists at Oxford University. The article provides detailed methodological information for other chemists to use to replicate the formulation of this medicine.

- **d** Based on your CRAP analysis in parts **b** and **c**, what can you conclude about the reliability of these sources of information? (2 marks)

14 Green chemistry practices are becoming increasingly common in the design of new chemical products and manufacturing processes.

- **a** Explain what is meant by the term 'green chemistry'. (1 mark)
- **b** List two ways in which a chemical manufacturing business could benefit from employing green chemistry practices. (2 marks)
- **c** One of the green chemistry principles is 'Design for energy efficiency'.
  - **i** Provide a description of this green chemistry principle. (1 mark)
  - **ii** What conditions could be used for a synthetic process to comply with this principle? (2 marks)
- **d** Which of the green chemistry principles are demonstrated in the following scenarios? Note that more than one principle may apply in some cases.
  - **i** The synthesis of an active pharmaceutical ingredient utilises an enzymatic process that eliminates the need for a metal catalyst and improves the overall product yield.
  - **ii** A company makes food containers using a method where microorganisms convert corn starch, sourced from agricultural waste, into a polymer. These food containers are biodegradable and designed to replace those typically formed from petroleum-based plastics.
  - **iii** Water-based acrylic paints are produced from recycled plastic bottles and soybean oil. Unlike oil-based paints, these paints do not give off volatile organic compounds as they dry and do not require petroleum-based solvents in their synthesis. (3 marks)
- **e** Moving from a linear economy to a circular economy can help to conserve resources that are at risk of becoming scarce. Draw a diagram that illustrates a simple circular economy. (3 marks)

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Unit 1 Revision exercise

Multiple-choice questions

1 A molecular substance such as methane, CH_4, has a low boiling temperature, −162°C, because
- **A** there are only weak dispersion forces between the methane molecules.
- **B** there are only weak dispersion forces within the methane molecules.
- **C** there are dipole–dipole forces acting between the methane molecules.
- **D** methane has a low molecular mass.

2 The gradual changes in boiling temperatures in families of hydrocarbon molecules can be explained by
- **A** ionic bonding increasing in strength as the size of the molecules increases.
- **B** hydrocarbons being polar molecules that boil at increasing temperatures as the size of the molecules increases.
- **C** hydrocarbons being non-polar molecules that boil at decreasing temperatures as the size of the molecules increases.
- **D** intermolecular dispersion forces increasing in strength as the size of the molecules increases.

3 A precipitate forms when barium chloride solution reacts with potassium sulfate solution. Which of the following are spectator ions in this reaction?
- **A** barium and chloride ions
- **B** potassium and chloride ions
- **C** potassium and sulfate ions
- **D** barium and sulfate ions

4 The atom of an element has the final ground state electron configuration of $[Ar]3d^1 4s^2$. Which of the following statements would describe the properties of this element?
- **A** It does not conduct electricity as a solid and has a low electronegativity.
- **B** It conducts electricity as a solid and has a low electronegativity.
- **C** It conducts electricity as a solid and has a high electronegativity.
- **D** It does not conduct electricity as a solid and has a high electronegativity.

5 How many lone pairs in total does the nitrogen molecule contain?
- **A** one
- **B** two
- **C** three
- **D** four

6 Which of the following contains both covalent and ionic bonds?
- **A** CH_3OH
- **B** Na_2CO_3
- **C** Al_2O_3
- **D** KCl

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Questions 7 and 8 are based on information provided in the following table.

Substance	Melting point (°C)	Boiling point (°C)	Electrical conduction of substance		
			In solid state	Molten liquid	As solution in water
P	–25	144	Poor	Poor	Insoluble
Q	–51	–35	Poor	Poor	Good
R	1453	2835	Good	Good	Insoluble
S	712	1412	Poor	Good	Good

7 Which substance, P, Q, R or S, would be a gas at room temperature?

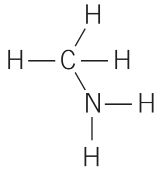
A substance P
B substance Q
C substance R
D substance S

8 Which of the substances has a structure made up of cations and anions?

A substance P
B substance Q
C substance R
D substance S

9 Consider the compounds shown.

I



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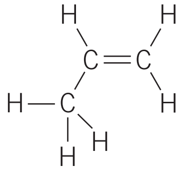

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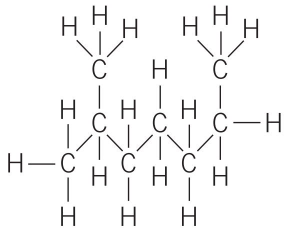
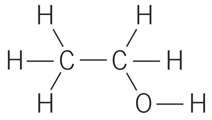


IV



Which of the following statements regarding these compounds is **incorrect**?

A Compounds II and III are non-polar.
B Compounds I and IV would be involved in hydrogen bonding.
C Compound I would have the lowest boiling point.
D Compound III would have larger dispersion forces than compound II.

10 Which of the following is a structural isomer of pentane?

A 2-methylpropane
B 2-ethylpentane
C 2,3-dimethylpropane
D 2,2-dimethylpropane

11 The percentage by mass of copper in chalcopyrite ($CuFeS_2$) would be

A 30.4%
B 34.6%
C 35.0%
D 41.9%

12 What is the amount of potassium sulfide, K_2S, present in 5.0 g of the substance?

A 0.045 mol
B 0.070 mol
C 22.1 mol
D 1.1×10^2 mol

13 Which of the following contains the lowest mass?

A 1.0 mol of sodium
B 0.50 mol of oxygen gas
C 17.6 g of $CuSO_4$
D 3.5 mol of $BaCl_2$

14 The list containing only empirical formulas is

A H_2O, $C_2H_4O_2$, HCl, $Na_2Cr_2O_7$
B H_2SO_4, H_2O, $Al(NO_3)_3$, $Al_2(SO_4)_3$
C HCl, $C_6H_{12}O_6$, NH_3, $Al(NO_3)_3$
D $Al_2(SO_4)_3$, O_2, HCl, H_2SO_4

15 Isomers are

A molecules with the same empirical formula but different arrangements of the atoms in the molecule.
B molecules with the same molecular formula but different arrangements of the atoms in the molecule.
C molecules with the same empirical formula but different molecular formulas.
D molecules that contain atoms that have the same atomic number but different mass numbers.

16 A sample of ethane, C_2H_6, contains 9.6×10^{22} atoms. How much ethane, in mol, is present in the sample?

A 0.020 mol
B 0.080 mol
C 0.027 mol
D 0.16 mol

17 The shape of a molecule is determined by

A the number of electrons lost and gained as ions are formed in the molecule.
B only the non-bonding electrons present in the molecule.
C both the bonding and non-bonding electrons in the atoms present in the molecule.
D only the bonding electrons present in the molecule.

Short-answer questions

1 Germanium has five naturally occurring isotopes, as shown in the table below.

Isotope	Percentage abundance
^{70}Ge	23.20
^{72}Ge	27.70
^{73}Ge	7.70
^{74}Ge	35.90
	7.40

a Using the periodic table provided in this textbook, determine the relative isotopic mass of the heaviest isotope not identified in the table. (3 marks)

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A sample of germanium weighs 5.34 g.

b What amount, in mol, of germanium is present in the sample? (1 mark)

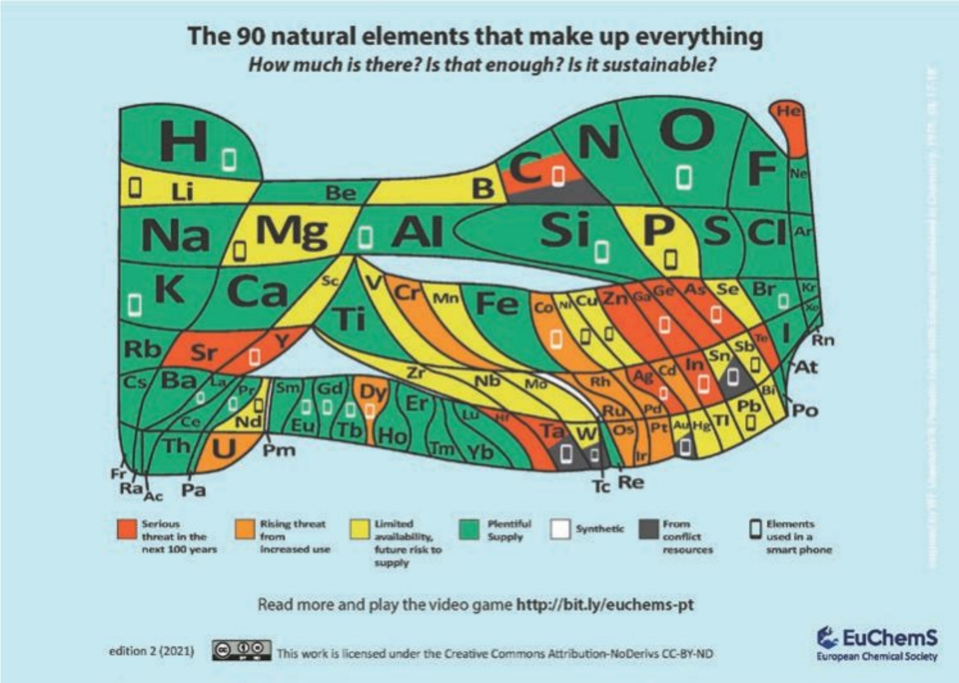
c How many atoms are present in this sample? (1 mark)

Germanium is classified as a metalloid. As germanium often combines with oxygen to form germanium oxides, it is mined in a way very similar to most post-transition metals using smelting.

d Identify one physical property that germanium has that confirms why it is not classified as a metal. (1 mark)

e Outline the basic steps involved in the smelting process. You may use a diagram as your answer. (2 marks)

According to the following image, germanium is listed as being in serious risk of running out in the next 100 years.



In the early 2000s, approximately 76% of germanium was recycled from a range of products and used in applications, such as night vision devices, fibre optics, solar cells and food containers made of plastic polymers.

f Using this information, explain what impact the absence of germanium could have in our society and how society may (or has) overcome this. (3 marks)

g Write both the shell (Bohr model) and subshell (Schrödinger model) electron configurations for germanium. (2 marks)

2 Hydrocarbons are a group of compounds containing hydrogen and carbon atoms covalently bonded to each other. One family of hydrocarbons are the alkenes, represented by the general formula C_nH_{2n}.

a Draw and name the valence structure for the members of this alkene family containing three, four and five carbon atoms respectively. (3 marks)

The two-carbon alkene, ethene, undergoes an addition reaction with chlorine gas, Cl_2.

b Write a structural chemical equation for this reaction, clearly representing and naming the product formed. (3 marks)

The production of polyethene involves a different addition reaction between ethene and other ethene molecules.

c Outline the meaning of the terms 'monomer' and 'polymer' in relation to polyethene. (2 marks)
d Draw a short section of polyethene containing at least six carbon atoms. (1 mark)

While hydrocarbons contain atoms of carbon and hydrogen bonded to each other, carbon and hydrogen can also separately bond to many other atoms. Two of these molecules are carbon dioxide (CO_2) and ammonia (NH_3). Carbon dioxide sublimes (changes directly from a solid to a gas) at −78°C, whereas ammonia boils at −33°C.

e Redraw and then complete the table below. (4 marks)

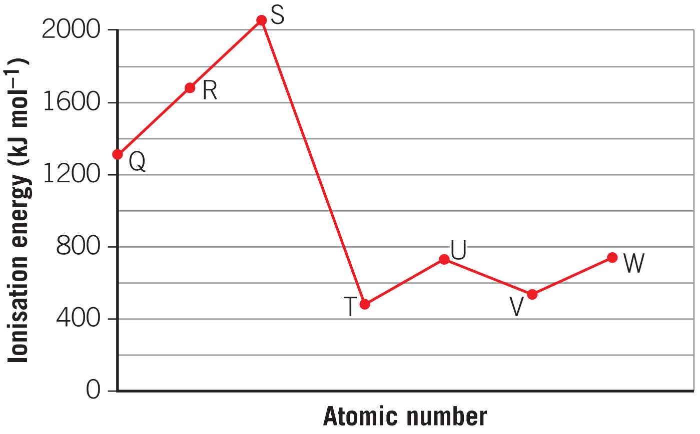
	Carbon dioxide (CO_2)	Ammonia (NH_3)
Structural diagram		
Name of shape		

f Explain why the boiling point of ammonia is much higher than the sublimation temperature of carbon dioxide. (2 marks)

One of these substances is polar and the other is non-polar, yet both are composed of polar bonds.

g With reference to these two molecules, explain why both have polar bonds and determine which is polar and which is non-polar. Also make sure you reference instantaneous and permanent dipoles in your answer. (4 marks)

3 The first ionisation energy of seven consecutive elements in the periodic table are shown in the graph below. The elements are identified by letters that have no relation to the correct elemental symbols.



a Define 'first ionisation energy'. (1 mark)
b Account for the large difference in ionisation energy between elements 'S' and 'T'. (2 marks)
c Would you expect to see the same trend for these elements in electronegativity? Explain. (2 marks)

4 Copper is a tough yet soft reddish transition metal. In its pure form, it has a high ductility and electrical conductivity. Copper is also often found combined with other elements (minerals) that are contained in rock samples, known as ores. A range of processes is used to extract various metals from their ores. One of these processes uses electricity to melt and electrolyse (separate or break down into ions) ore containing copper oxide to allow copper metal to be formed.

a Compare the structure and bonding in solid copper oxide and copper metal. Use fully labelled diagrams to support your written response. (4 marks)

b Using your diagrams from part **a**, explain how each supports the different properties for copper oxide and copper. (6 marks)

Copper can form different charged cations, a each of which can combine with oxygen.

c Redraw the table below and complete the gaps. (2 marks)

Name of ionic compound	Formula of ionic compound
Copper(I) oxide	
	CuO

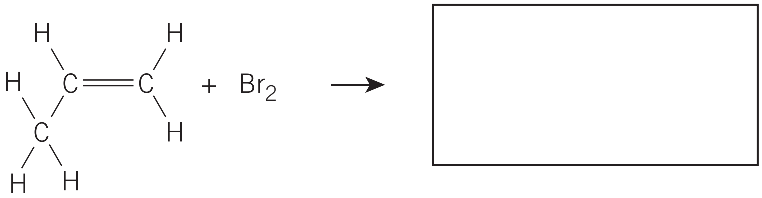
d The move from a linear to a circular economy for metals is becoming increasingly important to prevent metals from being exhausted or unnecessarily ending up in landfill. Using copper as an example, explain the difference between the stages of a linear and circular economy, and how copper may be re-purposed. (4 marks)

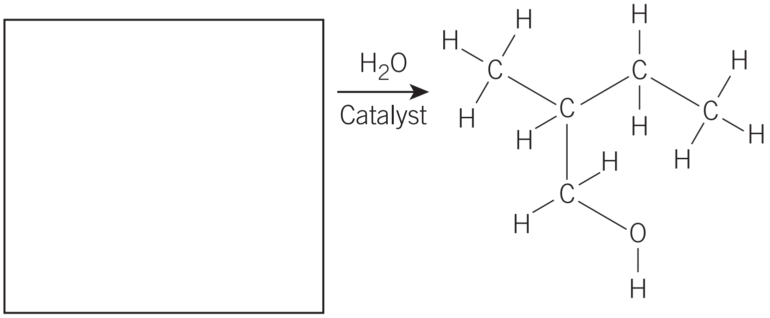
5 Organic halogen compounds are a large class of natural and synthetic chemicals that contain one or more halogens in combination with carbon and other elements. One of these is the compound 1-fluorobutane, which has the chemical formula C_4H_9F. There are **three** other structural isomers for 1-fluorobutane.

Draw the three isomers and give the systematic names for each of these molecules.

(6 marks)

6 Complete the reactions below by redrawing them and completing the missing gaps with the appropriate molecule(s).





(3 marks)

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7 Phosphorus is mined and then prepared industrially by heating calcium phosphate ($Ca_3(PO_4)_2$) with silicon dioxide (SiO_2) and carbon.

a What is the molar mass of calcium phosphate, $Ca_3(PO_4)_2$? (1 mark)

b What is the number of calcium ions present in 382 g of the substance? (3 marks)

Phosphate ions are generally insoluble in water, with very few exceptions. Calcium phosphate is an insoluble substance.

c Construct two hypothetical word and full balanced chemical precipitation reactions whereby calcium phosphate would be produced from the mixture of two different aqueous solutions. Ensure that you use a different combination of species and ions for each reaction. (4 marks)

d For each of the chemical equations you wrote in part **c** above, write the ionic equation and explain why you have not included some species. (4 marks)

Approximately 80% of phosphorus is lost in the supply chain from when it is mined to when it reaches you in your food, as most phosphate rock is used in food production. The large proportion of this percentage is lost to water supplies during agricultural runoff and in wastewater. It is estimated that phosphorus resources will be depleted by around 2080 due to the increase in global demand from developing countries for growing their own food to meet increasing population size.

e Identify one social and one environmental action that could be undertaken to slow the rate of phosphorus use. (2 marks)

8 Chromatography is a technique used to separate different components of a solution depending on their solubility. A student is conducting an experiment to determine the different components within a mixture.

a The student is required to draw a diagram of the setup of their thin-layer chromatography (TLC) experiment but has lost their diagram. Draw and fully label a diagram of the set-up for them. (4 marks)

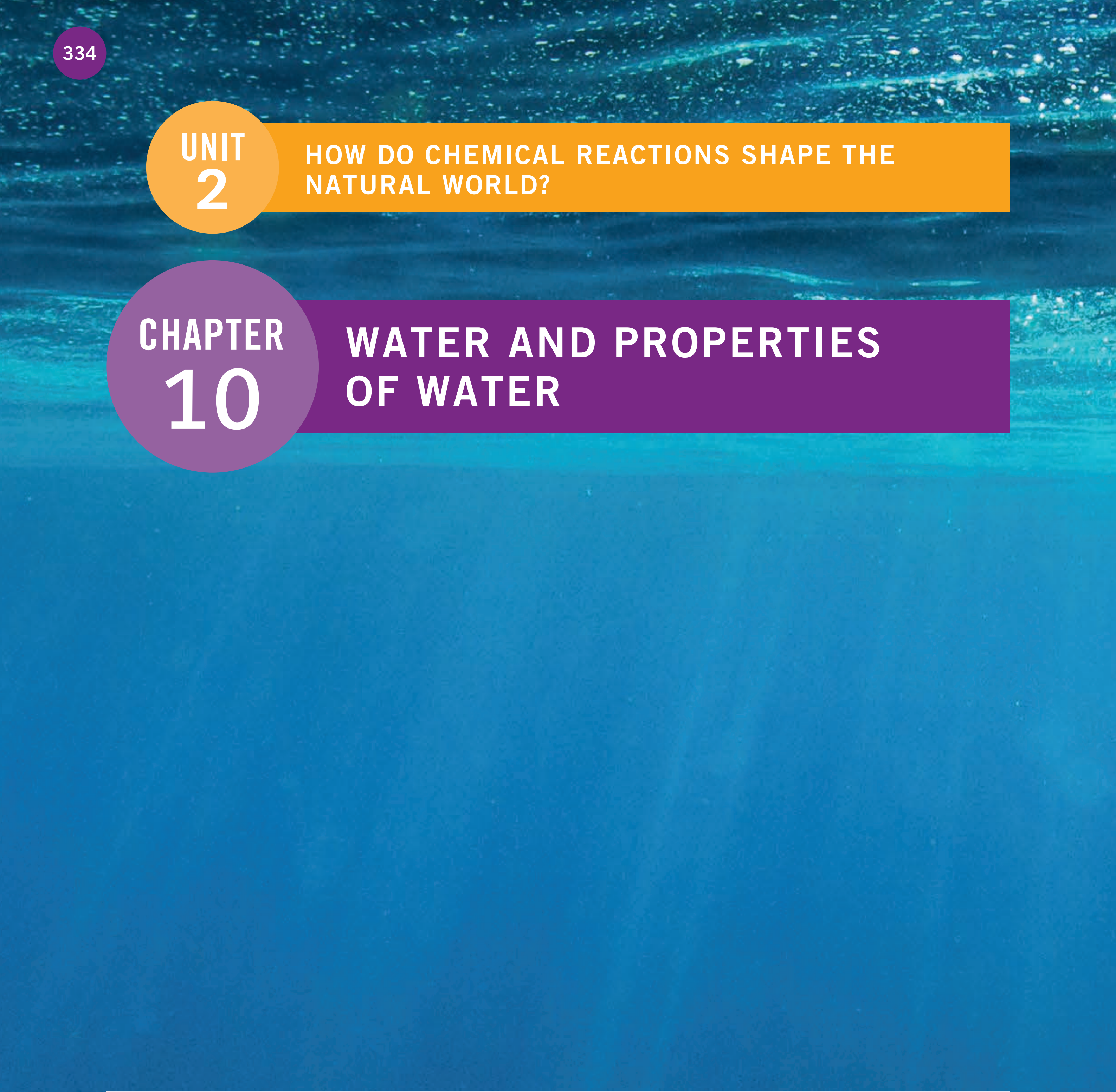
b If the TLC plate/gel is polar, explain how both substances that are either polar and non-polar would be separated on this plate/gel. (4 marks)

c If the mixture that the student is aiming to separate contains an alkane, an alcohol and a carboxylic acid, and the TLC plate/gel is polar, explain how would you expect the three components of the mixture to separate. (The three components are all of similar molecular mass.) (5 marks)

d With reference to what the R_f value is, explain how the identity of any of these components within the mixture can be determined. (3 marks)

UNIT 2
HOW DO CHEMICAL REACTIONS SHAPE THE NATURAL WORLD?

CHAPTER 10
WATER AND PROPERTIES OF WATER



Introduction

There is a lot of excitement and discussion around our ability to colonise other planets in our solar system. The obvious choice, based on its proximity to Earth, is Mars. One of the biggest challenges to overcome for this to happen is having a readily available source of water. While water has been discovered on Mars, it is present as ice. This is due to the fact that the average surface temperature on Mars is around –65°C. While this ice could be heated to provide liquid water, it would require large amounts of energy to do so. In this chapter, we are going to learn about the different states of water, as well as some of the special properties that water has.

INTRODUCTION VIDEO
WATER AND PROPERTIES OF WATER

Curriculum

Area of Study 1 Outcome 1
How do chemicals interact with water?

Study Design:	Learning intentions – at the end of this chapter I will be able to:
• The existence of water in all three states at Earth's surface, including the distribution and proportion of available drinking water	**10A Water on Earth** **10A.1** Compare the three states of water with reference to the arrangement of water molecules and bonds between them **10A.2** Describe the distribution of drinking water on Earth **10A.3** Explain why only a very small proportion of water on Earth is available as 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	**10B Properties of water** **10B.1** Describe the trend in boiling points observed for group 16 hydrides **10B.2** Explain why water doesn't follow the trend of boiling points observed for group 16 hydrides **10B.3** Explain the difference in density of ice and liquid water with reference to hydrogen bonding **10B.4** Define 'specific heat capacity' **10B.5** State the unit for specific heat capacity **10B.6** Explain why water has a relatively high specific heat capacity
• The relatively high latent heat of vaporisation of water and its impact on the temperature regulation of the oceans and aquatic life	**10B.7** Define 'latent heat of vaporisation' **10B.8** Describe how the high latent heat of vaporisation regulates ocean temperatures and allows for the survival of aquatic life

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Glossary

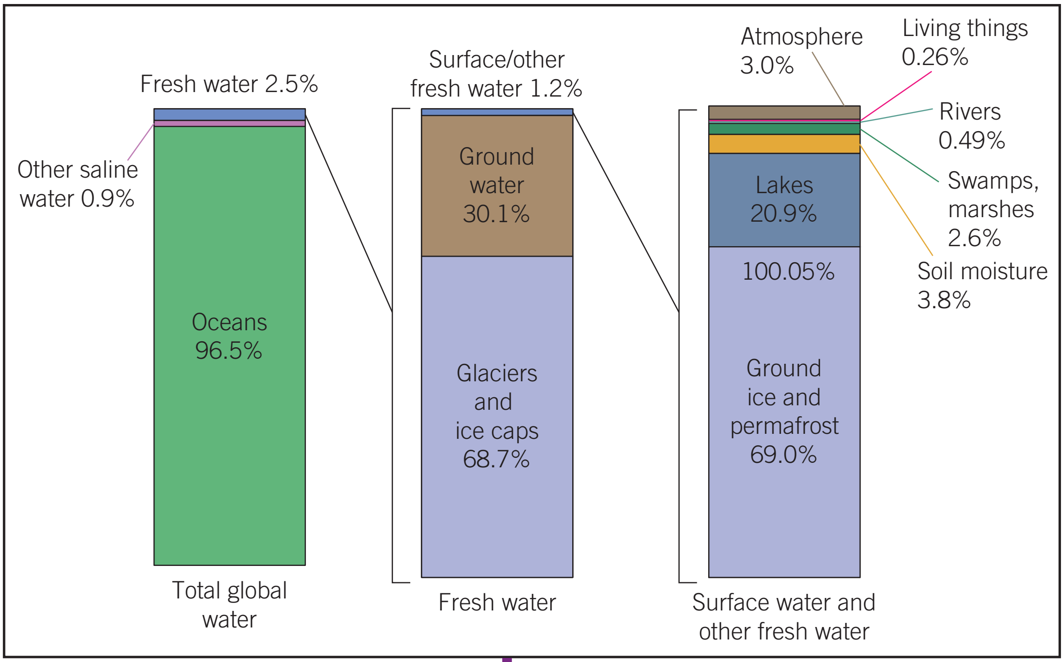
Dipole–dipole attraction
Dispersion force
Kinetic energy
Latent heat of vaporisation
Polar molecule
Specific heat capacity

Concept map

Chapter 10A Water on Earth

Solid
Liquid
Gas

Fresh water
Salt water



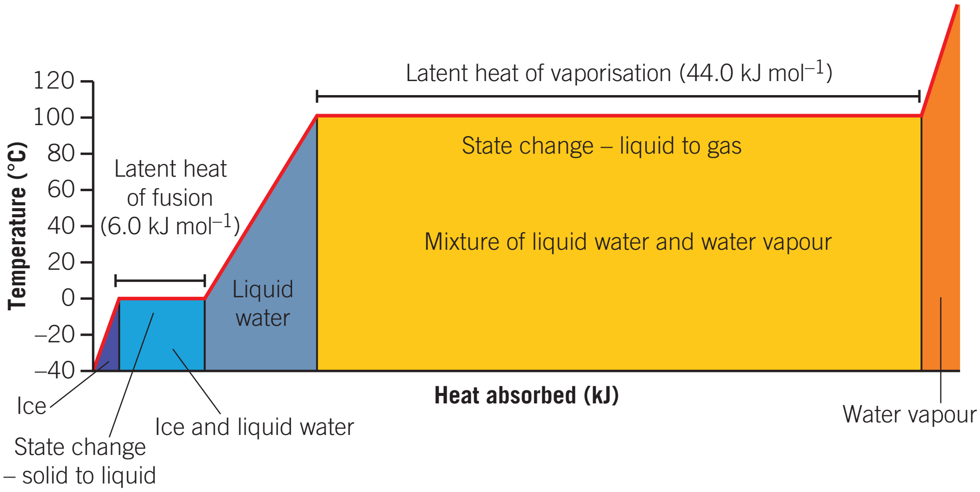
Chapter 10B Properties of water

Boiling points of hydrides

Density of ice and water

Specific heat capacity

Latent heat of vaporisation of water



See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

10A Water on Earth

Study Design:
The existence of water in all three states at Earth's surface, including the distribution and proportion of available drinking water

Glossary:
Kinetic energy

ENGAGE

Water is the most abundant liquid on Earth

When a block of ice melts in your drink, you don't think anything of it. However, when this happens on a much larger scale, it can be incredibly interesting and informative to scientists. Icebergs, blocks of ice that are greater than five metres in size, float in the oceans surrounding the Arctic and Antarctica. As these icebergs drift away from the poles into warmer water, they begin to melt and break up. This process is interesting to climate scientists, as it helps them better understand the impact that global warming is going to have. It has also been observed that the melting of icebergs releases not only liquid water back into the ocean but also many nutrients. For this reason, biologists enjoy studying melting icebergs, as they have found that there is an increase in the aquatic life found surrounding these melting blocks of ice. In this section, we are going to learn about the different states of water, as well as how it is distributed across the planet and what amount is available to drink.

Figure 10A–1 The melting of icebergs provides a lot of information to scientists in many different research areas.

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EXPLAIN

The three states of water

Water is one of the few compounds that we commonly encounter in all three states of matter. We refer to the liquid state as water, the solid state as ice and the gaseous state as water vapour.

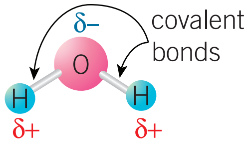


Figure 10A–2 Water is a polar molecule with partial positive charges on the hydrogen atoms and a partial negative charge on the oxygen atom, and intramolecular covalent bonds between the oxygen and hydrogen atoms.

2A REPRESENTING COVALENT MOLECULES LINK

It is important to realise that in each state there is no change to the compound: water, ice and water vapour are all composed of H_2O molecules. Although you have come across it in Chapter 2 already, the structure of water is shown in Figure 10A–2, indicating the covalent bonds holding the oxygen and hydrogen atoms together.

If there is no change to the bonds or atoms within a water molecule, how do these three states all look and behave so differently? To answer this, we must look at what is happening at the atomic level, as shown in Figure 10A–3.

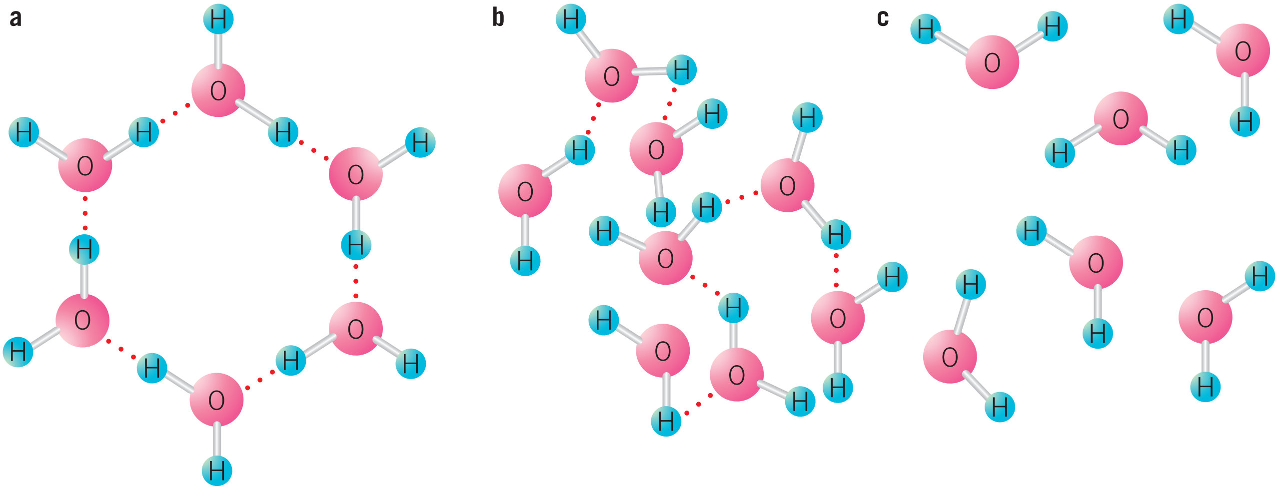


Figure 10A–3 Water molecules arrange very differently in **a** ice **b** liquid water and **c** water vapour. The red dotted lines indicate hydrogen bonds.

2C PROPERTIES OF COVALENT COMPOUNDS DETERMINED BY INTERMOLECULAR FORCES LINK

Kinetic energy the energy an object has due to its motion

In ice (solid state of water), the water molecules are arranged in a highly organised structure. In fact, if you study it closely, you will see that all the water molecules in a block of ice are arranged in a hexagonal pattern with hydrogen bonding occurring between the water molecules. As ice occurs and is formed at low temperatures, these water molecules only have a small amount of **kinetic energy**. This means that the water molecules will vibrate in place but are not moving significantly enough to alter this structure, so it maintains a rigid structure and holds its shape.

Liquid water, on the other hand, behaves very differently. We know that if you fill a container with liquid water, the liquid will adopt the shape of the container it is in. This is because in liquid water, the water molecules have more kinetic energy than they do in ice. This allows them to move further away from each other and break away from that hexagonal pattern. Despite this, hydrogen bonding still exists between the water molecules; it is just that these forces of attraction are constantly breaking and re-forming as the water molecules move. As such, water doesn't have a distinct structure and exists as a fluid.

WORKSHEET 10A–1 THE THREE STATES OF WATER

LINK **15A** GASES AND THE IDEAL GAS EQUATION

As the temperature increases even further, water will vaporise and turn into gaseous water (water vapour). At these temperatures, the water molecules have relatively high kinetic energy and are able to completely break the intermolecular hydrogen bonds between them. As with other gases, water vapour will expand to fill the space that it is in. This means that it also has a variable volume and can be condensed.

Check-in questions – Set 1

1 What are the names given to solid and gaseous water?
2 What is the name of the shape that water molecules arrange themselves into in ice?
3 Identify the name of the force that holds water molecules together in ice, water and water vapour.

Water on Earth

Earth is covered by oceans, lakes, rivers and seas. In fact, around 71% of Earth's surface is water, but very little of it is available as fresh drinking water. So why is this the case?

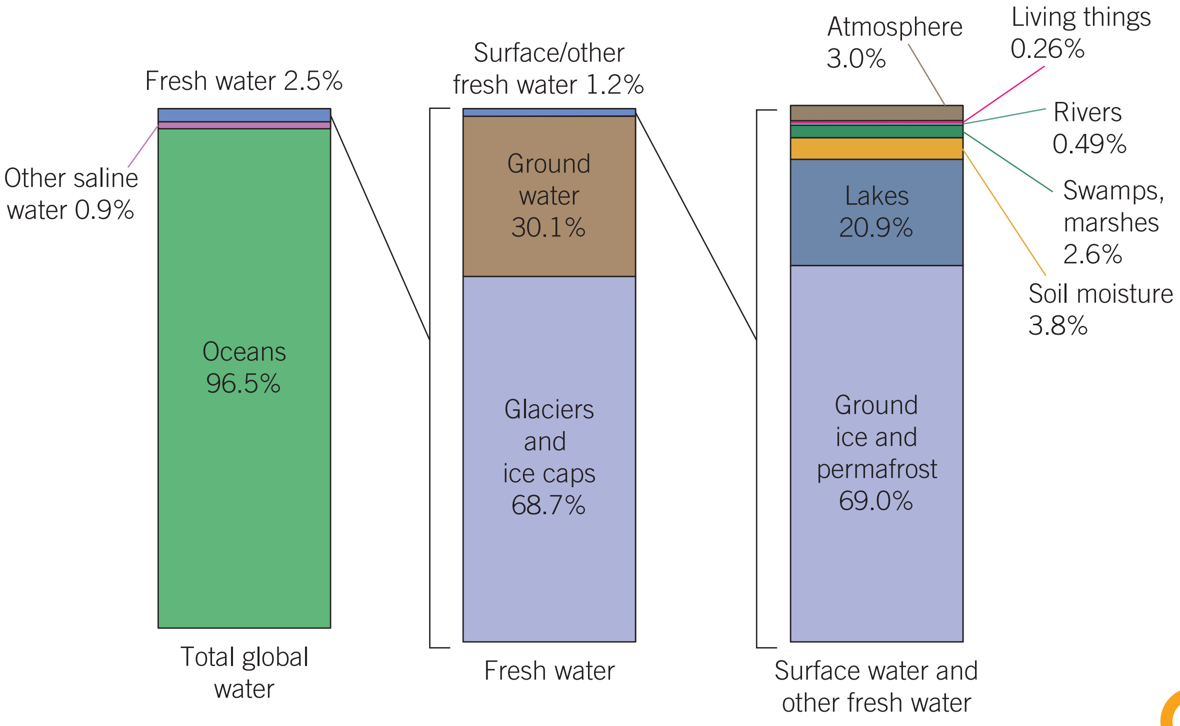


Figure 10A–4 Although Earth is covered with water, very little is available as fresh drinking water.

VIDEO 10A–1 WATER ON EARTH

As you can see from the left-hand column in Figure 10A–4, most of the total global water present on Earth contains salt and is found in oceans. Drinking salt water, especially at the concentrations found in the ocean, is incredibly harmful to human health. If too much is ingested, it can lead to organ failure and death. Therefore, this water is not safe to drink without first being properly desalinated.

NOTE

Analysis of the substances in water samples is one potential option for a practical investigation you or your school may choose to investigate further in Area of Study 3 of Unit 2.

Looking at the middle column in Figure 10A–4, you can see that only 31.3% of the fresh water on Earth exists as liquid water.

For the 1.2% present on the surface, there are other issues that prevent it all from being suitable as drinking water, as shown in the third column. As well as being in locations that are difficult to extract drinking water from (for example, water that is present in soil moisture), other sources contain many impurities or pollution that can be harmful to human health.

1D CRITICAL ELEMENTS AND RECYCLING PROCESSES LINK

VIDEO 10A–2 SKILLS: USING DIAGRAMS TO ANSWER QUESTIONS

10A SKILLS

Using diagrams to answer questions

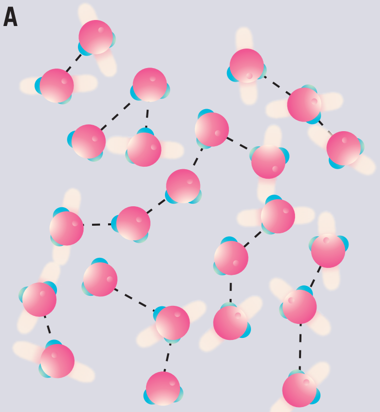
When providing examples of the different states of water, it is important to refer to specific features of the arrangement of water molecules in each as part of your justification. A key feature of water in the solid state (ice) is the presence of water molecules in fixed positions yet spread out. Each water molecule is also attracted to four other water molecules. This shouldn't be confused with water in the liquid state, where water molecules can be drawn both close together and further apart, as water in this state fills the shape of the container it is in. Therefore, it is important to remember how to represent water molecules in each state, as shown in Figure 10A–3 earlier in this section.

You could be asked to **identify** which of the examples represents water in the solid, liquid or gas state. When a question asks you to identify something, there is no need to include a justification or explanation. All you need to do is put the example into a category or label it as one of the three states, as in the following example.

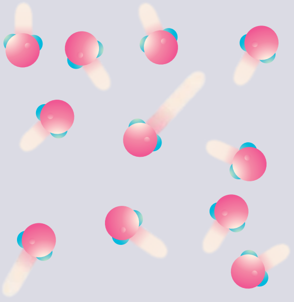
Question:

Identify which of the diagrams shows water molecules in the gaseous state.

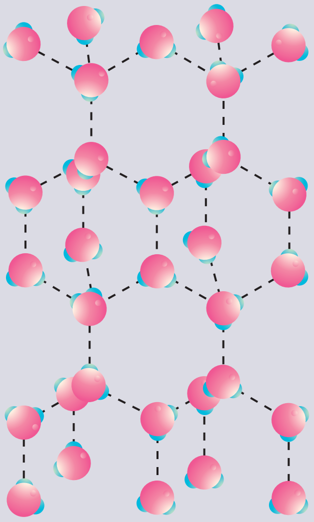
A



B



C



Answer:

B.

As you can see, the answer consists of just the letter associated with the correct diagram.

This is in contrast to a question that asks you to **justify** your choice, as in the following example.

Question:

Provide a justification for why you have selected the diagram chosen as representing water molecules in the gaseous state.

Answer:

Diagram B shows water molecules spaced out and in various positions, making it seem as though water molecules are moving very quickly, which they do at high temperatures due to an increase in kinetic energy. The other two diagrams, A and C, show dashed lines between water molecules. These dashed lines represent hydrogen bonds, which are more likely to be present in the liquid and solid states as the water molecules have less kinetic energy. Therefore, the gaseous state is represented by diagram B.

Note that the answer does not repeat the question, and it gives a specific piece of evidence as the reason for the choice. The response also finishes by coming back to the context of the question, restating that the option that represents water in the gaseous state is diagram B.

Section 10A questions

1. Compare the distribution of fresh water and salt water on Earth.
2. Describe the proportion of fresh water on Earth that is readily available as drinking water.
3. List three freshwater sources on or above Earth's surface.
4. Describe the arrangement of covalent bonds and hydrogen bonds in gaseous atmospheric H_2O.
5. Contrast the arrangement of molecules and their hydrogen bonding in liquid H_2O, found in the ocean, lakes and rivers, and solid H_2O found in permafrost and glaciers.
6. Explain why water is a polar molecule.
7. Outline how hydrogen bonding occurs between different molecules by referring to the molecular structure of water. Use a fully labelled diagram to support your answer.

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10B Properties of water

Study Design:
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 temperature regulation of the oceans and aquatic life

Glossary:
Dipole–dipole attraction
Dispersion force
Latent heat of vaporisation
Polar molecule
Specific heat capacity

ENGAGE

Water is crucial for life

When we exercise, we perspire. While it may not seem that useful to get soaking wet (and probably quite smelly) when we are running, the perspiration is serving a very important purpose. Perspiring is one of the mechanisms that our bodies have to maintain a constant body temperature, even when we are heating it up through activities such as exercise. As the perspiration evaporates off the skin, it leads to a decrease in body temperature. The reasons for this have to do with the amazing properties of water, which we will be discussing in this section.

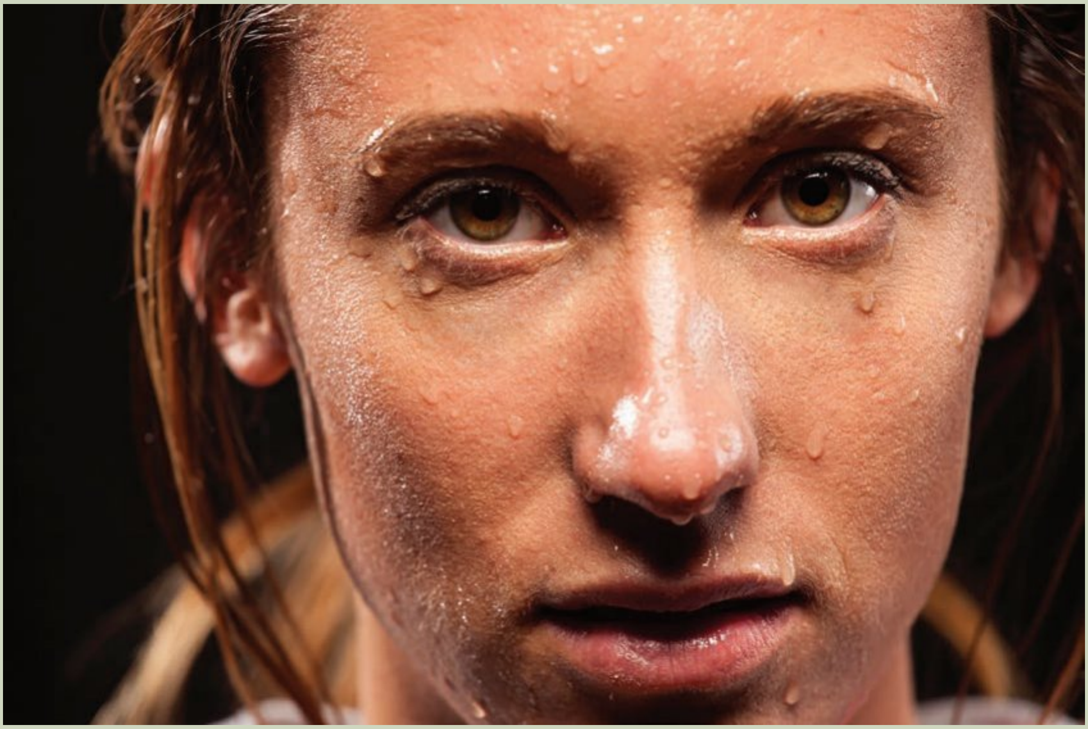

Figure 10B–1 The way that perspiration cools us down during exercise highlights just one of the many amazing properties of water.

ISBN 978-1-009-22964-7 Drummond et al

EXPLAIN

Properties of water

Despite just being an oxygen atom covalently linked to two hydrogen atoms, water has some unique and important properties. In this section, we are going to look at a number of ways in which water differs from other molecules. We will be doing this by examining some specific examples: the boiling points of group 16 hydrides (molecules with group 16 atoms covalently bonded to two hydrogen atoms), the density of liquid water and ice, and the high specific heat capacity of water.

Boiling points of group 16 hydrides

Each of the group 16 elements, which include oxygen (O), sulfur (S), selenium (Se), tellurium (Te) and polonium (Po) (Figure 10B-2), can covalently bond with hydrogen to form compounds known as hydrides.

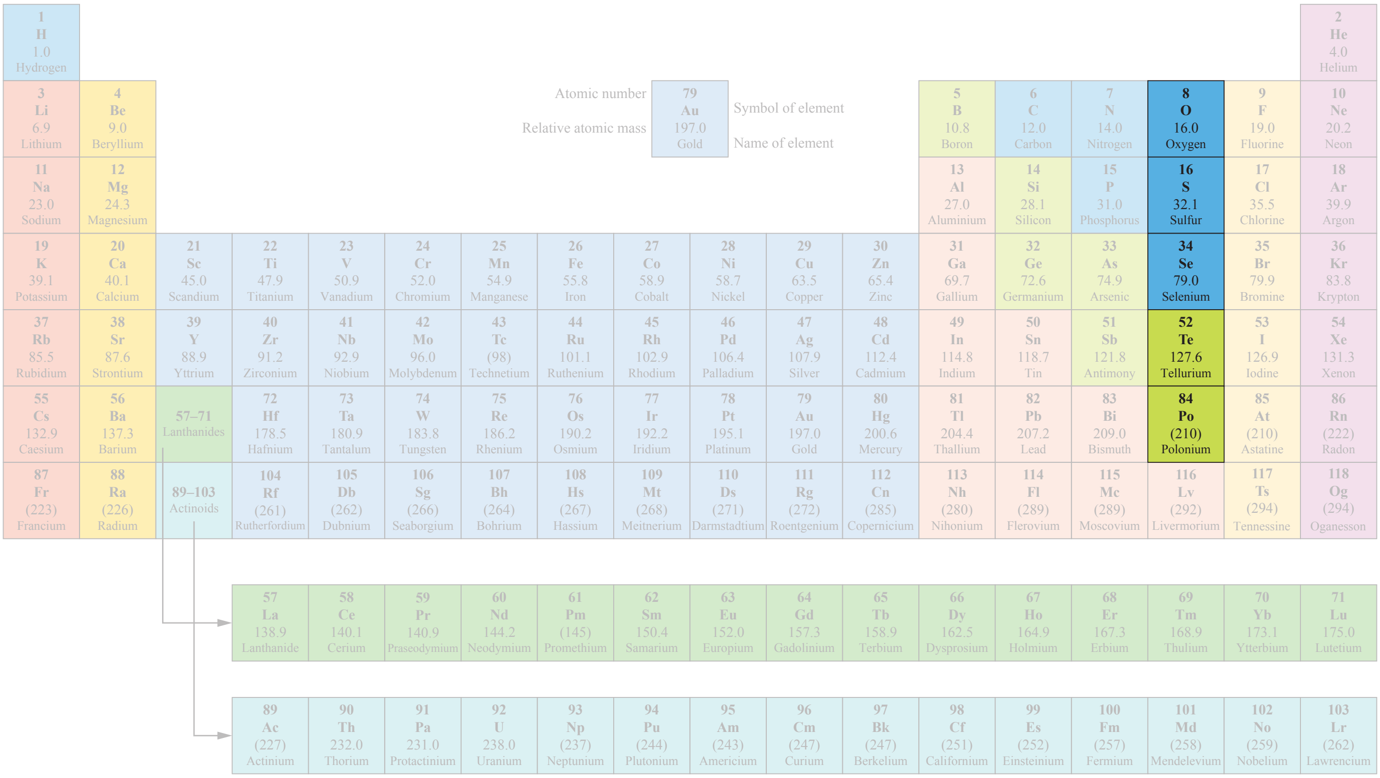

Figure 10B–2 The group 16 elements in periods 2–6

Figure 10B–3 shows a graph of the trend in boiling points for four of the group 16 hydrides: H_2O, H_2S, H_2Se and H_2Te. There are two main observations that we can make from this graph.

1 If we look at the trend for those hydrides with elements in period 3 to period 5, we can see that there is an increase in boiling point as we move down the group.
2 Water, H_2O, does not follow this trend and has a significantly higher boiling point than any of the other group 16 hydrides.

Polar
a molecule with two distinct poles of partial electric charge, a partial positive charge on one side of the molecule and a partial negative charge on the other side

Dipole–dipole attraction
an intermolecular force present between polar molecules that have permanent dipoles

Dispersion force
a weak intermolecular force present in all molecules created between instantaneous dipoles

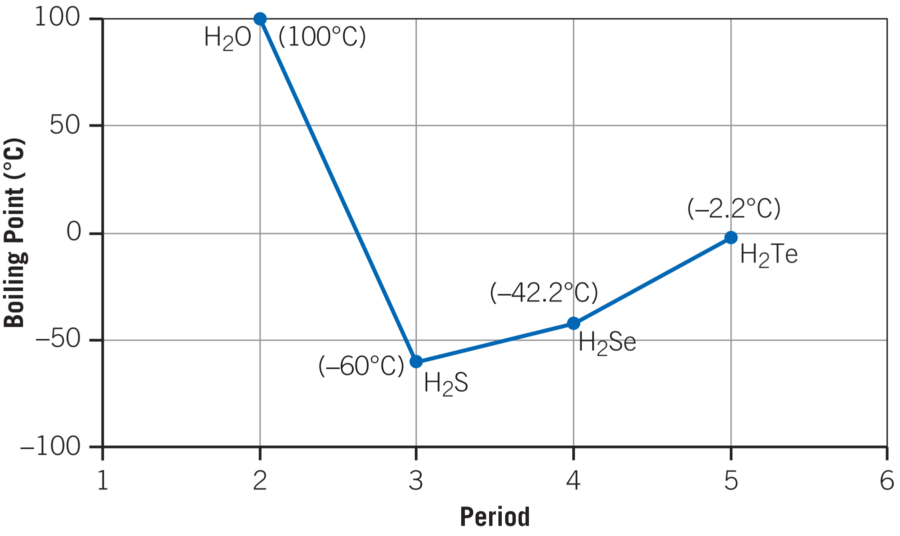

Figure 10B–3 The boiling points for the group 16 hydrides across periods 2–5

How do we explain these observations?

Firstly, it is important to note that all of these molecules are 'bent' molecules, as shown in Figure 10B–4.

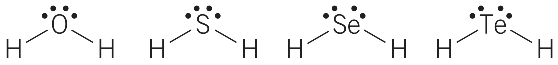

Figure 10B–4 Lewis structures for the group 16 hydrides have the same 'bent' molecular shape.

1C THE PERIODIC TABLE **LINK**

As they all have this shape, they are also all **polar** molecules. However, there is a big difference in how strongly polar these molecules are. Water, H_2O, is highly polar due to the large electronegativity difference that exists between the oxygen and hydrogen atoms. H_2S, H_2Se and H_2Te are all less polar than water. This is because there is a smaller electronegativity difference between the atoms that make up those molecules compared with those that make up water.

2C PROPERTIES OF COVALENT COMPOUNDS DETERMINED BY INTERMOLECULAR FORCES **LINK**

Having established the polarity of these molecules, we can determine the types of intermolecular forces that influence the boiling points of these molecules. The hydrides from periods 3–5 (H_2S, H_2Se and H_2Te) have both **dipole–dipole attractions** and **dispersion forces** holding together their molecules. The dispersion forces increase with increasing molecular mass. This increase in dispersion forces explains why H_2Te has a higher boiling point than both H_2Se and H_2S.

Water, on the other hand, has the lowest molecular mass of any of these molecules and therefore the weakest disperson forces. Why then does it have a significantly higher boiling point than any of the hydrides in group 16 and goes completely against the observed trend? The answer is that water molecules are capable of forming hydrogen bonds. As a reminder, hydrogen bonds can only be formed between hydrogen atoms and either a nitrogen, oxygen or fluorine atom of another molecule. This is why water is the only group 16 hydride that is capable of making these strong intermolecular forces. As these hydrogen bonds are stronger than both dipole–dipole attractions and dispersion forces, they require much more energy to break.

Density of ice and water

Another interesting and unique property of water is that the solid state, ice, has a lower density than the liquid form, water. It is for this reason that icebergs float in the ocean, instead of sinking to the sea floor. The explanation for this anomaly can be found in the information presented at the start of Chapter 10A, where we looked at the differences in atomic structures of ice, water and water vapour (Figure 10A–3). In ice, the hydrogen bonding between water molecules is arranged into a hexagonal pattern. This arrangement results in a large amount of space between the molecules, leading to a low density. In contrast, liquid water does not have this hexagonal arrangement, so the molecules are more tightly packed together with less space between them, leading to a higher density.

LINK **10A** WATER ON EARTH

Figure 10B–5 Unlike most other molecules, solid water (ice) is less dense than liquid water, which is why icebergs float in the ocean.

Specific heat capacity

Another property of water that makes it unique is that it has a relatively high specific heat capacity. **Specific heat capacity** is defined as the amount of energy required to raise the temperature of one gram of a substance by one degree Celsius.

Therefore, when we say that water has a high specific heat capacity, we mean that it requires a large amount of energy to raise its temperature. In fact, the specific heat capacity of water is 4.18 $\text{J g}^{-1}\,°\text{C}^{-1}$. This means that it requires 4.18 joules of energy to heat one gram of water by 1°C. To put this in perspective, cooking oil has a specific heat capacity of around 2.2 $\text{J g}^{-1}\,°\text{C}^{-1}$ and sand has a specific heat capacity of 0.48 $\text{J g}^{-1}\,°\text{C}^{-1}$. This means that it would take almost twice as much energy to heat the same mass of water compared to cooking oil, and 10 times as much energy to heat the same mass of water compared to sand!

VIDEO 10B–1 SPECIFIC HEAT CAPACITY

Specific heat capacity
the amount of energy required to raise the temperature of one gram of a substance by one degree Celsius

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Figure 10B–6 Water requires about twice as much energy to heat as oil and 10 times as much as sand.

The energy required to heat a certain mass of substance by a certain amount is linked to the specific heat capacity, and this relationship can be expressed through the equation

$$q = m \times c \times \Delta T$$

where:

q = The energy in joules (J)
m = The mass in grams (g)
c = The specific heat capacity (J g^{-1} $°C^{-1}$)
ΔT = Change in temperature (°C).

UNIT 3 LINK

It is important to note that this equation is being provided to allow for some extra understanding of these concepts, but you will not use it to perform calculations in Unit 2. Calculations using this equation will form part of your studies in Unit 3.

WORKSHEET 10B–1 UNIQUE PROPERTIES OF WATER

When heat is applied to molecules, their kinetic energy (the energy they have as a result of their motion) increases. It is this increase in kinetic energy that causes the temperature of the substance to increase. However, the molecules in solid and liquid water cannot freely move around due to the strong hydrogen bonds that hold them together. This means that a large amount of energy has to be applied before those strong hydrogen bonds break, which then allows for the molecules to move around with higher levels of kinetic energy. This explains why it takes a lot of energy for the temperature of water to increase, which is why water has a high specific heat capacity.

Check-in questions – Set 1

1. Excluding water, what is the general trend in boiling point for group 16 hydrides?
2. Define ‘specific heat capacity’.
3. What property of water molecules is mostly responsible for it having a much higher boiling point than other group 16 hydrides?

Latent heat of vaporisation of water

10A WATER ON EARTH LINK

As was discussed in Chapter 10A, a vast proportion of the planet is covered by oceans. These oceans are home to an abundance of plant and animal life. For these organisms to survive, they require a relatively constant external environment, especially in relation to temperature. Considering these oceans will be absorbing significant energy from the sun during the day, how is it possible that they don’t heat up in a manner that would cause problems for the life forms that inhabit the water?

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The answer to this question relates to two key properties of water: its high specific heat capacity and high latent heat of vaporisation. You learned about high specific heat capacity earlier in this chapter, and the same concepts can be applied to the oceans. This property of water means that a large amount of energy is required to heat a small mass of water by a small amount. When you consider how vast and deep the oceans are, it is understandable that it would take a huge amount of energy to cause a meaningful change in the temperature.

Figure 10B–7 The oceans are home to many different plant and animal species, which all require a relatively constant water temperature to survive.

As well as a high specific heat capacity, water also has a relatively high latent heat of vaporisation. The **latent heat of vaporisation** of water can be defined as the amount of energy required to convert one mole of liquid water into one mole of gaseous water (water vapour). The unit for the latent heat of vaporisation is $kJ\,mol^{-1}$. Like the specific heat capacity, the high latent heat of vaporisation can be explained by the strong hydrogen bonds that exist between water molecules. The amount of heat energy required to move one mole of liquid molecules into the gaseous state and overcome the net intermolecular force strength is large.

Latent heat of vaporisation the amount of energy required to convert one mole of liquid water into one mole of gaseous water (water vapour)

A high latent heat of vaporisation means that water in the ocean requires a large amount of energy to evaporate and move to the gaseous state. That energy is removed from the surface of the ocean as heat, which reduces the surface ocean temperature. This therefore affects the aquatic temperature for living organisms, including fish and coral life. However, the high latent heat of vaporisation also reduces the possibility of large changes in the temperature of the ocean, and therefore living organisms are protected by a consistent habitual temperature.

VIDEO 10B–2
SKILLS: LATENT HEAT OF VAPORISATION

10B SKILLS

Graphically representing the latent heat of vaporisation

Essentially, latent heat of vaporisation is the energy absorbed in order to change a substance from its liquid state to its gaseous state at its boiling point. Graphically, this is represented in Figure 10B–8.

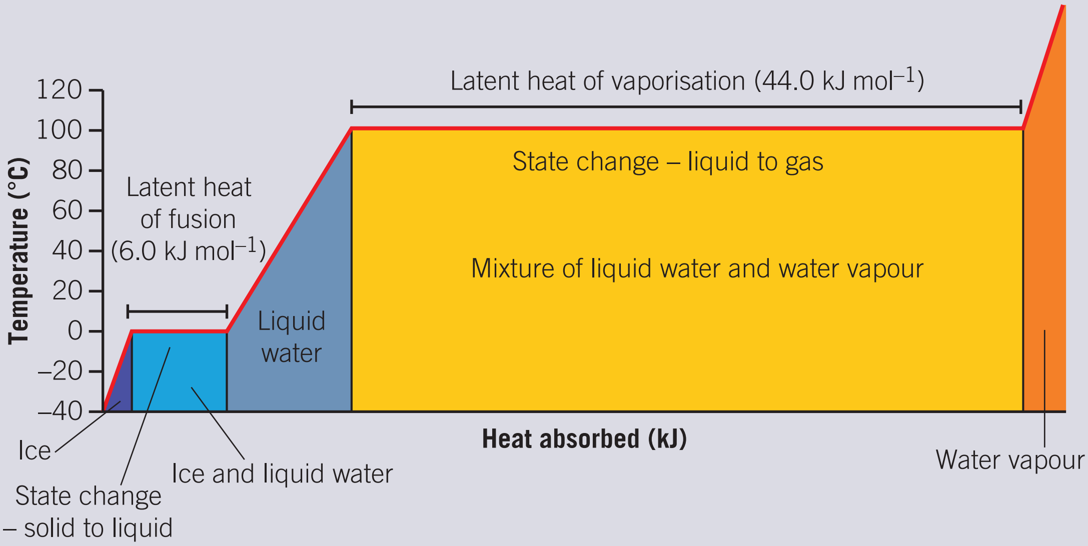

Figure 10B–8 Graphical representation of the state changes of water with latent heat of vaporisation and latent heat of fusion

The graph confirms that the melting point of water is 0°C and the boiling point of water is 100°C. You will notice that the latent heat of fusion is 6.0 kJ mol^{-1}, whereas the latent heat of vaporisation of water is 44.0 kJ mol^{-1}. This means that 44.0 kJ are required to change one mole of liquid water into a gas at 100°C, compared to only 6.0 kJ required to change one mole of ice into liquid water. Therefore, the amount of energy required to completely break the hydrogen bonds between liquid water molecules is more than the energy required to break the hydrogen bonds between water molecules in ice.

The importance of the high latent heat of vaporisation of water was discussed in reference to ocean aquatic life, but this property provides some other benefits, which include the following:

- Sweat on the surface of your skin requires a large amount of energy to vaporise, and therefore it is effective in cooling the body.
- Evaporation of water from water dams (like the ocean) is not large.

Calculations involving latent heat of vaporisation

As a chemical equation, the latent heat can be represented by:

$$q = n\,L$$

where:

q = amount of heat energy (kJ)
n = amount of substance (mol)
L = latent heat of vaporisation value (kJ mol^{-1}).

Therefore, questions could link this knowledge to the content about the mole, which you learned in Chapter 6B. Two example questions are outlined below. The first is straightforward and uses the equation as given; the second is slightly more complex and uses other mole formulas from Chapter 6B.

LINK 6B THE MOLE

Question 1
Calculate the heat energy in kJ required to evaporate 2.5 mol of water at 100°C.

Answer:

Step 1	Write down what information is known from the question and what value is to be determined. Recall that the value for the latent heat of vaporisation of water is 44.0 kJ mol^{-1}.	$q = ?$ $n = 2.5$ mol $L = 44.0$ kJ mol^{-1}
Step 2	Write down the formula to calculate latent heat and input values outlined in Step 1 above.	$q = nL$ $q = 2.5 \times 44.0$
Step 3	Calculate the answer.	$q = 110.0$ kJ

Question 2
A teacher decided to vaporise 2.107×10^{24} molecules of water at 100°C. Calculate the amount of heat energy, in kJ, required to change its state.

Answer:

Step 1	Write down what information is known from the question and what value is to be determined. Recall that the value for the latent heat of vaporisation of water is 44.0 kJ mol^{-1}.	$q = ?$ $N = 2.107 \times 10^{24}$ molecules $L = 44.0$ kJ mol^{-1}
Step 2	Determine what other information is required to answer this question.	$N_A = 6.02 \times 10^{23}$ $n = ?$
Step 3	Write down the formula that is required to find the amount in mol, then calculate this.	$n = \frac{N}{N_A}$ $n = \frac{2.107 \times 10^{24}}{6.02 \times 10^{23}}$ $n = 3.5$
Step 4	Write down the formula to calculate latent heat and input values determined in Steps 1 and 3 above.	$q = nL$ $q = 3.5 \times 44.0$
Step 5	Calculate the answer.	$q = 154.0$ kJ

Section 10B questions

1 The boiling points for the first four group 16 hydrides are shown in the table below.

Compound	Boiling point (°C)
H_2O	100
H_2S	–60
H_2Se	–42.2
H_2Te	–2.2

- **a** Describe the trend in boiling point for the first four group 16 hydrides.
- **b** Explain the trend in boiling point for H_2S, H_2Se and H_2Te.
- **c** Explain why H_2O does not follow this trend.

2 The density of liquid water and ice is 1.00 $g\,L^{-1}$ and 0.92 $g\,L^{-1}$ at 0°C, respectively. Account for the difference in these values with reference to the arrangement of molecules.

3 Define 'specific heat capacity', including its unit of measurement.

4 **a** Define 'latent heat of vaporisation' and state its unit of measurement.
- **b** Draw a fully labelled graph to represent the latent heat of vaporisation for water and explain what is occurring at different stages of the graph.
- **c** Calculate the heat energy required to evaporate 723 g of water at 100°C.

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Chapter 10 review

Summary

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

Success criteria – I am now able to:		Linked question
10A.1	Compare the three states of water with reference to the arrangement of water molecules and bonds between them	2☐, 10☐
10A.2	Describe the distribution of drinking water on Earth	4☐
10A.3	Explain why only a very small proportion of water on Earth is available as drinking water	7☐
10B.1	Describe the trend in boiling points observed for group 16 hydrides	3☐
10B.2	Explain why water doesn't follow the trend of boiling points observed for group 16 hydrides	12☐
10B.3	Explain the difference in density of ice and liquid water with reference to hydrogen bonding	8☐
10B.4	Define 'specific heat capacity'	6☐
10B.5	State the unit for specific heat capacity	1☐
10B.6	Explain why water has a relatively high specific heat capacity	11☐
10B.7	Define 'latent heat of vaporisation'	5☐
10B.8	Describe how the high latent heat of vaporisation regulates ocean temperatures and allows for the survival of aquatic life	9☐, 11☐

Multiple-choice questions

1 The unit for specific heat capacity is

A $\text{kJ g}^{-1}\,{}^{\circ}\text{C}^{-1}$.

B $\text{J kg}^{-1}\,{}^{\circ}\text{C}^{-1}$.

C $\text{J g}^{-1}\,{}^{\circ}\text{C}^{-1}$.

D J.

2 Which state of water has the least regular arrangement of molecules?

A ice

B water vapour

C water

D they are all the same

3 Which phrase best describes the trend in boiling points for the group 16 hydrides?
- **A** initially decreases, but then increases going down the group
- **B** initially increases, but then decreases going down the group
- **C** the boiling point doesn't change going down the group
- **D** the boiling points alternate between increasing and decreasing going down the group

4 Most of the fresh water on Earth is found in
- **A** lakes.
- **B** rivers.
- **C** soil moisture.
- **D** glaciers and ice caps.

5 The amount of energy required to transform water into water vapour is known as the
- **A** boiling point.
- **B** freezing point.
- **C** latent heat of vaporisation.
- **D** specific heat capacity.

6 A green container and a blue container, which are identical in size and shape, are placed in an 80°C water bath for 10 minutes. At the end of the incubation, the green container is too hot to touch, but the blue container can be easily held. One explanation for this is that the blue container has a
- **A** lower specific heat capacity than the green container.
- **B** higher specific heat capacity than the green container.
- **C** lower latent heat of vaporisation than the green container.
- **D** higher latent heat of vaporisation than the green container.

7 The main reason that most of the water on Earth is not readily available as drinking water is because it is
- **A** frozen.
- **B** polluted.
- **C** salty.
- **D** underground.

8 When an ice cube is placed in a glass of liquid water, the ice cube
- **A** floats because it is less dense.
- **B** floats because it is more dense.
- **C** sinks because it is less dense.
- **D** sinks because it is more dense.

9 The main reason that the high latent heat of vaporisation of water helps aquatic life survive is that it
- **A** provides nutrients that act as a food source.
- **B** maintains a relatively constant water temperature to live in.
- **C** allows communication signals to travel through water.
- **D** conducts electricity that is important for nerve signals.

10 The strongest intermolecular force found between water molecules is
- **A** covalent bonding.
- **B** dipole–dipole attraction.
- **C** dispersion forces.
- **D** hydrogen bonding.

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Short-answer questions

11 The latent heat of vaporisation of water is considered high, at approximately 40.8 kJ mol^{-1}.

a Explain why water would have a high latent heat of vaporisation, with reference to intermolecular forces in the liquid state. (3 marks)

b Consider the effect of a high latent heat of vaporisation on ocean temperature. (2 marks)

c Consider the consequences of a high latent heat of vaporisation on aquatic organisms. (2 marks)

d State another property of water that also has an effect on the regulation of ocean temperature. Explain the reasons for this property of water. (2 marks)

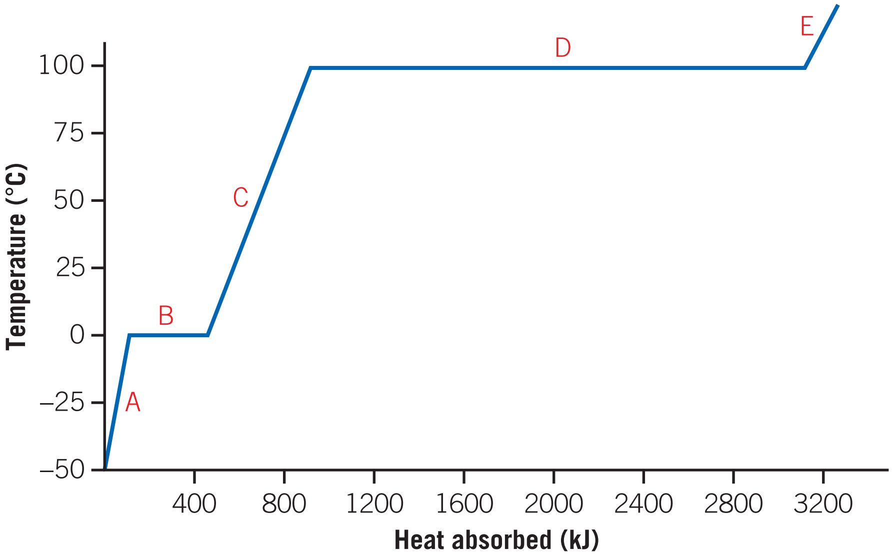
12 For the following pairs of hydrides, state which will have the higher boiling point and explain why. (4 marks)

a H_2Se and H_2Po

b H_2O and H_2S

13 How many molecules of liquid water can boil at 100°C when 150 000 J of heat energy is applied? (2 marks)

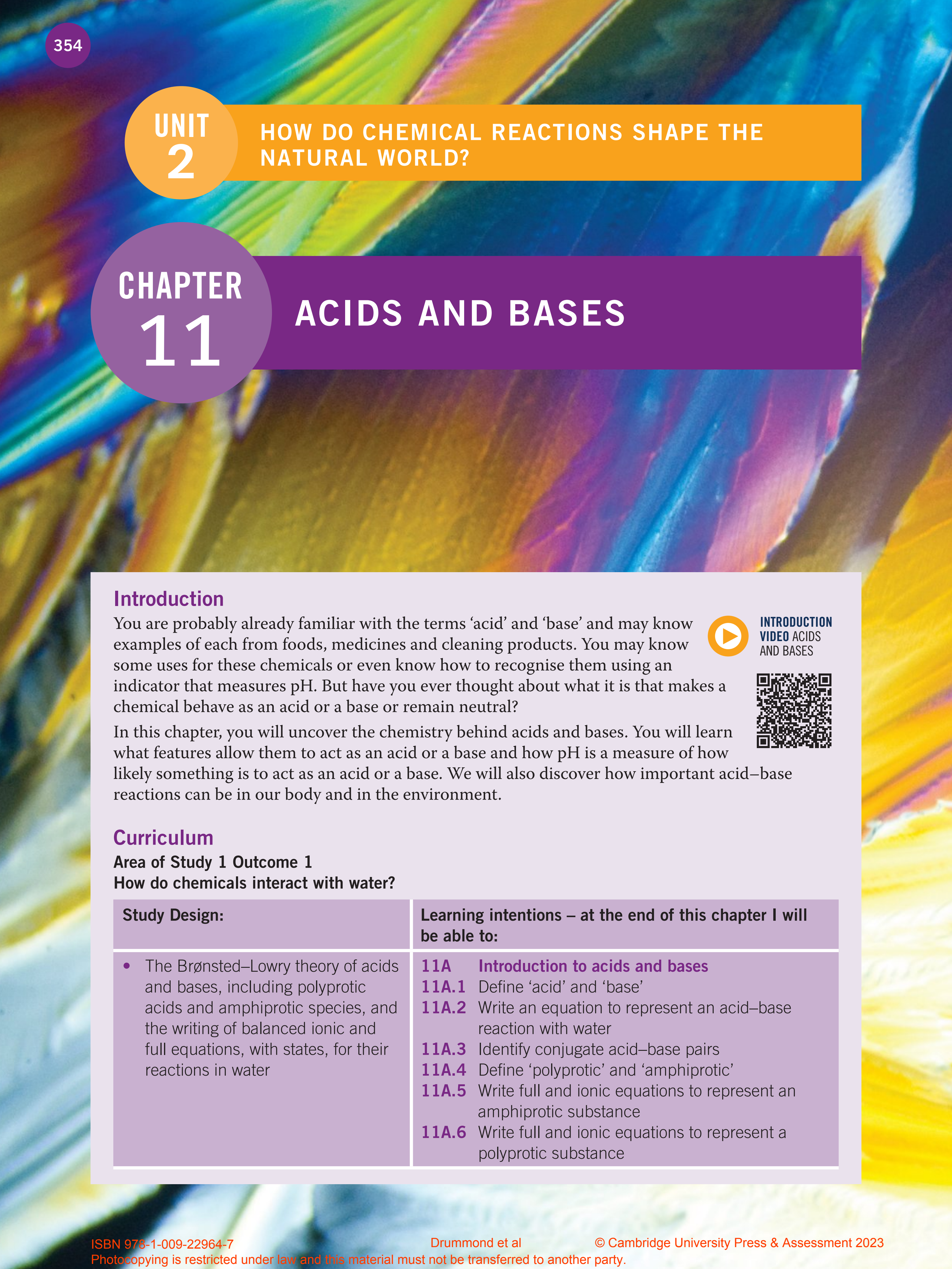
14 The graph below represents the energy required to change the state of water.



a At what temperature(s) is the liquid water being heated but the temperature is not rising? (1 mark)

b Using your understanding of water molecules and their intermolecular bonds, explain why there is no rise in temperature from your answer to part **a**. (2 marks)

c Describe what is happening at each of sections C, D and E on the graph. (3 marks)



UNIT 2

HOW DO CHEMICAL REACTIONS SHAPE THE NATURAL WORLD?

CHAPTER 11

ACIDS AND BASES

Introduction

You are probably already familiar with the terms 'acid' and 'base' and may know examples of each from foods, medicines and cleaning products. You may know some uses for these chemicals or even know how to recognise them using an indicator that measures pH. But have you ever thought about what it is that makes a chemical behave as an acid or a base or remain neutral?

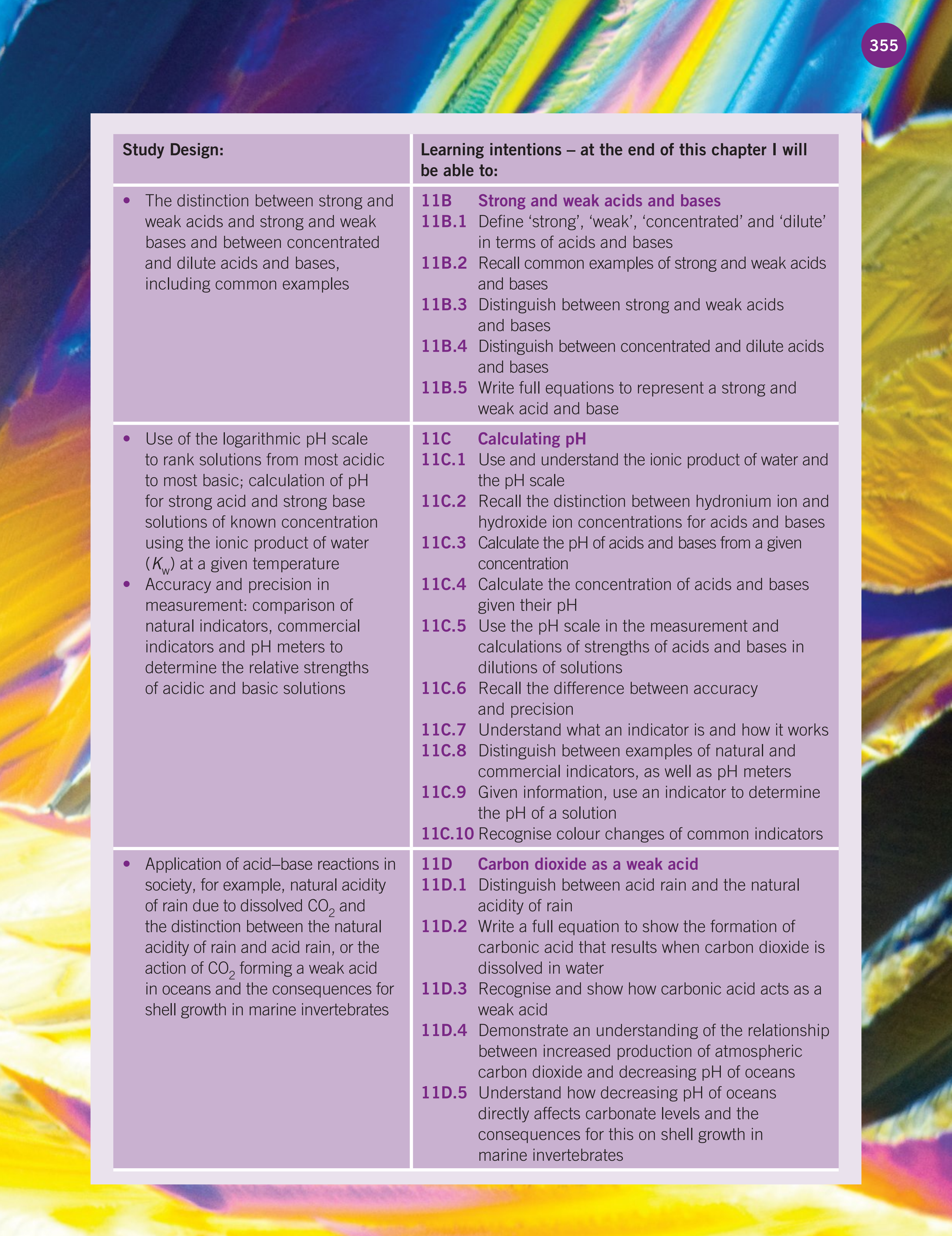
INTRODUCTION VIDEO ACIDS AND BASES

In this chapter, you will uncover the chemistry behind acids and bases. You will learn what features allow them to act as an acid or a base and how pH is a measure of how likely something is to act as an acid or a base. We will also discover how important acid–base reactions can be in our body and in the environment.

Curriculum

Area of Study 1 Outcome 1
How do chemicals interact with water?

Study Design:	Learning intentions – at the end of this chapter I will be able to:
• 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	**11A Introduction to acids and bases** **11A.1** Define 'acid' and 'base' **11A.2** Write an equation to represent an acid–base reaction with water **11A.3** Identify conjugate acid–base pairs **11A.4** Define 'polyprotic' and 'amphiprotic' **11A.5** Write full and ionic equations to represent an amphiprotic substance **11A.6** Write full and ionic equations to represent a polyprotic substance



Study Design:	Learning intentions – at the end of this chapter I will be able to:
- The distinction between strong and weak acids and strong and weak bases and between concentrated and dilute acids and bases, including common examples	**11B Strong and weak acids and bases** **11B.1** Define ‘strong’, ‘weak’, ‘concentrated’ and ‘dilute’ in terms of acids and bases **11B.2** Recall common examples of strong and weak acids and bases **11B.3** Distinguish between strong and weak acids and bases **11B.4** Distinguish between concentrated and dilute acids and bases **11B.5** Write full equations to represent a strong and weak acid and base
- 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: comparison of natural indicators, commercial indicators and pH meters to determine the relative strengths of acidic and basic solutions	**11C Calculating pH** **11C.1** Use and understand the ionic product of water and the pH scale **11C.2** Recall the distinction between hydronium ion and hydroxide ion concentrations for acids and bases **11C.3** Calculate the pH of acids and bases from a given concentration **11C.4** Calculate the concentration of acids and bases given their pH **11C.5** Use the pH scale in the measurement and calculations of strengths of acids and bases in dilutions of solutions **11C.6** Recall the difference between accuracy and precision **11C.7** Understand what an indicator is and how it works **11C.8** Distinguish between examples of natural and commercial indicators, as well as pH meters **11C.9** Given information, use an indicator to determine the pH of a solution **11C.10** Recognise colour changes of common indicators
- Application 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	**11D Carbon dioxide as a weak acid** **11D.1** Distinguish between acid rain and the natural acidity of rain **11D.2** Write a full equation to show the formation of carbonic acid that results when carbon dioxide is dissolved in water **11D.3** Recognise and show how carbonic acid acts as a weak acid **11D.4** Demonstrate an understanding of the relationship between increased production of atmospheric carbon dioxide and decreasing pH of oceans **11D.5** Understand how decreasing pH of oceans directly affects carbonate levels and the consequences for this on shell growth in marine invertebrates

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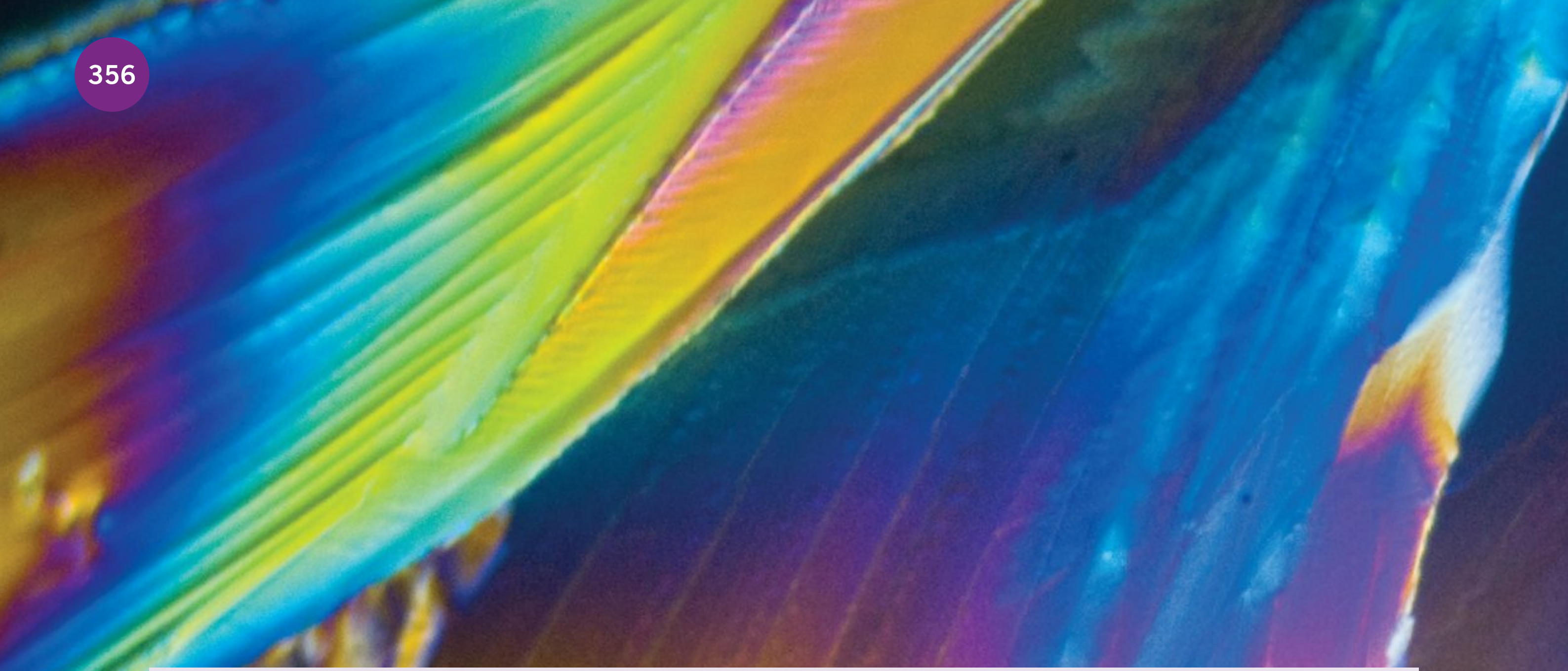

Study Design:	**Learning intentions – at the end of this chapter I will be able to:**
- Neutralisation reactions to produce salts: ▸ reaction of acids with metals, carbonates and hydroxides, including balanced full and ionic equations, with states ▸ types of antacids and their use in the neutralisation of stomach acid	**11E Neutralisation reactions** **11E.1** Recall what a neutralisation reaction is and determine products from a reaction **11E.2** Identify and write full and ionic equations for the reactions of acids with metals, identifying reactants and products **11E.3** Identify and write full and ionic equations for the reactions of acids with carbonates, identifying reactants and products **11E.4** Identify and write full and ionic equations for the reaction of acids with hydroxides, identifying reactants and products **11E.5** Recall types of antacids and explain their beneficial action in neutralising stomach acid

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Glossary

Acid
Acid rain
Alkaline
Amphiprotic
Aragonite
Base
Calibrated
Carbon dioxide
Concentrated
Conjugate acid–base pair
Dilute
Diprotic acid
Greenhouse gas
Indicator
Ionic product of water (K_w)
Monoprotic acid
Natural indicator
Neutralisation reaction
pH
Photosynthesis
pOH
Polyprotic
Spectator ion
Strong acid
Strong base
Systematic error
Triprotic acid
Weak acid
Weak base

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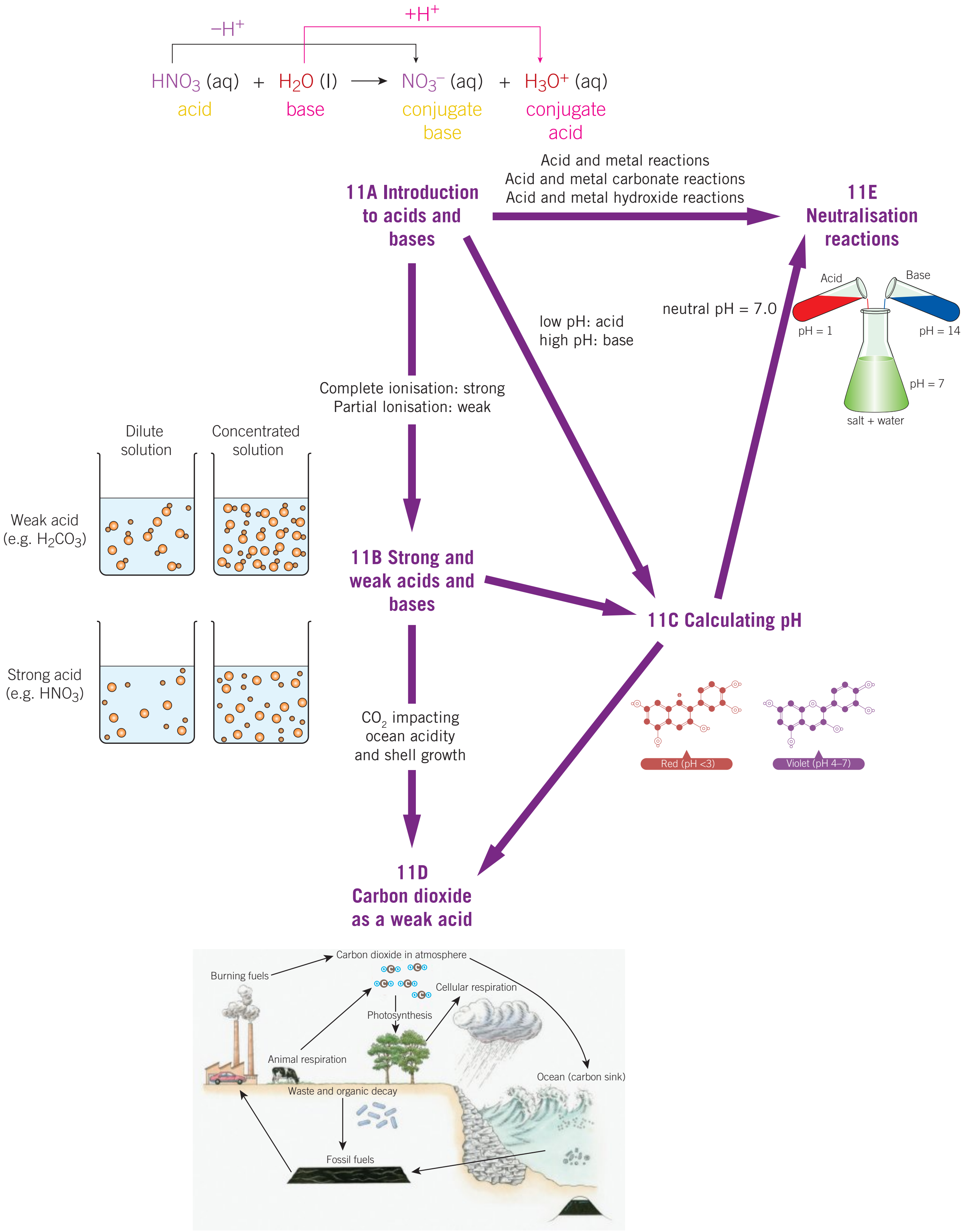
Concept map



$$\underset{\text{acid}}{HNO_3\,(aq)} + \underset{\text{base}}{H_2O\,(l)} \rightarrow \underset{\text{conjugate base}}{NO_3^-\,(aq)} + \underset{\text{conjugate acid}}{H_3O^+\,(aq)}$$

11A Introduction to acids and bases

Acid and metal reactions
Acid and metal carbonate reactions
Acid and metal hydroxide reactions

11E Neutralisation reactions

neutral pH = 7.0

low pH: acid
high pH: base

Complete ionisation: strong
Partial Ionisation: weak

11B Strong and weak acids and bases

11C Calculating pH

CO_2 impacting ocean acidity and shell growth

11D Carbon dioxide as a weak acid

See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

11A Introduction to acids and bases

Study Design:
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

Glossary:
Acid
Amphiprotic
Base
Conjugate acid–base pair
Diprotic acid
Monoprotic acid
Polyprotic
Triprotic acid

ENGAGE

Acids, bases and your body

Did you know that to help with digestion our body contains both acids and bases? Stomach acid has a pH between 1 and 2 – this is acidic. At high concentrations this can cause damage to some of the body's strongest materials, like your teeth. Meanwhile, a little further down the digestive system trail, bile produced by the liver and stored in the gall bladder is alkaline – this means basic. So, what is it about these chemicals produced by the body that make them acidic or basic, and why is it so important for the functioning of the body to have acids and bases released and acting in specific places?

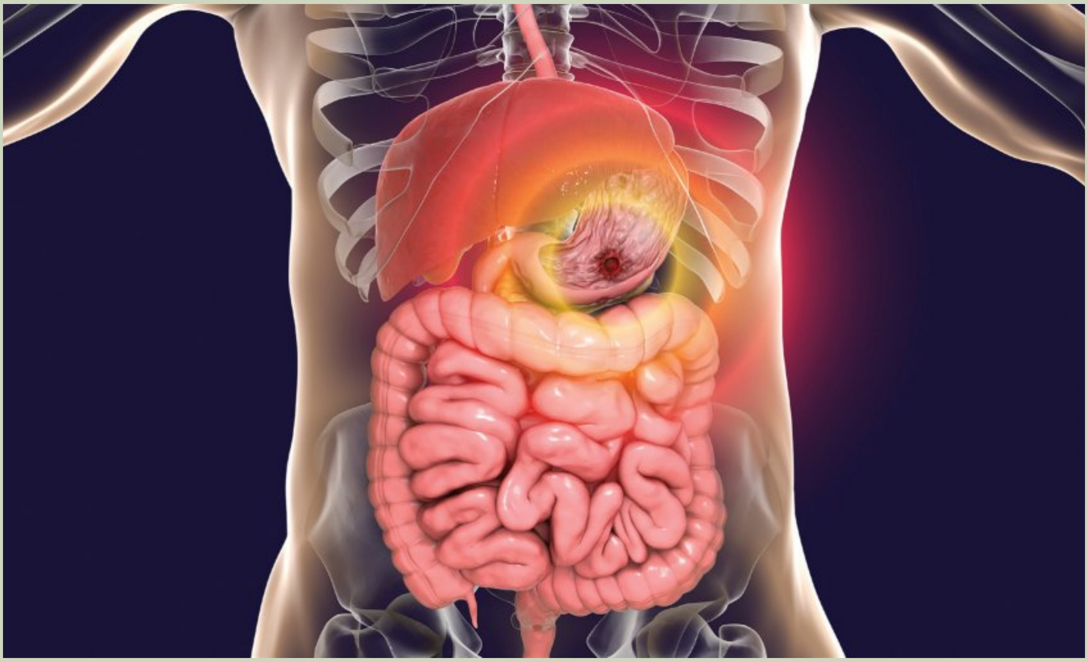


Figure 11A–1 As food is digested it passes through the highly acidic stomach. To neutralise it as it passes through the intestines, the gall bladder releases bile, which is alkaline in nature. The stomach is lined with a special mucous that protects the stomach tissue from the corrosive effects of acids. However, very small, painful ulcers can form in the lining of the stomach.

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EXPLAIN

What are acids and bases?

Table 11A–1 shows a list of some common acids and bases.

Table 11A–1 Some common acids and bases

Common acids	Common bases
Hydrochloric acid (HCl)	Sodium hydroxide ($NaOH$)
Nitric acid (HNO_3)	Potassium hydroxide (KOH)
Sulfuric acid (H_2SO_4)	Barium hydroxide ($Ba(OH)_2$)
Ethanoic acid (CH_3COOH)	Ammonia (NH_3)

What do you notice about these chemicals? What do all the acids have in common? What do all the bases have in common? Is there anything that both the acids and bases have in common?

You may have noticed that some of the examples of common bases given are soluble ionic salts, meaning that they will dissociate in water, which you will recall from Chapter 4. You may have also noticed that they all contain hydrogen (H) atoms.

LINK **4C** PRECIPITATION REACTIONS

The Brønsted–Lowry theory of acids and bases

According to Brønsted and Lowry (Figure 11A–2), an acid–base reaction involves the transfer of protons. In this definition, hydrogen ions (H^+ ions) are represented as protons.

VIDEO 11A–1 DISTINCTION BETWEEN ACIDS AND BASES

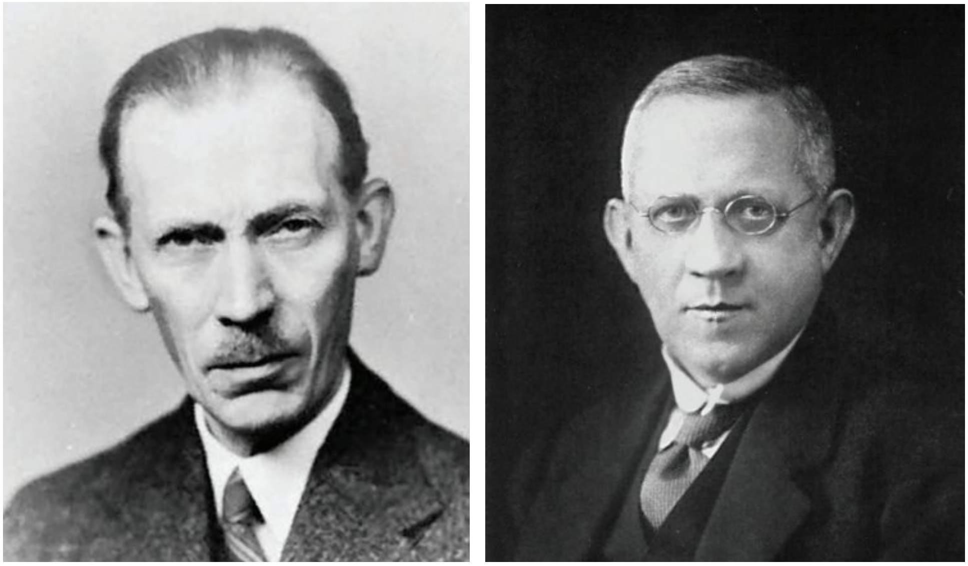

Figure 11A–2 The Brønsted–Lowry definition of acids and bases is named after two scientists: Johannes Brønsted (left) and Thomas Lowry (right). In 1923, they both independently developed the idea that acids are proton donors, while bases are proton acceptors.

Acid
an acidic substance that is a proton, H^+, donor. It has a pH value of less than 7 in aqueous solution at 25°C

Base
an alkaline substance that is a proton, H^+, acceptor. It has a pH value between 8 and 14 in aqueous solution at 25°C

An **acid** is a proton, H^+, donor, while a **base** is a proton, H^+, acceptor.

This means that to act as an acid, the solution must contain an H atom that can ionise in solution to form an H^+ ion, while a base must be able to accept an H^+ ion.

LINK **2A** REPRESENTING COVALENT COMPOUNDS

Let's look at a reaction between a common acid and base. Hydrogen chloride gas, HCl, will react with ammonia, NH_3, to form ammonium chloride, NH_4Cl. The hydrogen chloride acts as an **acid**, *donating* a proton, H^+, while the ammonia acts as a **base**, *accepting* a proton, H^+.

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This can be represented with an equation to show the transfer of protons as:

$$\overbrace{HCl(g) + NH_3}^{\text{H+}}(g) \longrightarrow NH_4Cl(s)$$

We can show the transfer of protons, H^+ ions, for Brønsted–Lowry acid–base reactions.

Worked example 11A–1: Identifying acids and bases in a reaction

VIDEO WORKED EXAMPLE 11A–1

Use the Brønsted–Lowry acid–base definition to identify the acid and the base in the following reaction.

$$HF(aq) + NH_3(aq) \rightarrow F^-(aq) + NH_4^+(aq)$$

Solution

Step 1	Identify the reactant–product pairs.	$HF(aq) + NH_3(aq) \rightarrow F^-(aq) + NH_4^+(aq)$ Two pairs: $HF \rightarrow F^-$ $NH_3 \rightarrow NH_4^+$
Step 2	Has the pair lost an H^+ ion? = acid Has the pair gained an H^+ ion? = base	$HF \rightarrow F^-$ lost an H^+ ion, therefore HF is an acid. $NH_3 \rightarrow NH_4^+$ gains an H^+ ion, therefore NH_3 is a base.

Worked example 11A–2: Writing equations for acid–base reactions

VIDEO WORKED EXAMPLE 11A–2

Show how the reaction between water and nitric acid (HNO_3) represents a Brønsted–Lowry acid–base reaction.

Solution

Step 1	Write the chemical formula for the reactants.	$HNO_3 + H_2O$
Step 2	Transfer the proton, H^+ ion, from the acid (proton donor) to the base (proton acceptor).	$\overbrace{\underset{\text{acid}}{HNO_3} + \underset{\text{base}}{H_2O}}^{\text{H+}}$
Step 3	Write the chemical formula of the products that are formed. Don't forget to include any charges.	$NO_3^- + H_3O^+$ **Note that because an H^+ ion is lost from the neutral HNO_3, it becomes a negatively charged ion. The neutral H_2O becomes positively charged with the addition of an H^+ ion.**
Step 4	Write the overall equation, ensuring it is balanced and includes states. **Note that Brønsted–Lowry theory requires a protonated solvent, for example, water.**	$HNO_3(aq) + H_2O(l) \rightarrow NO_3^-(aq) + H_3O^+(aq)$

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Worked example 11A–3: Writing equations for acid–base reactions

Show how the reaction between hydrochloric acid (HCl) and the base sodium hydroxide (NaOH) represents a Brønsted–Lowry acid–base reaction.

Solution

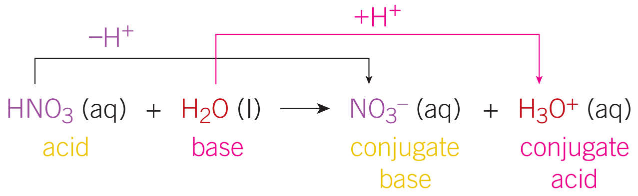
Step 1	Write chemical formula for the reactants.	HCl + NaOH
Step 2	Transfer the proton, H^+ ion, from the acid (proton donor) to the base (proton acceptor).	H+ $\overset{\text{H}^+}{\text{HCl} \rightarrow \text{NaOH}}$ — HCl + NaOH acid base $Cl^- + Na^+ + H_2O$
Step 3	Write the chemical formula of the products that are formed. Don't forget to include any charges.	**Note that because an H^+ ion is lost from the neutral HCl, it becomes a negatively charged ion; whereas, the OH^- ion of the NaOH acts as the base, gaining an H^+ ion to form water.**
Step 4	Write the overall equation, ensuring it is balanced and includes states.	$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$ **Note that the Na^+ ion and the Cl^- ion would form an ionic salt, which is represented in an overall equation as a soluble ionic compound.**

VIDEO
WORKED
EXAMPLE 11A–3

LINK
11E
NEUTRALISATION
REACTIONS

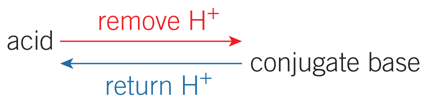
Conjugate acid–base pairs

In the worked examples 11A–1 to 11A–3, we saw how in the Brønsted–Lowry definitions of acids and bases an H^+ ion is transferred from the acid to the base. When we looked at the equations we used to show species acting as acids and bases, you will notice the reactants and the products they form differ by a single H^+ ion.



These are known as **conjugate acid–base pairs**. They represent the reactant–product pairs that differ by the loss or gain of a single hydrogen ion, H^+.

In the example above, pair one is HNO_3/NO_3^- and pair two is H_2O/H_3O^+.



Conjugate acid–base pair represents the reactant (acid)/product (base) pairs that differ by the loss or gain of a hydrogen ion, H^+, e.g. HNO_3/NO_3^-

Worked example 11A–4: Identifying conjugate acid–base pairs

VIDEO WORKED EXAMPLE 11A–4

Determine the conjugate acid–base pair for the following chemicals.

a Sulfuric acid, H_2SO_4

Solution

Step 1	Identify if the substance is an acid or a base.	H_2SO_4 = acid
Step 2	Apply the rule: acids are proton, H^+, donors. Lose one H^+ ion from the molecular formula. Don't forget to change the charge on the molecule.	HSO_4^-
Step 3	Write as a conjugate acid–base pair.	H_2SO_4/HSO_4^-

b Carbonate ion, CO_3^{2-}

Solution

Step 1	Identify if the substance is an acid or a base.	CO_3^{2-} = base
Step 2	Apply the rule: bases are proton, H^+, acceptors. Gain one H^+ ion from the molecular formula. Don't forget to change the charge on the molecule.	HCO_3^-
Step 3	Write as a conjugate acid–base pair.	HCO_3^-/CO_3^{2-}

WORKSHEET 11A–1 CONJUGATE ACID–BASE PAIRS

Check-in questions – Set 1

1 Write a definition of a Brønsted–Lowry acid, using HCl as an example.
2 Write a definition of a Brønsted–Lowry base, using NH_3 as an example.
3 Write an acid–base equation according to the Brønsted–Lowry definition to show the reaction when hydrogen fluoride gas is bubbled through water, forming an acidic solution.
4 Identify the conjugate acid–base pairs in the equation in Question 3 above.

Amphiprotic substances

Some substances have the ability to either donate or accept a proton, or H^+ ion. These substances are known as **amphiprotic**. Water is an excellent example of an amphiprotic substance.

Amphiprotic a substance that is able to donate or accept a proton, H^+ ion

Water can act as an acid: $H_2O(l) + NH_3(aq) \rightleftharpoons OH^-(aq) + NH_4^+(aq)$ Conjugate pair = H_2O/OH^-

Water can act as a base: $H_2O(l) + HCl(aq) \rightleftharpoons H_3O^+(aq) + Cl^-(aq)$ Conjugate pair = H_3O^+/H_2O

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Worked example 11A–5: Writing equations involving amphiprotic substances

VIDEO WORKED EXAMPLE 11A–5

Write equations to show how the hydrogen carbonate ion (HCO_3^-) is an amphiprotic substance when it reacts with water.

Solution

Equation 1: HCO_3^- reacting as an acid (losing an H^+) with water.

Step 1	**a** Write the chemical formulas of the reactants.	$HCO_3^-(aq) + H_2O(l)$
	b Transfer the proton, H^+ ion, from the acid (proton donor) to the base (proton acceptor).	$\underset{\text{acid}}{HCO_3^-(aq)} + \underset{\text{base}}{H_2O}$
	c Write the chemical formula of the products that are formed. Don't forget to include any charges.	Products: $CO_3^{2-}(aq) + H_3O^+(aq)$
	d Write the overall equation.	$HCO_3^-(aq) + H_2O(l) \rightleftharpoons CO_3^{2-}(aq) + H_3O^+(aq)$

Equation 2: HCO_3^- reacting as a base (gaining an H^+) with water

Step 2	**a** Write the chemical formulas of the reactants.	$HCO_3^-(aq) + H_2O(l)$
	b Transfer the proton, H^+ ion, from the acid (proton donor) to the base (proton acceptor).	H^+ $\underset{\text{base}}{HCO_3^-} + \underset{\text{acid}}{H_2O}$
	c Write the chemical formula of the products that are formed. Don't forget to include any charges.	Products: $H_2CO_3(aq) + OH^-(aq)$
	d Write the overall equation.	$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$

Polyprotic substances

Some substances have the ability to donate or gain more than one proton, or H^+ ion, per molecule. These are known as **polyprotic** substances.

Acids

Monoprotic acids have **one** proton, an H^+ ion, to donate. Hydrochloric acid, HCl, and ethanoic acid, CH_3COOH, are examples.

Diprotic acids have **two** protons, H^+ ions, that they can donate. Sulfuric acid, H_2SO_4, is an example. The protons are donated in a two-step process.

Triprotic acids have **three** protons, H^+ ions, that they can donate. Phosphoric acid, H_3PO_4, is an example. The protons are donated in a three-step process as shown in Worked example 11A–6.

Polyprotic
a substance that can act as a Brønsted–Lowry acid more than once, as it has several protons, H^+ ions, to donate. Diprotic and triprotic acids are examples

Monoprotic acid
a substance that can act as a Brønsted–Lowry acid once, as it has one proton, H^+ ion, to donate to a Brønsted–Lowry base

Diprotic acid
a substance that can act as a Brønsted–Lowry acid twice, as it has two protons, H^+ ions, to donate to a Brønsted–Lowry base

Triprotic acid
a substance that can act as a Brønsted–Lowry acid three times, as it has three protons, H^+ ions, to donate to a Brønsted–Lowry base

Worked example 11A–6: Writing equations involving triprotic substances

VIDEO
WORKED
EXAMPLE 11A–6

Write equations to show how phosphoric acid, $H_3PO_4(aq)$, is a triprotic substance.

Also list the conjugate acid–base pairs found in the equations.

Solution

Step 1	Write the step 1 equation showing H_3PO_4 reacting as an acid (losing an H^+) with water. **a** Write the chemical formula of the reactants. **b** Transfer the proton, H^+ ion from the acid (proton donor) to the base (proton acceptor). **c** Write the chemical formula of the products that are formed. Don't forget to include any charges.	$H_3PO_4(aq) + H_2O(l)$ $\overset{H^+}{\overline{H_3PO_4(aq) + H_2O(l)}}$ acid base $H_3PO_4(aq) + H_2O(l) \rightleftharpoons H_2PO_4^-(aq) + H_3O^+(aq)$
Step 2	Write the step 2 equation showing the product, $H_2PO_4^-$, reacting as an acid (losing an H^+) with more of the unreacted water molecules present.	$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$
Step 3	Write the step 3 equation showing the product, HPO_4^{2-}, reacting as an acid (losing an H^+) with more of the unreacted water molecules present.	$HPO_4^{2-}(aq) + H_2O(l) \rightleftharpoons PO_4^{3-}(aq) + H_3O^+(aq)$
Step 4	List the conjugate acid–base pairs. **Note that $H_2PO_4^-$ and HPO_4^{2-} are also examples of amphiprotic substances.**	$H_3PO_4/H_2PO_4^-$ H_3O^+/H_2O $H_2PO_4^-/HPO_4^{2-}$ HPO_4^{2-}/PO_4^{3-}

Bases

Bases can also be polyprotic substances. In these instances, the base is able to accept more than one proton.

In Worked example 11A–6 above, phosphoric acid is a triprotic acid. The final product, the phosphate ion (PO_4^{3-}), can accept up to three protons, making it a polyprotic base.

Check-in questions – Set 2

1 Copy and complete the table below summarising polyprotic acids.

Acid	Example	How many protons can it donate?	Equation(s) showing the substance reacting with water as an acid
Monoprotic	HCl		
Diprotic	H_2SO_4		
Triprotic	H_3PO_4		

WORKSHEET
11A–2
POLYPROTIC
ACIDS

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11A SKILLS

Determining the correct charge on molecules

As molecules act as acids or bases – donating or gaining protons – it is important that we also indicate the change in charge of the molecule.

Each time a substance *loses* an H^+ ion, its overall charge will decrease by 1. For example, a molecule like water, H_2O, which has no charge, will become a 1– charged ion, OH^-, if it donates an H^+ ion when acting as an acid.

Similarly, a molecule like HPO_4^{2-}, which already has a charge, will gain 1– again if it donates its final H^+ ion to become PO_4^{3-}.

Alternatively, each time a substance *gains* an H^+ ion it will gain a 1+ in charge. For example, a molecule like water, H_2O, which has no charge, will become a 1+ charged ion, H_3O^+, if it gains an H^+ ion when acting as a base.

Similarly, a molecule like HPO_4^{2-}, which already has a charge, will gain 1+ in charge if it gains an H^+ ion to become $H_2PO_4^-$.

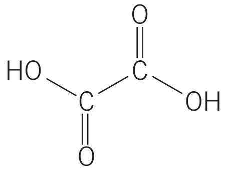
Section 11A questions

1 Identify the conjugate acid–base pairs in the following equations.
  a $H_2S(g) + H_2O(l) \rightarrow HS^-(aq) + H_3O^+(aq)$
  b $Ba(OH)_2(aq) + 2HCl(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$
  c $CH_3COOH(aq) + H_2O(l) \rightarrow CH_3COO^-(aq) + H_3O^+(aq)$

2 Identify the following substances as monoprotic acids, monoprotic bases, diprotic acids, triprotic acids, diprotic bases or triprotic bases.
  a H_2S
  b PO_4^{3-}
  c CH_3COOH
  d NH_3

3 $H_2PO_4^-$ ions are both *diprotic* and *amphiprotic*. Explain these terms.

4 Oxalic acid ($C_2H_2O_4$) is a diprotic acid.



  a Write equations to show how it behaves as a diprotic acid when reacting with water.
  b Highlight substances in your equations, other than water, that are amphiprotic.

5 Identify the acids and bases in the following equation and their conjugate pairs.
  a $NH_3(aq) + HF(aq) \rightarrow F^-(aq) + NH_4^+(aq)$
  b $H_2O(l) + CN^-(aq) \rightarrow OH^-(aq) + HCN(aq)$
  c $NH_4^+(aq) + CH_3COO^-(aq) \rightarrow NH_3(aq) + CH_3COOH(aq)$

11B Strong and weak acids and bases

Study Design:
The distinction between strong and weak acids and strong and weak bases and between concentrated and dilute acids and bases, including common examples

Glossary:

Concentrated	Strong base
Dilute	Weak acid
Strong acid	Weak base

ENGAGE

Acid or magic?

Did you know that the second strongest acid ever recorded is known as magic acid? While 100% sulfuric acid (with a pH of –12) is one of the strongest acids that is used in most chemistry labs, magic acid is 100 billion times stronger (around a pH of –23)! Interestingly, there is another acid that is 100 000 times stronger than magic acid – fluoroantimonic acid (a pH equivalent of –28!). These acids, known as superacids, are too corrosive to be stored in any sort of normal container. Instead, they are kept in containers made out of Teflon. In this section, we are going to learn about the differences between strong and weak acids and bases.

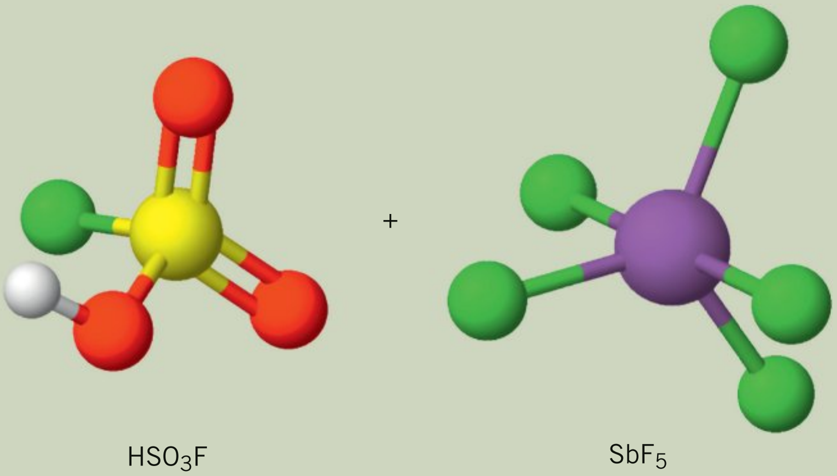


Figure 11B–1 Magic acid, the second strongest acid ever recorded with a pH of –23

Magic acid is typically composed of a 1 : 1 molar mixture of fluorosulfuric acid (HSO_3F) and antimony pentafluoride (SbF_5). Researchers working with the acid in the 1960s deemed it to be 'magic' when they discovered that it could dissolve paraffin wax – a substance that is usually unreactive due to its long, saturated hydrocarbon chains.

EXPLAIN

Strong versus weak

11A INTRODUCTION TO ACIDS AND BASES LINK

The strength of an acid or base is determined by how likely they are to lose or gain a proton, H^+ ion, respectively. Refer to Table 11B–1 on the following page for a list of common strong and weak acids and bases.

Table 11B–1 List of common strong and weak acids and bases

Common STRONG acids	Common WEAK acids	Common STRONG bases	Common WEAK bases
Hydrochloric acid HCl	Ethanoic acid CH_3COOH	Sodium hydroxide $NaOH$	Ammonia NH_3
Hydrobromic acid HBr	Methanoic acid $HCOOH$	Potassium hydroxide KOH	Sodium bicarbonate $NaHCO_3$
Hydroiodic acid HI	Carbonic acid H_2CO_3	Calcium hydroxide $Ca(OH)_2$	
Nitric acid HNO_3	Phosphoric acid H_3PO_4	Barium hydroxide $Ba(OH)_2$	
Perchloric acid $HClO_4$			
Sulfuric acid H_2SO_4			

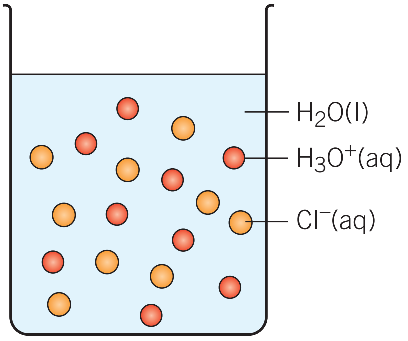
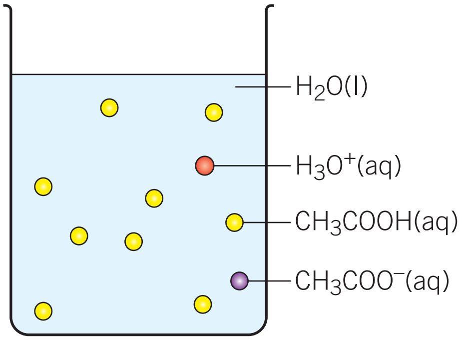
Strong and weak acids

Strong acids will completely ionise, meaning that every acid molecule will donate its H^+ ion, forming a conjugate base. In contrast, a **weak acid** will only partially ionise, meaning that not every acid molecule will donate an H^+ ion to form a conjugate base. This results in a solution that contains both reactant and product molecules. Table 11B–2 summarises how we can represent strong and weak acids by using hydrochloric acid (a strong acid) and ethanoic acid (a weak acid) as an example.

Strong acid an acid that will ionise completely in solution – i.e. every acid molecule will donate its H^+ ion to form a conjugate base

Weak acid an acid that will only partially ionise in solution – i.e. not every acid molecule will donate an H^+ ion to form a conjugate base, resulting in a solution that contains both undissociated acid molecules and conjugate base molecules

Table 11B–2 Comparison of strong and weak acids

Strong acids	Weak acids
Fully (100%) ionise	Partially ionise
React completely; EVERY acid molecule donates its H^+ ion.	Does NOT react completely; not every acid molecule donates its H^+ ion.
Example: Hydrochloric acid (HCl) reacting with water  	Example: Ethanoic acid (CH_3COOH) 
What's present: products only *(Note that excess water that has not gained a H^+ ion from an acid molecule is also present.)*	What's present: reactants and products
Equation: $\overbrace{HCl(aq) + H_2O(l)}^{H^+} \longrightarrow H_3O^+(aq) + Cl^-(aq)$ Note that to represent a strong acid in an equation, a full arrow is used. This shows that only products will be present.	Equation: $\overbrace{CH_3COOH(aq) + H_2O(l)}^{H^+} \rightleftharpoons CH_3COO^- + H_3O^+(aq)$ Note that to represent a weak acid in an equation, equilibrium/reversible arrows are used. This shows that both reactants and products will be present.

Strong and weak bases

Strong base
a base where every base molecule accepts an H^+ ion from an acid

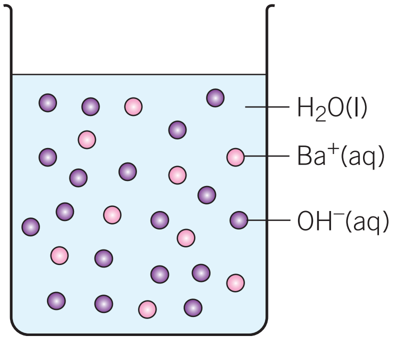
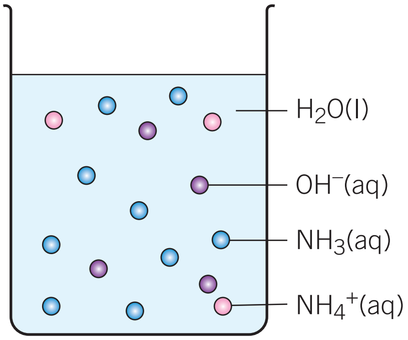
Weak base
a base where not every base molecule accepts an H^+ ion from an acid

The strength of a base is defined in the same way. **Strong bases** will completely ionise, which means that every base molecule will accept an H^+ ion, forming a conjugate acid. A **weak base** will only partially ionise, which means that not every base molecule will accept an H^+ ion to form a conjugate acid, resulting in a solution that contains both reactant and product molecules. Table 11B–3 summarises how we can represent strong and weak bases using sodium hydroxide, a strong base, and ammonia, a weak base, as an example.

VIDEO 11B–1
STRONG, WEAK, DILUTE OR CONCENTRATED?

WORKSHEET 11B–1
STRONG OR WEAK?

Table 11B–3 Comparison of strong and weak bases

Strong bases	Weak bases
Fully (100%) ionise	Partially ionise
React completely; EVERY base molecule accepts an H^+ ion	Does NOT react completely; not every base molecule will accept an H^+ ion.
Example: Barium hydroxide ($Ba(OH)_2$)  	Example: Ammonia (NH_3) 
What's present: products only *(Note that excess water that has not gained/ donated an H^+ ion is also present)*	What's present: reactants and products
Equation: $Ba(OH)_2(aq) \xrightarrow{H_2O(l)} Ba^{2+}(aq) + 2OH^-(aq)$ Note that to represent a strong base in an equation, a full arrow is used. This shows that only products will be present.	Equation: $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$ Note that to represent a weak base in an equation, equilibrium/reversible arrows are used. This shows that both reactants and products will be present.

Check-in questions – Set 1

1 Outline the difference between the terms 'strong' and 'weak' when discussing acids and what is present at the conclusion of a reaction.

2 Give one example for each of the following:
   a strong acid
   b strong base
   c weak acid
   d weak base

Concentrated versus dilute

Concentrated
a solution with a large number of molecules of the solute dissolved in a given volume

Another important factor in determining the safety of handling an acid or base is their concentration. It is important to realise this is different from being strong or weak. Acids and bases, like all solutions, can be **concentrated**, meaning that there are a lot of molecules

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of the solute present in a given volume, or **dilute**, meaning there are fewer molecules of the solute present in a given volume (see Table 11B–4). We can determine how concentrated a solution is using the equation

$$c = \frac{n}{V}$$

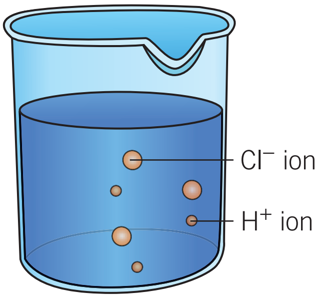
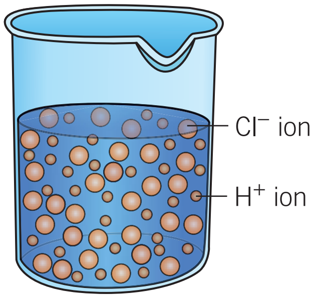
where:

c = The concentration of the solution measured in mol L^{-1} (or M)
n = The amount in moles (mol)
V = The volume in litres (L)

Dilute
a solution with a small number of molecules of the solute dissolved in a given volume

LINK **5A POLARITY**

Table 11B–4 Comparison between dilute and concentrated solutions

Dilute solution	Concentrated solution
 0.1 mole HCl in 250 mL water	 1 mole HCl in 250 mL water
$c = \frac{n}{V} = \frac{0.1}{0.25} = 0.4$ M	$c = \frac{n}{V} = \frac{1}{0.25} = 4.0$ M

Acid and base solutions can be strong OR weak AND concentrated OR dilute, as shown in Figure 11B–2. The top row compares weak (partially ionised) dilute and concentrated solutions, and the bottom row compares strong (completely ionised) dilute and concentrated solutions.

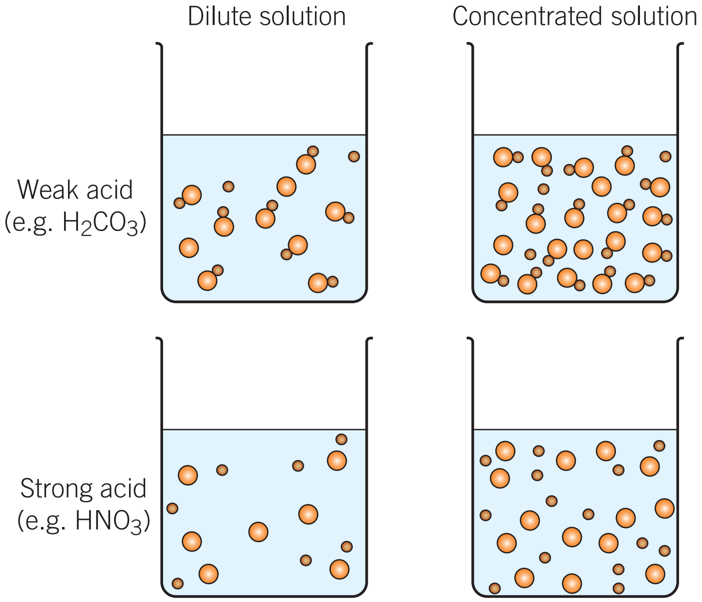


Figure 11B–2 The concentration of ions in an acid solution depends on its concentration and strength.

VIDEO 11B–2
SKILLS: KNOWING THE DIFFERENCE BETWEEN STRENGTH AND CONCENTRATION

11B SKILLS

Knowing the difference between strength and concentration

As you have seen in this chapter, it is important that you do not confuse strength of an acid or a base with the concentration of an acid or a base; in particular, strong and concentrated, or weak and dilute.

Strength is explained best as the extent to which an acid or a base ionises, whereas concentration refers to the ratio of solute to solvent.

Another way to look at strength is the ratio of ions (products of a reaction) to starting molecules (reactants of a reaction) as represented below:

number of ions : number of reactant molecules

If this ratio is:

many : none

then it is a strong acid or base.

Whereas, if this ratio is:

few : many

then it is a weak acid or base.

This is where a diagram, such as that shown in Figure 11B–2, is useful. A similar diagram in which these ions and molecules are labelled is shown in Figure 11B-3 below.

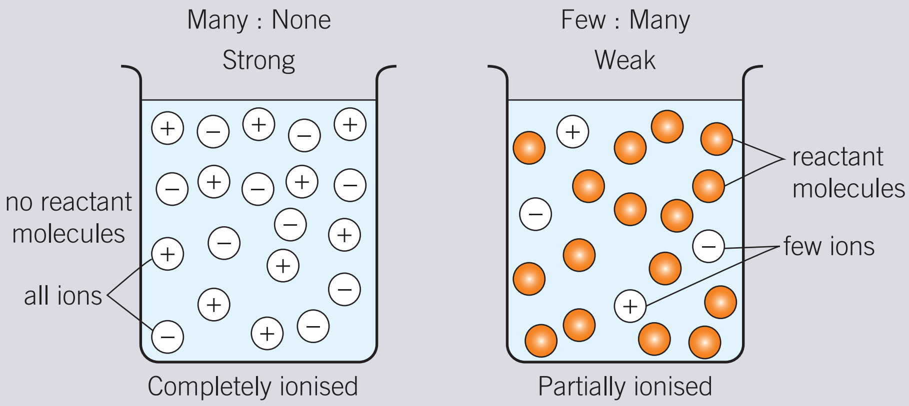


Figure 11B–3 Strong versus weak

The way to understand concentration is to view it as a ratio of *solute : solvent*. For acids and bases, this can also be better represented as:

acid/base : water

If this ratio is:

many : few

then it is a concentrated acid or base.

Whereas, if this ratio is:

few : many

then it is a dilute acid or base.

This can again be represented diagrammatically, as shown in Figure 11B–4.

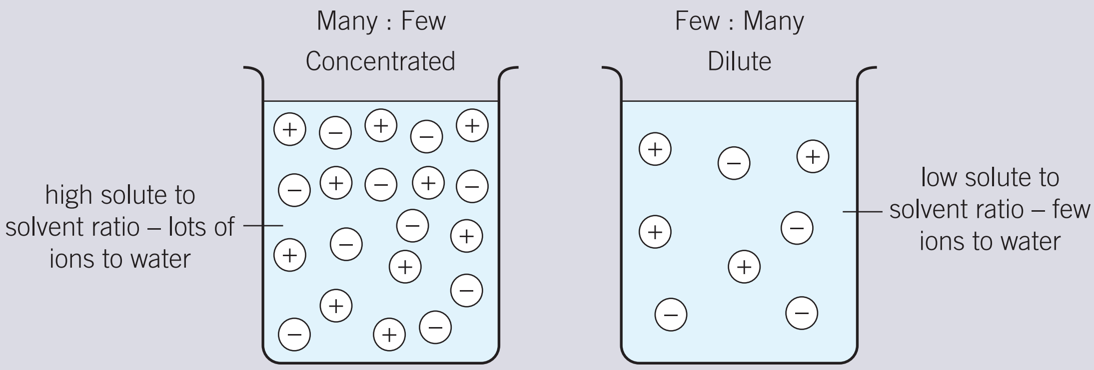


Figure 11B–4 Concentrated versus dilute

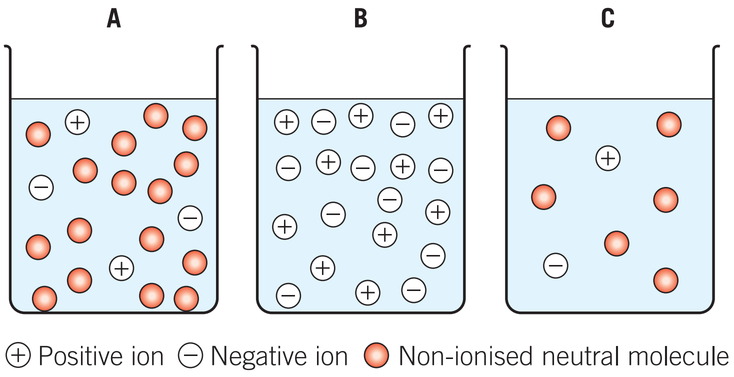
It is also expected that you are able to connect these two ideas and apply your knowledge to questions. As such, a question may give you a figure that is labelled (or unlabelled) like that above and ask you to identify and explain whether the solution represents a strong or weak acid. Looking at the diagram, because there are only ions present and no reactant molecules, we can determine that this represents a strong acid that has fully ionised in water. If it were a weak acid, then both of these concentrated or dilute solutions would still have reactant molecules present, indicating that it has only partially ionised.

Section 11B questions

1 Explain the difference between strong and concentrated acids.

2 Write an equation to represent the following solutions reacting with water.
  a methanoic acid (weak acid)
  b $HClO_4$ (strong acid)
  c methylamine, CH_3NH_2 (weak base)

3 a Identify the solutions in the beakers as:
    i strong or weak.
    ii concentrated or dilute.
  b Explain your choice for parts **i** and **ii** above for each beaker.



11C Calculating pH

Study Design:

- 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: comparison of natural indicators, commercial indicators and pH meters to determine the relative strengths of acidic and basic solutions

Glossary:

Alkaline
Calibrated
Indicator
Ionic product of water (K_w)
Natural indicator
pH
pOH
Systematic error

ENGAGE

Beware – corrosive!

Often you see a corrosive substance warning label on acids and bases (Figure 11C–1). Highly corrosive chemicals have a pH lower than 2 or higher than 12. ‘Corrosive’ means that something will cause irreversible damage to a surface it touches. If a corrosive substance comes in contact with the skin, it can cause a ‘chemical burn’. Sulfuric acid is an example of a corrosive substance, and it is commonly found in car batteries. With a pH of 0.5, it reacts violently with water, and if it touches your skin, it can cause severe burns and tissue damage. Being able to determine the pH of a substance is important in determining how to safely handle and store the substance.

Figure 11C–1 Hazchem signs are displayed on storage containers and vehicles to indicate that dangerous chemicals or substances are present.

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pH scale

The pH scale provides a guide for measuring how strongly acidic or basic an aqueous solution is. The scale typically ranges from 0 to 14, with 0 being highly acidic, 7 being neutral for aqueous solutions at 25°C and 14 being highly **alkaline** (basic) (Figure 11C-2).

Alkaline
a term for a water-soluble base. A substance that will gain an H^+ ion and has a pH greater than 7 in water at 25°C

Ionic product of water (K_w)
the equilibrium constant K_w, where $K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$ M^2 at 25°C. It represents the ionisation reaction of water with itself. K_w is the product of the concentration of ions in solution so is referred to as the ionic product of water

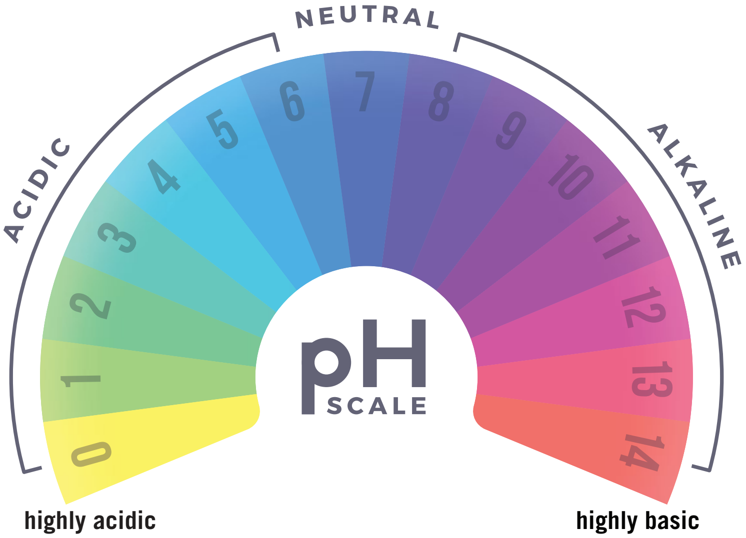

Figure 11C–2 The pH scale is colour coded and ranges from 0–14.

To understand how pH is calculated, we must first understand the ionic product of water, K_w.

Ionic product of water

Recall from Chapter 11A that water is an amphiprotic substance. This means that water can react with itself to self-ionise, with one molecule of water acting as an acid to donate an H^+ ion and another molecule of water acting as a base to accept an H^+ ion. This is shown by the equation

LINK **11A** INTRODUCTION TO ACIDS AND BASES

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

NOTE

Notice this equation has a reversible arrow separating the reactants and the products, indicating that water is a weak acid and a weak base, therefore only partially ionising.

From this equation, the **ionic product of water (K_w)** is derived as

$$K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}\ M^2 \text{ at } 25°C$$

where:
$[H_3O^+]$ represents the concentration of hydronium ions (acid – able to donate H^+)
$[OH^-]$ represents the concentration of hydroxide ions (base – able to accept H^+).

This is a representation of the relative amounts (concentrations) of the products present. In Unit 3 when learning about equilibrium, you will learn more about how this constant is derived.

LINK **UNIT 3**

The expression $[H_3O^+][OH^-]$ is the 'ionic product of water' or 'ionisation constant of water'. This means that the concentration of hydronium ions, H_3O^+, multiplied by the concentration of hydroxide ions, OH^-, always equals 1.00×10^{-14} M^2 at room temperature (25°C), as shown in Table 11C–1.

NOTE

- Square brackets [ ] are often used to represent molar concentration.
- 25°C is often considered the standard 'room temperature'.

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Table 11C–1 Calculating the ionic product of water for different solutions at 25°C

Neutral solutions	**$[H_3O^+]$** 10^{-7} M	**=**	**$[OH^-]$** 10^{-7} M	$[H_3O^+] \times [OH^-] = 10^{-14}\ M^2$ $10^{-7}\ M \times 10^{-7}\ M = 10^{-14}\ M^2$
Acidic solutions	**$[H_3O^+] > 10^{-7}$ M** e.g. 10^{-1} M	**>**	**$[OH^-] < 10^{-7}$ M** e.g. 10^{-13} M	$[H_3O^+] \times [OH^-] = 10^{-14}\ M^2$ $10^{-1}\ M \times 10^{-13}\ M = 10^{-14}\ M^2$
Basic solutions	**$[H_3O^+] < 10^{-7}$ M** e.g. 10^{-12} M	**<**	**$[OH^-] > 10^{-7}$ M** e.g. 10^{-2} M	$[H_3O^+] \times [OH^-] = 10^{-14}\ M^2$ $10^{-12}\ M \times 10^{-2}\ M = 10^{-14}\ M^2$

Worked example 11C–1: Calculating concentrations of ions in strong acids

VIDEO
WORKED
EXAMPLE 11C–1

For a 0.025 M aqueous solution of H_2SO_4 at room temperature, calculate the concentration of:

a $[H_3O^+]$ **b** $[OH^-]$

Solution

Step 1	Identify the solution as either a strong acid or strong base.	H_2SO_4 = strong acid
Step 2	If acid: [acid] = $[H_3O^+]$ If base: [base] = $[OH^-]$ **The concentration of H_3O^+ (for an acid) or OH^- (for a base) can be determined using this method only if they are strong acids or bases that fully ionise.**	$[H_3O^+] = [H_2SO_4] = 0.025$ M
Step 3	Is it monoprotic or polyprotic? If the solution is monoprotic, [acid] = $[H_3O^+]$ remains. If the solution is polyprotic, then it will act as an acid more than once, donating more than one H^+ ion as a result. [acid] = $[H_3O^+]$ × number of H^+ ions donated	**a** H_2SO_4 is diprotic. Therefore: $[H_3O^+] = [H_2SO_4] \times 2$ $= 0.025\ M \times 2$ $= 0.050$ M
Step 4	Calculate the concentration of the other ion using the ionic product of water. $[H_3O^+] \times [OH^-] = 10^{-14}\ M^2$	**b** Rearranging gives $[OH^-] = \frac{10^{-14}\ M^2}{[H_3O^+]}$ $[OH^-] = \frac{10^{-14}\ M^2}{0.050\ M}$ $[OH^-] = 2.0 \times 10^{-13}\ M$

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Worked example 11C–2: Calculating concentrations of ions in strong bases

For a 0.1 M aqueous solution of $Ba(OH)_2$ at room temperature, calculate the concentration of:

a $[H_3O^+]$ **b** $[OH^-]$

Solution

Step 1	Identify the solution as either a strong acid or a strong base.	$Ba(OH)_2$ = strong base
Step 2	Is it monoprotic or polyprotic? If the solution is monoprotic, [base] = $[OH^-]$ remains. If the solution is polyprotic, then it will act as a base more than once, accepting more than one H^+ ion. $[OH^-]$ = [base] × number of OH^- ions present	This is a diprotic substance, therefore will act as a base twice. $[OH^-] = 2 \times [Ba(OH)_2]$ $= 2 \times 0.1\ M$ $= 0.2\ M$
Step 3	Calculate the concentration of the other ion using the ionic product of water. $[H_3O^+] \times [OH^-] = 10^{-14}\ M^2$	$[H_3O^+] = \dfrac{10^{-14}\ M^2}{[OH^-]}$ $[H_3O^+] = \dfrac{10^{-14}\ M^2}{0.2\ M}$ $[H_3O^+] = 5 \times 10^{-14}\ M$

VIDEO
WORKED EXAMPLE 11C–2

Check-in questions – Set 1

1 What is the ionic product of water at 25°C?

2 For a 0.0001 M aqueous solution of nitric acid (HNO_3) at room temperature, calculate the concentration of:

  a $[H_3O^+]$ **b** $[OH^-]$

3 For a 0.05 M aqueous solution of calcium hydroxide ($Ca(OH)_2$) at room temperature, calculate the concentration of:

  a $[OH^-]$ **b** $[H_3O^+]$

Calculating pH

pH essentially stands for the 'potential of hydrogen'. The pH of an aqueous solution is a measurement of the concentration of the hydronium ions $[H_3O^+]$ in the solution using the formula

$$pH = -\log_{10} [H_3O^+]$$

As strong acids and bases completely ionise to form $[H_3O^+]/[OH^-]$ ions, the pH of these solutions can be determined if we know or can calculate the concentration of the H_3O^+ ions in solution. Table 11C–2 shows the relationship between concentration of H_3O^+ and OH^-, and pH.

pH
the quantitative value of how acidic or basic a substance is

VIDEO 11C–1
PERFORMING pH CALCULATIONS

NOTE

Using $\log_{10}$ in the equation for pH removes the need to write pH values to the power of 10, and instead keeps the numbers more manageable.

Also, a difference in pH of 1 means the concentration differs by a factor of 10. For example, lemon juice with a pH of 3 has 100 times the concentration of hydronium ions $[H_3O^+]$ compared to tomatoes, which have a pH of 5.

It is also important to ensure that when you enter this calculation into your calculator, you use $\log_{10}$ and not $\log_e$ (or ln).

Unlike concentration, pH values do not have a unit assigned to them.

Table 11C–2 The relationship between pH, $[H_3O^+]$ and $[OH^-]$ for different solutions and substances.

Solution	pH	$[H_3O^+]$ (M)	pOH	$[OH^-]$ (M)	$[H_3O^+]\,[OH^-]$
1.0 M $HCl(aq)$	0	1	14	10^{-14}	10^{-14}
Stomach acid	1	10^{-1}	13	10^{-13}	10^{-14}
Lemon juice	2	10^{-2}	12	10^{-12}	10^{-14}
Vinegar	3	10^{-3}	11	10^{-11}	10^{-14}
Saliva	6	10^{-6}	8	10^{-8}	10^{-14}
Deionised water	7	10^{-7}	7	10^{-7}	10^{-14}
Baking soda	8	10^{-8}	6	10^{-6}	10^{-14}
Soap	9	10^{-9}	5	10^{-5}	10^{-14}
Household bleach	13	10^{-13}	1	10^{-1}	10^{-14}
1.0 M $NaOH(aq)$	14	10^{-14}	0	1	10^{-14}

Worked example 11C–3: Calculating the pH of strong acids

VIDEO WORKED EXAMPLE 11C–3

Calculate the pH of a 0.0001 M HCl aqueous solution.

Solution

Step 1	Identify the solution as either a strong acid or base.	HCl = strong acid
Step 2	If acid: [acid] = $[H_3O^+]$ If base: [base] = $[OH^-]$	$[H_3O^+] = [HCl]$ $= 0.0001$ M or 1×10^{-4} M
Step 3	Calculate the pH using the formula: $pH = -\log_{10}[H_3O^+]$	$pH = -\log_{10}[H_3O^+]$ $= -\log_{10}[1 \times 10^{-4}]$ $= 4$

Worked example 11C–4: Calculating the pH of bases

VIDEO WORKED EXAMPLE 11C–4

11B STRONG AND WEAK ACIDS AND BASES LINK

Calculate the pH of a 0.0001 M NaOH aqueous solution.

Solution

Step 1	Identify the solution as either a strong acid or base.	NaOH = base
Step 2	If acid: [acid] = $[H_3O^+]$ If base: [base] = $[OH^-]$	$[OH^-] = [NaOH]$ $= 0.0001$ M or 1×10^{-4} M
Step 3	Calculate the concentration of the H_3O^+ ion using the ionic product of water. $[H_3O^+] \times [OH^-] = 10^{-14}\ M^2$	$[H_3O^+] = \dfrac{10^{-14}\ M^2}{[OH^-]}$ $[H_3O^+] = \dfrac{10^{-14}\ M^2}{1 \times 10^{-4}\ M}$ $[H_3O^+] = 1 \times 10^{-10}$ M
Step 4	Calculate the pH using the formula: $pH = -\log_{10}[H_3O^+]$	$pH = -\log_{10}[H_3O^+]$ $= -\log_{10}[1 \times 10^{-10}]$ $= 10$

pH and pOH

When calculating the pH of bases, it is convenient to use a measure of **pOH**. pOH is related to the concentration of hydroxide ions $[OH^-]$ present in an aqueous solution. It can be calculated using the formula

$$pOH = -\log_{10}[OH^-]$$

This can be related back to the ionic product of water to provide an alternative way to determine the pH of a basic aqueous solution at 25°C. The two equations below are useful for these calculations.

$$[H_3O^+] \times [OH^-] = 10^{-14}\ M^2$$

$$pH + pOH = 14$$

pOH
the quantitative value of how acidic or basic a substance is on a scale of 1–14. It is a measure of the concentration of hydroxide ions, $[OH^-]$

Worked example 11C–5: Calculating the pH of bases

VIDEO
WORKED EXAMPLE 11C–5

Calculate the pH of a 5.0×10^{-6} M $Ba(OH)_2$ aqueous solution at 25°C.

Solution

Step 1	Identify the solution as either a strong acid or base.	$Ba(OH)_2$ = base
Step 2	If acid: [acid] = $[H_3O^+]$ If base: [base] = $[OH^-]$. **Note that solution is diprotic.**	$[OH^-] = 2 \times [Ba(OH)_2]$ $= 2 \times 5.0 \times 10^{-6}$ M $= 1.0 \times 10^{-5}$ M
Step 3	Calculate the pOH using the formula: $pOH = -\log_{10}[OH^-]$	$pOH = -\log_{10}[OH^-]$ $pOH = -\log_{10}[1.0 \times 10^{-5}]$ $= 5$
Step 4	Calculate the pH using the formula: $pH + pOH = 14$	$pH = 14 - pOH$ $= 14 - 5$ $= 9$

pH and concentration

If the pH of a solution is known, we can use this to calculate the concentration of hydronium ions $[H_3O^+]$ and therefore the concentration of the solution.

It can be calculated using the following formula:

$$[H_3O^+] = 10^{-pH}$$

You may like to use the flow chart in Figure 11C–3 to assist you when completing calculation questions involving pH.

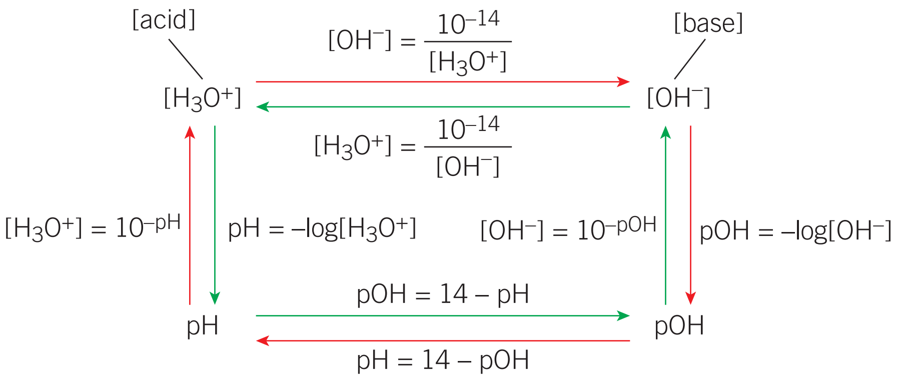

Figure 11C–3 This flow chart can help with calculating the pH of a solution based on what you know from the question, thereby determining what is required in order to find the pH.

VIDEO
WORKED EXAMPLE 11C–6

Worked example 11C–6: Calculating concentration of ions from the pH

An aqueous solution of ammonia (NH_3) has a pH of 9 at 25°C.

Calculate the concentrations of:

a hydronium ions $[H_3O^+]$
b hydroxide ions $[OH^-]$

Solution

Step 1	Calculate the concentration of $[H_3O^+]$ using the pH. $[H_3O^+] = 10^{-pH}$	$[H_3O^+] = 10^{-pH}$ $= 10^{-9}$ M
Step 2	Calculate the concentration of $[OH^-]$ using the ionic product of water. $[H_3O^+] \times [OH^-] = 10^{-14}$ M^2	$[OH^-] = \frac{10^{-14} \text{ M}^2}{[H_3O^+]}$ $[OH^-] = \frac{10^{-14} \text{ M}^2}{10^{-9}}$ $[OH^-] = 10^{-5}$ M

Check-in questions – Set 2

1 Calculate the pH of the following aqueous solutions at 25°C.
  a a solution in which $[H_3O^+] = 0.10$ M
  b a solution in which $[H_3O^+] = 1.0 \times 10^{-6}$ M
  c a 0.030 M solution of HCl
2 Calculate the pOH of the following aqueous solutions at 25°C.
  a a solution in which $[OH^-] = 0.01$ M
  b a solution in which $[OH^-] = 1 \times 10^{-3}$ M
  c a 0.025 M solution of $Ca(OH)_2$
3 The pH of a water sample is 6 at 25°C. Calculate the $[H_3O^+]$ in the sample.
4 The pH of body wash is 8 at 25°C. Calculate the $[OH^-]$ in the sample.
5 Determine the pH at 25°C of a 250 mL solution that contains 0.160 g of dissolved HNO_3.
6 Determine the pH at 25°C of a 50 mL solution that contains 0.060 g of dissolved NaOH.

Other calculations involving pH

pH calculations can be used along with other quantitative chemistry calculations, such as those involving dilutions. This is best expressed with the equation

$$c_1V_1 = c_2V_2$$

where:

c_1 = The starting concentration
V_1 = The starting volume
c_2 = The final concentration
V_2 = The final volume

NOTE

In this calculation, both the concentration and volume are present on the left- and right-hand sides of the equation. Therefore, it doesn't matter what units of measurement are used for each, as long as the same units are used on both sides of the equation (e.g. volume can be in mL instead of L).

Worked example 11C–7: Calculating the pH of diluted solutions from initial concentration

VIDEO
WORKED
EXAMPLE 11C–7

A 20.0 mL aqueous solution of 0.050 M HCl is diluted to 100.0 mL. What is the pH of the resulting solution at 25°C?

Solution

Step 1	From the information provided in the question, determine known and unknown quantities according to the following formula and solve for the unknown. $c_1V_1 = c_2V_2$ **It is at this step that you should also ensure units of measurement on each side of the equation for both concentration and volumes are the same.**	$c_1 = 0.050$ M $V_1 = 20.0$ mL $c_2 = ?$ $V_2 = 100.0$ mL $c_2 = \frac{c_1V_1}{V_2}$ $c_2 = \frac{0.050 \times 20.0}{100.0}$ $= 0.000500 = 5 \times 10^{-4}$ M
Step 2	Is the solution an acid or a base? If solution is monoprotic, acid: $[\text{acid}] = [H_3O^+]$ If base: $[\text{base}] = [OH^-]$	HCl is an acid. $[\text{acid}] = [H_3O^+]$ $\therefore [H_3O^+] = [HCl] = 5 \times 10^{-4}$ M
Step 3	Calculate the final pH using the final concentration. $pH = -\log_{10}[H_3O^+]$	$pH = -\log_{10}[H_3O^+]$ $= -\log_{10}[5 \times 10^{-4}]$ $= 3.3$

Worked example 11C–8: Calculating the pH of diluted solutions from initial pH

VIDEO
WORKED
EXAMPLE 11C–8

WORKSHEET
11C–1
pH
CALCULATIONS

A 50.0 mL solution of nitric acid of pH 1.0 is added to 250.0 mL of water. What is the pH of the resulting solution?

Solution

Step 1	Calculate the concentration of the initial solution. $[H_3O^+] = 10^{-pH}$	$[H_3O^+] = 10^{-1.0}$ $= 0.10$ M $\therefore [HNO_3] = [H_3O^+] = 0.10$ M
Step 2	Calculate the concentration of the final solution. $C_1V_1 = C_2V_2$	$C_1 = 0.10$ M $V_1 = 50.0$ mL $C_2 = ?$ $V_2 = 50.0 \text{ mL} + 250.0 \text{ mL} = 300.0$ mL $c_2 = \frac{c_1V_1}{V_2}$ $C_2 = \frac{0.10 \times 50.0}{300.0}$ $= 0.0167$ M

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Step 3	Is the solution an acid or a base? If solution is monoprotic, acid: [acid] = $[H_3O^+]$ If base: [base] = $[OH^-]$	HNO_3 is an acid. [acid] = $[H_3O^+]$ $\therefore [HNO_3] = [H_3O^+]$ = 0.0167 M
Step 4	Calculate final pH using the final concentration $pH = -\log_{10} [H_3O^+]$	$pH = -\log_{10}[H_3O^+]$ $= -\log_{10}[0.0167\ M]$ = 1.8

Acid–base indicators

Indicator
a substance that changes colour in the presence of an acid or a base

Another way we can quickly and easily establish if a solution is acidic or basic is by using an acid–base **indicator**. This is a substance that changes colour in the presence of an acid or a base. Knowing the colours of common indicators allows us to identify substances as an acid or a base. Most indicators themselves are actually weak acids or bases, where the conjugate they form is a different colour.

Figure 11C–4 Universal indicator paper contains a mixture of indicators: it uses a qualitative approach, changing colour in solution to indicate approximate pH. Using a scale, this can be translated into a quantitative pH number for the solution.

2B MOLECULAR SHAPES AND POLARITY LINK

Let's look at the commonly used indicator phenolphthalein as an example, as seen in Figure 11C–5. Phenolphthalein ($C_{20}H_{14}O_4$) is a water soluble (polar) organic compound. In the pH range 0–8, phenolphthalein is colourless but in the pH range 8–14, it is pink. Notice in the equation depicted on the following page that it is an example of a weak acid, where the acid is clear and the conjugate base ion is pink. When phenolphthalein is added to a basic solution in which the concentration of hydronium ions $[H_3O^+]$ is low, the conjugate base ion is formed in greater concentration, turning the solution pink. This colour change indicates that the solution is basic.

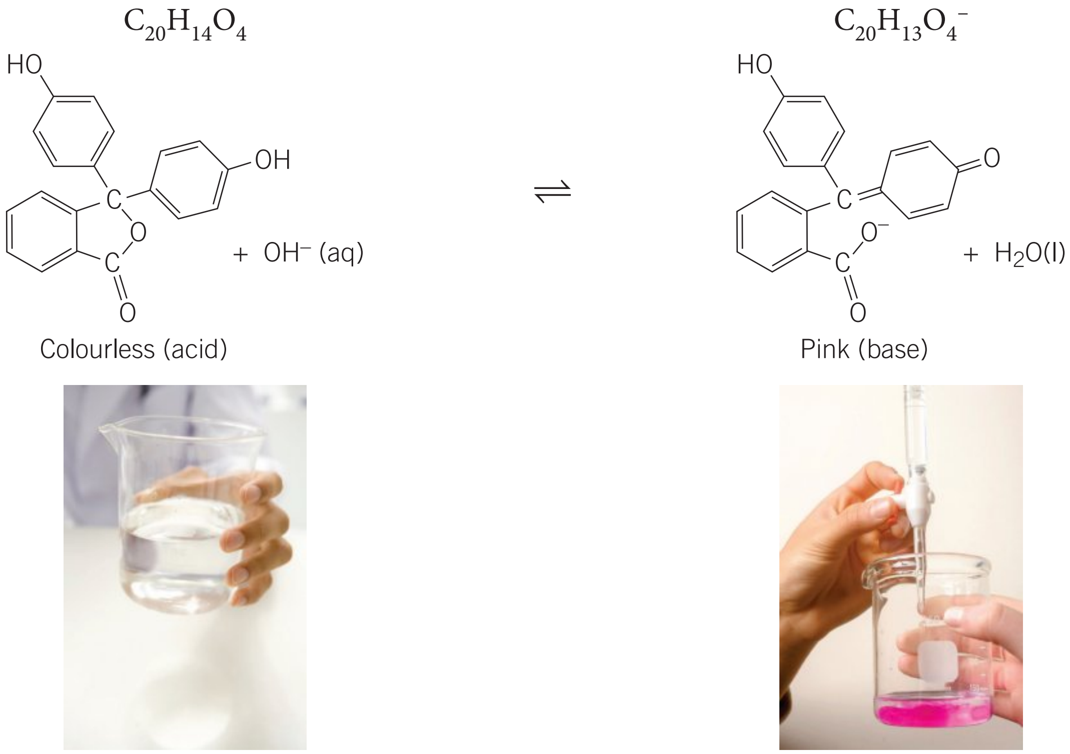

Figure 11C–5 Using phenolphthalein to test the pH of a solution

Commercial indicators

The colour change of an indicator can be used quickly to qualitatively determine if a substance is acidic or basic. In Chapter 13D, you will learn how indicators can be used during volumetric analysis to quantitatively determine the concentration of an acid or base. Table 11C–3 lists some commercial acid–base indicators commonly used in laboratories. Each indicator has a specific range in which it will change colour.

LINK **13D** VOLUMETRIC ANALYSIS OF ACIDS AND BASES

NOTE

You will have access to this list during your assessment tasks and will be expected to identify and select appropriate indicators for given reactions that produce products with a known pH.

Table 11C–3 Common indicators and their colours, as listed in the VCAA Data Booklet

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

Natural indicator a substance that occurs in nature and changes to different colours when in contact with acids or bases

Natural indicators

Outside of the chemistry laboratory, there are many examples of **natural indicators**. These substances are naturally occurring in nature and change to different colours when in contact with acids or bases. Have you ever admired the range of coloured flowers produced by the hydrangea? The colour of hydrangea flowers is caused by the pH of the soil in which they are planted. To grow pink flowers, gardeners will make the soil basic; however, for blue flowers, the soil is acidic. Hydrangeas are not the only plant to indicate differing pH. Lichens contain a substance called litmus that is red under acidic conditions and blue in basic conditions. The dye is extracted to create the commonly used litmus paper.

Figure 11C–6 Left: The colour of these hydrangea flowers is a result of the soil in which the plants were grown. Right: Litmus paper was developed after noticing that lichens were different colours in different pH ranges. They turn blue, as shown in this picture in the presence of a basic solution.

Table 11C–4 shows examples of natural indicators and the colours they turn in acidic and basic environments.

Table 11C–4 Many plant products can be used as indicators once the colour-changing part of them has been extracted.

Natural indicator	In nature	In acid	In base
Litmus solution	Purple	Red	Blue
Turmeric solution	Yellow	Yellow	Deep red
Beetroot juice	Purple	Purple	Yellow
Red cabbage juice	Purple	Pink/Red	Green/Yellow
China rose solution	Red	Magenta	Green

Digital pH meters

13C PURIFYING WATER LINK

In Chapter 13, you will learn that pH meters have a number of useful everyday applications, such as being used to measure the pH of water samples to determine if they are within safe levels. They can also be used to test the pH and suitability of soils. pH meters provide a quick, accurate and portable means of testing pH. Results can be collected by scientists in the field immediately. To ensure their accuracy, pH meters need to be **calibrated**. Calibration involves adjusting the pH meter by measuring solutions of a known pH value. This helps to minimise **systematic errors**.

LINK 16B SCIENTIFIC EVIDENCE

Calibrated any instrument for quantitative measurement that has been checked, verified and rectified if required

Systematic error when the readings obtained from measurements differ from the 'true' value consistently in one direction every time

Figure 11C–7 A pH meter can be used to test the pH of the soil. Some plants require a certain pH for optimal growth or to produce flowers of a particular colour.

A pH meter provides a quantitative measurement of pH, while the natural and commercial indicators provide only a qualitative measurement. While the natural and commercial indicators tend to be accurate, they are not as precise as a pH meter as they give a range in which the pH value of the substance may lie rather than a specific single numerical value.

Figure 11C–8 This pH meter is calibrated with a standard solution of pH 7.

Check-in questions – Set 3

1. What is a natural indicator?
2. What colour would phenolphthalein indicator be when added to a solution with pH 10?
3. What colour would methyl red indicator be when added to a solution with pH 1?
4. What colour would red cabbage indicator be when added to a solution with pH 12?

11C SKILLS

Indices

Mathematical skills are often required in quantitative chemistry.

In this chapter, you saw that the index laws are an important part of pH calculations and calculations of concentration using the ionisation constant of water. To calculate with indices, you need to be able to use the 'laws of indices', which are the rules applied for simplifying calculations involving powers of the same base. The rules that are relevant for this chapter are outlined below.

Rule 1: Multiplying indices

When multiplying indices with the same *base*, the *powers* are added together. An example is shown below:

power

$$10^{-8} \times 10^{-6} = 10^{-14}$$

base

Multiplying indices is useful for calculations involving the ionisation constant of water, K_w, where:

$$K_w = [H_3O^+] \times [OH^-] = 10^{-14}\ \text{M}^2 \text{ at } 25°\text{C}$$

Rule 2: Dividing indices

When dividing indices with the same *base*, the *powers* are subtracted. An example is shown below:

$$\frac{10^{-14}}{10^{-6}} = 10^{-8}$$

Dividing indices is useful for calculations involving the ionisation constant of water, K_w, where either the hydronium ion concentration, $[H_3O^+]$, or the hydroxide ion concentration, $[OH^-]$, needs to be determined.

$$[H_3O^+] = \frac{10^{-14}\ \text{M}^2}{[OH^-]\ \text{M}}$$

Rule 3: Power of 0

Any base (which is not a 0) raised to the power of 0 will equal 1. For example:

$$10^0 = 1$$

This rule is useful if you are trying to determine the hydronium ion concentration, $[H_3O^+]$, where the pH of a solution is 0, according to the following equation:

$$\begin{aligned}[H_3O^+] &= 10^{-\text{pH}}\\ &= 10^0\\ &= 1\end{aligned}$$

You need to treat completing calculations for pH or ion concentrations in the same way that you treat mathematics topics. The way to improve your handling of indices and these types of equations is to complete more questions, as repetitive as it may seem. Showing full working and writing out the formula each time you use it in a question is another important way to ensure this equation is recalled easily when you need it in assessment situations. The other aspect of completing more of these types of questions is that you become more familiar with the different types of acids and bases and their molecular formulas. In solving calculations involving these, you also unknowingly practice writing ionic or covalent formulas and can determine how many protons, H^+ ions, an acid will donate or a base will accept.

LINK **4A** FORMATION AND NAMING OF IONIC COMPOUNDS

Accuracy versus precision with indicators
In this section, the Study Design explicitly refers to the terms 'accuracy' and 'precision' of measurements involving indicators.

Accuracy is a measure of how close recorded data is to the 'true value' of the quantity being measured. For this reason, usually instruments that provide quantitative data that are not calibrated correctly lead to inaccuracy. As such, digital pH meters are most involved with a measure of accuracy.

On the other hand, precision is how close all measurements of data are to each other. While this is also commonly associated with quantitative data, qualitative data like that provided by commercial and natural indicators can give precise results if they are tested on the same solution and turn the same colour. However, what you need to know is that this data may be precise but not accurate.

These terms will be discussed in relation to different types of experimental errors in Chapter 16.

LINK **16B** SCIENTIFIC EVIDENCE

Section 11C questions

1 a Define 'pH'.
  b Calculate the pH of a 0.013 M solution of HCl.
2 Calculate the pH for each of the following solutions.
  a 0.0109 M NaOH
  b 3.1 M LiOH
  c 0.12 M $Mg(OH)_2$
3 Determine the $[H^+]$ and $[OH^-]$ concentration of a solution of HBr that has a pH of 5.5.
4 Calculate the pH of a 0.060 M solution of sulfuric acid, H_2SO_4.
5 Use a Venn diagram to compare the similarities and differences between natural and commercial indicators.
6 When conducting an experiment, a student collected 50 mL of acid, 50 mL of base and 50 mL of water. They forgot to label the beakers and decided to test the three solutions using indicators to determine which sample was which. Use the table below to predict which beaker contains acid, which contains base and which contains water.

Indicator	Beaker A	Beaker B	Beaker C
Phenolphthalein	Clear	Clear	Pink
Cabbage solution	Red	Purple	Green
Methyl red	Red	Yellow	Yellow

7 Determine the pH of a 100.0 mL solution containing 1.00 g of dissolved barium hydroxide, $Ba(OH)_2$. If this solution was tested with bromophenol blue, what colour would this indicator turn?

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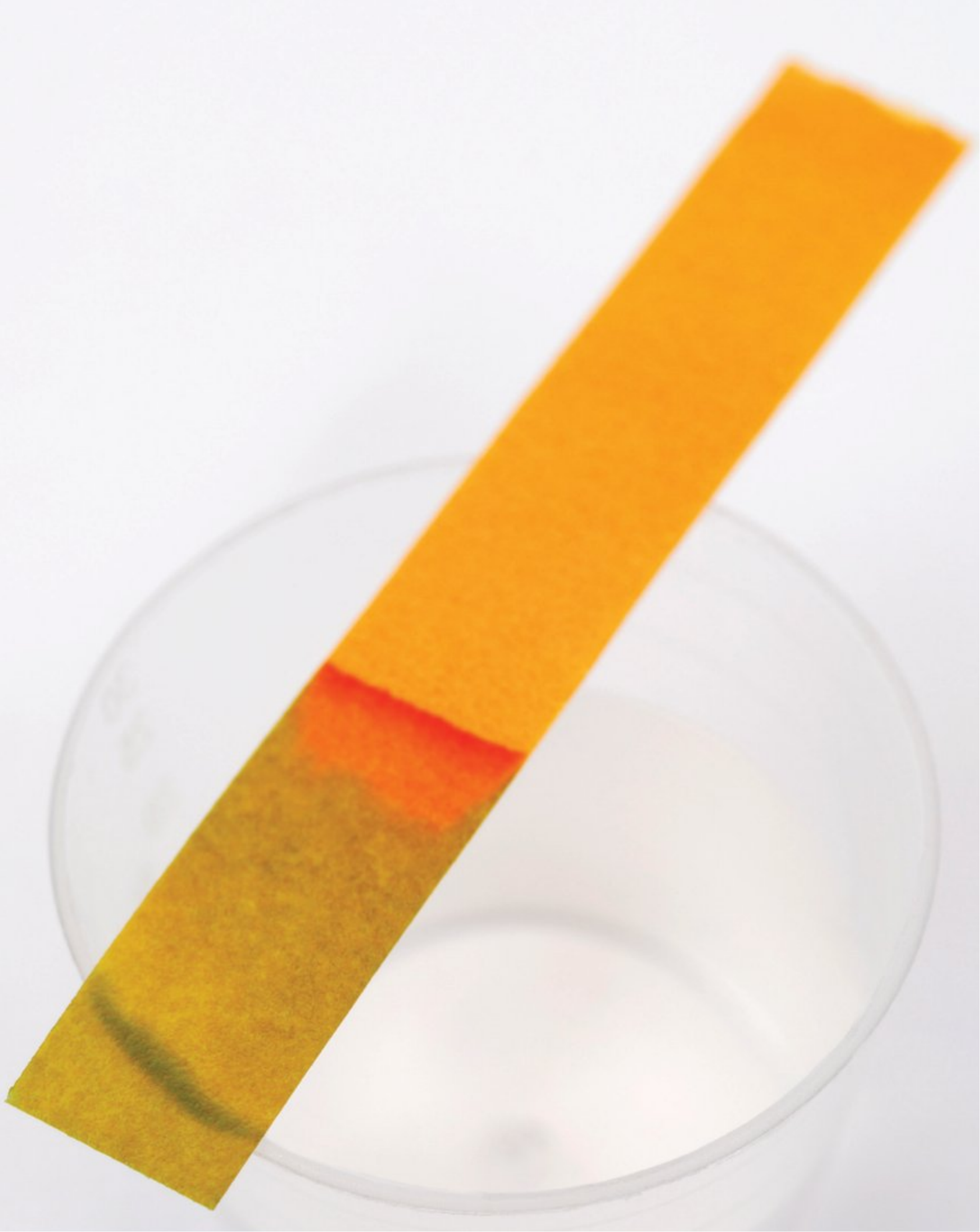
8 Calculate the final pH of the following solutions.

  a A 9.0 mL solution with pH of 2.1 is diluted to 90 mL.
  b A 10.0 mL solution with pH of 14.4 is diluted to 100 mL.
  c 2475 mL of water is added to 25.0 mL of a solution with pH 0.12.
  d 1229.6 mL of water is added to 8.7 mL of a solution with pH of –0.40.

9 What volume of water is required to dilute 30 mL of nitric acid (HNO_3) solution of pH 3 to pH 5?

10 Determine the pH of a solution containing 10 g of sodium hydroxide (NaOH) dissolved in 40 mL of water.



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11D Carbon dioxide as a weak acid

Study Design:
Application 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

Glossary:
Acid rain
Aragonite
Carbon dioxide
Greenhouse gas
Photosynthesis

ENGAGE

Carbon dioxide: the greenhouse gas that is acidifying the oceans

Carbon dioxide is well known for being the primary greenhouse gas emitted and for its contribution to global warming, changing climates and rising sea levels. Produced through human activities like burning fossil fuels, carbon dioxide levels have reached an all-time high worldwide, and these molecules can remain in the atmosphere for hundreds of years. Carbon dioxide pollution and the desperate need for countries to lower their emissions has been well documented in the media. However, there is a less well known and equally dangerous effect that carbon dioxide is having on Earth, and that is in the ocean.

Did you know that carbon dioxide forms an acid when dissolved in water? According to data collected by NASA, the oceans absorb approximately 25% of the carbon dioxide produced through the burning of fossil fuels. This is resulting in acidification of the oceans, which, if it continues, will have a large impact on many aquatic plant and animal species. In this section, you will learn more about this process, as well as the effects of other acids in the environment.

Figure 11D–1 The ocean is a huge carbon sink – a natural reservoir that absorbs and stores carbon.

EXPLAIN

Carbon dioxide as a weak acid

Sources of carbon dioxide

Carbon dioxide is a natural part of Earth's ecosystems and is essential to life. Plants require carbon dioxide in the atmosphere to undergo **photosynthesis**, the chemical reaction vital in making glucose – an energy source and a structural molecule in all plants. The equation for photosynthesis is

$$\underset{\text{Carbon dioxide}}{6CO_2(g)} + \underset{\text{Water}}{6H_2O(l)} \xrightarrow{\text{Light}} \underset{\text{Glucose}}{C_6H_{12}O_6(aq)} + \underset{\text{Oxygen}}{6O_2(g)}$$

Carbon dioxide
a gas, CO_2, that is an important part of the carbon cycle

Photosynthesis
a chemical reaction that plants utilise to create glucose from carbon dioxide

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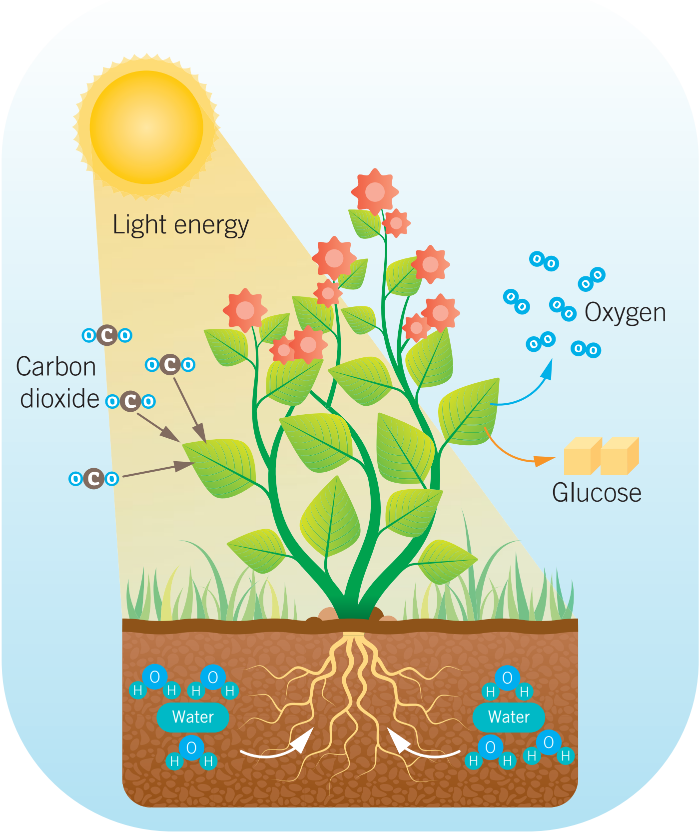

Figure 11D–2 Plant undergoing photosynthesis where carbon dioxide enters and oxygen and glucose are produced

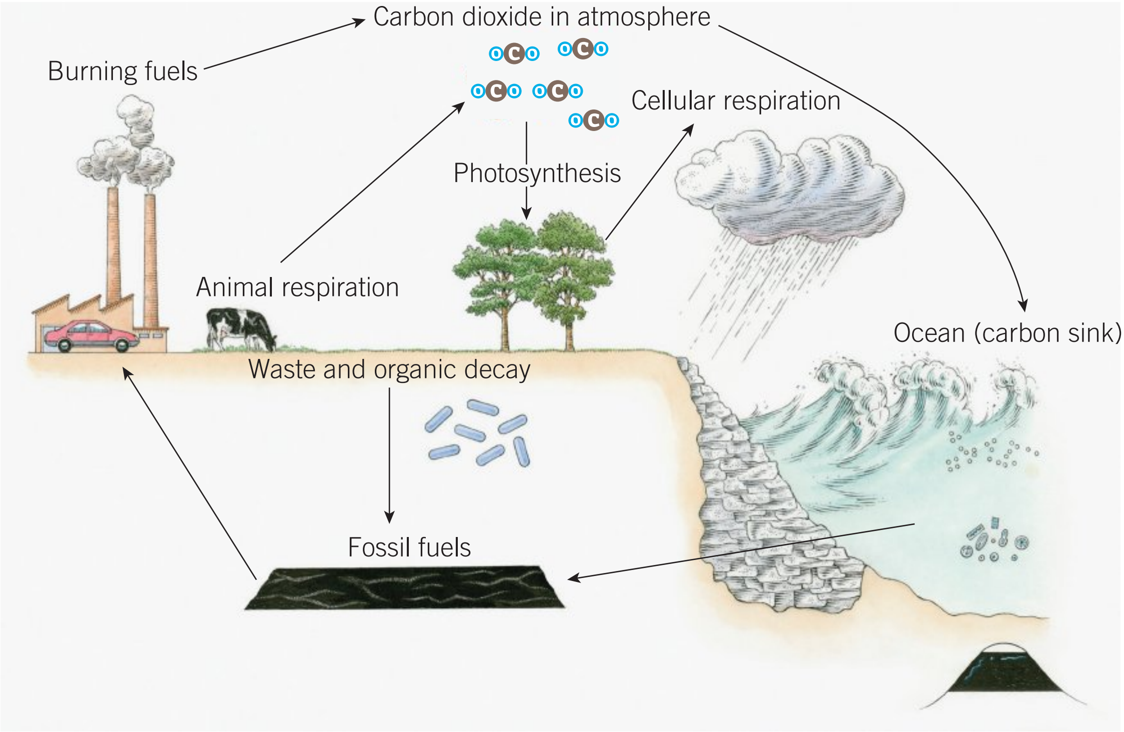

Figure 11D–3 The carbon cycle

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As the human population and its reliance on industry has increased over time, so too have the levels of carbon dioxide in the atmosphere. Carbon dioxide is known as a **greenhouse gas** – it causes climate change by trapping heat within Earth's atmosphere. Carbon dioxide is a by-product of complete combustion and is released when fossil fuels are burned. The world's reliance on the burning of coal, gas and oil to fuel industry and agriculture, along with the destruction and removal of forests, are the primary reasons carbon dioxide levels continue to rise.

Greenhouse gas a gas that contributes to the greenhouse effect by trapping heat close to the Earth's surface – e.g. carbon dioxide

VIDEO 11D–1 ENVIRONMENTAL IMPACTS OF CARBON DIOXIDE

Carbon dioxide in the ocean

Carbon dioxide from the sources mentioned above predominately stays in the atmosphere in gaseous form. However, approximately 25% of the carbon dioxide added to the atmosphere is absorbed by the ocean. As carbon dioxide levels increase, so too does the amount absorbed by oceans. This increasing absorption of carbon dioxide is altering the chemistry of ocean waters, most notably the pH. Ocean water naturally has a slightly basic pH of around 8.2.

In Figure 11D–4, we can clearly see the correlation between rising carbon dioxide levels and the decreasing pH of the North Pacific Ocean.

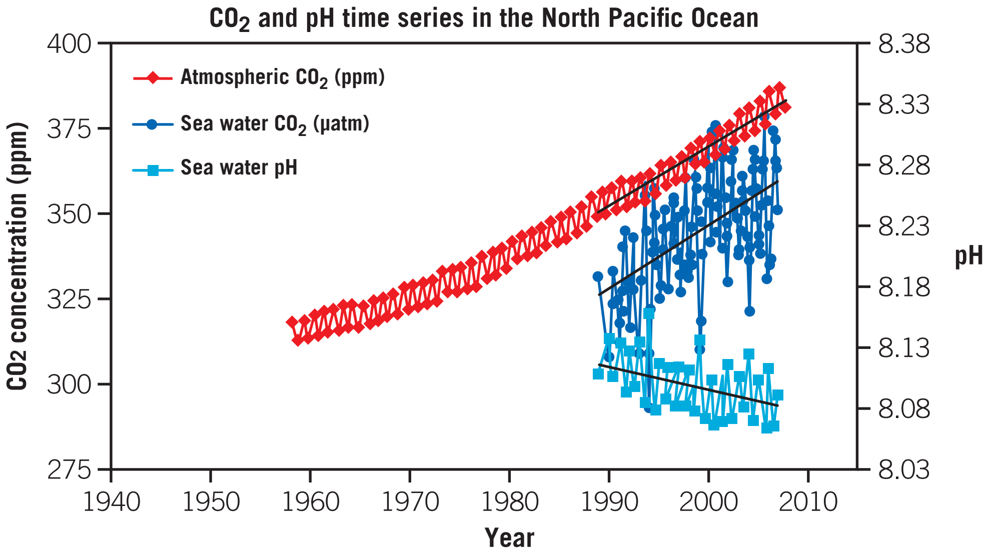

Figure 11D–4 Measurements of carbon dioxide levels in the atmosphere and in the ocean show that there is a correlation between increasing CO_2 levels and decreasing pH of ocean waters.

So how is it that carbon dioxide can cause the pH to decrease and become more acidic? This occurs because when carbon dioxide dissolves in sea water it reacts with water molecules to form carbonic acid (H_2CO_3), a weak acid.

$$H_2O(l) + CO_2(g) \rightleftharpoons H_2CO_3(aq)$$

As both a weak and diprotic acid, carbonic acid will further react with the water to produce hydronium ions (H_3O^+).

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H_3O^+$$

$$HCO_3^-(aq) + H_2O(l) \rightleftharpoons CO_3^{2-}(aq) + H_3O^+$$

Hydronium ion

Hydrogen carbonate ion

Carbonate ion

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Given that pH = $-\log_{10}[H_3O^+]$, increasing the concentration of hydronium ions (H_3O^+) results in a lower pH. Scientists who have been measuring ocean pH have found that pH of surface ocean water has dropped by approximately 0.1 pH units. Although this doesn't seem like a lot, we need to remember that pH is measured on a logarithmic scale and that this correlates to an increase in acidity of approximately 30%.

Impacts of ocean acidity

The decreasing pH levels of the ocean are of significant concern. The effects of changing pH levels are already being observed in many marine species and have the potential for devastating marine life as we know it, affecting ocean health and the economy the ocean provides.

Aragonite a source of calcium carbonate

Specifically, decreasing pH resulting from higher concentrations of carbon dioxide reduces the concentration of carbonate ions (CO_3^{2-}) and the availability of **aragonite** (a source of calcium carbonate) in the ocean. Carbonate ions are combined with calcium ions from sea water and used by many marine organisms, such as coral and shellfish, to build and maintain their shells and skeletons. In levels of low concentrations of available carbonate ions, marine animals are unable to fully form their shells. If ocean water becomes too acidic, shells and skeletons can begin to dissolve, making the species susceptible to predators. In the long term, this would result in not only damage to these marine animals but also other animals further up the food chain.

Sea water has a high concentration of dissolved carbonate and other materials like aragonite. As carbonic acid levels increase, the H^+ (H_3O^+) ions they produce by reacting with water react with available carbonate ions, thereby removing them from the sea water.

$$H^+(aq) + CaCO_3(s) \rightleftharpoons HCO_3^-(aq) + Ca^{2+}(aq)$$

In predictions made by scientists, it is believed that if pH levels drop to 7.8, the carbonate ion concentration in ocean water would decrease by 50%.

Reactions in the ocean sit in careful balance. Figure 11D–5 shows how the increase in concentration of atmospheric carbon dioxide is linked through a series of equations to the decrease of calcium carbonate in marine organisms. As the level of atmospheric carbon dioxide increases, so too does the concentration of dissolved carbon dioxide, resulting in increased levels of carbonic acid and [H^+ ions]. The increased, excess [H^+ ions] bond with free carbonate ions, resulting in fewer carbonate ions available for shell production. Calcium carbonate in shells and coral can also react with the excess [H^+ ions], breaking down the shell to release the carbonate to bind with the H^+ ions. Essentially, carbonic acid formed from dissolved carbon dioxide 'steals' the carbonate needed by marine organisms to form shells and skeletons.

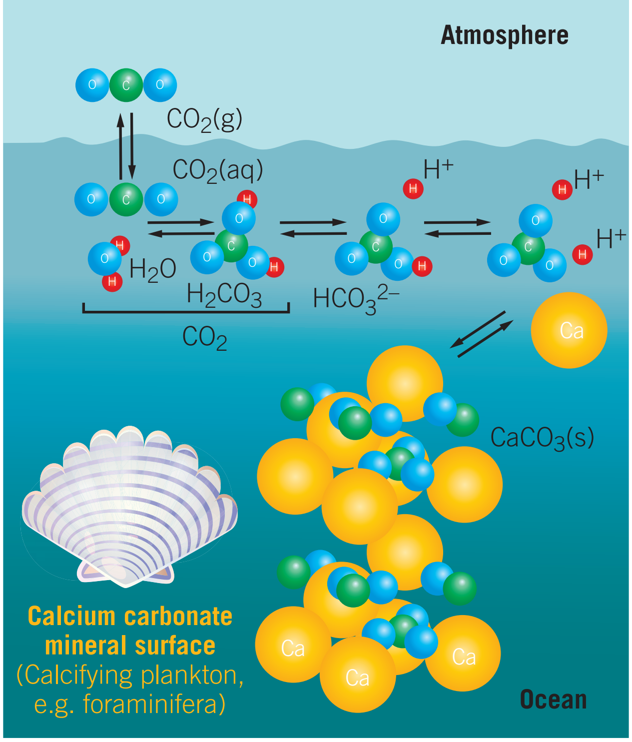

Figure 11D–5 As the level of atmospheric carbon dioxide increases, so too does the concentration of dissolved carbon dioxide.

Acid rain

Rain water, like sea water, does not have a purely neutral pH. When rain falls, it collects impurities, such as atmospheric carbon dioxide, thereby changing the pH and making it slightly acidic (Figure 11D–6). Naturally, this pH is approximately 5.6.

However, to be classified as **acid rain**, the pH needs to be below 5. One of the main causes of acid rain is the burning of fossil fuels, which you learned about in Chapter 7.

It is also important to realise that Earth's water is in a closed system, called the water cycle. Water, which is stored in the oceans and lakes, evaporates and forms clouds of water vapour. As this water vapour condenses, it falls to the ground as rain. If this water is already slightly acidic due to the absorbance of atmospheric CO_2 by ocean water, then this water, when evaporated and combined with further atmospheric CO_2 and other impurities, can lead to an increased likelihood of acid rain.

Acid rain precipitation with a pH less than 5

LINK **7A** ORGANIC CHEMISTRY IN SOCIETY

WORKSHEET 11D–1 ACID RAIN

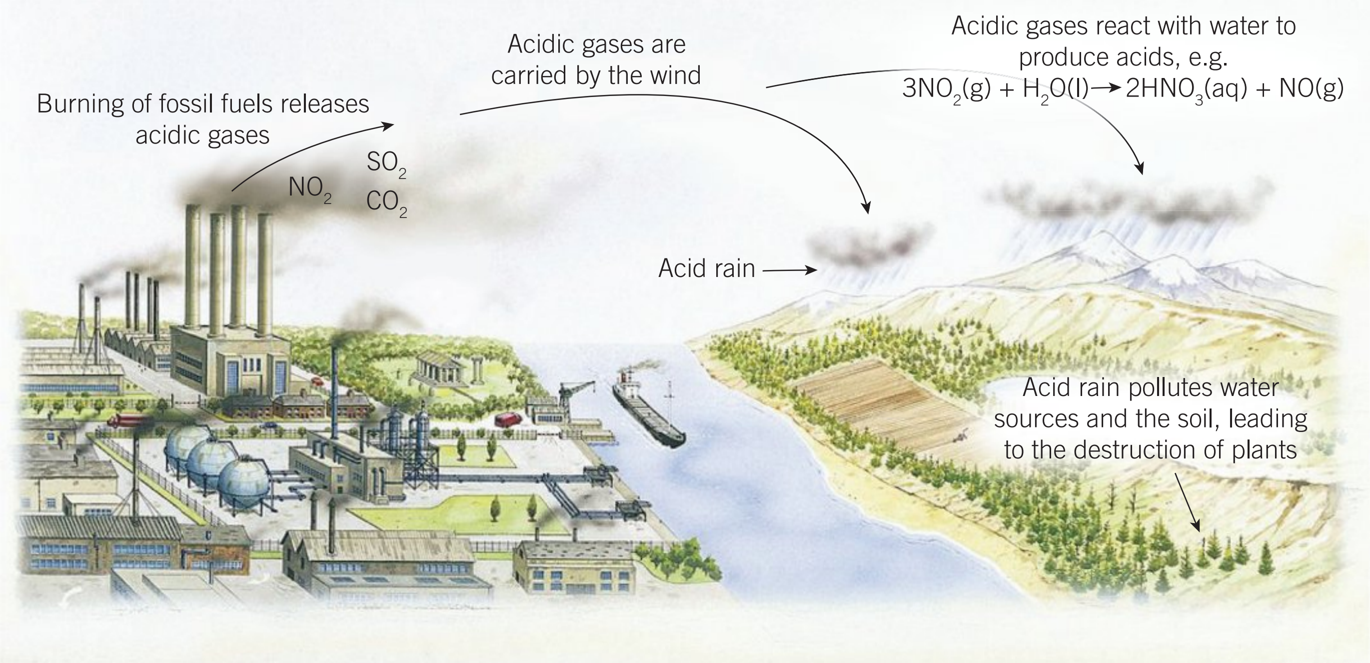

Figure 11D–6 Primary causes of acid rain

11D SKILLS

Considering other implications of ocean acidity

This section covered the environmental (specifically biological) impact of ocean acidity on marine organisms, particularly the formation and maintenance of healthy shells and skeletons.

While the current ocean environments can vary greatly in their pH over an organism's lifetime, if an organism was exposed to this decreased pH throughout its entire life, then more significant effects will obviously be visible. Mollusc shells, for example, could completely dissolve in acidic conditions. And their demise may further lead to reduced food sources for fish, birds and marine mammals.

However, it is not all doom and gloom for all species. Algae and seagrasses may actually benefit from higher carbon dioxide concentrations, as this would increase their abilities to photosynthesise.

There are also other significant non-environmental implications of changes in the acidity of the ocean. Humans rely on the health of the ocean and organisms within it for a number of socioeconomic reasons, including:

- Food security
  - Molluscs, like oysters and mussels, will decline in availability, costing billions of dollars globally to businesses and fisheries.
- Sheltering of land
  - Coral reefs protect the shorelines from the destructive nature of storms and cyclones. Without coral reefs, we would see more property damage, erosion and potentially loss of lives.
- Tourism
  - Activities such as snorkelling and scuba diving, for example, in the Great Barrier Reef will decline.
- Fishing industry
  - Fisheries and independent fishers may have regulations imposed on them to reduce overfishing, especially if some organisms become endangered or extinct.

Section 11D questions

1 Write the equation for the removal of carbon dioxide from the atmosphere by plants.

2 Write an equation to show how carbon dioxide, when dissolved in ocean water, forms an acid.

3 What does the term 'weak acid' mean in chemistry?

4 What is the pH of ocean water currently estimated to be? What is predicted to happen to this value if carbon dioxide emissions into the atmosphere continue to increase?

5 Scientists are concerned that the pH of ocean water has decreased from 8.2 to 8.1 over time. Explain why this small change in pH is of concern. Use the [H^+ ions] in your reasoning.

6 Oyster farmers have reported a significant decline in oyster beds, which is thought to be due to incomplete formation of their shells. Explain how carbonic acid affects the formation of oyster shells.

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11E Neutralisation reactions

Study Design:
Neutralisation reactions to produce salts:

- reactions of acids with metals, carbonates and hydroxides, including balanced full and ionic equations, with states
- types of antacids and their use in the neutralisation of stomach acid

Glossary:
Neutralisation reaction
Spectator ion

ENGAGE

What's that fizzing?

Have you ever eaten Wizz Fizz® Sherbet? Remember the taste combination of sweet and sour, while your tongue feels like it is fizzing? Sherbet contains citric acid and sodium hydrogen carbonate. When they react on your tongue the fizzing you feel is carbon dioxide bubbles being produced. It is also a neutralisation reaction.

Figure 11E–1 Acids typically have a sour taste.

EXPLAIN

Neutralisation reactions

Neutralisation reactions involve an acid reacting with a base. It is known as a neutralisation reaction because the products formed are neutral water and salt (Figure 11E–2).

These reactions are represented by the following ionic equation:

$$\underset{\text{(acid)}}{H^+(aq)} + \underset{\text{(base)}}{OH^-(aq)} \rightarrow H_2O(l)$$

Neutralisation reaction a type of reaction involving an acid reacting with a base. The products formed are neutral water and salt

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4A FORMATION AND NAMING OF IONIC COMPOUNDS LINK

Salt is a product of all neutralisation reactions. Remember that salt is an ionic compound formed from cations and anions. In neutralisation reactions, the cation comes from the base and the anion comes from the acid. Knowing this helps to write the products of acid–base neutralisation reactions.

$$\underset{\text{Acid}}{HA} + \underset{\text{Base}}{BOH} \rightarrow \underset{\text{Salt}}{BA} + \underset{\text{Water}}{H_2O}$$

14A SOURCES OF SALT IN WATER LINK

It is possible to group neutralisation reactions together based on the similarity of the reactants present, and this allows us to accurately predict the products of these neutralisation reactions.

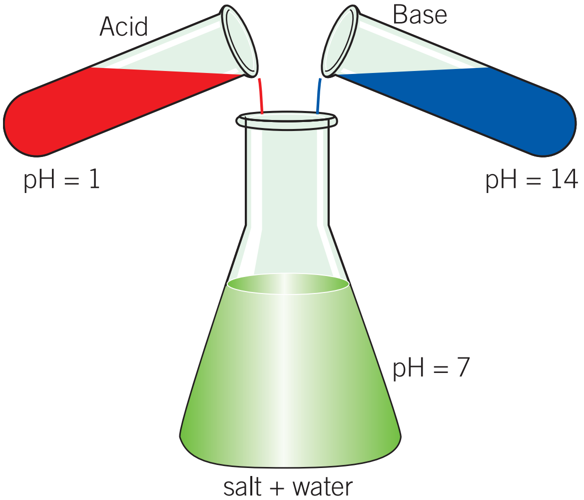


Figure 11E–2 A simplified representation of a neutralisation reaction

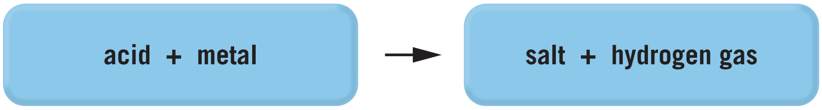
Reactions of acids

Although the Study Design dot point for this section classifies the different types of reactions of acids as being neutralisation reactions, they are in fact not. Only acids reacting with some metal hydroxides (which are alkaline in their nature) are an example of a neutralisation reaction.

Here, the focus will be on three types of reactions involving acids:

- acids + metals
- acids + metal carbonates
- acids + metal hydroxides.

Acid + metal reactions



$2HCl(aq) + Mg(s) \longrightarrow MgCl_2(aq) + H_2(g)$
(salt)

Figure 11E–3 Metals reacting with an acid. Bubbles of hydrogen gas can be seen forming on the surface of the metal.

When hydrochloric acid is added to a strip of magnesium ribbon, bubbles form rapidly, the temperature increases and the magnesium metal appears to disappear, and if the gas is collected and tested with a match, a loud pop is heard. The products of this reaction are hydrogen gas and magnesium chloride salt, as shown in Figure 11E–3.

Worked example 11E–1: Writing equations for acid–metal reactions

VIDEO WORKED EXAMPLE 11E–1

Write a fully balanced equation for the reaction that results when aqueous sulfuric acid is added to a strip of aluminium.

Step 1	Write out the reactants.	$H_2SO_4 + Al$
Step 2	Deduce the products formed following the rule: acid + metal → salt + hydrogen gas Deduce the salt formed: metal cation and acid anion. **Remember to check the charge on the ions and make sure you balance them when forming the salt.**	Hydrogen gas (H_2) + Salt Salt: Al^{3+} SO_4^{2-} = $Al_2(SO_4)_3$
Step 3	Write out the reactants and products into a chemical equation format.	$H_2SO_4 + Al \rightarrow Al_2(SO_4)_3 + H_2$
Step 4	Assign states to all reactants and products. **Remember to check the solubility rules for the salt to determine if it will be in a solid or aqueous state.**	$H_2SO_4(aq) + Al(s) \rightarrow Al_2(SO_4)_3(aq) + H_2(g)$
Step 5	Balance the equation.	$3H_2SO_4(aq) + 2Al(s) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$

LINK **4A** FORMATION AND NAMING OF IONIC COMPOUNDS

LINK **4C** PRECIPITATION REACTIONS

Worked example 11E–2: Writing equations for acid–metal reactions with spectator ions

VIDEO WORKED EXAMPLE 11E–2

Write a balanced ionic equation for the reaction that results when aqueous sulfuric acid is added to a strip of aluminium.

Step 1	Write any dissolved ionic compounds (aq) as its separate ions. Remember: (s), (l) and (g) substances won't break into ions.	$H^+(aq) + SO_4^{2-}(aq) + Al(s) \rightarrow Al^{3+}(aq) + SO_4^{2-}(aq) + H_2(g)$
Step 2	Cross out any ions that appear in the same state on both sides. These ions have not participated in the reaction and are termed **spectator ions**.	$H^+(aq)$ + ~~SO_4^{2-} (aq)~~ + $Al(s) \rightarrow Al^{3+}(aq)$ + ~~SO_4^{2-} (aq)~~ + $H_2(g)$
Step 3	Rewrite the equation, deleting the spectator ions. Ensure the equation is still balanced and states are included on all species. Note: The charges on both sides of the reaction also need to be equal. Here, there is a 6+ total charge on each side.	$6H^+(aq) + 2Al(s) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

Spectator ion an ion that does not take part in the precipitation of the solid but remains soluble in solution

Acid + metal carbonate reactions

acid + metal carbonate → **salt + water + carbon dioxide**

$C_4H_6O_6(aq) + 2NaHCO_3(s) \longrightarrow Na_2C_4H_4O_6(aq) + 2H_2O(l) + 2CO_2(g)$

In baking, self-raising flour is used to help cakes rise. Self-raising flour contains tartaric acid ($C_4H_6O_6$) and sodium hydrogen carbonate ($NaHCO_3$). These react together to form bubbles of carbon dioxide (CO_2) gas in the dough, which is what causes the cake to rise.

Figure 11E–4 Cake rising in the oven is an example of an acid–metal carbonate reaction.

Worked example 11E–3: Writing equations for acid–metal carbonate reactions

VIDEO
WORKED
EXAMPLE 11E–3

Write a fully balanced equation for the reaction that results when aqueous nitric acid reacts with solid barium carbonate.

Step 1	Write out the reactants.	$HNO_3 + BaCO_3$
Step 2	Deduce the products formed by following the rule: acid + metal carbonate → salt + water + carbon dioxide Deduce the salt formed: metal cation and acid anion. **Remember to check the charge on the ions and make sure you balance them when forming the salt.**	Water: H_2O Carbon dioxide: CO_2 Salt: Ba^{2+} NO_3^- = $Ba(NO_3)_2$
Step 3	Write out the reactants and products into a chemical equation format.	$HNO_3 + BaCO_3 \rightarrow H_2O + CO_2 + Ba(NO_3)_2$
Step 4	Assign states to all reactants and products. **Remember, for the salt, to check the solubility rules to determine if it will be in a solid or aqueous state.**	$HNO_3(aq) + BaCO_3(s) \rightarrow H_2O(l) + CO_2(g) + Ba(NO_3)_2(aq)$
Step 5	Balance the equation.	$2HNO_3(aq) + BaCO_3(s) \rightarrow H_2O(l) + CO_2(g) + Ba(NO_3)_2(aq)$

Worked example 11E–4: Writing equations for acid–metal carbonate reactions with spectator ions

VIDEO
WORKED
EXAMPLE 11E–4

Write a balanced ionic equation for the reaction that results when aqueous hydrochloric acid reacts with solid calcium carbonate.

Step 1	Write any dissolved ionic compound (aq) as their separate ions. Remember: (s), (l) and (g) substances won't break into ions.	$CaCO_3(s) + H^+(aq) + Cl^-(aq) \rightarrow Ca^{2+}(aq) + Cl^-(aq) + H_2O(l) + CO_2(g)$
Step 2	Cross out any ions that appear in the same state on both sides. These are spectator ions that have not participated in the reaction.	$CaCO_3(s) + H^+(aq) +$ ~~$Cl^-(aq)$~~ $\rightarrow Ca^{2+}(aq) +$ ~~$Cl^-(aq)$~~ $+ H_2O(l) + CO_2(g)$
Step 3	Rewrite the equation. Ensure the equation is still balanced and states are included on all species.	$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$

Acid + metal hydroxide reactions

acid + metal hydroxide → **salt + water**

The final reaction of acids we will learn about is an acid reacting with a metal hydroxide. As metal hydroxides usually behave as bases, this type of reaction can be an example of a neutralisation reaction, covered at the beginning of this section. However, in this reaction no bubbling is observed because a gas is not produced. Here the products are salt and water.

Worked example 11E–5: Writing equations for acid–metal hydroxide reactions

VIDEO
WORKED
EXAMPLE 11E–5

Write a fully balanced equation for the neutralisation reaction that results when aqueous sulfuric acid reacts with solid potassium hydroxide.

Step 1	Write out the reactants.	$H_2SO_4 + KOH$
Step 2	Deduce the products formed following the rule: acid + metal hydroxide → salt + water Deduce the salt formed: metal cation and acid anion. **Remember to check the charge on the ions and make sure you balance them when forming the salt.**	Water: H_2O Salt: $K^+\ SO_4^{2-} = K_2SO_4$
Step 3	Write out the reactants and products into a chemical equation format.	$H_2SO_4 + KOH \rightarrow H_2O + K_2SO_4$
Step 4	Assign states to all reactants and products. **Remember to check the solubility rules to determine if the salt will be in a solid or aqueous state.**	$H_2SO_4(aq) + KOH(s) \rightarrow H_2O(l) + K_2SO_4(aq)$
Step 5	Balance the equation.	$H_2SO_4(aq) + 2KOH(s) \rightarrow 2H_2O(l) + K_2SO_4(aq)$

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Worked example 11E–6: Writing equations for acid–metal hydroxide reactions with spectator ions

VIDEO
WORKED
EXAMPLE 11E–6

Write a balanced ionic equation for the neutralisation reaction that results when aqueous sulfuric acid reacts with solid potassium hydroxide.

Step 1	Write any dissolved ionic compound (aq) as their separate ions. Remember: (s), (l) and (g) substances won't break into ions.	$2H^+(aq) + SO_4^{2-}(aq) + 2KOH(s) \rightarrow 2H_2O(l) + 2K^+(aq) + SO_4^{2-}(aq)$
Step 2	Cross out any ions that appear in the same state on both sides. These are spectator ions, which have not participated in the reaction.	$2H^+(aq) + \cancel{SO_4^{2-}(aq)} + 2KOH(s) \rightarrow 2H_2O(l) + 2K^+(aq) + \cancel{SO_4^{2-}(aq)}$
Step 3	Rewrite the equation. Ensure that the equation is still balanced and that states are included on all species. **Note that here, that the final balanced equation must be simplified by changing all of the '2' coefficients to '1', which is then not written in the final balanced equation.**	$H^+(aq) + KOH(s) \rightarrow H_2O(l) + K^+(aq)$

WORKSHEET
11E–1
REACTIONS OF
ACIDS

Check-in questions – Set 1

1 What is a neutralisation reaction?

2 Complete and balance the following equations:
  a $Zn(s) + HNO_3(aq) \rightarrow$
  b $MgCO_3(s) + HCl(aq) \rightarrow$
  c $NaOH(aq) + H_2SO_4(aq) \rightarrow$
  d $K_2CO_3(s) + HBr(aq) \rightarrow$
  e $CaCO_3(s) + H_2SO_4(aq) \rightarrow$
  f $Pb(s) + H_3PO_4(aq) \rightarrow$

3 Write ionic equations for each of the reactions above.

Neutralisation reactions in the body

11B STRONG
AND WEAK
ACIDS AND
BASES
LINK

In the introduction to this chapter, we saw that the body contains both acids and bases to help with digestion. In the stomach, digestion is aided by stomach acid, hydrochloric acid (HCl), which is a strong acid. Hydrochloric acid helps to break down proteins, destroy bacteria and provides the optimal conditions required for the digestive enzyme pepsin to operate. However, if this acid does not stay in the stomach, it can be very painful and cause damage. Heartburn is a relatively common condition: it is a burning pain in the chest that occurs when stomach acid rises into the oesophagus and irritates the lining of the oesophagus. Luckily, our knowledge of acids and bases provides a relatively simple solution for heartburn – a neutralisation reaction.

Antacids, which you may have heard of or previously taken, assist with this neutralisation.

Antacids contain at least one of the following neutralising ingredients:

- magnesium hydroxide ($Mg(OH_2)$)
- aluminium hydroxide ($Al(OH)_3$)
- sodium bicarbonate ($NaHCO_3$)
- calcium carbonate ($CaCO_3$).

These ingredients act to neutralise stomach acid, ultimately taking away the pain associated with its presence in the oesophagus. We can use the acid reaction rules to write equations to represent how antacids work.

acid + metal carbonate → **salt + water + carbon dioxide**

$$\underset{\text{Stomach acid}}{2HCl(aq)} + \underset{\text{Antacid ingredient}}{CaCO_3(s)} \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

acid + metal hydroxide → **salt + water**

$$\underset{\text{Stomach acid}}{2HCl(aq)} + \underset{\text{Antacid ingredient}}{Mg(OH)_2(s)} \longrightarrow MgCl_2(aq) + 2H_2O(l)$$

Figure 11E–5 Antacid tablets contain basic ingredients to help neutralise stomach acid and relieve the symptoms of heartburn.

Check-in questions – Set 2

1 Explain how taking an antacid is an example of a neutralisation reaction.

2 Write a balanced full and ionic equation for the reaction between the active ingredient of an antacid, aluminium hydroxide ($Al(OH)_3$), with stomach acid (HCl).

11E SKILLS

Correct formulas for ionic compounds

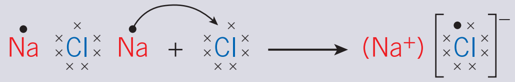
Although many of the acids you have seen in this chapter are covalent compounds and some bases are ionic compounds, one of the products of reactions involving acids is always an ionic salt (ionic compound).

It is important to double check that the formula of the ionic compound you have written is balanced – meaning that it has no overall charge and the number of electrons lost from the cations equals the number of electrons gained by the anions.

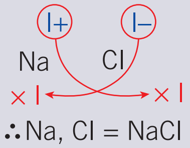
4A FORMATION AND NAMING OF IONIC COMPOUNDS **LINK**

There are several ways you can check this.

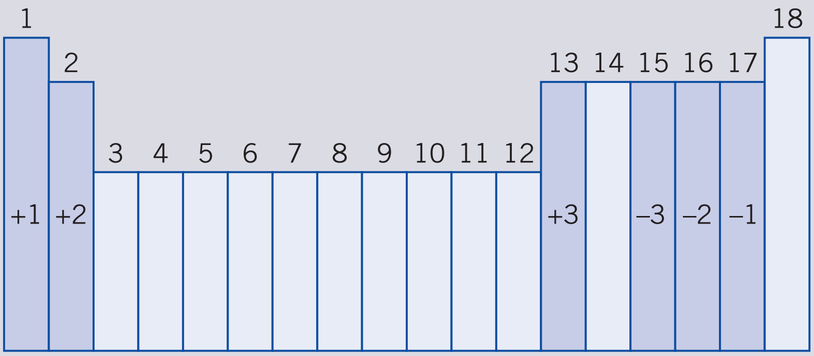
1 Draw an electron transfer diagram as learned in Chapter 4A.



2 Use a valency table and the crossover method.



However, you won't always have access to a valency table, so a key skill to develop is the ability to use the periodic table to determine the charge on an ion. Remember, the group that an element is positioned in tells you what type of ion it will form.



Ions and the periodic table

Section 11E questions

1 Complete the table below, summarising the different types of reactions with acids studied in this chapter.

Acid	+		→		+	Hydrogen gas		
Acid	+	Metal carbonate	→		+		+	
Acid	+	Metal hydroxide	→		+			

2 Name and give the formula of the salt formed when the following acids come into contact with copper(II) metal.

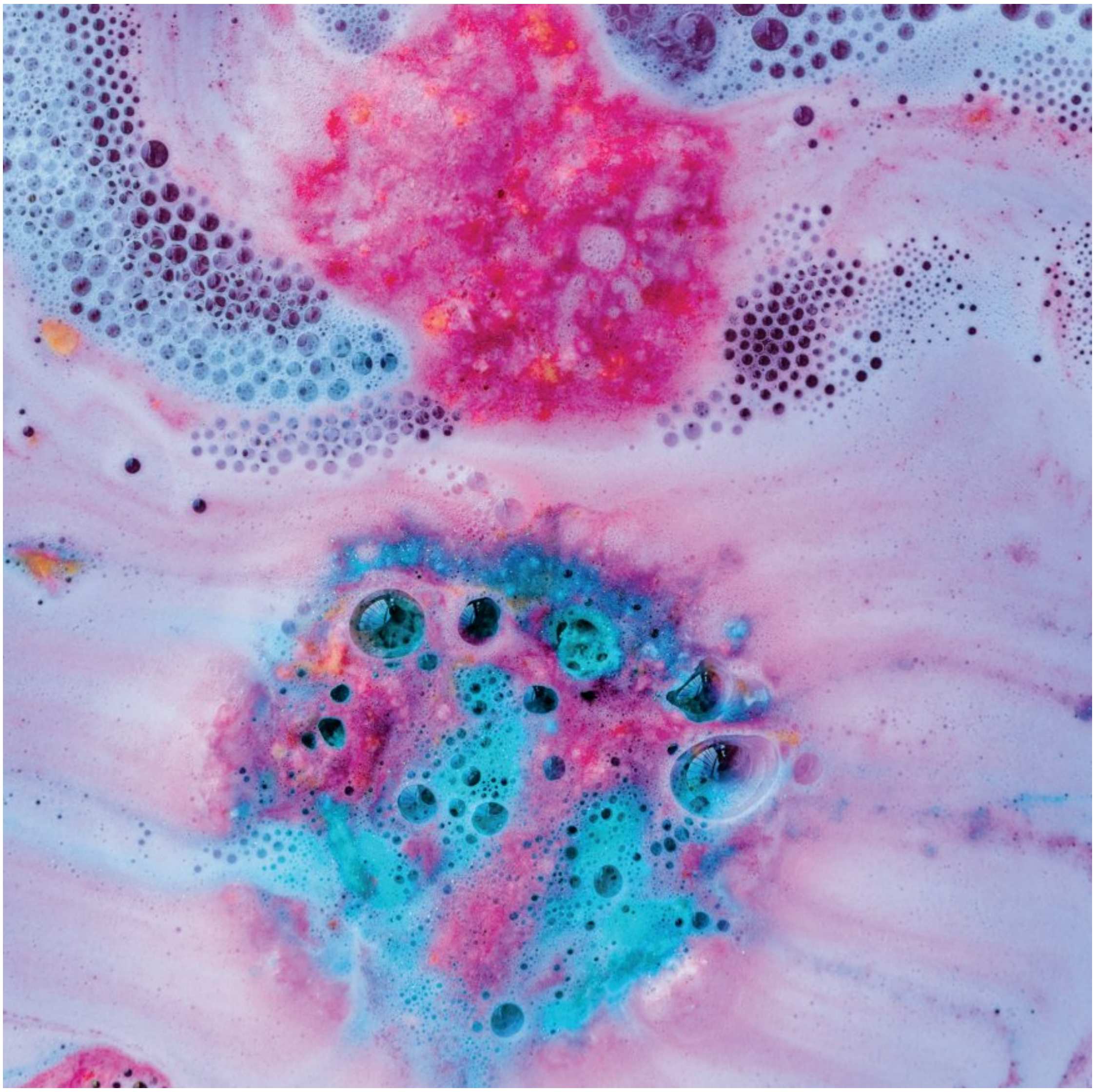
- a nitric acid
- b phosphoric acid
- c ethanoic acid

3 Sherbet is a confectionery that contains citric acid and sodium hydrogen carbonate. Explain what causes the 'fizzing' sensation on your tongue when you consume sherbet.

4 In some countries, air pollution has resulted in acid rain. Acid rain is precipitation that contains high levels of sulfuric and nitric acid. When acid rain falls it has many adverse effects, one of which is damage to marble statues.
   - a Write a full balanced equation to represent the reaction between acid rain (H_2SO_4) and marble ($CaCO_3$).
   - b Identify the spectator ions in this reaction.
   - c Explain the term 'spectator ion'.

5 Formic acid, also known as methanoic acid (HCOOH), is a major component of ant venom. To relieve itching, calamine lotion, which contains zinc carbonate, is often used.
   - a Write a full balanced equation for this neutralisation reaction.
   - b Write an ionic equation for this neutralisation reaction.

6 Write balanced full and ionic equations for the following reactions:
   - a magnesium and sulfuric acid
   - b copper(II) carbonate and nitric acid
   - c magnesium hydroxide + phosphoric acid



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Chapter 11 review

Summary

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

Success criteria – I am now able to:		Linked question
11A.1	Define ‘acid’ and ‘base’	1☐
11A.2	Write an equation to represent an acid–base reaction with water	5☐
11A.3	Identify conjugate acid–base pairs	9☐, 12☐
11A.4	Define ‘polyprotic’ and ‘amphiprotic’	2☐, 3☐
11A.5	Write full and ionic equations to represent an amphiprotic substance	2☐
11A.6	Write full and ionic equations to represent a polyprotic substance	2☐, 3☐
11B.1	Define ‘strong’, ‘weak’, ‘concentrated’ and ‘dilute’ in terms of acids and bases	4☐, 11☐
11B.2	Recall common examples of strong and weak acids and bases	4☐, 5☐
11B.3	Distinguish between strong and weak acids and bases	4☐, 11☐
11B.4	Distinguish between concentrated and dilute acids and bases	7☐
11B.5	Write full equations to represent a strong and weak acid and base	7☐
11C.1	Use and understand the ionic product of water and the pH scale	8☐
11C.2	Recall the distinction between hydronium ion and hydroxide ion concentrations for acids and bases	16☐
11C.3	Calculate the pH of acids and bases from a given concentration	6☐, 10☐, 13☐, 19☐
11C.4	Calculate the concentration of acids and bases given their pH	14☐, 16☐
11C.5	Use the pH scale in the measurement and calculations of strengths of acids and bases in dilutions of solutions	15☐
11C.6	Recall the difference between accuracy and precision	20☐
11C.7	Understand what an indicator is and how it works	20☐
11C.8	Distinguish between examples of natural and commercial indicators, as well as pH meters	20☐
11C.9	Given information, use an indicator to determine the pH of a solution	20☐
11C.10	Recognise colour changes of common indicators	18☐
11D.1	Distinguish between acid rain and the natural acidity of rain	
11D.2	Write a full equation to show the formation of carbonic acid that results when carbon dioxide is dissolved in water	16☐

Success criteria – I am now able to:		Linked question
11D.3	Recognise and show how carbonic acid acts as a weak acid	16☐
11D.4	Demonstrate an understanding of the relationship between increased production of atmospheric carbon dioxide and decreasing pH of oceans	16☐
11D.5	Understand how decreasing pH of oceans directly affects carbonate levels and the consequences for this on shell growth in marine invertebrates	17☐
11E.1	Recall what a neutralisation reaction is and determine products from a reaction	21☐
11E.2	Identify and write full and ionic equations for the reactions of acids with metals, identifying reactants and products	19☐
11E.3	Identify and write full and ionic equations for the reactions of acids with carbonates, identifying reactants and products	17☐
11E.4	Identify and write full and ionic equations for the reaction of acids with hydroxides, identifying reactants and products	21☐
11E.5	Recall types of antacids and explain their beneficial action in neutralising stomach acid	21☐

Multiple-choice questions

1 Which of the following is **not** a Brønsted–Lowry acid–base reaction?

A $2NH_3(aq) + H_2SO_4(aq) \rightleftharpoons (NH_4)_2SO_4(aq)$

B $O^{2-}(aq) + H_2O(l) \rightleftharpoons 2OH^-(aq)$

C $HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_4^{2-}(aq)$

D $Zn(s) + H_2SO_4(aq) \rightleftharpoons ZnSO_4(aq) + H_2(g)$

2 Which of the following substances is both amphiprotic and polyprotic?

A H_2SO_4

B HCO_3^-

C $H_2PO_4^-$

D H_3PO_4

3 Which of the following ions is an example of a diprotic base?

A HSO_4^-

B CO_3^{2-}

C $H_2PO_4^-$

D PO_4^{3-}

4 Which of the following best describes a weak acid?

A an acid that contains only one hydrogen atom in its molecular formula

B an acid of low concentration

C a dilute acid

D an acid that dissociates only partially in aqueous solution

5 The hydrogen carbonate ion, HCO_3^-, may act as an acid or a base in aqueous solution. In which one of the equations below is it acting as an acid?

A $HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$

B $HCO_3^-(aq) + H_3O^+(aq) \rightleftharpoons H_2CO_3(aq) + H_2O(l)$

C $HCO_3^-(aq) + H_2O(l) \rightleftharpoons CO_3^{2-}(aq) + H_3O^+(aq)$

D $HCO_3^-(aq) + OH^-(aq) \rightleftharpoons H_2CO_3(aq) + O^{2-}(aq)$

6 Which of the following aqueous solutions has the lowest pH at 25°C?

A 0.1 mol L^{-1} NaOH
B 0.1 mol L^{-1} HNO_3
C 0.1 mol L^{-1} H_2SO_4
D 0.1 mol L^{-1} HF

7 In a chemistry laboratory, two reagent bottles are labelled 1 M sulfuric acid and 10.0 M ethanoic acid. Based on this information and your knowledge of acids, which of the following statement/s uses correct terminology?

A The ethanoic acid is more concentrated than the sulfuric acid.
B The ethanoic acid is stronger than the sulfuric acid.
C The sulfuric acid solution is weaker than the ethanoic acid solution.
D The ethanoic acid solution is more dilute than the sulfuric acid solution.

8 The $[H^+]$ concentration in pure water at 25°C is equal to

A 1 mol L^{-1}.
B 10^{-14} mol L^{-1}.
C 7 mol L^{-1}.
D 10^{-7} mol L^{-1}.

9 For the reaction, $HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$, the conjugate base of HF is

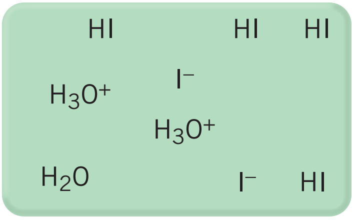
A H_2O.
B H_3O^+.
C F^-.
D OH^-.

10 What is the pH of a 0.001 mol L^{-1} solution of nitric acid (HNO_3)?

A 0.001
B 2
C 3
D 10

Short-answer questions

11 The solution shown in the image below represents hydroiodic acid (HI) reacting with water.



a Select which of the following best describes the hydroiodic acid as represented by the image: weak acid, strong acid, weak base, strong base. (1 mark)
b Referring to the image, explain your answer to part **a**. (2 marks)
c Write a fully balanced equation to represent this solution. (2 marks)

12 Copy and fill in the table below to complete the conjugate acid–base pairs. (5 marks)

HF	
	NO_3^-
	HPO_4^{2-}
CH_3COOH	
HS^-	

13 For a solution of 0.25 M HCl, calculate the following, showing all working.

a concentration of H_3O^+ ions (1 mark)
b concentration of OH^- ions (1 mark)
c pH of the solution (1 mark)
d pOH of the solution (1 mark)

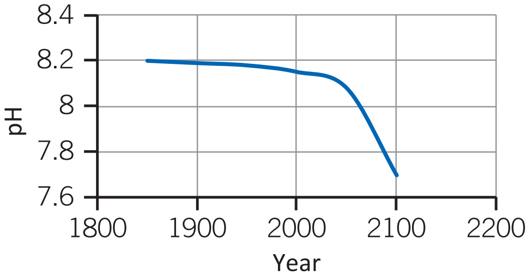
14 **a** Determine the final concentration of a solution of HCl when 50 mL of 0.025 M HCl is added to 150 mL of water. (2 marks)

b What was the initial pH and what is the final pH? (2 marks)

15 Calculate the concentration of the following solutions.

a HCl with pH of 5.0 (2 marks)
b NaOH with pH of 10 (3 marks)

16 The graph below shows the predicted changes to the pH of ocean waters.



a State whether ocean water is currently acidic or alkaline. (1 mark)

b Use the graph to determine the predicted change in the hydronium ion concentration $[H_3O^+]$ of sea water between the years 1850 and 2100. (3 marks)

c Increased levels of carbon dioxide dissolving in oceans is responsible for the predicted change in pH of ocean water.

Write two fully balanced equations to show how carbon dioxide reduces the pH of sea water. (4 marks)

17 Many organisms found in the ocean have shells and skeletons made of calcium carbonate. These organisms obtain carbonate ions from the sea water to make and maintain their shells and skeletons.

a Outline how the concentration of carbonate ions in sea water will change if pH continues to decrease as predicted. (2 marks)

b Given that these organisms form an important part of the marine food chain and are classed as an edible seafood, state and explain the consequences of ocean acidification. In your answer, refer to more than just the environmental/biological implications. (5 marks)

18 A 42.5 g sample of barium hydroxide ($Ba(OH)_2$) is dissolved in 850 mL of deionised water.

- **a** Calculate the pH of the solution. (6 marks)
- **b** The solution was tested with methyl red. What colour would be observed? (1 mark)
- **c** Would the pH increase or decrease if 900 mL of water was accidentally used? Explain your decision. (2 marks)

19 Acid rain is formed when sulfur dioxide reacts with water in the atmosphere, resulting in the formation of sulfuric acid that falls in rain.

- **a** When acid rain (H_2SO_4) falls on marble statutes ($CaCO_3$), it reacts.
  - **i** Write a full balanced equation to represent this reaction. (2 marks)
  - **ii** Write an ionic equation to represent this reaction. (2 marks)
- **b** When acid rain (H_2SO_4) falls on iron metal (Fe) structures, it begins to react and corrode the iron.
  - **i** Write a full balanced equation to represent this reaction. (2 marks)
  - **ii** Write an ionic equation to represent this reaction. (2 marks)

20 A student tested an acidic sample with a range of indicators and observed the following results.

Indicator	Colour observed in three trials	pH predicted
Turmeric solution	Yellow	
Methyl orange	Red	
Phenolphthalein	Clear	
Litmus solution	Red	

- **a** What is an acid–base indicator? (1 mark)
- **b** List which indicators above are natural indicators. (1 mark)
- **c** Use the colour observed for each indicator to predict the pH of the sample. (1 mark)
- **d** A calibrated pH probe is then used to test the sample. After three readings, the pH is determined to be 2.1. Explain if this reading is an example of quantitative or qualitative data. (2 marks)
- **e** Is it more accurate or more precise than the predicted pH readings given by the indicators? Explain your selection by referring to both precision and accuracy in your answer. (3 marks)

21 Explain, with the aid of an equation, how taking an antacid containing the active ingredient magnesium hydroxide, $Mg(OH)_2$, to relieve the symptoms of acid reflux is an example of both a neutralisation reaction and a Brønsted–Lowry acid–base reaction. (4 marks)

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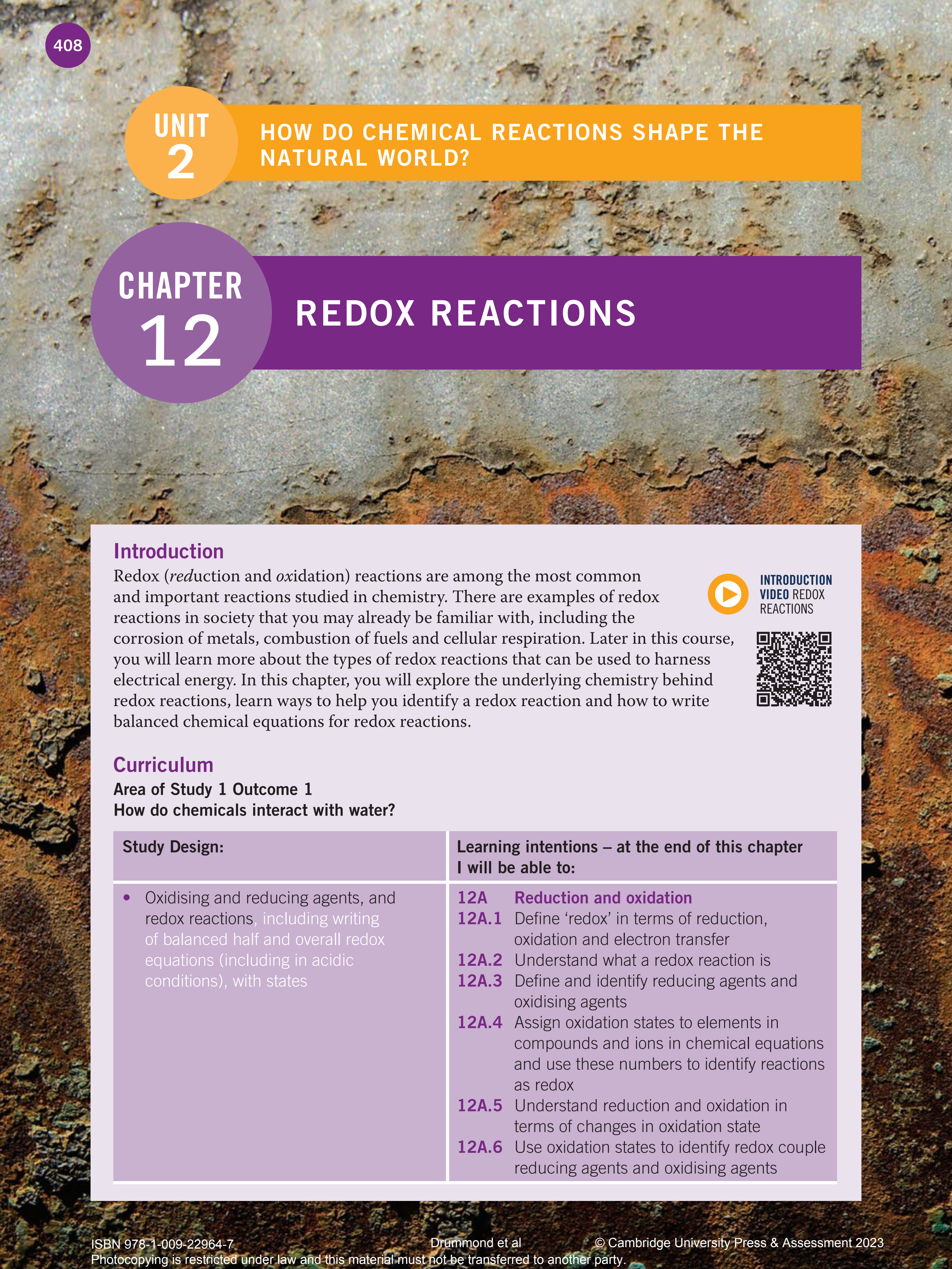

UNIT 2 HOW DO CHEMICAL REACTIONS SHAPE THE NATURAL WORLD?

CHAPTER 12 REDOX REACTIONS

Introduction

Redox (*red*uction and *ox*idation) reactions are among the most common and important reactions studied in chemistry. There are examples of redox reactions in society that you may already be familiar with, including the corrosion of metals, combustion of fuels and cellular respiration. Later in this course, you will learn more about the types of redox reactions that can be used to harness electrical energy. In this chapter, you will explore the underlying chemistry behind redox reactions, learn ways to help you identify a redox reaction and how to write balanced chemical equations for redox reactions.

INTRODUCTION VIDEO REDOX REACTIONS

Curriculum

Area of Study 1 Outcome 1
How do chemicals interact with water?

Study Design:	Learning intentions – at the end of this chapter I will be able to:
• Oxidising and reducing agents, and redox reactions, including writing of balanced half and overall redox equations (including in acidic conditions), with states	**12A Reduction and oxidation** **12A.1** Define 'redox' in terms of reduction, oxidation and electron transfer **12A.2** Understand what a redox reaction is **12A.3** Define and identify reducing agents and oxidising agents **12A.4** Assign oxidation states to elements in compounds and ions in chemical equations and use these numbers to identify reactions as redox **12A.5** Understand reduction and oxidation in terms of changes in oxidation state **12A.6** Use oxidation states to identify redox couple reducing agents and oxidising agents

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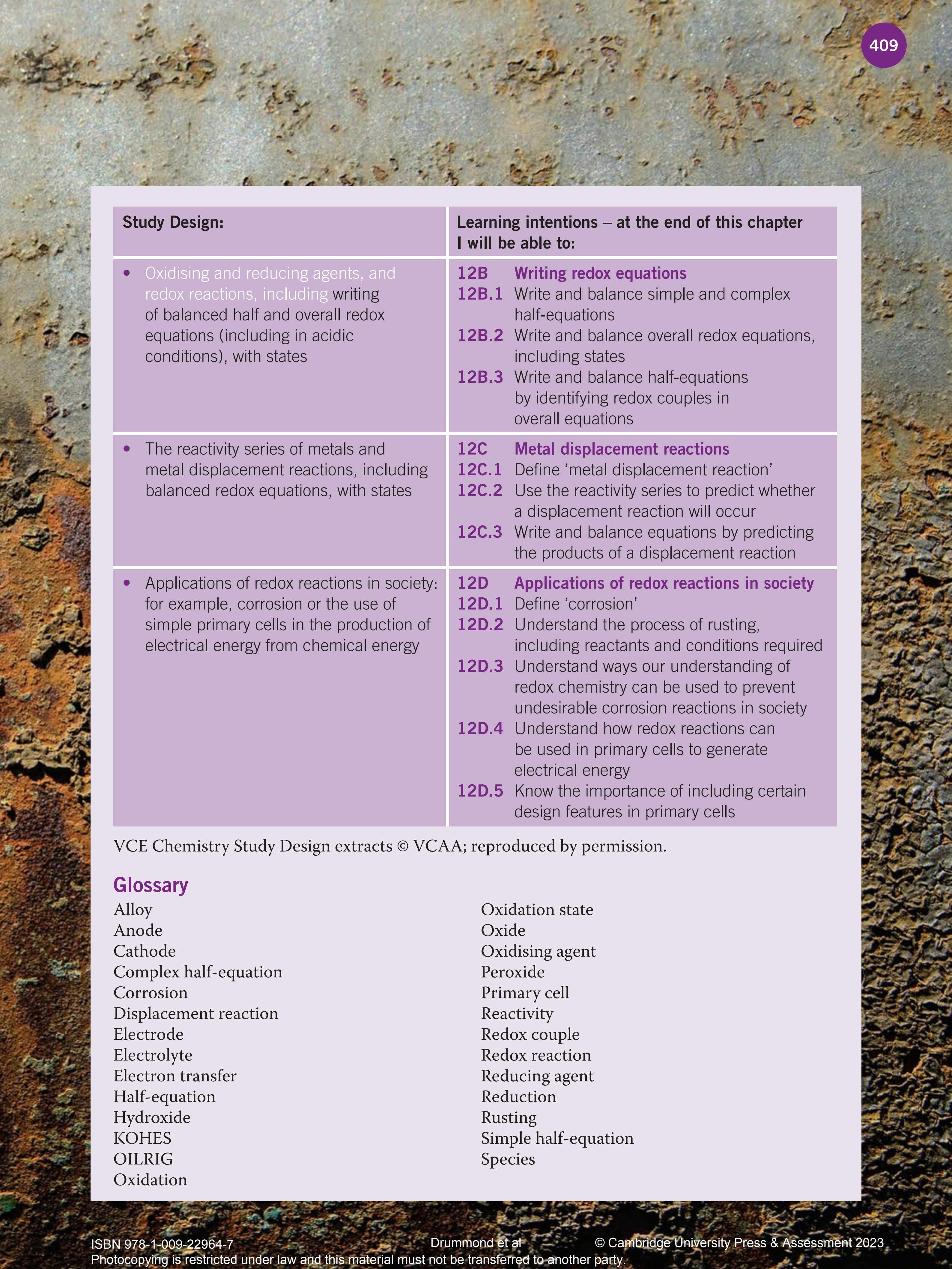

Study Design:	Learning intentions – at the end of this chapter I will be able to:
• Oxidising and reducing agents, and redox reactions, including writing of balanced half and overall redox equations (including in acidic conditions), with states	**12B Writing redox equations** **12B.1** Write and balance simple and complex half-equations **12B.2** Write and balance overall redox equations, including states **12B.3** Write and balance half-equations by identifying redox couples in overall equations
• The reactivity series of metals and metal displacement reactions, including balanced redox equations, with states	**12C Metal displacement reactions** **12C.1** Define 'metal displacement reaction' **12C.2** Use the reactivity series to predict whether a displacement reaction will occur **12C.3** Write and balance equations by predicting the products of a displacement reaction
• 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	**12D Applications of redox reactions in society** **12D.1** Define 'corrosion' **12D.2** Understand the process of rusting, including reactants and conditions required **12D.3** Understand ways our understanding of redox chemistry can be used to prevent undesirable corrosion reactions in society **12D.4** Understand how redox reactions can be used in primary cells to generate electrical energy **12D.5** Know the importance of including certain design features in primary cells

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Glossary

Alloy
Anode
Cathode
Complex half-equation
Corrosion
Displacement reaction
Electrode
Electrolyte
Electron transfer
Half-equation
Hydroxide
KOHES
OILRIG
Oxidation
Oxidation state
Oxide
Oxidising agent
Peroxide
Primary cell
Reactivity
Redox couple
Redox reaction
Reducing agent
Reduction
Rusting
Simple half-equation
Species

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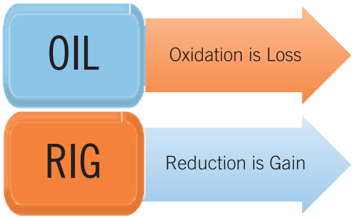
Concept map

12A Reduction and oxidation

Redox reactions



Reducing agent

Oxidising agent

Oxidation states

12B Writing redox equations

$$\frac{\begin{array}{r} 2Ag^{+}\,(ag) + 2e^{-} \rightarrow 2Ag(s) \\ Cu(s) \rightarrow Cu^{2+}(aq) + 2e^{-} \end{array}}{2Ag^{+}(aq) + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}(aq)}$$

Simple half-equations

Complex half-equations

Combining half-equations into full redox equations

12C Metal displacement reactions

Reactivity series

12D Applications of redox reactions in society

See the interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

12A Reduction and oxidation

Study design:
Oxidising and reducing agents, and redox reactions, including writing of balanced half and overall redox equations (including in acidic conditions), with states indicated

Glossary:

Electron transfer	Peroxide
Half-equation	Redox couple
OILRIG	Redox reaction
Oxidation	Reducing agent
Oxidation state	Reduction
Oxidising agent	Species

ENGAGE

What is oxidation?

You have most likely heard of reduction and oxidation during your study of chemistry. You may be wondering, where did these terms come from? When and how were redox reactions first discovered?

The terms 'reduction' and 'oxidation' stem from the historical exploration of combustion, which was the first type of oxidation to be studied. In 1775, English scientist, Joseph Priestley, who had discovered oxygen about a year earlier, met with French chemist, Antoine Lavoisier, who went on to revolutionise the field of chemistry by investigating the role of oxygen in combustion reactions. His work helped to later establish that combustion involved materials combining with oxygen, and these reactions were classified as oxidation reactions. Those reactions in which oxygen was produced were classified as reduction reactions.

Often, when oxygen is involved in a reaction, it will be acting as an oxidising agent. However, this is not always the case; but when redox reactions were first discovered by Priestley and Lavoisier, they believed that oxygen was always involved, which is another reason why the term 'oxidation' was adopted.

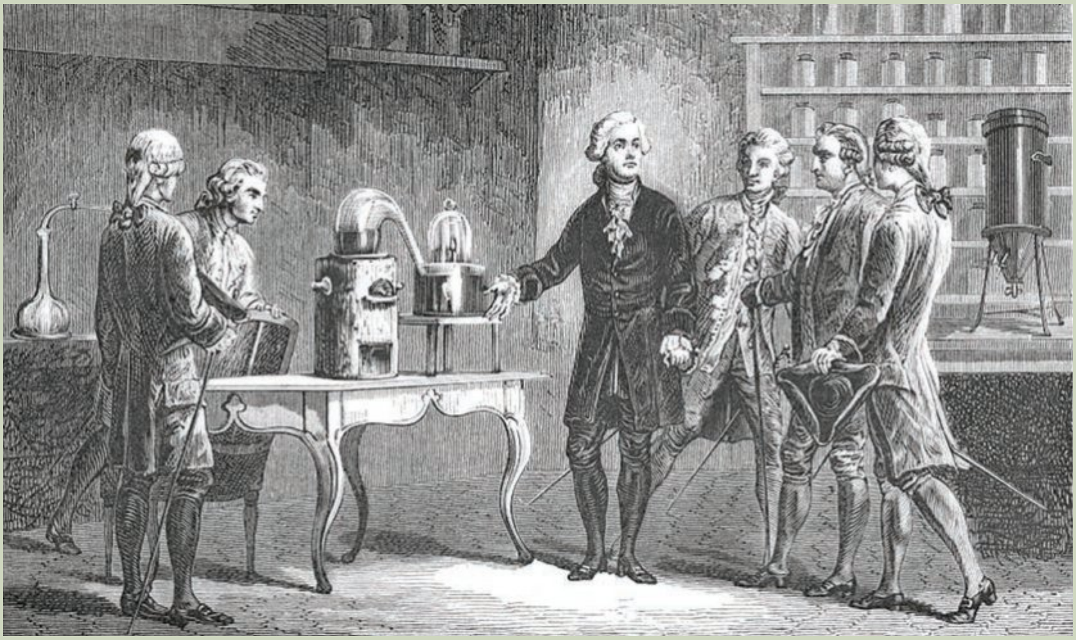


Figure 12A–1 French chemist Antoine Lavoisier revolutionised the field of chemistry with his discovery of the role oxygen plays in combustion reactions.

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EXPLAIN

What is redox?

The word **redox** is a combined abbreviation for ***red***uction and ***ox***idation, which are two different processes that occur simultaneously during a redox reaction. A **redox reaction** involves an **electron transfer** from one chemical **species** to another.

Therefore, **reduction** can be defined as the *gain* of electrons and **oxidation** as the *loss* of electrons. A nice way to remember this is using the mnemonic **OILRIG** – **O**xidation **I**s **L**oss (of electrons), **R**eduction **I**s **G**ain (of electrons).

Redox reaction a reaction involving a transfer of electrons. Both reduction and oxidation occur simultaneously

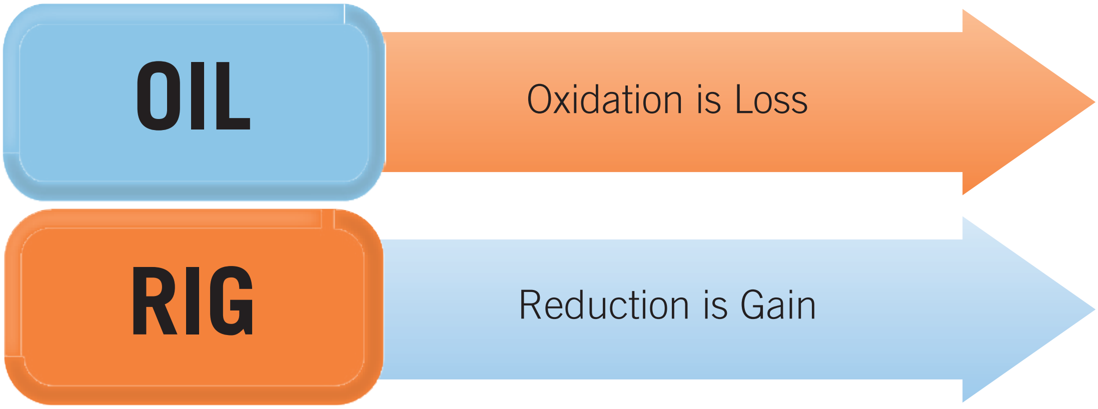
Electron transfer the exchange of electrons that occurs during a redox reaction. Electrons are transferred from the reducing agent to the oxidising agent

Species any reactant or product in a chemical reaction

Reduction a gain of electrons

Oxidation a loss of electrons

OILRIG a mnemonic to recall redox – **O**xidation **I**s **L**oss (of electrons), **R**eduction **I**s **G**ain (of electrons)



NOTE

There is also a 'rule of thumb', most often used by organic chemists, which identifies a species in a reaction as being oxidised due to a gain of oxygen or loss of hydrogen, or being reduced due to a gain of hydrogen or loss of oxygen. Note, however, that redox reactions do not always involve hydrogen or oxygen.

Figure 12A–2 shows the reaction between solid magnesium ribbon and oxygen gas, which, if performed at high temperatures, will produce solid magnesium oxide. The redox equation for this reaction can be written as

$$2Mg(s) + O_2(g) \rightarrow 2Mg^{2+}(s) + 2O^{2-}(s)$$

Figure 12A–2 In this redox reaction, magnesium ribbon burns in the presence of oxygen to produce a bright white flame.

Reducing agents and oxidising agents

You should recall that ionic compounds are formed where there is a large difference in the electronegativities of the bonding atoms. This commonly occurs between a non-metal and a group 1 metal, a group 2 metal or a transition metal with a low oxidation state. In a similar way, the redox reaction between magnesium and oxygen involves a transfer of electrons from the magnesium atoms to the oxygen atoms. Given that oxidation is the loss of electrons, in this reaction magnesium has undergone oxidation; it has been oxidised. Similarly, the oxygen has undergone reduction; it has been reduced, since it is the molecule that gained the electrons.

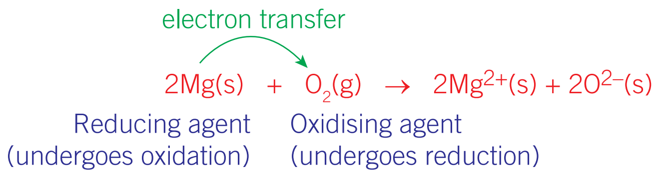
LINK **4A** FORMATION AND NAMING OF IONIC COMPOUNDS

It is also useful to be able to identify the reducing agent and the oxidising agent. The species that causes oxidation of the other reactant is termed the **oxidising agent**. The species that causes reduction to occur to the other species is termed the **reducing agent**.

Oxidising agent the reactant in a redox reaction that causes oxidation. It is itself reduced and therefore will gain electrons – its oxidation state will decrease

Reducing agent the reactant in a redox reaction that causes reduction. It is itself oxidised and therefore will lose electrons – its oxidation state will increase

In the example above, magnesium causes the reduction of oxygen gas and at the same time undergoes oxidation itself, losing electrons. Thus, magnesium can be identified as the reducing agent in this reaction. Similarly, oxygen causes the oxidation of magnesium and at the same time undergoes reduction itself, gaining electrons. Thus, oxygen can be identified as the oxidising agent in this reaction.



NOTE

It is helpful to remember that the species that undergoes reduction will be the oxidising agent, and the species that undergoes oxidation will be the reducing agent.

Check-in questions – Set 1

1. Explain what occurs during a redox reaction.
2. Write a definition for reduction.
3. Write a definition for oxidation.
4. Describe the role of the reducing agent in a redox reaction.

Oxidation states

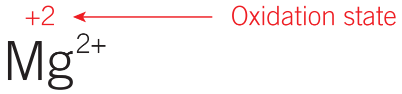
It is not always obvious to identify whether electrons are being transferred or not in a chemical reaction, nor is it obvious which species is gaining or losing them. To help identify whether a species is gaining or losing electrons and consequently whether it is undergoing reduction or oxidation, we can use **oxidation states**.

VIDEO 12A–1 USING OXIDATION STATES

Oxidation state a measure of the degree of oxidation of an atom in a substance. It is defined as the charge an atom might have when electrons are counted to help identify redox reactions, oxidising agents and reducing agents

Writing oxidation states

While not official IUPAC convention, oxidation states are easily represented when written directly above the element as –1, 0, +1, +2 and so on. It is important to distinguish between this format and that of an ion. For example, the magnesium ion, Mg^{2+}, has an ionic charge of 2+, but its oxidation state is written as +2, usually like this:



Assigning oxidation states

Peroxide a class of chemical compounds in which two oxygen atoms are linked together by a single covalent bond

In most cases, you will be able to work out the oxidation state of an unknown element by applying several rules to identify the oxidation states in other elements present in a species. Although many elements will often have the same oxidation state, some elements can have many different oxidation states. One example of this is nitrogen, which can have oxidation states ranging from –3 to +5!

The rules to follow when assigning oxidation states, as well as exceptions to some rules that are particularly important to remember, are summarised in Table 12A–1.

Table 12A–1 Oxidation states rules and examples

Category	Rule	Example(s)
Free elements Oxidation state = 0 (zero)	A free element is any element that exists on its own and that is not a charged ion. It may be monoatomic or polyatomic but must contain only one type of element.	Solid iron metal, Fe(s) Hydrogen gas, $H_2(g)$ Oxygen gas, $O_2(g)$ Sulfur, $S_8(s)$ Carbon, C(s)
Hydrogen Typical oxidation state = +1 Alternative oxidation state = –1	The alternative oxidation state only occurs when hydrogen is part of a metal hydride. This is because metal hydrides are ionic compounds, and as such, the metal ion will have an oxidation state equal to its charge. To achieve an overall neutral charge for metal hydrides, hydrogen must have an oxidation state of –1.	Methane, CH_4 In this neutral molecule, hydrogen will have its typical oxidation state of +1. Exception: Sodium hydride, NaH. In this case, the sodium ion will have an oxidation state of +1, and the hydrogen will have an oxidation state of –1.
Oxygen Typical oxidation state = –2 Typical alternative oxidation state = –1	Oxygen will typically have an oxidation state of –2 when it is not a free element. The alternative oxidation state occurs only when oxygen is part of a **peroxide**, such as hydrogen peroxide, H_2O_2.	Carbon dioxide, CO_2 In this neutral molecule, oxygen will have an oxidation state of –2. Exception: Hydrogen peroxide, H_2O_2. In this molecule, hydrogen will have an oxidation state of +1, and oxygen will have an oxidation state of –1.
Group 1 metals Oxidation state = +1	All group 1 metals that are not free elements	Lithium ions, Li^+ Sodium ions, Na^+ *Solid sodium, Na(s), would still have an oxidation state of 0 as it is a free element.*
Group 2 metals Oxidation state = +2	All group 2 metals that are not free elements	Magnesium ions, Mg^{2+} Calcium ions, Ca^{2+}
Pure ions Oxidation state = ionic charge	Pure ions have an oxidation state equal to their ionic charge.	Aluminium ions, Al^{3+} These ions will have an ionic charge of 3+ and an oxidation state of +3.

Table 12A–1 *Continued*

Category	Rule	Example(s)
Neutral compounds The sum of the oxidation states will be 0 (zero).	Neutral compounds, as well as having an overall neutral charge, will also be neutral in their overall oxidation state total.	Water, H_2O In the neutral molecule of water, hydrogen has an oxidation state of +1 and oxygen has an oxidation state of –2. There are two hydrogen atoms, so overall: $(2 \times +1) + (-2) = 0$.
Polyatomic ions The sum of the oxidation states = charge on the ion	Polyatomic ions have an overall oxidation state equal to the charge on the polyatomic ion.	Sulfate ions, SO_4^{2-} These polyatomic ions have an overall ionic charge of 2–. Oxygen has an oxidation state of –2, and there are four atoms of oxygen, so $4 \times -2 = -8$. Sulfur must have an oxidation state of +6, to match the overall charge on the ion. $-8 + (+6) = -2$.

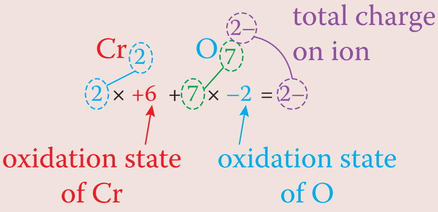
NOTE

The most electronegative element in a compound must be assigned a negative oxidation state. Therefore, fluorine, being the most electronegative element, will always have an oxidation state of –1. There are no exceptions to this scenario. More generally, group 17 halogens *usually* have an oxidation state of –1, but not always. Chlorine, in some cases, can have a positive oxidation state. An example of this is chlorine dioxide, ClO_2, where chlorine has an oxidation state of +4 because oxygen is more electronegative than chlorine and hence is assigned the oxidation state of –2.

Worked example 12A–1: Assigning oxidation states

If you are asked to assign oxidation states to each of the elements in the dichromate ion, $Cr_2O_7^{2-}$, you would apply the following rules.

Solution

Step 1	Oxygen has an oxidation state of –2 (unless in a peroxide, which the dichromate ion is not). Then determine the total number of oxygen atoms.	O = –2 Seven oxygen atoms $7 \times -2 = -14$
Step 2	Need to recognise that it is a polyatomic ion, so the sum of the oxidation states must add to the corresponding charge on the ion.	Charge on dichromate ion = 2–, so sum of oxidation states must equal –2.
Step 3	Determine the oxidation state of the unknown atom, taking into consideration the number of these.  	There are two Cr atoms. Use algebra and assign Cr as ‘x’. $2x + (-14) = -2$ $2x = +12$ $x = +6$.
Step 4	Write down the oxidation states of all elements.	Cr = +6 O = –2 +6 –2 $Cr_2O_7^{2-}$

VIDEO
WORKED EXAMPLE 12A–1

Check-in questions – Set 2

1 Assign oxidation states to each element in the species below.

a	Mg	**b**	HCl
c	CO_2	**d**	SO_4^{2-}
e	$KBrO_3$		

2 Determine the oxidation state for the element nitrogen, **N**, in each of the following compounds.

a	NO	**b**	N_2O
c	NO_2	**d**	N_2O_3
e	NH_3	**f**	N_2H_4
g	NH_4^+	**h**	NO_3^-

Using oxidation states

Identifying reducing agents and oxidising agents

A change in oxidation state can be used to identify reducing agents and oxidising agents. If the oxidation state *increases* for an element during a reaction, it indicates that it has undergone oxidation, losing electrons (hence becoming 'more positive'), and is therefore the **reducing agent**. Similarly, if the oxidation state *decreases* during a reaction, it indicates that it has undergone reduction, gaining electrons (hence becoming 'less positive'), and thus can be identified as the **oxidising agent**.

$$\overset{+2}{Fe^{2+}}(aq) + 2e^- \rightarrow \overset{0}{Fe}(s)$$

In the equation shown above, Fe^{2+}(aq) ions are reduced because the oxidation state of iron decreases from +2 to 0. Fe(s) has gained electrons.

Half-equation either the reduction or oxidation part of a redox reaction

In **half-equations** such as these, an alternative way to determine whether reduction or oxidation is taking place is to identify on which side of the equation the electrons are present. If electrons are on the left-hand side, as shown in Figure 12A–3, then it is a reduction half-equation. If electrons are on the right-hand side, then it is an oxidation half-equation.

$$Cu^{2+}(aq) + \underbrace{\boxed{2e^-}}_{\text{electrons gained}} \rightarrow Cu(s) \quad \text{Reduction} \qquad Zn(s) \rightarrow Zn^{2+}(aq) + \underbrace{\boxed{2e^-}}_{\text{electrons lost}} \quad \text{Oxidation}$$

Figure 12A–3 Reduction and oxidation half-equations showing location of electrons

WORKSHEET 12A–1 OXIDATION STATES

This is useful when dealing with half-equations where electrons are shown. However, for more complex redox reactions, oxidation states are necessary, as shown in Worked example 12A–2.

Worked example 12A–2: Identifying reducing and oxidising agents

Determine the reducing agent and oxidising agent in the reaction between copper ions, Cu^{2+}, and zinc metal, Zn, that produces copper metal, Cu, and zinc ions, Zn^{2+}.

Solution

Step 1	Write out the full balanced equation, including states.	$Cu^{2+}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq)$
Step 2	Assign oxidation states to each species present in the equation.	$\overset{+2}{Cu^{2+}}(aq) + \overset{0}{Zn}(s) \rightarrow \overset{0}{Cu}(s) + \overset{+2}{Zn^{2+}}(aq)$
Step 3	Answer the question by using the change in oxidation states of the atoms/ions.	$Cu^{2+}(aq)$ ions have been reduced since the oxidation state decreases from +2 to 0. Zn(s) has been oxidised to $Zn^{2+}(aq)$ ions because its oxidation state has increased from 0 to +2. Zn(s) can be identified as the reducing agent, and $Cu^{2+}(aq)$ ions are therefore the oxidising agent.

VIDEO
WORKED EXAMPLE 12A–2

Important considerations when using oxidation states

In many redox reactions, the same elements are often involved as reactants and as products. It is particularly important to note the difference in states and charges when you are describing changes in a redox reaction or when you are asked to identify reducing agents or oxidising agents. For example, copper ions, $Cu^{2+}(aq)$, are a different chemical species when compared to copper solid, Cu(s).

In addition, where the element whose oxidation state changes during a redox reaction belongs to a polyatomic species, you should include the *entire reactant compound* that contains the element when naming the reducing agent or oxidising agent. Consider Worked example 12A–3, where oxidation states are shown.

Worked example 12A–3: Identifying reducing and oxidising agents

Determine the reducing agent and oxidising agent in the reaction between iron(III) oxide and carbon monoxide.

Solution

Step 1	Write out the full balanced equation, including states.	$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$
Step 2	Assign oxidation states to each species present in the equation.	reduction oxidation $\overset{+3\ -2}{Fe_2O_3}(s) + \overset{+2\ -2}{3CO}(g) \rightarrow \overset{0}{2Fe}(s) + \overset{+4\ -2}{3CO_2}(g)$
Step 3	Answer the question by using the change in oxidation states of the atoms/ions.	Iron, Fe, shows an oxidation state change from +3 to 0, meaning that it has been reduced (gained electrons). However, if you were stating the oxidising agent, you would state iron oxide, $Fe_2O_3(s)$, as this is the chemical species causing the oxidation of the carbon monoxide. Similarly, the reducing agent for this reaction would be identified as carbon monoxide, CO(g), even though only the carbon undergoes an oxidation state increase (due to oxidation, loss of electrons).

LINK **4A** FORMATION AND NAMING OF IONIC COMPOUNDS

VIDEO
WORKED EXAMPLE 12A–3

NOTE

Recall from Chapter 4A that roman numerals are used for oxidation states.

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Identifying redox couples

During a redox reaction, the species acting as a reducing agent on the reactant side donates (loses) electrons and in doing so forms a product that is an oxidising agent. At the same time, the reactant acting as the oxidising agent is converted into a product that is a reducing agent. The association of each reactant and the product it forms during a redox reaction is referred to as a **redox couple**.

Redox couple a reducing agent, which donates electrons, and its corresponding oxidising agent, which accepts electrons

For example, when copper metal undergoes oxidation, losing electrons, the copper ions produced could act as an oxidising agent, gaining electrons.

$Cu(s)$ □ $Cu^{2+}(aq)$ $+ 2e^-$
Reducing agent — Redox couple oxidising agent

In a similar way, when zinc ions undergo reduction to zinc solid, the zinc solid could then act as a reducing agent.

$Zn^{2+}(aq)$ $+ 2e^-$ □ $Zn(s)$
Oxidising agent — Redox couple reducing agent

NOTE

The complete electrochemical series can be found in the Unit 3 & 4 VCE Exam Data Booklet and shows an extensive list of these redox couples, arranged in order of increasing strength of the reducing agent. You will explore this further in Unit 3 when predicting which redox reactions will take place.

UNIT 3 LINK

12B WRITING REDOX EQUATIONS LINK

Your ability to identify redox couples will be equally important in the next section of this chapter when you explore how to write half-equations from full redox equations. In most cases, you will be required to identify redox couples in full redox reactions rather than in half-equations. Each redox reaction will therefore have *two* redox couples.

To identify a redox couple, the following are required.

1 There must be one species on the reactant side and one species on the product side of the equation containing the same element.
2 There must be a change in oxidation state.

Figure 12A–4 shows that copper has an oxidation state of +2 in copper oxide, CuO, on the reactant side, which decreases to 0 in copper (solid) on the product side. This is the first redox couple. In a similar way, the hydrogen gas and water can be identified as the second redox couple, where hydrogen shows an increase in oxidation state from 0 to +1.

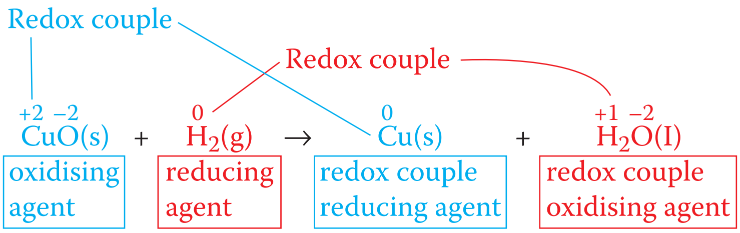

Figure 12A–4 Each redox reaction has two redox couples.

NOTE

You may have noticed that Figure 12A–4 also includes an element that does not change oxidation state. In this case, it is oxygen. If an element does not have a change in oxidation state during a redox reaction, it means that element is not undergoing reduction or oxidation.

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Check-in questions – Set 3

1 State whether the following are redox reactions or not and explain why.
  a $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$
  b $Na_2CO_3(aq) + 2HCl(aq) \rightarrow CO_2(g) + H_2O(l) + 2NaCl(aq)$
2 Identify whether the following are reduction or oxidation reactions.
  a $Ag(s) \rightarrow Ag^{+}(aq) + e^{-}$
  b $2Cl^{-}(aq) \rightarrow Cl_2(g) + 2e^{-}$
3 Assign oxidation states to each element in the following reactions.
  a $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$
  b $2Fe_2O_3(s) + 3C(s) \rightarrow 4Fe(s) + 3CO_2(g)$
4 For each reaction in Question **3**, state which element has been reduced and which has been oxidised.
5 For each reaction in Question **3**, identify which species is the reducing agent and which species is the oxidising agent.

12A SKILLS

Importance of oxidation states

In this section, you have seen the value of being able to assign and use oxidation states. The following important steps help you to do this.

1 Identify reactions as redox reactions (or not).
2 Identify reduction.
3 Identify oxidation.
4 Identify reducing agents.
5 Identify oxidising agents.
6 Identify redox couples.
7 Identify the movement of electrons.

Concept-mapping from key glossary terms

This section has explored much of the underlying chemistry and key definitions for redox reactions, including how to identify reduction, oxidation, reducing agents, oxidising agents, redox couples and ways of assigning and using oxidation states. Table 12A–2 and Figure 12A–5 (both on the following page) summarise the key concepts covered but also highlight how the language you need to know and understand has a lot of similarities. This can sometimes cause confusion initially, so it is important to use the glossary terms throughout this chapter to help you strengthen your understanding of the language used. A useful activity to consolidate your ability to correctly use these terms is to create your own concept map showing connections between them. You could also annotate the concept map at the beginning of the chapter to consolidate the connections between the related terminology used.

Table 12A–2 Comparing oxidation and reduction

Oxidation	Reduction
Loss of electrons	Gain of electrons
Electrons on RHS of half-equations	Electrons on LHS of half-equations
Oxidation state increases	Oxidation state decreases
Caused by oxidising agent	Caused by reducing agent
Reducing agent undergoes oxidation	Oxidising agent undergoes reduction

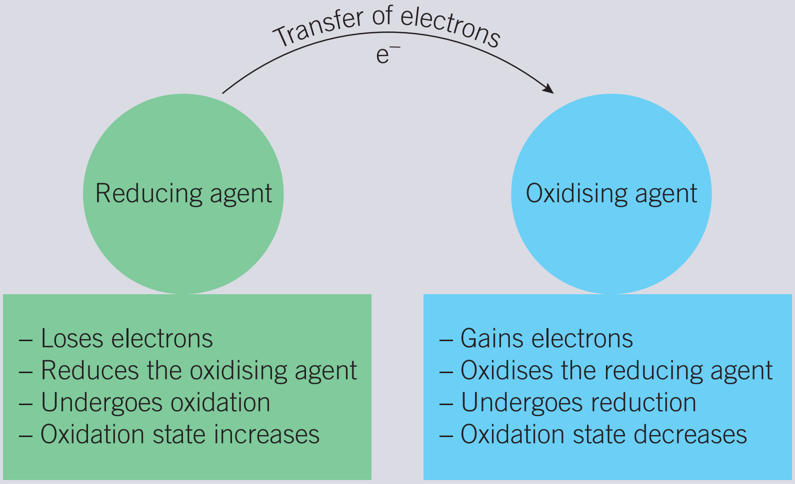

Figure 12A–5 Summary of key terms and concepts for reducing agents and oxidising agents

Section 12A questions

1 Use the following terms to correctly complete the sentences below:

decrease
gain
identify
increase
oxidation
reducing
reduction
transfer

Redox reactions involve a __________ of electrons. During a redox reaction, both ______________ and oxidation occur simultaneously. Reduction is the ________ of electrons. ______________ is the loss of electrons. Oxidation states can be used to _____________ the reducing agent and oxidising agent. The _______________ agent can be identified by a decrease in oxidation state. The oxidising agent can be identified by a/an _____________ in oxidation state. During reduction, the oxidising agent is reduced, and its oxidation state will _______________.

2 Write a definition for reduction in terms of oxidation states.

3 Use oxidation states to identify which of the following reactions is **not** a redox reaction.

A $Ag(s) + H_2S(g) \rightarrow Ag_2S(g) + H_2(g)$
B $Fe(s) + O_2(g) \rightarrow Fe_2O_3(g)$
C $2Ag^+(aq) + Ni(s) \rightarrow Ni^{2+}(aq) + Ag(s)$
D $HBr(aq) + KOH(aq) \rightarrow H_2O(l) + KBr(aq)$

4 Classify each of the following half-equations as reduction or oxidation.
  a $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$
  b $2Br^-(aq) \rightarrow Br_2(aq) + 2e^-$
  c $Cu^{2+}(aq) + e^- \rightarrow Cu^+(aq)$

5 For each of the reactions below, explain how you would determine oxidation states to help you identify the reducing agent and the oxidising agent.
  a $Pb^{2+}(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + Pb(s)$
  b $Ca(s) + Cl_2(g) \rightarrow CaCl_2(s)$

6 Consider the reaction $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$.
  a Identify the two redox couples.
  b State which species is being reduced and which is being oxidised.

7 For each of the following reactions, state the reducing agent and oxidising agent.
  a $2NaBr(aq) + Cl_2(g) \rightarrow 2NaCl(aq) + Br_2(l)$
  b $2As(s) + 3Cl_2(g) \rightarrow 2AsCl_3(l)$

8 For each reaction in Question 7, state which species is undergoing reduction and which is undergoing oxidation.

12B Writing redox equations

Study Design:
Oxidising and reducing agents, and redox reactions, including writing of balanced half and overall redox equations (including in acidic conditions), with states

Glossary:
Complex half-equation
KOHES
Simple half-equation

11A INTRODUCTION TO ACIDS AND BASES LINK

11C CALCULATING pH LINK

UNIT 3 LINK

ENGAGE

How do redox reactions happen?

What conditions are needed for a redox reaction to take place? Do they even matter? Redox reactions can take place in a range of conditions, including neutral, basic or acidic conditions. These conditions do have some influence over what species are involved during a redox reaction and are important to consider when writing balanced equations for a given scenario.

In basic conditions, the OH^- ions play a role in the redox reactions that can take place. You might already be familiar with alkaline batteries: these are examples of redox reactions taking place in basic conditions. Car batteries, on the other hand, are an example of redox reactions taking place in acidic conditions. You will learn more about batteries and cells and their conditions in Unit 3 Chemistry.

Figure 12B–1 AA batteries (on the left) are an example of redox reactions occurring under basic conditions. Car batteries (on the right) are an example of redox reactions occurring under acidic conditions.

Simple half-equation either the reduction or oxidation part of a redox reaction involving only one element

EXPLAIN

Writing redox half-equations

VIDEO 12B–1 HALF-EQUATIONS

12A REDUCTION AND OXIDATION LINK

Simple half-equations

You should recall that redox reactions involve *both* reduction and oxidation. As you have already seen in Chapter 12A, reduction and oxidation reactions can be written as separate half-equations, which show the electrons either as reactants in reduction half-equations or as products in oxidation half-equations. These are referred to as **simple half-equations**.

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When writing simple half-equations, which contain only one type of element, there are a few steps that you need to follow. Applying the acronym **ECS** will help you to remember these three critical steps.

1. Write the species on the correct side of the equation and then balance the **E**lement involved.
2. Balance the **C**harges in the equation. This is done by adding the appropriate number of electrons to the *more positive* side of the half-equation. Note that you can apply your understanding of oxidation states to each element in an equation to determine this.
3. Finally, ensure that you have included correct **S**tates for each element listed.

LINK **12A** REDUCTION AND OXIDATION

NOTE

When assigning states for redox equations in this course, note the following points.

1. Ions are always aqueous, (aq).
2. Electrons have **no** state.
3. Water, H_2O, is always liquid, (l).

Worked example 12B–1: Writing oxidation half-equations

VIDEO WORKED EXAMPLE 12B–1

Write the half-equation to represent the oxidation of solid iron, Fe(s), to iron ions, $Fe^{2+}(aq)$.

Step 1 **E**	From the description provided, write the species on the appropriate side of the equation. In this example, the moles of Fe is the same on each side, so the **E**lements are balanced.	$Fe \rightarrow Fe^{2+}$
Step 2 **C**	Identify which side is more positive and by how many. This can be done by applying oxidation states. Add the appropriate number of electrons to that side. In this example, the right-hand side is more positive by 2, meaning that two electrons need to be added to this side, which balances the **C**harges.	$\overset{0}{Fe} \rightarrow \overset{+2}{Fe^{2+}} + 2e^-$
Step 3 **S**	Add **S**tates to each element as required.	$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Figure 12B–2 In the above reaction, solid iron (Fe) is converted to Fe^{2+} ions. These ions then go on to react with oxygen, forming iron oxide, commonly known as rust.

LINK **12D** APPLICATIONS OF REDOX REACTIONS IN SOCIETY

ISBN 978-1-009-22964-7

NOTE

The charges on each side of an equation *do not* necessarily have to equal zero. You must only balance the charges so that they are the *same* on each side of the half-equation.

Worked example 12B–2: Writing reduction half-equations

VIDEO
WORKED
EXAMPLE 12B–2

Write the redox half-equation to represent the reduction of chlorine gas, $Cl_2(g)$ to chloride ions, $Cl^-(aq)$.

Step 1 **E**	From the description provided, write the species on the appropriate side of the equation. In this example, there are two chlorines on the LHS, but only one on the RHS. To balance the elements, we need to add an extra chloride ion, written as a coefficient in front of the species.	$Cl_2 \rightarrow Cl^-$ $Cl_2 \rightarrow 2Cl^-$
Step 2 **C**	Identify which side is more positive and by how many. Add the appropriate number of electrons to that side. In this example, the LHS is more positive by 2, meaning that two electrons need to be added to this side.	$Cl_2 + 2e^- \rightarrow 2Cl^-$
Step 3 **S**	Add states to each element as required.	$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$

Check-in questions – Set 1

1 Balance the following half-equations and identify each as either reduction or oxidation.
  a $F_2(g) \rightarrow F^-(aq)$
  b $Sn^{2+}(aq) \rightarrow Sn(s)$
  c $Na(s) \rightarrow Na^+(aq)$

Complex half-equations

Oxidation or reduction reactions in acidic conditions contain more than one element and so require a different approach to balance them correctly. Applying the acronym **KOHES**, shown below, to these **complex half-equations** will help ensure you follow the appropriate order when writing and balancing them.

KOHES acronym for remembering how to write and balance complex half-equations

Complex half-equation either the reduction or oxidation part of a redox reaction involving more than one element

K	Balance the **K**ey element.	This will be the element that is not oxygen or hydrogen. This is done by placing the correct coefficient in front of the required species.
O	Balance the **O**xygen.	This is done by adding water, H_2O, to the side of the equation with fewer oxygen atoms.
H	Balance the **H**ydrogen.	This is done by adding hydrogen ions, H^+, to the appropriate side of the equation.
E	Balance the **E**lectrons.	Add the appropriate number of electrons to the more positive side.
S	Include **S**tates on all species.	Water should be given a liquid state, (l), and the hydrogen ions will be written as aqueous, (aq).

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Worked example 12B–3: Writing half-equations in acidic conditions

VIDEO
WORKED
EXAMPLE 12B–3

Dichromate ions, $Cr_2O_7^{2-}$, are a strong oxidising agent and, as such, can be reduced to chromium ions, Cr^{3+}. Write the half-equation under acidic conditions to represent this.

Step 1 **K**	Start by writing each species on the appropriate side of the equation. Balance the **K**ey element, in this case the chromium, Cr. There are two chromium atoms on the LHS, but only one on the RHS. Balance by placing the coefficient 2 in front of the Cr^{3+}.	$Cr_2O_7^{2-} \rightarrow Cr^{3+}$ $Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$
Step 2 **O**	Balance the **O**xygen. Do this by adding seven water molecules to the RHS.	$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$
Step 3 **H**	Balance the **H**ydrogen. Do this by adding 14 H^+ ions to the LHS.	$Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$
Step 4 **E**	Balance the ionic charges by adding **E**lectrons to the more positive side. In this example, the total of the ionic charges is 12+ on the LHS and 6+ on the RHS. As such, six electrons are needed on the LHS.	$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$
Step 5 **S**	Add **S**tates to the half-equation.	$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$

Check-in questions – Set 2

1 Use KOHES to balance the following half-equations in acidic conditions and identify each half-equation as reduction or oxidation.

   a $MnO_4^-(aq) \rightarrow Mn^{2+}(aq)$

   b $H_2SO_3(aq) \rightarrow SO_4^{2-}(aq)$

Combining half-equations

You should recall that redox reactions involve a transfer of electrons where reduction and oxidation both occur simultaneously. That is, you can't have one without the other, and the number of electrons that the oxidising agent gains, *must be equal to* the number of electrons that the reducing agent loses. When considering half-equations, this means that the number of electrons on the left-hand side of a reduction half-equation *must be equal to* the number of electrons on the right-hand side of the corresponding oxidation half-equation.

In situations where the number of electrons differ in each half-equation, you must find the lowest common multiple for the number of electrons. Do this by multiplying one or both half-equations to ensure that the numbers of electrons are equal to this lowest common multiple value. The half-equations can then be added or combined to give the full, overall redox equation. Worked example 12B–4 shows you how to do this.

Worked example 12B–4: Combining redox half-equations

VIDEO
WORKED
EXAMPLE 12B–4

Write the combined redox equation for the reaction between silver ions and copper metal.

Solution

Step 1	Write each of the balanced half-equations that take place.	$Ag^+(aq) + e^- \rightarrow Ag(s)$ $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
Step 2	Balance the electrons in each equation by finding the lowest common multiple. The silver half-equation only has one electron, whereas the copper half-equation contains two electrons. To balance the electrons, we need to multiply the silver half-equation by 2.	$2 \times (Ag^+(aq) + e^- \rightarrow Ag(s))$ $2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$
Step 3	Combine the half-equations by adding them together. A simple way of doing this is to write each species that is on the left of the arrow of the half-equations together on the left and write each species that is on the right of the arrow of the half-equations together on the right. Note that it is important to place the arrows directly below each other.	$2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$ $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ $2Ag^+(aq) + 2e^- + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}(aq) + 2e^-$
Step 4	The electrons on the left-hand side of the arrow cancel out those on the right-hand side, leaving the full balanced redox equation.	$2Ag^+(aq) + \cancel{2e^-} + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}(aq) + \cancel{2e^-}$ $2Ag^+(aq) + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}(aq)$

4C PRECIPITATION
REACTIONS
LINK

The final equation in the solution to Worked example 12B–4 is an example of a redox equation that can also be classified as an ionic equation. It shows the species reacting and their charges and states but does not show the electrons that are being transferred.

Writing balanced half-equations from full redox reactions

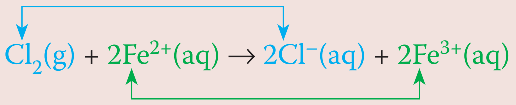
Sometimes you may be given an ionic equation, or full redox equation, from which you need to identify what the half-equations are. This involves applying most of the skills that you have been looking at in this chapter. You will need to identify the *redox couples* and follow the steps to balance each half-equation that were covered in Worked examples 12B–1, 12B–2 and 12B–3.

Worked example 12B–5: Writing half equations from full equations

VIDEO
WORKED
EXAMPLE 12B–5

Use the following redox reaction to write the two half-reactions. Also identify which half-reaction is reduction and which is oxidation.

$$Cl_2(g) + 2Fe^{2+}(aq) \rightarrow 2Cl^-(aq) + 2Fe^{3+}(aq)$$

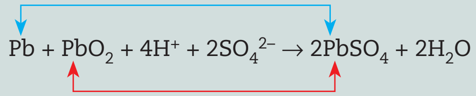
Step 1	Identify the redox couples and write each pair on appropriate sides of the half-equation.	 $Cl_2(g) + 2Fe^{2+}(aq) \rightarrow 2Cl^-(aq) + 2Fe^{3+}(aq)$ $Cl_2 \rightarrow Cl^-$ $Fe^{2+} \rightarrow Fe^{3+}$
Step 2 **E**	Balance the **E**lements by adding coefficients where necessary. Only the Cl⁻ ion requires the coefficient 2 in this example.	$Cl_2 \rightarrow 2Cl^-$ $Fe^{2+} \rightarrow Fe^{3+}$
Step 3 **C**	Balance the **C**harges by adding electrons to the more positive side of each half-equation.	$Cl_2 + 2e^- \rightarrow 2Cl^-$ $Fe^{2+} \rightarrow Fe^{3+} + e^-$
Step 4 **S**	Include **S**tates.	$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$ $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$
Step 5	Identify which is reduction and which is oxidation. The simplest way to do this is to identify on which side of the half-equation the electrons are shown. An alternative way to do this would be to use oxidation states.	Reduction (electrons on LHS): $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$ Oxidation (electrons on RHS): $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$

NOTE

There is sometimes a rare case where the reduction and oxidation reactions both result in the *same* product being formed. An example of this is shown below.

$$Pb(s) + PbO_2(s) + 4H^+(aq) + 2SO_4^{2-}(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$$

The conjugate pairs for such a reaction would each have the same product as indicated below.



$$Pb + PbO_2 + 4H^+ + 2SO_4^{2-} \rightarrow 2PbSO_4 + 2H_2O$$

Sometimes it is possible for the opposite to happen, where the same species in a redox reaction acts as both the reducing agent and the oxidising agent. In one half-reaction, it would undergo reduction, and in the other half-reaction it would undergo oxidation. These are uncommon circumstances but are important to be aware of.

WORKSHEET
12B–1
WRITING
SIMPLE AND
COMPLEX HALF-
EQUATIONS

ISBN 978-1-009-22964-7 Drummond et al

Check-in questions – Set 3

1 Write the full ionic equation by combining the following half-equations.
  - a $Al^{3+}(aq) + 3e^- \rightarrow Al(s)$ and $Ca(s) \rightarrow Ca^{2+}(aq) + 2e^-$
  - b $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$ and $K(s) \rightarrow K^+(aq) + e^-$

2 Write the reduction and oxidation half-equations for the full redox equations shown below. Label each half-equation as either reduction or oxidation.
  - a $2H^+(aq) + 2Li(s) \rightarrow 2Li^+(aq) + H_2(g)$
  - b $2MnO_4^-(aq) + 6H^+(aq) + 5H_2O_2(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5O_2(g)$

12B SKILLS

Key things to check when writing balanced equations

It is important that you are able to write balanced equations, particularly in this section of the course. Below are a few points to remember that you may find helpful.

1 Never change the subscripted numbers in a formula.

2 Only ever change coefficients to balance elements.

3 A balanced equation must have the same number of each element on both the reactant and product sides.

4 Charges on both the reactant and product sides must always be the same. One common mistake is adding electrons to each side of a redox equation to make the charges equal zero on both sides. Sometimes the charges will be zero, other times positive or negative. They simply need to be the same on each side.

5 Finally, always make sure that you add states to any equations you write.

Section 12B questions

1 Explain the difference between a redox half-equation and a full redox equation.

2 Balance the following half-equations.
  - a $H_2(g) \rightarrow H^+(aq)$
  - b $Cr^{3+}(aq) \rightarrow Cr(s)$
  - c $Ag(s) \rightarrow Ag^+(aq)$

3 Balance the following half-equation under acidic conditions.

$$HNO_2(aq) \rightarrow NO_3^-(aq)$$

4 Write the full redox equation by combining the following half-equations.
  - a $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$ and $Co(s) \rightarrow Co^{2+}(aq) + 2e^-$
  - b $F_2(g) + 2e^- \rightarrow 2F^-(aq)$ and $K(s) \rightarrow K^+(aq) + e^-$
  - c Explain the importance of multiplying the potassium, K, half-equation in part **b** when determining the overall equation.

5 Write the reduction and oxidation half-equations for the full redox equations shown below. Label each half-equation as either reduction or oxidation.
  - a $Zn^{2+}(aq) + 2Li(s) \rightarrow 2Li^+(aq) + Zn(s)$
  - b $3CH_3CH_2OH(aq) + 2Cr_2O_7^{2-}(aq) + 16H^+(aq) \rightarrow 4Cr^{3+}(aq) + 11H_2O(l) + 3CH_3COOH(aq)$

12C

Metal displacement reactions

Study Design:
The reactivity series of metals and metal displacement reactions, including balanced redox equations, with states

Glossary:
Displacement reaction
Reactivity

ENGAGE

Will it react?

Will it react? This is a simple question that scientists are often asking themselves before they combine two or more chemicals together. When solid pieces of metal are placed in solutions containing different ions, sometimes a reaction will take place, but other times nothing happens at all.

In this section, you will look at how the reactivity series can be used to predict whether a reaction will take place.

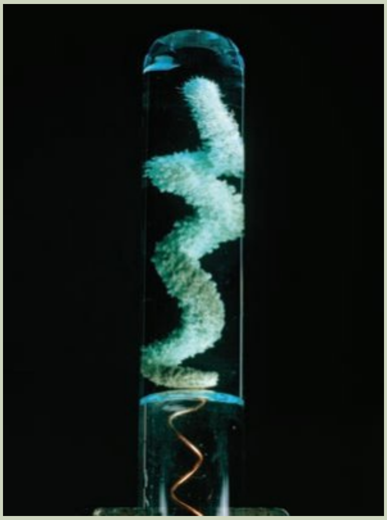


Figure 12C–1 In this simple reaction, silver ions are displaced by the copper wire, forming silver crystals on its surface.

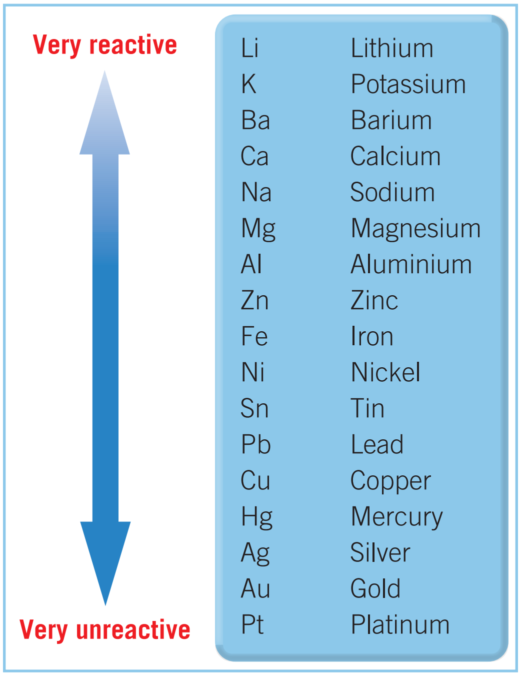
EXPLAIN

Metal displacement reactions

Displacement reactions occur when a more reactive element replaces another. There are two different types of displacement reactions: single displacement reactions involving one change, and double displacement reactions involving two changes. This section of the course looks at single displacement redox reactions involving metals.

Figure 12C–2 shows the reactivity series of metals. Metals are arranged in their order of **reactivity**, meaning those listed at the top can be considered more highly reactive than those towards the bottom, which may react more slowly, or not at all.

A metal displacement reaction will occur when a metal higher in the reactivity series is combined with a compound of a different, less active reactive metal below it. You should recall that active metals combine with non-metals to form ionic compounds. In a



Li	Lithium
K	Potassium
Ba	Barium
Ca	Calcium
Na	Sodium
Mg	Magnesium
Al	Aluminium
Zn	Zinc
Fe	Iron
Ni	Nickel
Sn	Tin
Pb	Lead
Cu	Copper
Hg	Mercury
Ag	Silver
Au	Gold
Pt	Platinum

Figure 12C–2 Reactivity series of metals. Elements listed at the top of this series are more reactive than those below them.

VIDEO 12C–1
DISPLACEMENT REACTIONS

Displacement reaction
when a more reactive element replaces another.
In a metal displacement reaction, the ions of the less reactive element are displaced from the solution and form a solid

Reactivity
how likely a chemical is to take part in a chemical reaction

solution, the ionic compound ions dissociate and move around freely. Two scenarios exist when a metal is placed into an ionic solution.

1 When a more reactive metal is placed into an ionic solution containing a less reactive metal ion, the less reactive metal ions are displaced from the solution and are converted into the pure solid metal element.
2 When a less reactive metal is placed into an ionic solution containing a more reactive metal ion, no reaction takes place.

Worked example 12C–1: Writing equations for displacement reactions

Predict whether a reaction will take place when solid copper wire is placed into a solution of silver nitrate. If a reaction does take place, write both the word equation and the balanced chemical equation.

Solution

Step 1	Determine if the metal is more reactive than the metal compound and hence whether a displacement reaction will occur.	Using the reactivity series in Figure 12C–3, copper is listed above silver, so a displacement reaction occurs. The copper will cause the displacement of silver ions out of the solution of silver nitrate, converting them into solid silver atoms, while the solid copper atoms are converted into copper ions. Fe Iron Ni Nickel Sn Tin Pb Lead Cu Copper Hg Mercury Ag Silver Au Gold Pt Platinum **Figure 12C–3** Part of the reactivity series
Step 2	Write a word equation to represent the reaction occurring.	copper + silver nitrate → silver + copper nitrate
Step 3	Now is a good opportunity to review your skills for writing balanced ionic formulas, so that you can balance the equation correctly. Take your time to get this correct because if you balance the ionic formula incorrectly it will lead to the entire equation being incorrectly balanced.	Ionic formula for silver nitrate: $Ag^+ \ NO_3^- = AgNO_3$ Ionic formula for copper nitrate: $Cu^{2+} \ NO_3^-$, therefore two nitrate ions are needed to balance the 2+ charge on the copper ion = $Cu(NO_3)_2$
Step 4	Write a balanced full equation to represent the reaction occurring. Remember that all ionic solutions are given aqueous (aq) states and that the pure metals need to be assigned solid (s) states.	$Cu(s) + AgNO_3(aq) \rightarrow Ag(s) + Cu(NO_3)_2(aq)$ We can see that the nitrate ions are not balanced so need to place a 2 in front of the silver nitrate: $Cu(s) + 2AgNO_3(aq) \rightarrow Ag(s) + Cu(NO_3)_2(aq)$ Now balance the silver by placing a 2 in front of the Ag(s): Balanced equation: $Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Cu(NO_3)_2(aq)$

VIDEO
WORKED EXAMPLE 12C–1

4A FORMATION AND NAMING OF IONIC COMPOUNDS
LINK

WORKSHEET 12C–1
DISPLACEMENT REACTIONS

NOTE

It is important to recognise that this reaction arrow is written in one direction and cannot occur in reverse. Silver cannot displace copper from copper nitrate because silver is *less* reactive than copper.

Displacement reactions as an example of redox

It is important to recognise that a metal displacement reaction is an example of a redox reaction, involving a transfer of electrons. In the displacement reaction, the species higher up in the reactivity series has the capacity to react with the ions of anything below them, displacing the ions into solid metal. The species higher up in the series is more reactive, but what this really means is that it is a stronger reducing agent and will therefore undergo oxidation more easily, losing its electrons to become a positive ion. The species lower in the series is a weaker reducing agent, or put another way, it is a stronger oxidising agent, and therefore will undergo reduction more easily, gaining electrons.

Summary of a metal displacement reaction

These important summary points will assist you in working with the redox chemistry involved in a metal displacement reaction.

1 The species located higher in the reactivity series will be the strongest reducing agent, undergo oxidation easily and lose electrons to become positively charged ions.

$M(s) \rightarrow M^+ + e^-$

2 The species located lower in the reactivity series will be the strongest oxidising agent, undergo reduction and be displaced from its ions to form solid metal.

$M^+ + e^- \rightarrow M(s)$

Check-in questions – Set 1

1 Predict whether a reaction will take place in each situation. If a reaction does take place, write both the word equation and the balanced chemical equation, including states.
  a solid magnesium placed in a solution of copper(II) sulfate
  b solid zinc placed in a solution of sodium chloride
  c solid copper placed in a solution of magnesium sulfate
  d solid aluminium placed in a solution of lead(II) nitrate

12C SKILLS

VIDEO 12C–2
SKILLS: USING THE REACTIVITY SERIES TO PREDICT REDOX REACTIONS

Using the reactivity series to predict redox reactions

This section has looked at how to predict whether a reaction will take place, using the reactivity series to identify the strongest reducing agent and strongest oxidising agent involved in a metal displacement reaction. The species higher up is a stronger reducing agent and will undergo oxidation. The species below is a stronger oxidising agent and will undergo reduction. Once you identify the reducing agent and oxidising agent, you can write the half-equations and then the equation for the full balanced redox reaction.

As the word suggests, *displacement* means ‘places have to be changed’. For metal displacement reactions, it means one metal has to be replaced (or displaced) by the other:

$$\text{Metal } \boxed{\text{A}} + \text{Salt solution of B} \rightarrow \text{Salt solution of } \boxed{\text{A}} + \text{Metal B}$$

For the above reaction to occur, A must be more reactive than B

To check if metal A is more reactive, you need to use the reactivity series, like the one shown here:

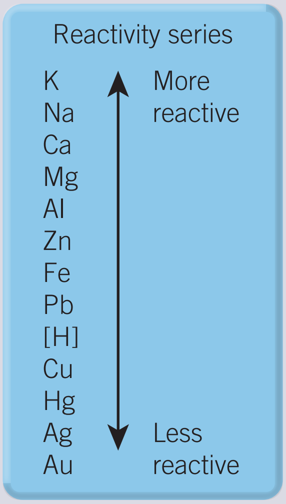


Figure 12C–4 Example of a reactivity series showing more reactive metals at the top and less reactive metals at the bottom

If the metals A and B in the reaction on the previous page were swapped, then no reaction would occur:

$$\underset{\text{Less reactive}}{\text{Metal }\boxed{\text{B}}} + \text{Salt solution of A} \rightarrow \underset{\text{(No reaction)}}{\text{X X}}$$

Section 12C questions

1 Explain what happens during a metal displacement reaction.

2 Explain why magnesium is more reactive than copper.

3 Predict whether a reaction will take place if solid tin is placed into a solution of potassium nitrate. If a reaction does take place, write both the word equation and the balanced chemical equation, including states.

4 Predict whether a reaction will take place if solid zinc is placed into a solution of silver chloride. If a reaction does take place, write both the word equation and the balanced chemical equation, including states.

5 Predict whether a reaction will take place if solid gold is placed into a solution of copper(II) sulfate. If a reaction does take place, write both the word equation and the balanced chemical equation, including states.

12D Applications of redox reactions in society

Study Design:
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

Glossary:

Alloy	Electrolyte
Anode	Hydroxide
Cathode	Oxide
Corrosion	Primary cell
Electrode	Rusting

ENGAGE

Corrosion and rust

Have you ever noticed how easily your bike, scooter or any other metal objects tend to rust over time? You put your bike away in the shed for a few months over winter and by the time you get it out again, the frame, spokes, handlebars and other metal parts of the bike are covered in a coating of brown rust. This rusty appearance is a sign that a chemical reaction is taking place. Unfortunately, as well as appearing unsightly, it can weaken the physical strength and integrity of metals. The undesirable nature of rusting metals means that society is always exploring ways of preventing or minimising rust.

Solid metal structures, such as the Sydney Harbour Bridge, Empire State Building and the Eiffel Tower need to be strong and long lasting (even though the Eiffel Tower was a temporary structure built for the 1889 World Fair). These structures, and many others like them around the world, are prone to rusting and corrosion, which could potentially pose a significant safety hazard should their structure become compromised, resulting in a collapse. In 2018, a section of the Ponte Morandi, a road viaduct in Genoa, Italy, collapsed killing 43 people. The cause of the collapse was corrosion of the southern stay, which played a critical structural role. Over time, the stay had suffered from corrosion, and the crack increased in size, ultimately leading to the failure and buckling of sections within the bridge.

In this section, you will look at the process of corrosion, including rust, and connect that back to the work covered in the previous sections of this chapter. You will explore how redox chemistry can be applied to help minimise or even prevent the undesirable rusting of certain materials used in society.

Figure 12D–1 Corrosion, including rust, can be found on common metal objects used in society, including bikes and bridges.

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EXPLAIN

Corrosion and rust

Corrosion is a destructive process that occurs when the exposed surface of a substance, usually a metal, undergoes electrochemical reactions between it and the surrounding liquid or gases in their environment. It results in the gradual deterioration of materials, as they are converted into new products that are more chemically stable, such as **oxides**, **hydroxides**, sulfides, and carbonates. When metals are involved, these chemical changes often result in a flaky, weakened substance with very different, and often undesirable, properties.

One of the most common types of corrosion is **rusting**, which occurs when iron metal is oxidised to oxides of iron in the presence of both air (oxygen) and moisture (water). The reaction takes place slowly over time, with the oxygen combining with two different versions of iron ions to form two different oxides. The result is a decrease in the thickness of the steel, resulting in a reduction of strength.

Corrosion
an irreversible reaction at the surface of a material with its environment, which results in deterioration of the material

Oxide
a chemical compound with one or more oxygen atoms combined with another element

Hydroxide
a chemical compound containing a hydroxide group (an oxygen and hydrogen covalently bonded together) combined with another element

Rusting
an example of corrosion, where the process specifically involves iron reacting with oxygen and moisture in the air, converting the iron into iron oxides and iron hydroxides

Oxidation of iron: forming rust

The process of rusting is quite complex and involves a number of chemical reactions that are modelled as the production of iron(II) oxide, FeO, and iron(III) oxide, Fe_2O_3. These products are also commonly known as ferrous oxide and ferric oxide, respectively. Compared with many other transition metals, iron is a good reducing agent, and oxygen is an excellent oxidising agent, so the redox reaction proceeds quite readily. This process can be summarised in the following reactions.

Iron, Fe, is oxidised to Fe^{2+} ions.

$$Fe \rightarrow Fe^{2+} + 2e^-$$

These Fe^{2+} ions then undergo oxidation to Fe^{3+} ions, while oxygen is reduced to oxide O^{2-} ions.

$$4Fe^{2+} + O_2 \rightarrow 4Fe^{3+} + 2O^{2-}$$

The Fe^{2+} ions and Fe^{3+} ions then react with water to form hydroxides.

$$Fe^{2+} + 2H_2O \,\square\, Fe(OH)_2 + 2H^+$$

$$Fe^{3+} + 3H_2O \,\square\, Fe(OH)_3 + 3H^+$$

The hydroxides react further, producing water and the iron oxides that constitute rust. Two simplified examples of such reactions are shown below, although it should be noted that there are a number of other reactions that take place to form iron oxide compounds.

$$Fe(OH)_2 \,\square\, FeO + H_2O$$

$$2Fe(OH)_3 \,\square\, Fe_2O_3 + 3H_2O$$

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Factors that influence the rate of rusting

How quickly iron metal rusts depends largely on environmental factors. Table 12D–1 lists some conditions that speed up the rusting of iron.

Table 12D–1 Some conditions that result in an increased rate of corrosion of iron

Condition	Reasoning
High moisture levels	Water is a required reactant in the formation of iron oxides.
High concentrations of salt	Aquatic areas are often prone to a faster rate of rusting because many of the ions from salt water act to speed up the redox process.
Low pH environment	Acidity speeds up the rusting process. This can be seen in areas that have higher levels of acid rain.
High surface area of iron metal	In any environment, the more surface exposed to oxygen and moisture, the faster the rate of rusting.

Preventing corrosion

Preventing corrosion in structures throughout society is of paramount importance. Perhaps the most obvious way to minimise corrosion would be to use a metal that is less reactive than iron. Another option is to consider using other metals that have a natural resistance to corrosion. Some metals, for example, copper, can form a thin oxide film on their surface when the metal reacts with oxygen in the air. This layer then 'protects' the rest of the metal underneath.

However, there are a number of reasons that iron is still preferred over these alternatives. Iron is the second-most abundant metal on Earth, relatively cheap and most importantly, it can be made into an **alloy** with carbon to form steel. Steel has specific and unique properties that make it more suitable than other metals for wide use in society.

Alloy a metallic substance composed of two or more elements

Iron can be protected in a number of ways to help minimise corrosion and rusting.

- Non-metallic coatings, such as paint, plastic, varnish, grease, wax or oils, can be added to the surface, preventing corrosion of the iron underneath.
- Iron can be combined with other metals (alloying) to alter its properties. Some alloys, such as stainless steel in which iron is combined with chromium, are resistant to rusting.
- Galvanisation is a process that is used to coat steel with a more reactive metal, usually zinc, which is relatively cheap and forms strong bonds with the iron or steel. Being located higher on the reactivity series and hence more reactive, zinc protects the iron from reacting. Instead, the zinc layer is corroded rather than the iron until the zinc has corroded away. Galvanised steel is commonly used for trailers, roofing materials, ladders, signposts, stairways and a lot more.

LINK **3C** METAL RECYCLING

LINK **12C** METAL DISPLACEMENT REACTIONS

LINK **UNIT 3**

Figure 12D–2 Grain storage bins made of galvanised steel are a common sight in Australia.

UNIT 3 LINK

- Cathodic protection can be used where an electrochemical cell is created with a piece of more reactive metal attached, known as a sacrificial anode, as shown in Figure 12D–3. You will learn more about electrochemical cells in Unit 3, but put simply, this results in the more reactive metal – usually magnesium, zinc or aluminium – corroding instead of the iron or steel. Electrical currents can be applied, which drive electrons out of a sacrificial anode to protect the metal (which is the cathode).
- On an industrial scale, metals can undergo a laser oxidation treatment that increases corrosion resistance.
- Nitriding (diffusing nitrogen into the surface of the metal) is another process that makes the material, such as steel, harder.

Figure 12D–3 Sacrificial anodes (the white blocks) are placed around a ship's motor to protect it against corrosion.

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Check-in questions – Set 1

1 What are the two critical conditions/reactants required for solid iron to form rust?

2 During the rusting process, iron is converted into Fe^{2+} ions according to the following reaction.

$$Fe \rightarrow Fe^{2+} + 2e^-$$

State whether this is oxidation or reduction.

3 Another one of the reactions that occurs during the rusting process is

$$4Fe^{2+} + O_2 \rightarrow 4Fe^{3+} + 2O^{2-}$$

State the oxidation states of each reactant and product and use this to determine whether iron has been reduced or oxidised.

Primary cells

A **primary cell** is used to generate electrical energy from stored chemical energy by using redox reactions. Single-use batteries are examples of primary cells, and there are many different types of batteries in use today that you may already be familiar with. Despite differences in their size and shape, all of them have a design enabling a redox chemical reaction to take place that converts stored chemical energy into usable, portable electrical energy.

VIDEO 12D–1
PRIMARY CELLS

Primary cell a device that converts chemical energy into electrical energy while discharging but cannot be recharged

Electrode a connecting terminal that allows current to enter or leave

Electrolyte a non-reactive solution or paste that enables the flow of charged particles and helps to facilitate the redox reactions

All primary cells have some common design features that allow for the harnessing of electrical energy.

- They have physically separated **electrodes**. This allows reduction and oxidation reactions to take place but forces the electrons involved to flow through an external circuit, providing usable electrical energy.
- They contain **reactants**. These will spontaneously undergo reduction and oxidation reactions when the circuit is connected.
- They are **unable to be recharged**. When the reactants eventually reach a state of equilibrium, no further electrical energy can be produced, and the battery is described as being 'flat' or out of charge.
- They contain an **electrolyte**. This is a non-reactive solution or paste that enables the flow of charged particles and helps to facilitate the redox reactions occurring.

Figure 12D–4 Batteries come in a range of shapes and sizes and involve redox chemical reactions.

You will learn more about primary cells and the redox reactions involved in these in Unit 3.

LINK UNIT 3

Alkaline cell batteries

Alkaline cell batteries are some of the most widely used batteries in society because they are cheap, have a long shelf-life and provide a reliable amount of power for a range of devices. They are called alkaline cell batteries because they use an alkaline electrolyte paste (often potassium hydroxide), which minimises how much water or liquid is needed to facilitate the reactions. This electrolyte allows for the flow of ions, which are part of the redox reactions taking place inside the battery. There are also electrodes inside that are separated to allow for the flow of electrons during the redox reactions. The **anode**, where oxidation takes place, is made from zinc powder and is dispersed within the electrolyte. The **cathode**, where reduction takes place, is made from manganese dioxide and it forms a layer on the inside of the battery.

Anode the electrode at which oxidation occurs

Cathode the electrode at which reduction occurs

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Redox reactions inside an alkaline cell battery

Although there are other, more complex steps involved in the overall reaction, for simplicity, we will focus on just the reduction and oxidation reactions that take place in the dry cell.

- At the anode, zinc particles on the surface undergo oxidation to zinc ions.

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$$

- At the cathode, the manganese dioxide reacts with water in the reduction step.

$$2MnO_2(s) + H_2O(l) + 2e^- \rightarrow Mn_2O_3(s) + 2OH^-(aq)$$

- The overall reaction is

$$Zn(s) + 2MnO_2(s) + H_2O(l) \rightarrow Zn^{2+}(aq) + Mn_2O_3(s) + 2OH^-(aq)$$

Examples of primary cells

One very common example of the dry cell battery is the zinc–carbon battery, shown in Figure 12D–5. It consists of a central carbon rod that acts as the cathode and a zinc case, which acts as the anode. The carbon rod is surrounded by a paste of manganese dioxide, MnO_2. Between the two electrodes there is an electrolyte paste made of ammonium chloride, NH_4Cl. Each battery, or cell, has the potential to produce 1.50 volts.

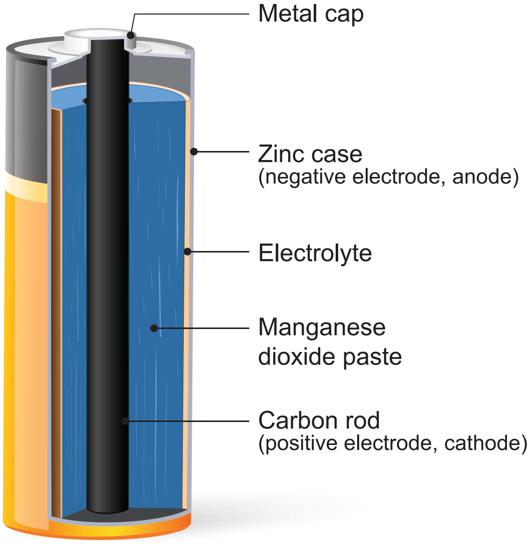

Figure 12D–5 Structural features of a zinc–carbon dry cell battery

Other common types of dry cell battery that will be well known to you include:

- **Alkaline batteries**. Alkaline batteries are a second example of a dry cell battery. They are called alkaline batteries because they contain an alkaline electrolyte, potassium hydroxide, KOH, instead of ammonium chloride.

11A INTRODUCTION TO ACIDS AND BASES LINK

- **Button cell batteries**. Button cells, as shown in Figure 12D–6, are commonly used in watches, calculators, hearing aids, cameras and other small electrical devices.

Figure 12D–6 Button cells come in a range of sizes and are commonly used in small electronic devices.

- **Silver button cell**. Silver oxide batteries are currently produced with flat circular cathodes made of silver oxide, Ag_2O, and an anode gel substance that contains amalgamated zinc powder and an electrolyte (either sodium hydroxide, NaOH, or potassium hydroxide, KOH). These electrodes are separated, and the batteries are insulated and sealed with nylon. Nickel or stainless steel is often used around the exterior of the battery to resist corrosion. Each button cell can produce around 1.60 volts.

The reactions occurring inside the silver oxide button cell are summarised below.

- Reduction of silver oxide, Ag_2O, at the cathode:

$$Ag_2O(s) + H_2O(l) + 2e^- \rightarrow 2Ag(s) + 2OH^-(aq)$$

- Oxidation of zinc, Zn, at the anode:

$$Zn(s) + 2OH^-(aq) \rightarrow ZnO(s) + H_2O(l) + 2e^-$$

- Overall redox reaction:

$$Zn(s) + Ag_2O(s) \rightarrow ZnO\ (s) + 2Ag(s)$$

Check-in questions – Set 2

1 State the type of energy conversion in a primary cell.
2 State which electrode reduction and oxidation each occur at in a primary cell.
3 Explain the purpose of the electrolyte in a primary cell.

12D SKILLS

Applying your understanding of redox reactions to unfamiliar contexts

Exams will often include questions about specific examples that you have not seen or studied before in the course. They could be about certain types of batteries or examples of chemical corrosion and require you to apply your knowledge and understanding of topics that you have studied. The key skill when dealing with these types of questions is to take a moment to think about the specific chemistry that you have studied and apply it to the question.

Experiment-based questions linked to corrosion

You will have noticed throughout chemistry that practical investigations can be conducted to support the theoretical knowledge you are developing. This could also mean that you should expect questions based on experiments or using data generated from experiments linked to specific topics. Examples of these are provided below for corrosion.

Example 1: Investigating factors that affect the rate of corrosion of an iron nail

Five test tubes were set up, each containing an iron nail exposed to different conditions, as shown in Figure 12D–7. The test tubes were left undisturbed for two days before viewing to record results.

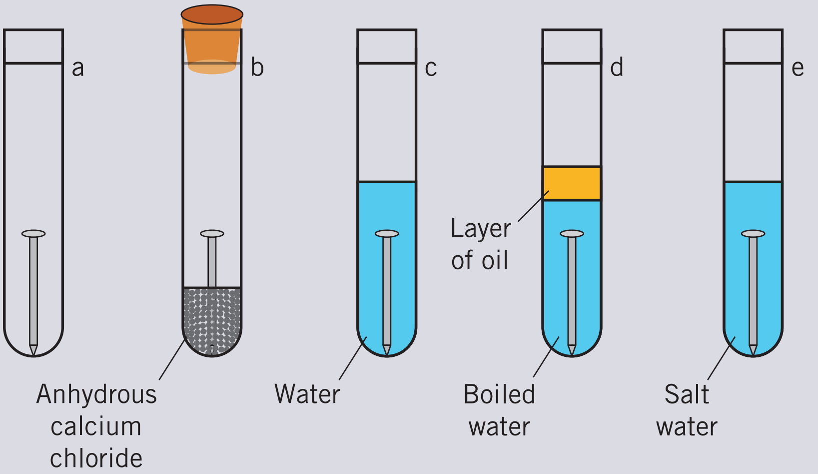

Figure 12D–7 Iron nails exposed to different conditions to compare their relative rates of corrosion

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Some of the types of questions you could expect to be asked in relation to either theory or experimental investigation design are outlined in Table 12D–2.

Table 12D–2 Example questions for corrosion relating to content or experimental investigation design

Corrosion content	Experimental investigation design
What is the purpose of the stopper in test tube b?	What is the purpose of test tube a?
Define corrosion and determine which test tubes will experience corrosion.	Identify both the independent and dependent variables in this experiment.
Write the chemical equation(s) for the reactions that would occur in test tube c that result in the production of iron(II) oxide and iron(III) oxide.	Identify two controlled variables in this experiment and state the importance of controlling each of these.
Identify one approach that is currently used by manufacturers to prevent the corrosion of nails.	Explain whether qualitative or quantitative data is being collected after the two days.
Use the results (or expected results) to explain why car bodies parked near the beach rust faster than those parked inland.	Determine a possible source of error in this experiment and outline one approach that could be employed to minimise the effect of this.
Write a hypothesis outlining what you think will occur in this experiment. In your answer, be sure to support your hypothesis with theory.	

Example 2: Using data to support theoretical knowledge and explanations

Table 12D–3 shows the mass of steel (an alloy of iron and carbon) pieces under five different conditions.

Table 12D–3 Effect of corrosion on mild steel in five different environments

	Mass of steel (g)					
Media	Initial	Week 1 168 hours	Week 2 336 hours	Week 3 504 hours	Week 4 672 hours	Week 5 840 hours
0.1 M HCl	76.5325	75.2104	74.9812	73.1423	71.2532	68.7823
Soil	76.5325	76.5237	76.4972	75.3928	73.3316	72.4259
Atmosphere	76.5325	76.5281	76.5121	76.4857	76.0619	76.0212
Salt water	76.5325	76.5016	75.5536	74.7882	73.8015	73.7856
Fresh water	76.5325	76.5253	76.4612	75.8054	74.8223	73.9441

Using the data, explain which of the media (type of condition) you would place steel rods in to limit the most amount of corrosion. Be sure to refer to the data in your response.

Both of these question scenarios are likely to appear in your own assessments. It is worthwhile taking time to think about what questions could be asked of you for different topics and how you might go about answering them. It is also important to consider where other content you have previously learned connects with new knowledge as you progress through each chapter in this textbook. The 'links' in the sidebars help to make this more obvious. Lastly, it is good to practise questions like these. Examples are included in the Chapter reviews, Unit Revision exercises and digital resources. You will also find more experimental investigation knowledge and example questions in Chapter 16.

16A
INVESTIGATIVE PLANNING AND DESIGN
LINK

Section 12D questions

1 Define ‘corrosion’.

2 The rusting of iron involves a number of steps, including the oxidation of iron.
  a What is the oxidising agent involved in the rusting process?
  b What species is required for the formation of iron hydroxides during the rusting process?

3 List two environmental conditions that might enhance the rate of rust formation.

4 List two chemical modifications that may prevent or slow the rate of rust formation.

5 Explain why magnesium is used as a sacrificial anode in cathodic protection of iron but nickel is not.

6 State two advantages of using batteries in society.

7 Five test tubes were set up, each containing an iron nail exposed to different conditions, as shown in the Figure 12D–7 in the Skills section. The test tubes were left undisturbed for two days before viewing to record results.

Note: Results are not shown for this experiment.

Use Figure 12D–7 to answer the following questions.

  a Identify the independent and dependent variables in this experiment.
  b Identify two controlled variables in this experiment and state the importance of controlling each of these.
  c Determine which test tubes will experience corrosion.
  d Write a hypothesis outlining what you think will occur in this experiment. In your answer, be sure to support your hypothesis with theory.
  e What is the purpose of test tube a?
  f What is the purpose of the stopper in test tube b?
  g Write the chemical equation(s) for the reactions that would occur in test tube c that result in the production of iron(II) oxide and iron(III) oxide.
  h Explain whether qualitative or quantitative data is being collected after the two days.
  i Determine a possible source of error in this experiment and outline one approach that could be employed to minimise the impact of this.
  j Identify one approach that is currently used by manufacturers to prevent the corrosion of nails.
  k Use the expected results to explain why car bodies parked near the beach rust faster than those parked inland.

8 Table 12D–3 in the Skills section shows the mass of steel (an alloy of iron and carbon) pieces under five different conditions.

Using the data, explain which of the media (type of condition) you would place steel rods in to limit the most amount of rust. Be sure to refer to the data in your response.

Chapter 12 review

Summary

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

Success criteria – I am now able to:		Linked question
12A.1	Define ‘redox’ in terms of reduction, oxidation and electron transfer	11☐
12A.2	Understand what a redox reaction is	1☐, 12☐
12A.3	Define and identify reducing agents and oxidising agents	16☐
12A.4	Assign oxidation states to elements in compounds and ions in chemical equations and use these numbers to identify reactions as redox	2☐, 3☐, 4☐, 13☐, 14☐
12A.5	Understand reduction and oxidation in terms of changes in oxidation state	5☐, 6☐
12A.6	Use oxidation states to identify redox couple reducing agents and oxidising agents	15☐, 16☐, 17☐, 18☐
12B.1	Write and balance simple and complex half-equations	19☐
12B.2	Write and balance overall redox equations, including states	20☐
12B.3	Write and balance half-equations by identifying redox couples in overall equations	21☐
12C.1	Define ‘metal displacement reaction’	22☐
12C.2	Use the reactivity series to predict whether a displacement reaction will occur	7☐
12C.3	Write and balance equations by predicting the products of a displacement reaction	23☐
12D.1	Define ‘corrosion’	24☐
12D.2	Understand the process of rusting, including reactants and conditions required	8☐, 13☐
12D.3	Understand ways our understanding of redox chemistry can be used to prevent undesirable corrosion reactions in society	9☐, 25☐
12D.4	Understand how redox reactions can be used in primary cells to generate electrical energy	10☐, 26☐
12D.5	Know the importance of including certain design features in primary cells	22☐

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Multiple-choice questions

1 Which of the following reactions is **not** a redox reaction?

A $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$

B $Pb(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow 2NaNO_3(aq) + PbSO_4(s)$

C $2Na(s) + Cl_2(g) \rightarrow 2NaCl(aq)$

D $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

2 What is the oxidation state of chromium, Cr, in Cr_2O_3?

A −3

B +2

C +3

D +4

3 In which of the following species does nitrogen have the lowest oxidation state?

A N_2O

B N_2

C NH_4^+

D NO_3^-

4 Which of the following is a redox reaction?

A $KCl(aq) + AgNO_3(aq) \rightarrow KNO_3(aq) + AgCl(s)$

B $2HCl(aq)+ Ca(OH)_2(aq) \rightarrow H_2O(l) + CaCl_2(aq)$

C $3Mg(s) + 2AlCl_3(aq) \rightarrow 3MgCl_2(aq) + 2Al(s)$

D $CaCO_3(s) + H_2SO_4(aq) \rightarrow CO_2(g) + H_2O(l) + CaSO_4(aq)$

5 Which species is the oxidising agent in the following reaction?

$$4HCl(aq) + MnO_2(s) \rightarrow Cl_2(g) + MnCl_2(aq) + 2H_2O(l)$$

A $Cl_2(g)$

B $HCl(aq)$

C $MnO_2(s)$

D $MnCl_2(aq)$

6 Consider the following redox reaction:

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$$

Identify which of the following represents the reduction half-equation.

A $I_2(aq) + 2e^- \rightarrow 2I^-(aq)$

B $2S_2O_3^{2-}(aq) \rightarrow S_4O_6^{2-}(aq) + 2e^-$

C $2I^-(aq) \rightarrow I_2(aq) + 2e^-$

D $S_4O_6^{2-}(aq) + 2e^- \rightarrow + 2S_2O_3^{2-}(aq)$

7 Identify which of the following combinations would **not** result in a reaction.

A Solid magnesium is placed into a solution of copper(II) sulfate.

B Solid tin is placed into a solution of potassium nitrate.

C Solid zinc is placed into a solution of silver chloride.

D Solid copper is placed into a solution of silver nitrate.

8 Which of the following conditions would enhance the rate of rust formation?

A low concentration of oxygen

B low pH

C low moisture levels

D low salt levels

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9 Which of the following treatments would **not** reduce the rate of rust formation in steel?

A painting the metal surface
B using nitrides to make the metal harder
C coating the steel with zinc
D using tin as a sacrificial anode

10 Which of the following is **not** a design feature of a primary cell?

A separated anode and electrode
B rechargeable
C contains an electrolyte
D converts chemical energy to electrical energy

Short-answer questions

11 State two definitions that can be used for the term 'oxidation'. (2 marks)

12 With reference to oxidation states, outline whether the following reactions are reduction or oxidation. (6 marks)

a $Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$
b $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$
c $C_2H_6O(aq) + H_2O(l) \rightarrow C_2H_4O_2(aq) + 4H^+(aq) + 4e^-$

13 State whether the following reactions are redox or not. (2 marks)

a $2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + SO_2(g)$
b $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

14 Assign oxidation states to each element in the following reactions. (4 marks)

a $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
b $CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$

15 For each reaction in Question **14**, state which element has been reduced and which has been oxidised. (4 marks)

16 For each reaction in Question **14**, identify which species is the reducing agent and which species is the oxidising agent. (4 marks)

17 With reference to oxidation states, identify which species is the reducing agent and the oxidising agent in each of the following reactions. (12 marks)

a $4Li(s) + O_2(g) \rightarrow 2Li_2O(s)$
b $Mg(s) + Fe^{2+}(aq) \rightarrow Mg^{2+}(aq) + Fe(s)$
c $Cu(s) + 4HNO_3(l) \rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$
d $2FeSO_4(aq) + H_2SO_4(aq) + H_2O_2(aq) \rightarrow Fe_2(SO_4)_3(aq) + 2H_2O(l)$

18 For each reaction in Question **17**, state which species is undergoing reduction and which is being oxidised. (8 marks)

19 Write and balance the half-equation for the production of carbon dioxide gas from carbon monoxide gas. Identify the reaction as reduction or oxidation. (2 marks)

20 Write the redox equation by combining the following zinc and copper half-equations.

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s) \text{ and } Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$$ (1 mark)

21 Consider the following unbalanced redox reaction:

$$C_2O_4^{2-}(aq) + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + CO_2(g)$$

a Write the balanced reduction and oxidation half-equations for the redox equation shown above, under acidic conditions. (2 marks)
b Combine the half-equations to give the full balanced redox equation. (2 marks)

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22 Explain what occurs during a metal displacement reaction. (2 marks)

23 Predict whether a reaction will take place in the following situations. If a reaction does take place, write both the word equation and the balanced chemical equation, including states.

- **a** Solid nickel is placed in a solution of tin(II) sulfate. (2 marks)
- **b** Solid lead is placed in a solution of aluminium nitrate. (2 marks)

24 Define 'corrosion'.

25 Referring to the reactivity series, explain why stainless steel is less prone to rust than carbon steel. (2 marks)

26 The reduction reaction for an alkaline cell battery is shown below:

$$2MnO_2(s) + H_2O(l) + 2e^- \rightarrow Mn_2O_3(s) + 2OH^-(aq)$$

- **a** Write the reaction for the oxidation of zinc to Zn^{2+} ions occurring at the anode. (1 mark)
- **b** Write the overall balanced redox reaction. (1 mark)
- **c** State two important design features that allow this cell to produce electrical energy from stored chemical energy. (2 marks)

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UNIT 2 HOW DO CHEMICAL REACTIONS SHAPE THE NATURAL WORLD?

CHAPTER 13 MEASURING SOLUBILITY AND CONCENTRATION

Introduction

What do toothpaste and drinking water have in common? The answer is that they both contain fluoride. Large amounts of scientific evidence have shown that fluoride is important for reducing tooth decay. For this reason, fluoride has been added to drinking water for decades to help improve dental health. In Australia, the first program to add fluoride to water occurred in Tasmania in 1953. However, it has been a common practice in almost all large Australian cities since the 1970s. The National Health and Medical Research Council (NHMRC) states that it 'supports Australian states and territories fluoridating their drinking water supplies within the range of 0.6 to 1.1 milligrams of fluoride per litre (mg/L)'.

INTRODUCTION VIDEO
MEASURING SOLUBILITY AND CONCENTRATION

While fluoride is beneficial for healthy teeth, other substances sometimes found in water can be harmful, particularly when at high concentrations. In this chapter, you are going to learn about concentration and solubility and how both can be measured. You will also learn about reactions that can be performed to determine an unknown concentration, referred to as volumetric analysis, as well as other reactions that can help remove unwanted substances from water, known as precipitation reactions.

Curriculum

Area of Study 2 Outcome 2
How are chemicals measured and analysed?

Study Design:	Learning intentions – at the end of this chapter I will be able to:
• 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	**13A Measures of solubility** **13A.1** Define 'solubility', 'solute' and 'solvent' **13A.2** Calculate the concentration of a solution **13A.3** Convert between units of concentration

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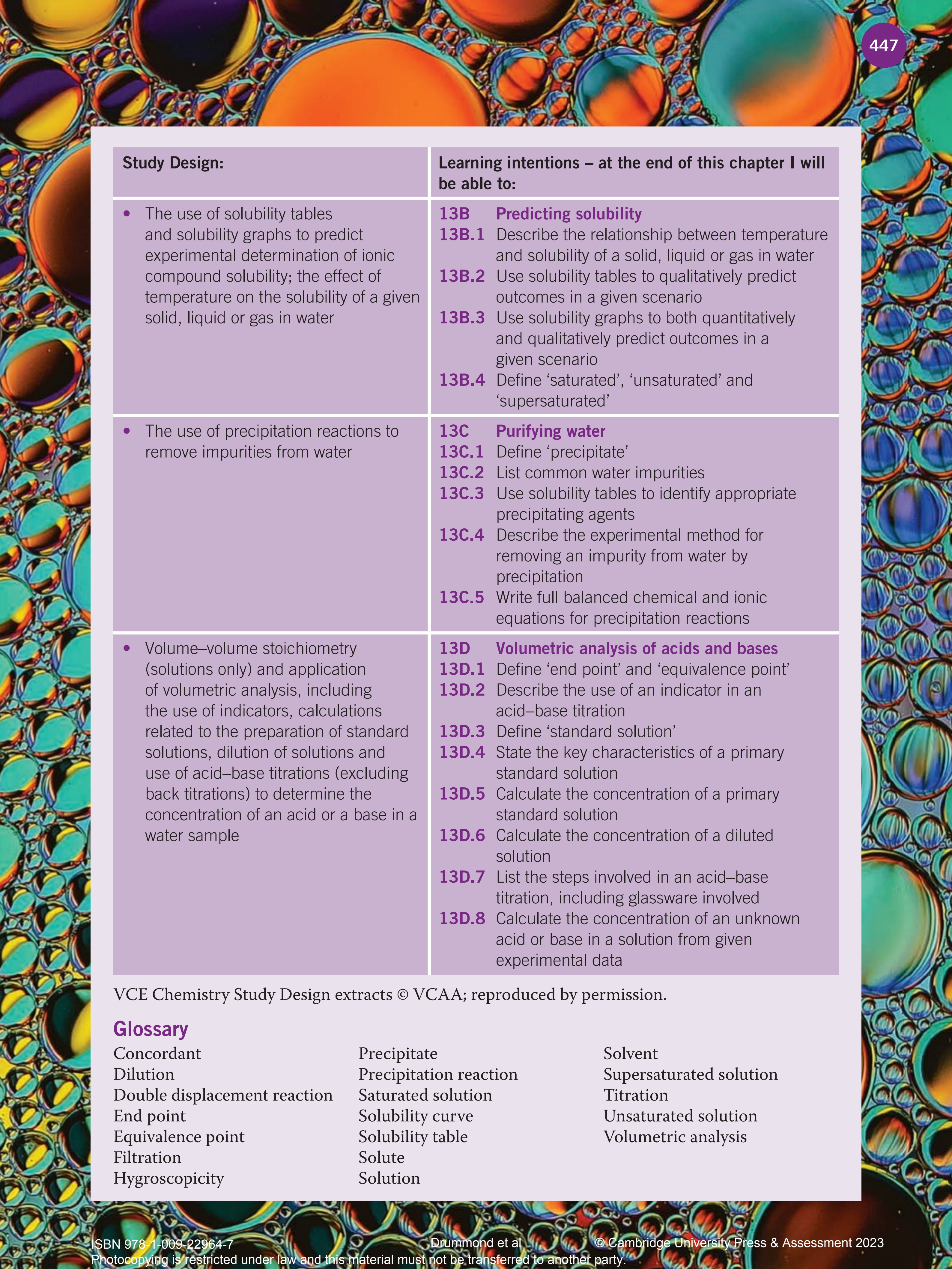

Study Design:	**Learning intentions – at the end of this chapter I will be able to:**
- The use of solubility tables and solubility graphs to predict experimental determination of ionic compound solubility; the effect of temperature on the solubility of a given solid, liquid or gas in water	**13B Predicting solubility** **13B.1** Describe the relationship between temperature and solubility of a solid, liquid or gas in water **13B.2** Use solubility tables to qualitatively predict outcomes in a given scenario **13B.3** Use solubility graphs to both quantitatively and qualitatively predict outcomes in a given scenario **13B.4** Define ‘saturated’, ‘unsaturated’ and ‘supersaturated’
- The use of precipitation reactions to remove impurities from water	**13C Purifying water** **13C.1** Define ‘precipitate’ **13C.2** List common water impurities **13C.3** Use solubility tables to identify appropriate precipitating agents **13C.4** Describe the experimental method for removing an impurity from water by precipitation **13C.5** Write full balanced chemical and ionic equations for precipitation reactions
- 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	**13D Volumetric analysis of acids and bases** **13D.1** Define ‘end point’ and ‘equivalence point’ **13D.2** Describe the use of an indicator in an acid–base titration **13D.3** Define ‘standard solution’ **13D.4** State the key characteristics of a primary standard solution **13D.5** Calculate the concentration of a primary standard solution **13D.6** Calculate the concentration of a diluted solution **13D.7** List the steps involved in an acid–base titration, including glassware involved **13D.8** Calculate the concentration of an unknown acid or base in a solution from given experimental data

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Glossary

Concordant
Dilution
Double displacement reaction
End point
Equivalence point
Filtration
Hygroscopicity
Precipitate
Precipitation reaction
Saturated solution
Solubility curve
Solubility table
Solute
Solution
Solvent
Supersaturated solution
Titration
Unsaturated solution
Volumetric analysis

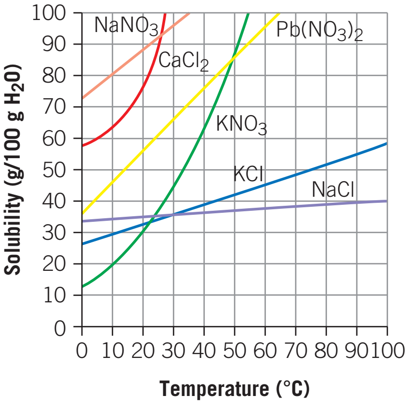
Concept map

Calculating the concentration of solutes in a solution using a variety of different units and converting between them

13A Measures of solubility

Using tables and graphs of experimental data to predict the relative solubility of different ionic compounds at various temperatures

13B Predicting solubility

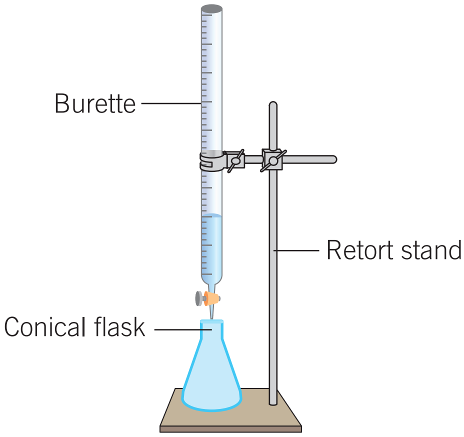


Using precipitation reactions to remove common impurities from water

13C Purifying water

Experimentally calculating the unknown concentration of an acid or base by titration

13D Volumetric analysis of acids and bases



See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

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13A Measures of solubility

Study Design:
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

Glossary:
Solute
Solution
Solvent

ENGAGE

The Dead Sea

The Dead Sea isn't the most appealing name for a tourist destination, but many people travel there every year to be able to experience the unique water. As you can see in the image, people don't sink in the Dead Sea, they float! How is this possible? It is because the water in the Dead Sea has an exceptionally high salt concentration. In fact, the salt concentration is around 34%, which is about 10 times higher than the normal salt concentration in sea water (approximately 3.5%). At this much higher concentration, the water in the Dead Sea is incredibly dense, even more dense than the human body. It is for this reason that people float in its water. However, the other consequence of such a high salt concentration is that not much can survive in it – which is why it is called the Dead Sea. In this section, you are going to learn more about solubility and concentration and how they can be measured and calculated.

Figure 13A–1 The Dead Sea is so salty and dense that people can float easily due to positive buoyancy.

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EXPLAIN

Concentration of solutions

Solution
a mixture of a solute dissolved in a solvent

Solute
the solid, liquid or gas that dissolves in a solvent

Solvent
the liquid that the solute dissolves in

VIDEO 13A–1 SOLUTES, SOLVENTS AND SOLUTIONS

11B STRONG AND WEAK ACIDS AND BASES LINK

A **solution** is created by dissolving a **solute** in a **solvent** as shown in Figure 13A–2.

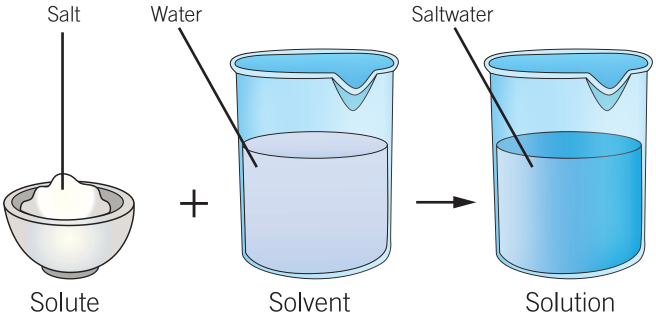

Figure 13A–2 When the salt (solute) is dissolved in the water (solvent), a solution is formed.

Solutions can be made with varying amounts of the solute added to the solvent, as demonstrated in Figure 13A–3. If only a small amount of solute is present in the solution, it may be referred to as a dilute solution. If a large amount of solute is present in the solution, it may be referred to as a concentrated solution.

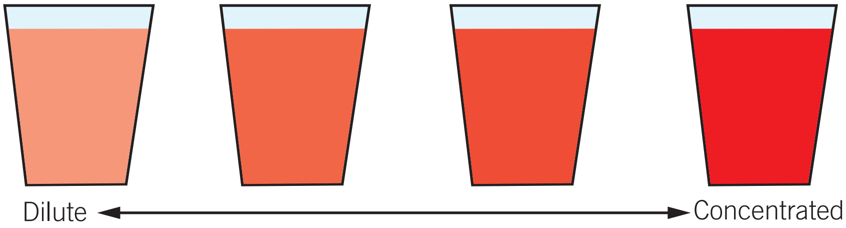

Figure 13A–3 In this diagram, the intensity of colour indicates the concentration of the solution.

To describe how concentrated a solution is, we can quantify the amount of the solute present in a certain volume of the solution. There are different units that can be used to describe the concentration of a solution.

Units of concentration

The decision regarding which units to choose to describe a solution may depend on a number of different factors, including how concentrated the solution is or whether or not you are comparing it to another solution where the concentration has been provided in given units. A description of some common concentration units is listed below in bold.

- **Moles per litre, mol L^{-1}**: The amount in moles of the solute per one litre of solution. The concentration in mol L^{-1} is given by the formula

$$c = \frac{n}{V}$$

where:

c = The concentration of the solution (in mol L^{-1}, or M)
n = The amount in moles of dissolved solutes (in mol)
V = The volume of solution (in litres)

This can be calculated if the mass of the solute in a volume is given, by using the relationship $n = m/M$, as in the formula on the following page.

NOTE

The concentration in mol L^{-1} describes the molarity (M) of the solution. A solution that has a concentration of 1 mol L^{-1} could be described as a 1 molar solution and be written as 1 M.

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$$c = \frac{m}{M}$$

where:

c = The concentration (in mol L^{-1}, or M)
m = The mass of dissolved solute (in grams)
V = The volume (in litres)
M = The molar mass (in g mol^{-1})

Worked example 13A–1: Calculating concentration from mass of solute

VIDEO
WORKED EXAMPLE 13A–1

A solution is made by dissolving 75.2 g of sodium carbonate, Na_2CO_3, in 500 mL of water. Calculate the concentration of the solution in mol L^{-1}.

Solution

Step 1	Extract the information from the question.	$m\,(Na_2CO_3) = 75.2$ g $V\,(Na_2CO_3) = 500$ mL
Step 2	Determine the number of moles of Na_2CO_3 using $n = \frac{m}{M}$. Recall, molar mass is calculated using the relative atomic masses of elements found on the periodic table.	$n(Na_2CO_3) = \frac{m}{M}$ $= \frac{75.2}{106.0}$ $= 0.709$ mol
Step 3	Convert the volume to litres.	$V\,(Na_2CO_3) = 500$ mL $= 0.500$ L
Step 4	Substitute into the equation $c = \frac{n}{V}$ and solve for $c(Na_2CO_3)$.	$c(Na_2CO_3) = \frac{n}{V}$ $= \frac{0.709}{0.500}$ $= 1.42$ mol L^{-1}

Worked example 13A–2: Calculating concentration from amount of solute

VIDEO
WORKED EXAMPLE 13A–2

A solution contains 2.4 mol of potassium chloride, KCl, dissolved in 4.0 L of water. Calculate the concentration of the potassium chloride solution in mol L^{-1}.

Solution

Step 1	Extract the information from the question.	$n\,(KCl) = 2.4$ mol $V\,(KCl) = 4.0$ L
Step 2	Substitute the values into the equation $c = \frac{n}{V}$ and solve for $c(KCl)$.	$c(KCl) = \frac{n}{V}$ $= \frac{2.4}{4.0}$ $= 0.60$ mol L^{-1}

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- **Grams per litre, g L^{-1}**: The number of grams of the solute per 1 litre of solution. The concentration in gL^{-1} is given by the formula

$$c = \frac{m}{V}$$

where:

c = The concentration of solution (in g L^{-1})
m = The mass of dissolved solute (in grams)
V = The volume of solution (in litres)

Worked example 13A–3: Calculating concentration in grams per litre

VIDEO WORKED EXAMPLE 13A–3

A 36.7 g sample of lithium hydroxide, LiOH, is dissolved in 2.35 L of water. Calculate the concentration of the lithium hydroxide solution in g L^{-1}.

Solution

Step 1	Extract the information from the question.	$m(\text{LiOH}) = 36.7$ g $V(\text{LiOH}) = 2.35$ L
Step 2	Determine the concentration in g L^{-1} by dividing the mass (in grams) by the volume (in litres).	$c(\text{LiOH}) = \frac{m}{V}$ $= \frac{36.7}{2.35}$ $= 15.6$ g L^{-1}

- **Percentage mass per volume, %(m/v)**: The mass of the solute in grams that is present in 100 millilitres of solution. It is a suitable unit for describing a solution that is a solid or gas dissolved in a liquid. It can be calculated by

$$\%\,(\text{m/v}) = \frac{\text{volume of solute (g)}}{\text{volume of solution (mL)}} \times 100$$

Worked example 13A–4: Calculating concentration in percentage mass per volume

VIDEO WORKED EXAMPLE 13A–4

A 300 mL solution contains 95.0 g of sodium phosphate, Na_3PO_4. Calculate the concentration of the sodium phosphate solution in %(m/v).

Solution

Step 1	Extract the information from the question.	$m(Na_3PO_4) = 95.0$ g $V(Na_3PO_4) = 300$ mL
Step 2	Substitute values into the equation $c = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100$ and solve for %(m/v)	$c(Na_3PO_4) = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100$ $= \frac{95.0}{300} \times 100$ $= 31.7\%$ (m/v)

- **Percentage volume per volume, %(v/v)**: The volume of the solute in millilitres that is present in 100 millilitres of solution. It is a suitable unit for describing a solution that is made by dissolving a liquid in another liquid. It can be calculated by

$$\%\,(\text{v/v}) = \frac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}} \times 100$$

- **Parts per million, ppm**: The number of parts of the solute per million parts of the solution. For example, the number of grams of the solute per million grams of the solution or the number of micrograms (μg, a millionth of a gram) per gram of the solution.

NOTE

For water, one gram is equal to one millilitre. This means that if water is the solvent, the number of micrograms per millilitre of the solution is also a ppm unit. Parts per million is a suitable unit for solutions that are very dilute.

Worked example 13A–5: Calculating concentration in ppm

A 2.00 L sample taken from a river near an industrial site was found to contain 54.2 mg of mercury, Hg. Calculate the concentration of mercury in ppm.

Solution

Step 1	Extract the information from the question.	$m(\text{Hg}) = 54.2\ \text{mg}$ $V(\text{Hg}) = 2.00\ \text{L}$
Step 2	Recall that μg/mL is a ppm unit. Convert the units of m(Hg) to μg and the units of V(Hg) to mL.	$m(\text{Hg}) = 54.2\ \text{mg} = 5.42 \times 10^4\ \mu\text{g}$ $V(\text{Hg}) = 2.00\ \text{L} = 2.00 \times 10^3\ \text{mL}$
Step 3	Determine the concentration in μg/mL. This is numerically equivalent to the concentration in ppm.	$c(\text{Hg}) = \dfrac{5.42 \times 10^4}{2.00 \times 10^3}$ $= 27.1\ \mu\text{g mL}^{-1}$ $= 27.1\ \text{ppm}$

VIDEO
WORKED EXAMPLE 13A–5

Converting between units of concentration

There are times when it is useful to convert the concentration of a solution into different units. This can be required to compare the concentration of different solutions or to make the number more reasonable rather than a very large or very small value.

WORKSHEET 13A–1
UNIT CONVERSIONS

Worked example 13A–6: Converting units of concentration

A magnesium nitrate, $Mg(NO_3)_2$, solution has a concentration of 25.3 g L^{-1}. What is the concentration of the solution in mol L^{-1}?

Solution

Step 1	Extract the information from the question.	$c(\text{Mg(NO}_3)_2) = 25.3\ \text{g L}^{-1}$
Step 2	Given there is 25.3 g in a litre, determine the number of moles in a litre, using $n = \dfrac{m}{M}$.	$n(\text{Mg(NO}_3)_2) = \dfrac{m}{M}$ $= \dfrac{25.3}{148.3}$ $= 0.170\ \text{mol}$
Step 3	With the information about the number of moles in a litre, express the concentration in mol L^{-1}, using the formula $c = \dfrac{n}{V}$.	$c(\text{Mg(NO}_3)_2) = \dfrac{n}{V}$ $= \dfrac{0.170}{1}$ $= 0.170\ \text{mol L}^{-1}$

VIDEO
WORKED EXAMPLE 13A–6

Worked example 13A–7: Calculating concentration in percentage volume per volume

VIDEO
WORKED
EXAMPLE 13A–7

An alcohol-based hand sanitiser is reported to have an ethanol, C_2H_6O, concentration of 7.5×10^5 ppm. What is the concentration of ethanol in %(v/v)?

Solution

Step 1	Extract the information from the question.	$c(C_2H_6O) = 7.5 \times 10^5$ ppm
Step 2	You are converting to a v/v-based unit, so recall that µL/L is a ppm unit. Express the concentration in µL/L.	$c(C_2H_6O) = 7.5 \times 10^5$ µL/L This can also be written more clearly as a fraction to show the volume of solvent is 1 L: $c(C_2H_6O) = \dfrac{7.5 \times 10^5 \mu L}{1\ L}$
Step 3	The formula for %(v/v) requires that both the solute and solvent volumes are in units of mL. Convert both volumes to mL.	$c(C_2H_6O) = \dfrac{7.5 \times 10^5 \mu L}{1\ L}$ $= \dfrac{7.5 \times 10^2\ mL}{1000\ mL}$
Step 4	Determine the concentration in %(v/v) using the formula: $c = \dfrac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}} \times 100$	$c(C_2H_6O) = \dfrac{7.5 \times 10^2\ mL}{1000\ mL} \times 100$ $= 75\%$ (v/v)

Check-in questions – Set 1

1 Calculate the concentration of each of the following solutions.
  a 2.57 g of KNO_3 in 2.5 L of water, in mol L^{-1}
  b 8.2 g of H_2SO_4 in 1.2 L of water, in ppm
  c 80.1 g of HNO_3 in 17.1 L of water, in %(m/v)

2 Describe how concentrations of solutions measured in %(v/v) are calculated.

13A SKILLS

Final check of answers

It is important that with all questions, you thoroughly read over your answers to check your use of key terminology and also that you have explained everything in a chemically correct manner. For calculation-based questions, it is also important that you show your full working for the Worked examples in this section. Equally as crucial is ensuring that you have written your final answer, correct or incorrect, with the correct number of significant figures and that it includes the correct units as required from the question.

Let's look at Worked example 13A–4 from this section as an example. In this question, two figures were provided:

- 300.0 mL (which represents four significant figures)
- 95.0 g (which represents three significant figures).

Remember, your final answer needs to be given to the least number of significant figures from the values provided in the stem of the question. In this case it is three significant figures, so the answer was given as 31.7, as opposed to 31.67.

Lastly, the units required need to be correct. This question asked for the concentration to be given in %(m/v). Even though using the data provided in this question the units in the calculation could be g mL^{-1}, you need to recall that this is not the same as %(m/v). Therefore, be sure to use this at the end of your final answer.

Section 13A questions

1 A 50 g sample of glucose is dissolved in water. What is the solute, solvent and solution in this scenario?

2 A solution is made by dissolving 2.38 g of sodium nitrate, $NaNO_3$, in 428 mL of water. Calculate the concentration of sodium nitrate, in mol L^{-1}.

3 Determine the mass of potassium carbonate, K_2CO_3, required to make 250 mL of a 0.40 mol L^{-1} solution.

4 An unlabelled salt solution was made up to 500.0 mL with a concentration of 0.1000 mol L^{-1}. It is known that 4.765 g of the salt was used to prepare the solution.
  a Calculate the molar mass of the unknown salt.
  b The salt has a formula of $\mathbf{X}Cl_2$. Suggest an identity for element **X**.

5 Determine the concentration of a 1.25 mol L^{-1} HBr solution in g L^{-1}.

6 A 55.0 g sample of ammonium phosphate, $(NH_4)_3PO_4$, was dissolved in 835 mL of water. Determine the concentration of ammonium ions, NH_4^+, in this solution in g L^{-1}. (Hint: Write out the equation for the dissolution of ammonium phosphate first.)

7 A 5.00 mL groundwater sample is found to contain 96.5 ppm of chloride ions, Cl^-. Calculate the amount, in mol, of chloride ions in this sample.

8 A copper(II) sulfate, $CuSO_4$, solution has a concentration of 31.6%(m/v). Determine the concentration of copper(II) sulfate in ppm.

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13B Predicting solubility

Study Design:
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 gas in water

Glossary:
Saturated solution
Solubility curve
Solubility table
Supersaturated solution
Unsaturated solution

ENGAGE

Why does Milo® float?

Many people will argue about the best way to drink Milo® – is it better hot or cold? If you've tried both options for yourself, you've probably noticed one key difference between the two. When you add cold milk to the Milo® powder, a lot of it seems to just float on the top. However, if you add hot milk, you will notice that most of the powder 'disappears' into the milk. In this section, you will be exploring concepts, such as solubility and the effect of temperature, which help to explain these observations. You will then be applying those principles to experimental data to predict what you might observe in other situations.

Figure 13B–1 Why does Milo® mix more easily into hot milk but not cold?

EXPLAIN

Solubility graphs

4C PRECIPITATION REACTIONS LINK

In Chapter 4C, you learned that ionic compounds have a range of solubilities in water. You saw that this information could be represented in **solubility tables** and that the acronyms SNAPE, BICS and CHPS are useful to help you remember which ions are soluble and which are insoluble.

Solubility table
a reference table used to predict the solubility of ionic compounds

VIDEO 13B–1
UNDERSTANDING SOLUBILITY GRAPHS

The solubility of a solid in water varies with temperature. For most solids, as the temperature increases, their solubility in water also increases. The data for the solubility of ammonium chloride in water at different temperatures is shown in Figure 13B–2. The data are presented as the maximum mass of ammonium chloride that can be dissolved in 100 g of water at a range of different temperatures.

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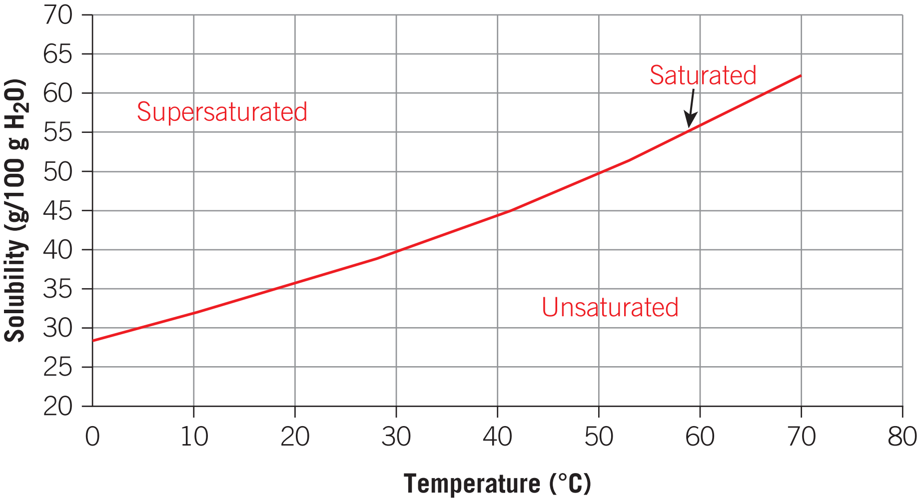


Figure 13B–2 The solubility of ammonium chloride in water at different temperatures. The solubility increases from 28 g/100 g H_2O at 0°C to 62 g/100 g H_2O at 70°C.

A solution of ammonium chloride that has a concentration represented by any point along the line would be called a **saturated solution**. A saturated solution contains the maximum mass of solute that can be dissolved in that volume of solvent at that specific temperature. A saturated solution of ammonium chloride at 30°C would contain 40 g/100 g H_2O.

Any point below the line represents a solution of ammonium chloride that is unsaturated. An **unsaturated solution** contains less than the maximum mass of solute that can be dissolved in that volume of solvent at that specific temperature. Therefore, if extra solute is added to an unsaturated solution, it can dissolve up to a point where the solution becomes saturated.

With care, it is possible to prepare a solution with a concentration higher than that of a saturated solution at a given temperature. These solutions are referred to as **supersaturated**, and they contain more dissolved solute than is usually possible at that specific temperature. Any point above the line on the **solubility curve** represents a supersaturated solution. A supersaturated solution is unstable. Disturbing the solution by shaking it or adding a small amount of extra solute will cause some of the solute to precipitate. The concentration of dissolved solute will return to that of a saturated solution, and any extra solute will be present as an undissolved solid.

Different solids will have different-shaped solubility curves. Examples of some of these are shown in Figure 13B–3.

All the substances on this solubility graph are ionic solids, and it is visible that their solubility increases with increasing temperature. Temperature affects the solubility of these solids to different degrees. As an example, the solubility of NaCl does not change significantly between 0°C and 100°C. On the other hand, the solubility of KNO_3 increases dramatically between 0°C and 100°C.

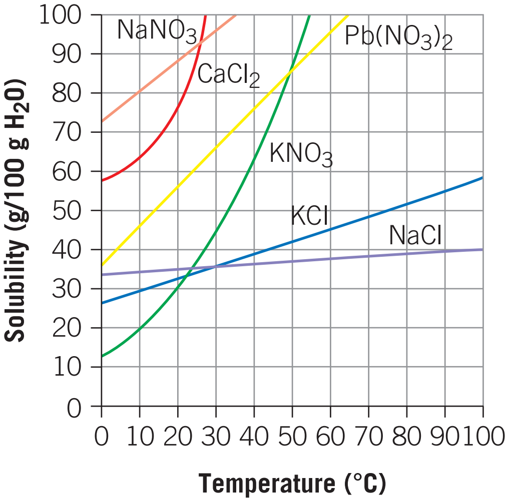


Figure 13B–3 Solubility curves for a variety of compounds between 0°C and 100°C

Saturated solution
a solution that cannot dissolve any more solute in the volume of solvent at a given temperature

Unsaturated solution
a solution that contains less solute than a saturated solution in the volume of solvent at a given temperature

Supersaturated solution
a solution that contains more solute than a saturated solution in a certain volume at a given temperature

Solubility curve
a graph of solubility versus temperature for a certain solute dissolved in a certain solvent

VIDEO
WORKED EXAMPLE 13B–1

Worked example 13B–1: Interpreting solubility graphs

Using Figure 13B–3, what mass of potassium chloride, KCl, should be added to 100 g of water at 60°C to create a saturated solution?

Solution

Step 1	Extract the information from the question.	Locate the solubility curve for KCl.
Step 2	Read from the graph the solubility (*y*-axis value) when the temperature (*x*-axis value) is 60°C.	Solubility $= \frac{46\text{ g}}{100\text{ g }H_2O}$
Step 3	Use the solubility to determine the mass of potassium chloride in 100 g of H_2O.	$m(\text{KCl}) = 46\text{ g}$

VIDEO
WORKED EXAMPLE 13B–2

Worked example 13B–2: Determining saturation using solubility graphs

A solution of calcium chloride ($CaCl_2$) is known to contain 40 g of $CaCl_2$ in 50 g of H_2O at 15°C. Use Figure 13B–3 to determine if this solution is unsaturated, saturated or supersaturated.

Solution

Step 1	To read data from the graph, the concentration needs to be expressed in units of g / 100 g of H_2O. The information about the $CaCl_2$ solution is given in units of g / 50 g of H_2O, so needs to be multiplied by 2.	$c(CaCl_2) = \frac{40\text{ g}}{50\text{ g }H_2O} \times 2$ $= 80\text{ g}/100\text{ g }H_2O$
Step 2	Use the graph to locate the position of this solution, 80 g / 100 g H_2O at 15°C, relative to the solubility curve for $CaCl_2$.	This solution is located at a point above the solubility curve for $CaCl_2$.
Step 3	Recall that solutions represented by points above the solubility curve are supersaturated.	This solution is supersaturated.

Check-in questions – Set 1

1 Distinguish between the terms ‘saturated’, ‘unsaturated’ and ‘supersaturated’.
2 What mass of NaCl would dissolve in 20.0 g of H_2O at 20°C to make a saturated solution? Use the solubility curve in Figure 13B–3 to assist your calculation.

WORKSHEET 13B–1
SOLUBILITY DATA ANALYSIS

Effect of temperature on the solubility of solids, liquids and gases

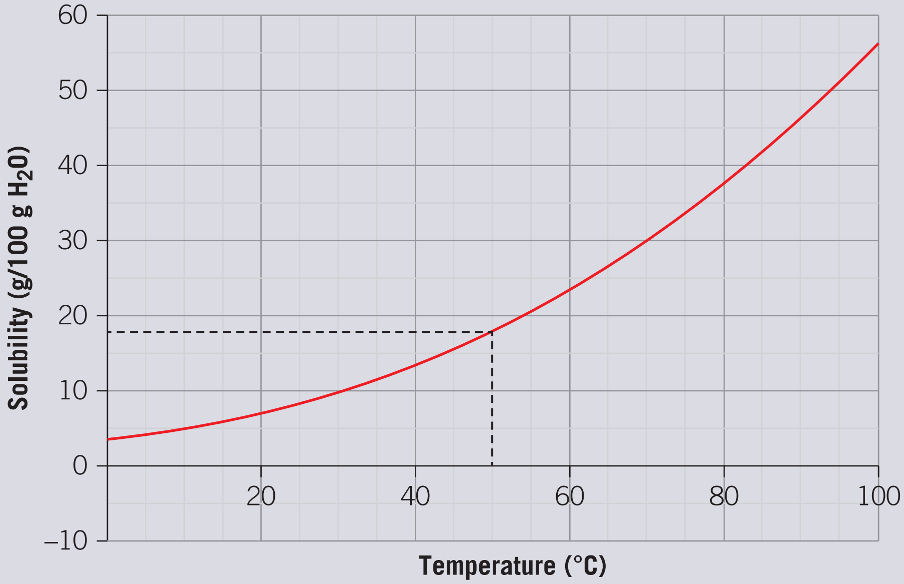
We have seen solubility curves for a number of solids dissolved in water, and they follow the same trend of increasing solubility with increasing temperature. The trend seen for liquids dissolving in water is the same. With increasing temperature, the solubility of liquids in water increases. However, the same is not true of gases. For most gases, as the temperature increases, the solubility of the gas in water decreases.

13B SKILLS

Tips for reading data from graphs

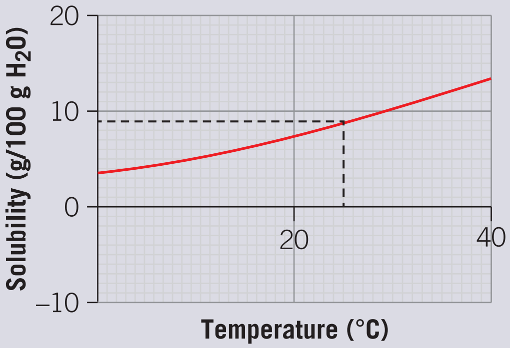
Questions that ask you to interpret data from graphs are not always going to be kind and seek a value that is perfectly on the curve coinciding with a grid line. Thus, there is going to be some margin of uncertainty in the value you select on the graph. This is dependent on the scale of the grid line axes.

Let's look at the example graph below, which shows the solubility curve of potassium chlorate.



Here the grid lines on both the *x*- and *y*-axes are five units of measurement apart. So, if a question asks you to determine the saturation point of compound at 50°C, you rule a dashed line up from the 50°C mark on the *x*-axis until you reach the curve. From there, you rule another dashed line across to the left until you reach the *y*-axis. As you can see on the graph above, this does not read perfectly to one of the grid lines. Therefore, you need to read to the nearest grid line, which is '20' and then estimate from there. You can see from this it is approximately halfway between '15' and '20'. So, you would estimate the next place value here as being either 17 or 18 g/100 g water.

However, you may be given a graph like the one below.



The gridlines on both the *x*- and *y*-axes are now one unit of measurement apart. So, if a question asked you to determine the saturation point of potassium chlorate at 25°C, you would follow the same procedure as outlined above. Again, you need to read to the nearest grid line, which is ‘10’, and then estimate from there. You can see from this that it is approximately ‘9’. So, you would estimate the next place value here as being approximately 9 g/100 g water.

Of course, with reading any graph like this, there is some imprecision or uncertainty. The degree of this uncertainty depends on the scale of your gridlines, and you should be given some margin for this uncertainty in your final answer.

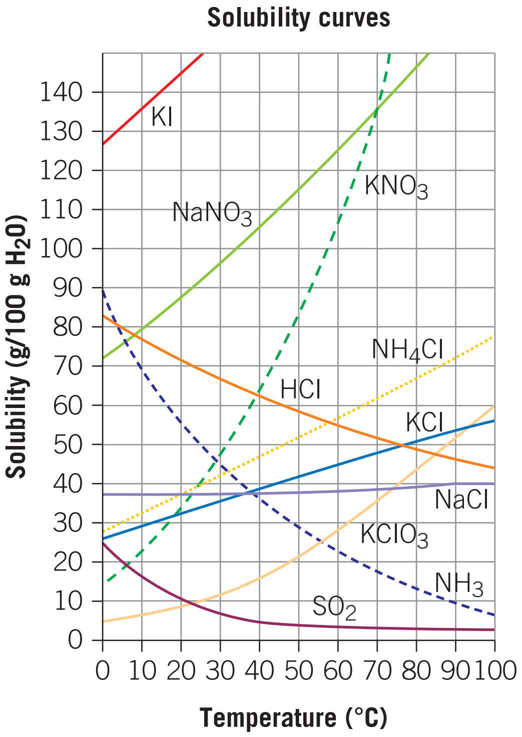
Section 13B questions

1 State the general relationship between temperature and the solubility of solids and gases in water. Refer to the solubility curves shown here in your answer.

2 Refer to the solubility curves shown here to answer the following questions.
   a State the solubility of ammonium chloride (NH_4Cl) at 65°C.
   b Rank the following salts in order of increasing solubility in water at 30°C:

   KCl, NH_4Cl, $KClO_3$, KNO_3, NaCl.
   c At 20°C, 20 g of SO_2 was dissolved in 100 g of H_2O. Describe this solution with reference to the saturation of SO_2.

3 A scientist measured the solubility of $MgSO_4$ in 100 g of water at different temperatures. The data was recorded as shown below in the table.



Temperature (°C)	Solubility (g) per 100 g of water
0	31
10	39
20	43
30	50
40	56
50	62
60	70
70	82
80	91
90	110
100	123

a Construct a plot for this data including a line of best fit.
b Predict the solubility of $MgSO_4$ at 15°C.
c Label this diagram with the terms ‘saturated’, ‘unsaturated’ and ‘supersaturated’.
d Explain why you have placed the three labels in their positions on the graph.

13C Purifying water

Study Design:
The use of precipitation reactions to remove impurities from water

Glossary:
Double displacement reaction
Filtration
Precipitate
Precipitation reaction

ENGAGE

Melbourne water – great from the tap!

Thomson Dam is the largest reservoir supplying a major city in Victoria and the third largest in Australia (Figure 13C–1). It is responsible for providing water throughout the Melbourne metropolitan region. When full, it holds 1068 billion litres of water! To put the size of Thomson Dam into perspective, that is the equivalent of filling 628 Melbourne Cricket Grounds (MCG) or 427 000 Olympic-sized swimming pools! Considering the importance of Thomson Dam for supplying water to such a large number of people in Victoria, you can imagine how important it is that it is kept clean. In this section, we are going to discuss situations where water may contain impurities and discuss what reactions can be undertaken to help solve this issue.

Figure 13C–1 Thomson Dam supplies most of metropolitan Melbourne with clean water.

EXPLAIN

Impurities in water

Water can contain several different types of impurities that may make it unsuitable for drinking. Some of these are listed in Table 13C–1 on the following page.

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Table 13C–1 Different examples of impurities that may be present in water

Type of impurity	Description	Example
Biological	Living microscopic organisms, such as bacteria and/or protozoa	*Giardia*, which causes gastrointestinal problems 
Colloidal	Small insoluble particles found throughout the water	Organic waste products or iron 
Dissolved	Common but in high concentrations can become dangerous for consumption Levels constantly measured to ensure they remain at safe levels	Gases, such as ammonia and carbon dioxide Salts of metals, such as calcium, magnesium, potassium and sodium 

The National Health and Medical Research Council (NHMRC) has published water drinking guidelines, which outline the limits for some metal impurities (Table 13C–2).

Table 13C–2 Common water impurities and the concentrations at which they become unsafe

Impurity	Unsafe level (mg L^{-1})
Calcium carbonate (water hardness)	200
Sodium	180
Fluoride	1.5
Iron	0.3
Manganese	0.1

What can be done when these impurities reach concentration levels that are too high and start to affect the appearance, taste or safety of drinking water? The answer is that they can be removed through precipitation reactions. The general principles underpinning precipitation reactions and the writing of ionic equations can be revisited in Chapter 4.

LINK **4A** FORMATION AND NAMING OF IONIC COMPOUNDS

LINK **4C** PRECIPITATION REACTIONS

Soluble (aqueous) ionic compounds are able to react in solution to form new ionic compounds, in which cations and anions exchange their respective ionic counterparts. For example, solutions of sodium chloride and silver nitrate can be mixed to form a mixture of solid silver chloride and a solution of sodium nitrate. In this case, the Na^+ and Ag^+ cations have swapped their Cl^- and NO_3^- anion counterparts:

$$\text{sodium chloride} + \text{silver nitrate} \rightarrow \text{sodium nitrate} + \text{silver chloride}$$

$$NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$$

$$NaCl(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)$$

This type of reaction is typically known as a **double displacement reaction**, as two sets of ions have been displaced. The general form is shown in Figure 13C–2.

A double displacement reaction that produces a solid from two aqueous solution reactants is called a **precipitation reaction**. When writing chemical equations for precipitation reactions, knowledge of the solubility rules is important (Table 13C–3). Thus, all resulting products should be considered as to whether they would be soluble or insoluble in water to determine a possible **precipitate** and the appropriate state symbols.

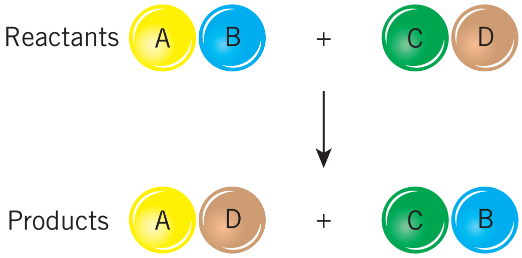

Figure 13C–2 Pictorial representation of a double displacement reaction

Double displacement reaction a chemical reaction in which ionic compounds swap their cations and anions

Precipitation reaction a chemical reaction in which two or more solutions of soluble salts are mixed and one of the products formed is an insoluble solid – known as a **precipitate**

Precipitate the solid formed in a reaction in which two or more solutions of soluble salts are mixed

Table 13C–3 Solubilities of common ionic compounds and their exceptions

Soluble in water	Exceptions – insoluble	Exceptions – slightly insoluble
All nitrates (NO_3^-)	No exceptions	No exceptions
All ammonium salts (NH_4^+)	No exceptions	No exceptions
All sodium (Na^+) and potassium (K^+) salts	No exceptions	No exceptions
All ethanoates (CH_3COO^-)	No exceptions	No exceptions
Most chlorides (Cl^-), bromides (Br^-) and iodides (I^-)	AgCl, AgBr, AgI, PbI_2	$PbCl_2$, $PbBr_2$
Most sulfates (SO_4^{2-})	$BaSO_4$, $PbSO_4$, $SrSO_4$	Ag_2SO_4, $CaSO_4$
Insoluble in water	**Exceptions – soluble**	**Exceptions – slightly soluble**
Most carbonates (CO_3^{2-})	Na_2CO_3, K_2CO_3, $(NH_4)_2CO_3$	No exceptions
Most sulfides (S^{2-})	Na_2S, K_2S, $(NH_4)_2S$	No exceptions
Most phosphates (PO_4^{3-})	Na_3PO_4, K_3PO_4, $(NH_4)_3PO_4$	No exceptions
Most hydroxides (OH^-)	NaOH, KOH, $Ba(OH)_2$, NH_4OH	$Ca(OH)_2$, $Sr(OH)_2$

Check-in questions – Set 1

1 List three different examples of impurities that may be present in water.
2 Explain what a double displacement reaction is, with reference to precipitation reactions.
3 What is the mnemonic used to recall the names of the ions that are always soluble?

4C PRECIPITATION REACTIONS LINK

VIDEO 13C–1 REMOVING IMPURITIES FROM WATER

Purifying water using precipitation reactions

When removing soluble metal ions and phosphates from water, the same overall process is followed, as shown in Figure 13C–3. Firstly, the precipitation reagent is added to the water to enable it to react with the dissolved impurity. This reaction will form an insoluble solid product, which is known as a precipitate. The water sample is then left for a while, as the solid product settles to the bottom of the container. The precipitate is then removed from the water through **filtration**.

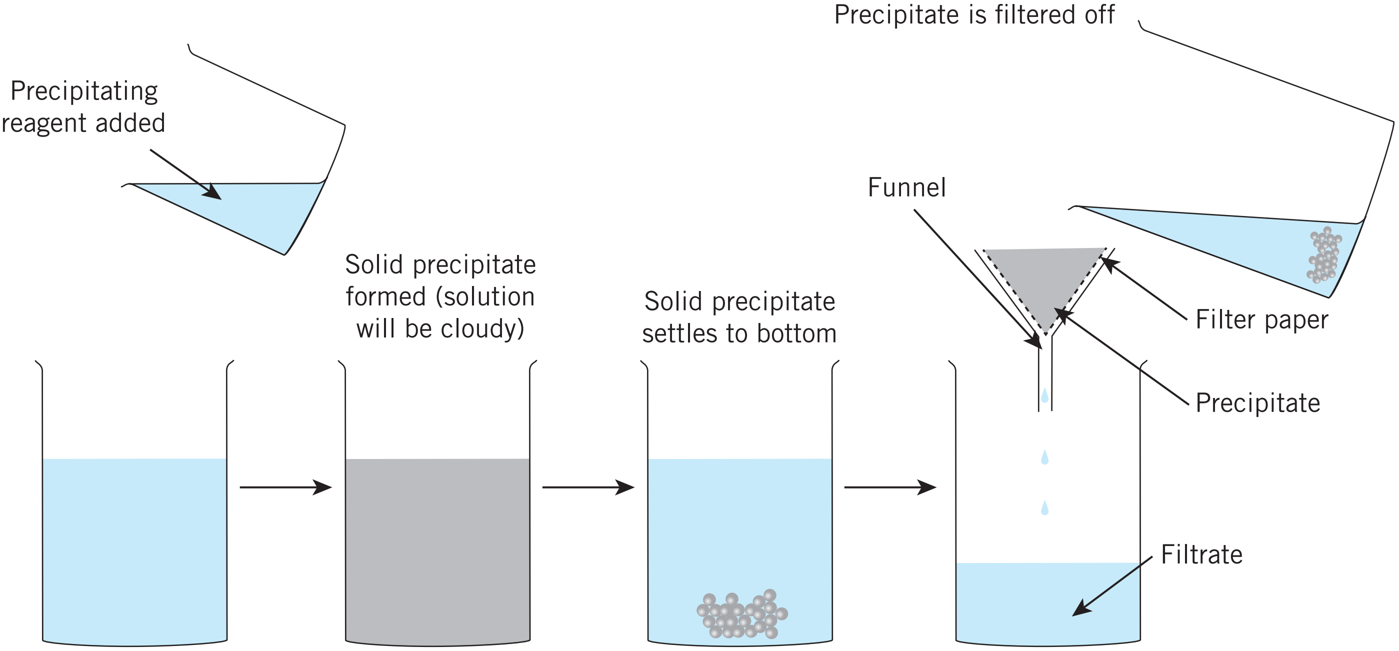

Figure 13C–3 Precipitation reactions are a valuable method for removing impurities from water.

Filtration the process by which a solid precipitate is removed from a solution by passing it through a medium that only allows fluid to pass through, such as filter paper

Removal of calcium and magnesium ions

Dissolved salts of both calcium and magnesium lead to water hardening. While these don't tend to have any negative impacts on health, they can lead to deposits throughout the water supply chain that can cause a variety of issues. Therefore, it is advantageous to remove some of these ions in a process called water softening. To do this, precipitation reactions are performed in two steps.

In the first step, the magnesium ions (Mg^{2+}) are removed through a reaction with calcium hydroxide, ($Ca(OH)_2$, also known as lime. The net ionic equation for this reaction is

$$Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$$

The calcium ions in the lime act as spectator ions in this instance. However, this means that we now have the original Ca^{2+} ions that were present in the water, as well as the additional calcium ions that are now present from the lime, that need to be removed. In the second step, this removal is achieved through reacting with sodium carbonate, Na_2CO_3 (sometimes called soda ash). The net ionic equation for this reaction is

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$$

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Removal of phosphates

LINK **1D** CRITICAL ELEMENTS AND RECYCLING PROCESSES

One of the biggest problems that can affect the suitability of drinking water is the uncontrolled growth of plant material, such as algae. When this occurs to a large extent, it is known as an algal bloom, as shown in Figure 13C–4. Algal blooms can potentially have harmful consequences for marine life, make consumption dangerous and halt any fishing industries that may be working in those waterways. Phosphates have been identified as a key impurity that results in an aquatic environment that is suited for algal growth. Thus, ensuring that the concentration of phosphates remains low is important, and so precipitation reactions have been designed to help remove them from wastewater.

Figure 13C–4 A lake that has been affected by algal bloom

Commonly, removal of phosphates is done with precipitating reagents containing either iron or aluminium. The net ionic equations for reactions utilising Fe^{2+}, Fe^{3+} and Al^{3+} are

$$3Fe^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Fe_3(PO_4)_2(s)$$

$$Fe^{3+}(aq) + PO_4^{3-}(aq) \rightarrow FePO_4(s)$$

$$Al^{3+}(aq) + PO_4^{3-}(aq) \rightarrow AlPO_4(s)$$

As you can see, all of these reactions lead to the formation of an insoluble, solid product that, after settling, can be removed from the water by filtration (Figure 13C–5).

Figure 13C–5 The reaction of phosphate ions with iron(II) ions leads to the production of a solid precipitate, which can then be easily removed from the water.

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Check-in questions – Set 2

WORKSHEET 13C–1 PURIFYING WATER USING PRECIPITATION REACTIONS

1 Recall which ions should be present in reagents to help with the removal of phosphates from water.

2 Recall which ions should be present in reagents to help with the removal of calcium ions from water.

13C SKILLS

Using precipitation reactions to remove impurities from water

When removing impurities from water, it is important to be able to calculate the appropriate amount of a precipitating agent to add to the water. To do this effectively, you also need to know the expected concentration range of that impurity.

For example, as was discussed in this section, a solution of sodium carbonate can be added to remove calcium ions from water. The combination of hydrated calcium ions and hydrated carbonate ions results in a solid precipitate of calcium carbonate. If the estimated concentration range of calcium ions present in a water sample was between 5 and 10 mg L^{-1}, there are two approaches you could take.

13D VOLUMETRIC ANALYSIS OF ACIDS AND BASES LINK

14B GRAVIMETRIC ANALYSIS OF SALTS LINK

The first would be to plan for the highest concentration in this range to be present and add a bit more of the sodium carbonate precipitating agent, doubling the amount expected, for example. This can be determined using stoichiometry, which you will explore in Chapters 13D and 14B, and is shown below.

Ionic equation:

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$$

Calculation:

Concentration of calcium ions = 10 mg L^{-1} (0.010 g L^{-1})

13A MEASURES OF SOLUBILITY LINK

Using concentration conversions from Chapter 13A, this is equivalent to 0.010 g L^{-1} divided by the molar mass of Ca^{2+} ions, which is 40.1 g mol^{-1}.

Concentration = 2.49×10^{-4} mol L^{-1}

As this ionic equation shows a mole ratio of 1 : 1 between the calcium and carbonate ions, the concentration required of the carbonate ions (and hence the sodium carbonate) being added is the same. So if we were to double this amount, we would add a sodium carbonate solution with a concentration of 4.99×10^{-4} mol L^{-1}.

The disadvantage of this method is that adding too large of an excess of precipitating agent results in a volume increase. This can be practically difficult. It also means that an excess of carbonate ions may remain in the water. So, while attempting to remove an impurity, more impurities have been added in the process.

13D VOLUMETRIC ANALYSIS OF ACIDS AND BASES LINK

The second would be planning for the lowest concentration in this range, or a concentration within this range. As a result, not enough of the precipitating agent is added to remove all the calcium ion impurities from the water. You then need to perform other tests to determine this amount remaining and potentially add more precipitating agent. These tests include options such as using a photometer or titration. The latter is explored in Chapter 13D but only in the context of acids and bases.

Section 13C questions

1 Define the term ‘precipitate’ in chemistry.

2 Write balanced equations showing full symbol formulas and states for the following precipitation reactions.
   - a A solution of silver nitrate is added to a solution of sodium chloride.
   - b A solution of barium hydroxide is added to a solution of potassium sulfate.
   - c A solution of aluminium ethanoate is added to a solution of ammonium phosphate.
   - d A solution of nickel(II) sulfate is added to a solution of sodium carbonate.

3 Write balanced net ionic equations, with states, for the following precipitation reactions.
   - a $AlCl_3(aq) + K_2CO_3(aq)$
   - b $NH_4Br(aq) + AgCH_3COO(aq)$
   - c A solution of sodium sulfide is added to a solution of vanadium(V) iodide
   - d A solution of copper(II) nitrate is added to a solution of potassium hydroxide.

4 A sample of wastewater was collected containing dissolved phosphorus in the form of phosphate ions.
   - a Refer to the solubility table Table 13C–3 to identify a precipitating agent that could be used to remove phosphate from this wastewater sample.
   - b Write an experimental method for removing dissolved phosphate from a 100 mL wastewater sample using a saturated iron(III) nitrate solution.
   - c Write a net ionic equation, with states, for the precipitation reaction that takes place in part **b**.
   - d Write a net ionic equation, with states, for the precipitation reaction that would take place if iron(II) sulfate was used in place of iron(III) nitrate.
   - e Write a net ionic equation, with states, for the precipitation reaction that would take place if aqueous aluminium ethanoate was used in place of iron(III) nitrate.

5 Chloride pollution in water is common and purification requires removal of chloride ions via methods such as precipitation and filtration. A source of chloride in natural waterways might be the erosion of rocks and minerals, such as those containing sodium chloride.
   - a Describe, using an equation, how solid sodium chloride salt can form ions when dissolved in water.
   - b A sample of chloride-containing water is collected. State a precipitating agent that could be added to the sample to remove chloride ions.
   - c Write an experimental method for removing dissolved chloride from this sample using the precipitating agent identified in part **b**.
   - d Write a net ionic equation, with states, for the precipitation reaction that takes place in part **b**.

13D Volumetric analysis of acids and bases

Study Design:
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

Glossary:
- Concordant
- Dilution
- End point
- Equivalence point
- Hygroscopicity
- Titration
- Volumetric analysis

ENGAGE

The Ribena experiment

In 2007, two high school students in New Zealand conducted an experiment to measure the vitamin C content in a popular blackcurrant drink called Ribena®. Ribena® marketed itself as having four times as much vitamin C as orange juice. However, when the students performed their experiment, they found that they could barely detect any vitamin C at all. After the results were confirmed, the company that made Ribena were fined a considerable amount of money for having falsely advertised their product. The experiment that the students had done to test for vitamin C was a titration, also known as volumetric analysis. This technique is the focus of this section.

Figure 13D–1 Volumetric analysis performed by New Zealand high school students Jenny Suo (left) and Anna Devathasan (right) in 2007 showed that Ribena® contained almost no vitamin C, despite the marketing claims of the company. The company was fined NZ$3 million for misleading advertising.

EXPLAIN

Volumetric analysis

Being able to accurately calculate the concentration of an acid or a base is very important in many different industries. For food manufacturers, it is used to work out the levels of citric acid in orange juice, phosphoric acid in cola products or acetic acid in vinegar. It is also useful in the development of pharmaceuticals (such as the measurement of acetylsalicylic acid in aspirin) and in making cosmetics, where shampoo, hair dyes and skin creams all contain various acids and bases. The technique that is commonly used for this purpose is called **volumetric analysis**.

VIDEO 13D–1 PERFORMING VOLUMETRIC ANALYSIS

Volumetric analysis analysis of solutions using measurements of volumes

NOTE

In this section, the focus of volumetric analysis is purely on acid–base neutralisation reactions. However, there are many other instances where this process is used in industry.

Volumetric analysis takes advantage of the neutralisation reactions that occur between acids and bases. As a reminder, the general equation for these reactions is

$$\text{acid} + \text{base} \rightarrow \text{salt} + \text{water}$$

As an example, we can look at the equation for the reaction of hydrochloric acid with sodium hydroxide.

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

LINK **11E** NEUTRALISATION REACTIONS

The most common experiment used to perform volumetric analysis is known as a **titration**. The general set-up of a titration experiment can be seen in Figure 13D–2.

As can be seen, titrations involve the use of a burette (held up by a retort stand) and a conical flask. The burette and the conical flask will contain separate solutions, one of which will be an acid and the other will be a base. Importantly, the concentration of one of these solutions will be unknown and the other will be known. The solution with a known concentration is known as the primary standard solution.

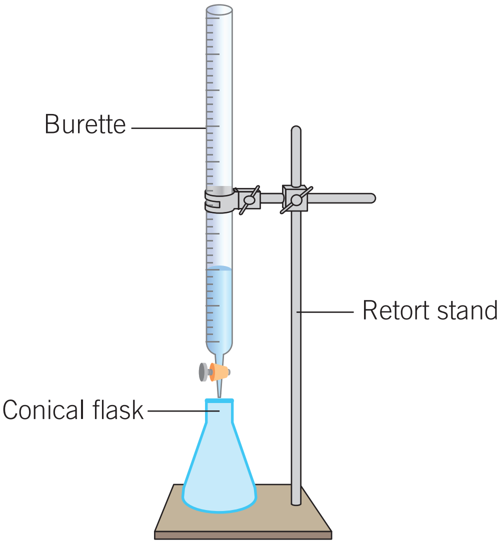


Figure 13D–2 The general set-up of a titration experiment

Titration a type of volumetric analysis whereby a measured volume of a standard solution is added to a measured volume of a solution whose concentration is unknown

Preparation of standard solutions

A critical part of any titration experiment is the preparation of the standard solution. To be able to accurately calculate the concentration of the unknown acid or base, the standard solution needs to be made correctly. For a compound to be considered suitable to be used as a primary standard, it needs to fulfil the following criteria:

- be available with a high level of purity
- have a high stability/low level of reactivity
- have low **hygroscopicity**.

To prepare a primary standard for a titration, a precise mass of the compound needs to be weighed out on an electronic balance and made up to a precise volume in a volumetric flask. Mole equations are used to determine the known concentration, or the mass required

Hygroscopicity the ability to attract and hold water from the environment to be measured, of this primary standard, as shown in Worked example 13D–1. Primary standards can be used to standardise other solutions. These other solutions are often referred to as secondary standards and can also be used in titrations.

Worked example 13D–1: Calculating the mass of solute from concentration

VIDEO WORKED EXAMPLE 13D–1

You require 200 mL of a 0.100 mol L^{-1} solution of sodium carbonate (Na_2CO_3) to use as a primary standard. Calculate the mass of sodium carbonate that must be weighed out.

Solution

Step 1	Use $n = c \times V$ to calculate the number of moles of sodium carbonate required.	$n(Na_2CO_3) = c \times V$ $= 0.100\ \text{mol L}^{-1} \times 0.200\ \text{L}$ $= 0.0200\ \text{mol}$
Step 2	Calculate the molar mass of sodium carbonate (Na_2CO_3).	$M(Na_2CO_3) = [(23 \times 2) + 12 + (16 \times 3)]$ $= 106.0\ \text{g mol}^{-1}$
Step 3	Use $m = n \times M$ to calculate the mass of sodium carbonate required.	$m(Na_2CO_3) = n \times M$ $= 0.0200\ \text{mol} \times 106.0\ \text{g mol}^{-1}$ $= 2.12\ \text{g}$

Performing dilutions

In some instances, the solution you are analysing may have an estimated concentration range that is too high to achieve an accurate measurement of its concentration by titration. Generally, this can lead to one of two problems, outlined below.

- If the solution you are analysing is in the burette and has a high concentration, a very small volume will be required for the titration. Using small volumes for calculations can lead to higher amounts of error.
- If the solution you are analysing is in the conical flask and has a high concentration, a large volume of your standard will need to be added. This may be in excess of the 50 mL volume that is able to be held in the burette, which will mean that you will need to refill it during the experiment. While technically this is possible, it is not ideal and should be avoided.

Dilution the process of adding a solvent to a solution to reduce its concentration

To lower the concentration, a **dilution** can be performed. A dilution involves reducing the concentration of the solute by adding more of the solvent. The amount by which the substance is diluted is known as the dilution factor. A dilution can be explained and expressed in a number of different ways. The following points all describe exactly the same scenario.

- A 10 mL sample was diluted with 90 mL of water.
- A 1/10 dilution was performed on a 10 mL sample.
- A 10 mL sample was diluted 1 : 9.
- A 10 mL sample was diluted 10-fold.
- The concentration of a 10 mL sample was reduced by a factor of 10.

11C CALCULATING pH LINK

To calculate the resulting concentration after a dilution, or the volume of samples required to achieve appropriate concentrations, the following formula can be used:

$$c_1 \times V_1 = c_2 \times V_2$$

where:

c_1 = The concentration before dilution (initial concentration)
V_1 = The volume before dilution (initial volume)
c_2 = The concentration after dilution (final concentration)
V_2 = The volume after dilution (final volume)

Worked example 13D–2: Calculating the volume of diluted solutions

VIDEO WORKED EXAMPLE 13D–2

You require 500 mL of a 0.0500 mol L^{-1} solution of hydrochloric acid. What volume of a 0.200 mol L^{-1} solution and water would be required to prepare this dilution?

Solution

Step 1	Extract the information about c_1, V_1, c_2 and V_2 from the question.	$c_1 = 0.200\ \text{mol L}^{-1}$ $V_1 = ?$ $c_2 = 0.0500\ \text{mol L}^{-1}$ $V_2 = 500\ \text{mL}$
Step 2	Rearrange $c_1 \times V_1 = c_2 \times V_2$ to make V_1 the subject.	$c_1 \times V_1 = c_2 \times V_2$ $V_1 = \frac{c_2 \times V_2}{c_1}$
Step 3	Substitute in the values to solve for V_1.	$V_1 = \frac{c_2 \times V_2}{c_1}$ $= \frac{0.0500\ \text{mol L}^{-1} \times 500\ \text{mL}}{0.200\ \text{mol L}^{-1}}$ $= 125\ \text{mL}$
Step 4	Determine the volume of water required to make the diluted solution.	$V(\text{water}) = V_2 - V_1$ $V(\text{water}) = 500\ \text{mL} - 125\ \text{mL}$ $V(\text{water}) = 375\ \text{mL}$

Worked example 13D–3: Calculating the concentration of diluted solutions

VIDEO WORKED EXAMPLE 13D–3

A 500 mL sample of a 0.10 mol L^{-1} solution of sodium hydroxide is added to 1.5 L of water. What is the concentration of the resulting solution?

Solution

Step 1	Extract the information about c_1, V_1, c_2 and V_2 from the question. Remember that V_2 is the total volume of the diluted solution, not just the volume of water added.	$c_1 = 0.10\ \text{mol L}^{-1}$ $V_1 = 500\ \text{mL}$ $c_2 = ?$ $V_2 = 2.0\ \text{L}$
Step 2	Convert volume units to be identical.	$V_1 = 500\ \text{mL} = 0.500\ \text{L}$ $V_2 = 2.0\ \text{L}$
Step 3	Rearrange $c_1 \times V_1 = c_2 \times V_2$ to make c_2 the subject.	$c_1 \times V_1 = c_2 \times V_2$ $c_2 = \frac{c_1 \times V_1}{V_2}$
Step 4	Substitute in the values to solve for c_2.	$c_2 = \frac{c_1 \times V_1}{V_2}$ $= \frac{0.10\ \text{mol L}^{-1} \times 0.500\ \text{L}}{2.0\ \text{L}}$ $= 0.025\ \text{mol L}^{-1}$

Check-in questions – Set 1

1 Outline what is meant by 'a 20 mL sample was diluted 10-fold'. Then, give the volume of water that would be added to this 20 mL sample and the total final volume of the solution.

2 Determine the concentration of the final solutions prepared by diluting 20.00 mL of a 5.00 M cobalt(II) sulfate solution to 0.5 L, and then 10.00 mL of that solution to 100.0 mL.

Performing a titration

As mentioned previously, the purpose of performing a titration is to determine the unknown concentration of an acid or a base. To do this, that solution undergoes a neutralisation reaction with a solution of known concentration, which we call the primary standard solution. Let's go through the steps involved in conducting this experiment.

For this scenario, we are going to have an acid with an unknown concentration in the conical flask and a base as our primary standard in the burette. It is important to note that you could perform this experiment with the solutions in either location or with a base of unknown concentration and an acid of known concentration.

Concordant in titration, the volume of two or more titres that are within 0.1 mL of each other

The main steps in performing a titration are shown in Figure 13D–3. The figure shows these steps in detail and explains what is happening during each one.

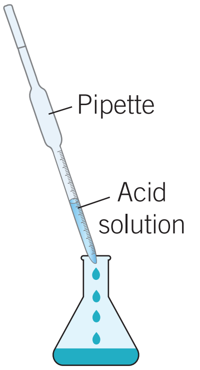
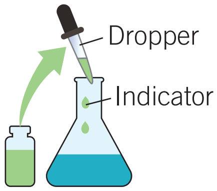
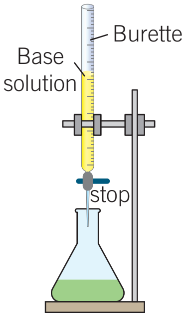
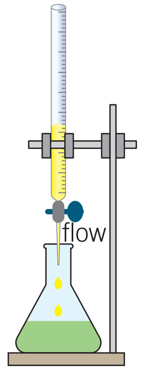
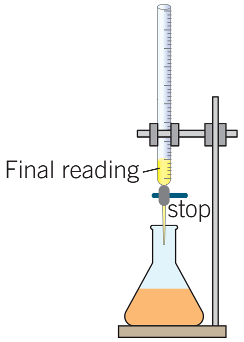
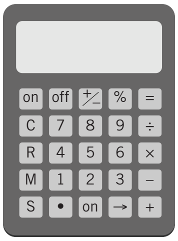
Step 1	Step 2	Step 3	Step 4	Step 5	Step 6
					
A precise volume of the acid solution, which has an unknown concentration, is added to the conical flask using a pipette. This solution may or may not have been diluted prior to adding to the conical flask, depending on what the estimated concentration range was.	An indicator is added to the solution in the conical flask. The use of indicators will be discussed in more detail later in this chapter.	The base solution of known concentration, acting as the primary standard, is added to the burette. The initial volume reading on the burette is recorded. The tap of the burette is in the stop position at this time.	The burette tap is turned to the flow position and the base solution is added in a dropwise fashion into the conical flask.	As soon as the indicator changes colour, the burette tap is moved back to the stop position. The final volume reading on the burette is recorded. The total volume of base solution added to the conical flask (initial volume minus the final volume on the burette) is recorded. This process should be repeated until three **concordant** results are obtained.	The total volume of primary standard solution (base in this example) required to neutralise the solution is used to calculate the unknown concentration of the acid. The process for conducting these calculations is outlined in depth later in this chapter.

Figure 13D–3 The main steps in performing an acid–base titration

In step 5, it is important to note that the procedure should be repeated until three concordant results are obtained. In Table 13D–1, five example **titre volumes** have been recorded for an experiment. You can see that volumes 1, 2 and 5 are concordant, and so the average of only these values should be used for performing any subsequent calculations.

Titre volume
the volume of solution required to reach the end point in a titration

Table 13D–1 Determining concordance

Titre volumes (mL)				
1	2	3	4	5
17.25	17.30	17.45	17.05	17.20

Indicators

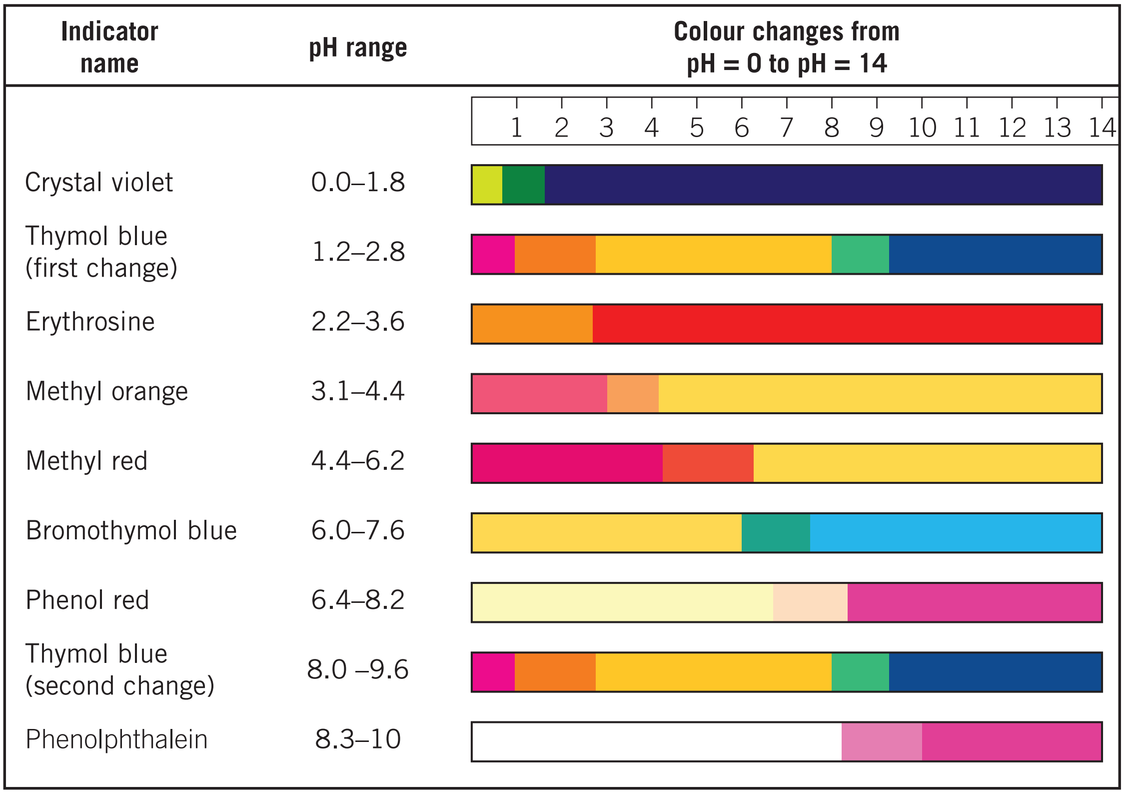
A titration reaction is performed until the number of moles of acid and base are in the correct stoichiometric ratio to have neutralised each other. This is known as the **equivalence point**. However, it is not something that can be observed when performing the experiment. Therefore, we use a compound known as an indicator to help us visualise this moment in the reaction. Acid–base indicators change colour in response to changes in the pH of the solution. The point at which the indicator changes colour is known as the **end point**. In a well-designed titration, the end point will be as close as possible to the equivalence point.

LINK **11C** CALCULATING pH

Equivalence point
the point in a titration when the number of moles of the acid and the base are present in the stoichiometric ratio that achieves neutralisation

End point
the point in a titration when the indicator changes colour

There are many different indicators that can be used that differ in the pH range at which they change colour (Figure 13D–4).



Indicator name	pH range	Colour changes from pH = 0 to pH = 14
Crystal violet	0.0–1.8	
Thymol blue (first change)	1.2–2.8	
Erythrosine	2.2–3.6	
Methyl orange	3.1–4.4	
Methyl red	4.4–6.2	
Bromothymol blue	6.0–7.6	
Phenol red	6.4–8.2	
Thymol blue (second change)	8.0 –9.6	
Phenolphthalein	8.3–10	

Figure 13D–4 Some common acid–base indicators. Indicators change colour in different pH ranges, making it critical that the right one is chosen for any given titration experiment.

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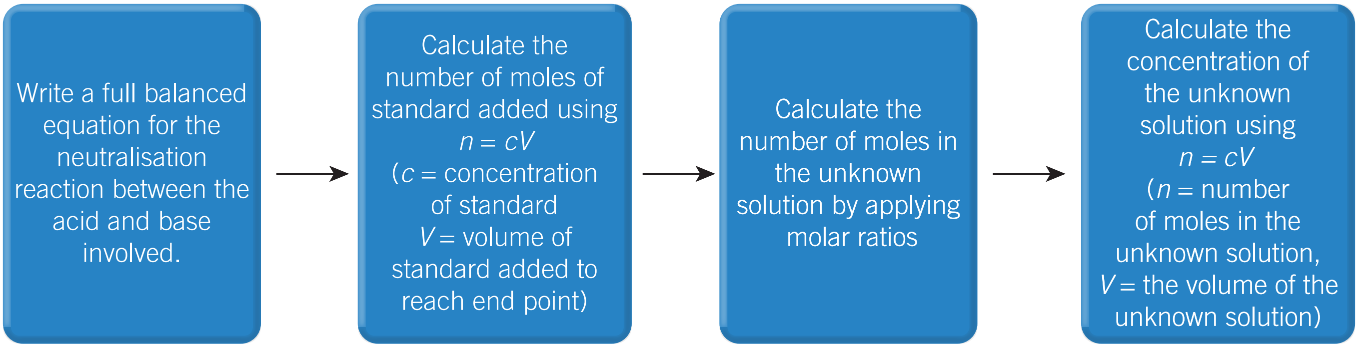
WORKSHEET 13D–1
VOLUMETRIC ANALYSIS CALCULATIONS

Titration calculations

The process to calculate the unknown concentration of a solution in any acid–base titration is illustrated in the flow chart in Figure 13D–5. An example of these steps is then shown in Worked example 13D–4.



Write a full balanced equation for the neutralisation reaction between the acid and base involved.

Calculate the number of moles of standard added using $n = cV$ (c = concentration of standard V = volume of standard added to reach end point)

Calculate the number of moles in the unknown solution by applying molar ratios

Calculate the concentration of the unknown solution using $n = cV$ (n = number of moles in the unknown solution, V = the volume of the unknown solution)

Figure 13D–5 Calculating the unknown concentration of a solution in any acid–base titration

Worked example 13D–4: Calculating concentration in an acid–base titration

VIDEO
WORKED EXAMPLE 13D–4

A 30.00 mL sample of a hydrochloric acid solution of unknown concentration is titrated with a secondary standard solution of sodium hydroxide. It takes 22.68 mL of the 0.1340 mol L^{-1} sodium hydroxide solution to reach the end point. What is the concentration of the hydrochloric acid solution?

Solution

Step 1	Write a balanced equation for the neutralisation reaction.	$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
Step 2	Extract the relevant information from the question and convert the volumes to litres.	$V(HCl) = 30.00\text{ mL} = 0.03000\text{ L}$ $c(HCl) = ?$ $V(NaOH) = 22.68\text{ mL} = 0.02268\text{ L}$ $c(NaOH) = 0.1340\text{ mol L}^{-1}$
Step 3	Calculate the number of moles of NaOH using $n = cV$.	$n(NaOH) = cV$ $n(NaOH) = 0.1340\text{ mol L}^{-1} \times 0.02268\text{ L}$ $= 3.039 \times 10^{-3}\text{ mol}$
Step 4	Calculate the number of moles of HCl using molar ratios.	$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$ Molar ratio of NaOH to HCl is 1:1 $n(HCl) = n(NaOH) = 3.039 \times 10^{-3}\text{ mol}$
Step 5	Rearrange $n = cV$ to make c the subject.	$n = cV$ $c(HCl) = \frac{n}{V}$
Step 6	Substitute in the values to solve for c.	$c(HCl) = \frac{n}{V}$ $= \frac{3.039 \times 10^{-3}\text{ mol}}{0.03000\text{ L}}$ $= 0.101304\text{ mol L}^{-1}$
Step 7	Finalise your answer using the appropriate number of significant figures.	$c(HCl) = 0.1013\text{ mol L}^{-1}$

Check-in questions – Set 2

1 A nitric acid solution of unknown concentration is titrated with a 0.100 M solution of magnesium hydroxide of known concentration. A 25.00 mL aliquot (using a pipette) of the nitric acid solution is transferred into a conical flask. The indicator phenolphthalein is added. The burette is filled to the 2.25 mL mark with magnesium hydroxide. The magnesium hydroxide solution is delivered into the conical flask until the indicator changes colour. The final volume on the burette is recorded as 20.30 mL.

What is the concentration of the nitric acid solution?

13D SKILLS

VIDEO 13D–2
SKILLS:
NOTATION IN
CHEMISTRY
CALCULATIONS

Notation in chemistry calculations

In this section, you learned about performing calculations in a volumetric analysis. As you saw with these calculations, the concentration of one reactant needs to be determined from information known about the other reactant. Questions for volumetric analysis include many numerical values. Therefore, it is important when writing symbols for given equations to assign more information as to which chemical species the value is referring to. Here, this is termed 'notations'.

Let's look at the example question below.

Question:

An aliquot of 20.00 mL of an unknown concentration of sulfuric acid solution is reacted with potassium hydroxide in a titration. The burette is filled with a 0.450 M solution of potassium hydroxide. The potassium hydroxide solution is delivered into the conical flask until the indicator changes colour. The mean titre volume is 30.45 mL. The reaction occurs according to the following equation:

$$H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$$

What is the concentration of the sulfuric acid solution?

Answer:

Written-out steps	Notation
Use information in the stem of the question to write down what is known and unknown. Clearly include notation about which species each value refers to, as opposed to just c or V.	$V(H_2SO_4) = 20.00$ mL $= 0.02000$ L $c(H_2SO_4) = ?$ $V(KOH) = 30.45$ mL $= 0.03045$ L $c(KOH) = 0.450$ M
Determine the amount, in mol, of potassium hydroxide. Again, use specific notation, $n(KOH)$, as opposed to just n. *Remember to keep the full number in the calculator in this calculation.	$n(KOH) = cV$ $n(KOH) = 0.450 \times 0.03045$ $n(KOH) = 0.0155$ mol

Written-out steps	Notation
Use the mole ratio in the balanced equation provided to determine the amount in mol of sulfuric acid. Again, use specific notation, $n(H_2SO_4)$, as opposed to just 'n' *Remember to keep the full number in the calculator in this calculation.	$n(H_2SO_4) = \frac{1}{2}n(KOH)$ $n(H_2SO_4) = \frac{1}{2} \times 0.0155$ $n(H_2SO_4) = 0.00776$ mol
Calculate the unknown concentration of sulfuric acid, using notation as mentioned above.	$c(H_2SO_4) = \frac{n(H_2SO_4)}{V(H_2SO_4)}$ $c(H_2SO_4) = \frac{0.00776}{0.02000}$ $c(H_2SO_4) = 0.388$ M

In these types of questions, you are not only trying to find the answer but have your teacher or examiner follow each step in your process of working. As these questions involve a number of steps, they are also not worth only one mark. Therefore, process marks would be awarded for each (or most) steps, so remember to label exactly what you are doing in each step with correct notation of the chemical species you are dealing with.

Section 13D questions

1 A number of different solutions are used in volumetric analysis.
   - a What is the name of the solution that has a known concentration in a titration?
   - b List three key characteristics of a compound that will be used to prepare the solution mentioned in part **a**.
   - c Potassium hydrogen iodate (HI_2KO_6) is a good primary standard for acid–base titrations. Calculate the mass of potassium hydrogen iodate required to make 350 mL of a 0.200 $mol\,L^{-1}$ solution.

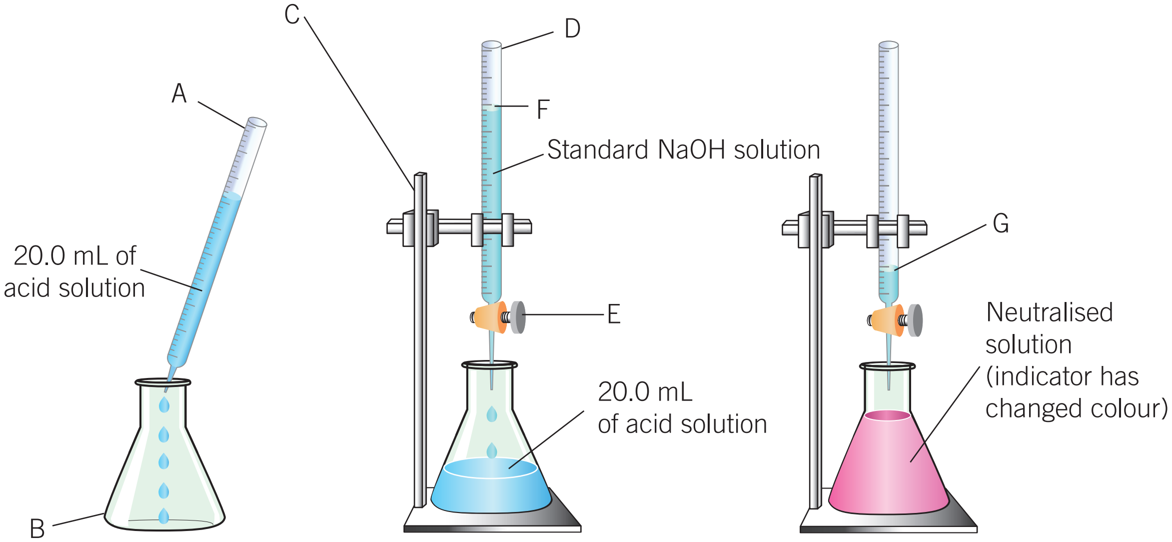
2 A titration is performed to determine the unknown concentration of a solution of barium hydroxide, $Ba(OH)_2$. For the reaction, 25.0 mL of 0.100 $mol\,L^{-1}$ hydrochloric acid is transferred to a conical flask with a pipette.
   - a Draw the experimental setup of this titration. Ensure that your diagram is fully labelled.
   - b If 50.0 mL of barium hydroxide is required to reach the end point, what is the concentration of the barium hydroxide solution?

3 Indicators are used in titrations to help guide when the reaction should be stopped.
   - a Define 'equivalence point' and 'end point' and describe how they relate to the use of an indicator.
   - b You are conducting a titration that has an equivalence point at pH 7. Using Figure 13D–4, choose two appropriate indicators for this titration and describe the colour change that will occur for each.

4 Calcium hydroxide and nitric acid can be used in a titration if the concentration of one of them is known.
   - a Write a balanced equation for the reaction between calcium hydroxide and nitric acid.
   - b What mass of calcium hydroxide would be required to neutralise 425 mL of 2.05 $mol\,L^{-1}$ nitric acid?

5 You are working in a laboratory and find a bottle labelled 1.50 mol L^{-1} sulfuric acid (H_2SO_4). Considering you didn't make this solution yourself, you want to perform a titration to check that the listed concentration is accurate. You have a 1.20 mol L^{-1} solution of sodium hydroxide, which you have just standardised. What volume, in mL, of the sulfuric acid would you need to titrate 30.00 mL of your sodium hydroxide if the label is correct?

6 Identify each of the following labelled pieces of equipment and important features (A–G) used in a titration.



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Chapter 13 review

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

Success criteria – I am now able to:		Linked question
13A.1	Define 'solubility', 'solute' and 'solvent'	10☐, 12☐
13A.2	Calculate the concentration of a solution	9☐, 11☐
13A.3	Convert between units of concentration	3☐
13B.1	Describe the relationship between temperature and solubility of a solid, liquid or gas in water	10☐
13B.2	Use solubility tables to qualitatively predict outcomes in a given scenario	4☐
13B.3	Use solubility graphs to both quantitatively and qualitatively predict outcomes in a given scenario	12☐
13B.4	Define 'saturated', 'unsaturated' and 'supersaturated'	8☐, 12☐
13C.1	Define 'precipitate'	13☐
13C.2	List common water impurities	1☐
13C.3	Use solubility tables to identify appropriate precipitating agents	13☐
13C.4	Describe the experimental method for removing an impurity from water by precipitation	13☐
13C.5	Write full balanced chemical and ionic equations for precipitation reactions	13☐
13D.1	Define 'end point' and 'equivalence point'	2☐, 11☐
13D.2	Describe the use of an indicator in an acid–base titration	6☐
13D.3	Define 'standard solution'	11☐
13D.4	State the key characteristics of a primary standard solution	7☐
13D.5	Calculate the concentration of a primary standard solution	11☐
13D.6	Calculate the concentration of a diluted solution	9☐
13D.7	List the steps involved in an acid–base titration, including glassware involved	5☐
13D.8	Calculate the concentration of an unknown acid or base in a solution from given experimental data	11☐

Multiple-choice questions

1 Which of the following is **not** a common type of impurity found in water?

A allergenic
B biological
C colloidal
D dissolved

2 The equivalence point of a titration is when

A the solution in the conical flask changes colour.
B there are the same number of moles of acid and base.
C the unknown solution has been neutralised.
D the pH is neutral.

3 The units $\mu g\,mL^{-1}$ are directly equivalent to

A $g\,L^{-1}$.
B $mol\,L^{-1}$.
C ppm.
D %(v/v).

4 Compounds containing which of the following ions are least likely to be soluble in water?

A ammonium (NH_4^+)
B carbonate (CO_3^{2-})
C ethanoate (CH_3COO^-)
D nitrate (NO_3^-)

5 A titration experiment should be repeated until

A there is no more primary standard solution remaining.
B the colour in the conical flask cannot get any darker.
C the pH of the unknown solution has reached 7.
D three concordant results have been obtained for the volume of standard solution required to neutralise the unknown solution.

6 Which of the following statements is correct about the use of an indicator in an acid–base titration?

A Any indicator will be suitable for any titration; they will just change colour at different times.
B The indicator helps visualise the point at which the reaction should be stopped.
C The indicator is added to the primary standard solution.
D The indicator should be added after all of the primary standard solution has been added to the unknown solution.

7 Which of the following characteristics is **not** important when selecting a compound to act as a primary standard in volumetric analysis?

A It must have high purity.
B It must have low hygroscopicity.
C It must have generally low reactivity.
D It must react with water in the atmosphere.

8 Compared to a saturated solution, the amount of solute in a supersaturated solution is

A less.
B more.
C the same.
D dependent on the temperature.

9 A 50 mL sample of a 0.15 mol L^{-1} solution of magnesium chloride is prepared. A 20 mL sample of this is transferred to a new flask and 40 mL of water is added to it. What is the concentration of the final solution?

A 0.050 mol L^{-1}
B 0.075 mol L^{-1}
C 0.15 mol L^{-1}
D 0.30 mol L^{-1}

10 The solubility of ionic solids will generally

A be unaffected by the temperature of the solvent.
B decrease as the temperature of the solvent increases.
C increase as the temperature of the solvent increases.
D only be possible once the temperature of the solvent is over 40°C.

Short-answer questions

11 You are completing work experience at a company that produces a popular brand of window cleaner. The active ingredient in the window cleaner is ammonia. You are asked to do some random quality control on the product and measure the concentration of ammonia present in the cleaner.

a You are told to first prepare a primary standard solution of sodium carbonate. If you need to make 300 mL of a 0.180 mol L^{-1} solution, how much sodium carbonate do you need to weigh out? (2 marks)

b The primary standard solution of sodium carbonate is used to make a secondary standard solution of 0.120 mol L^{-1} sulfuric acid. This will be used to titrate ammonia in a reaction that produces ammonium sulfate and water. Write a balanced chemical equation for this reaction. (2 marks)

The experiment is performed as described below and the following results obtained.

A 10.0 mL sample of window cleaner is diluted in 140 mL of water and then 25.0 mL of this is transferred into a conical flask using a pipette. The sulfuric acid solution, which is placed in the burette, is added drop by drop to the conical flask until a constant colour change is observed. The results for four trials are listed in the following table.

Trial number	Initial reading (mL)	Final reading (mL)	Volume added
1	48.4	39.2	
2	39.2	30.1	
3	30.1	20.2	
4	20.2	11.1	

c Complete the results table to determine the volume of sulfuric acid added (the titre) in each trial. (1 mark)

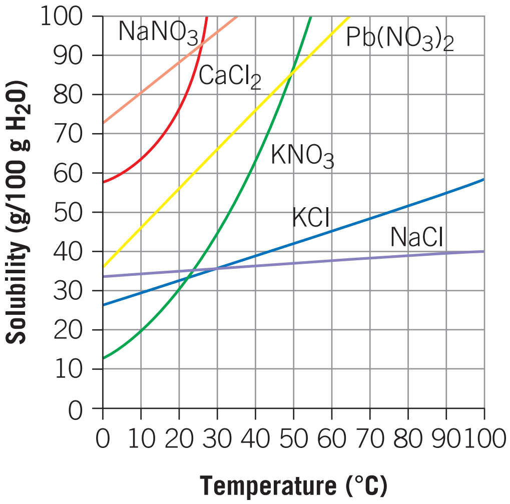
d Using the results above, calculate the average titre volume that should be used to determine the concentration of ammonia. (2 marks)

e Calculate the concentration of ammonia in the conical flask. (2 marks)

f Calculate the concentration of ammonia in the window cleaner. (1 mark)

g Someone repeats your experiments; however, they accidentally continue to add sulfuric acid after the indicator colour change has occurred. Using the terms 'equivalence point' and 'end point', explain whether the calculated concentration of ammonia that they obtain would likely be an overestimation or underestimation of the true value. (2 marks)

12 Solubility curves for six solutions are show in the graph below.



- **a** How many compounds have increased solubility as the temperature increases? (1 mark)
- **b** The solubility of which compound is least affected by a change in temperature? (1 mark)
- **c** How many grams of potassium chloride can be dissolved in 100 g of water at 40°C? (1 mark)
- **d** What is the minimum temperature required to dissolve 30 g of potassium nitrate in 100 g of water? (1 mark)
- **e** What mass of calcium chloride can be dissolved in 150 g of water at 25°C? (2 marks)
- **f** A solution is prepared with 80 g of lead nitrate dissolved in 200 g of water at 40°C. Is this solution unsaturated, saturated or supersaturated? Explain. (2 marks)

13 Water can have many impurities that may need to be removed for it to be safe for human consumption.

- **a** Describe, using a diagram, the process by which precipitation reactions can be used to remove an impurity from water. (3 marks)
- **b** Write an ionic equation for a precipitation reaction that could remove calcium ions from water. (2 marks)

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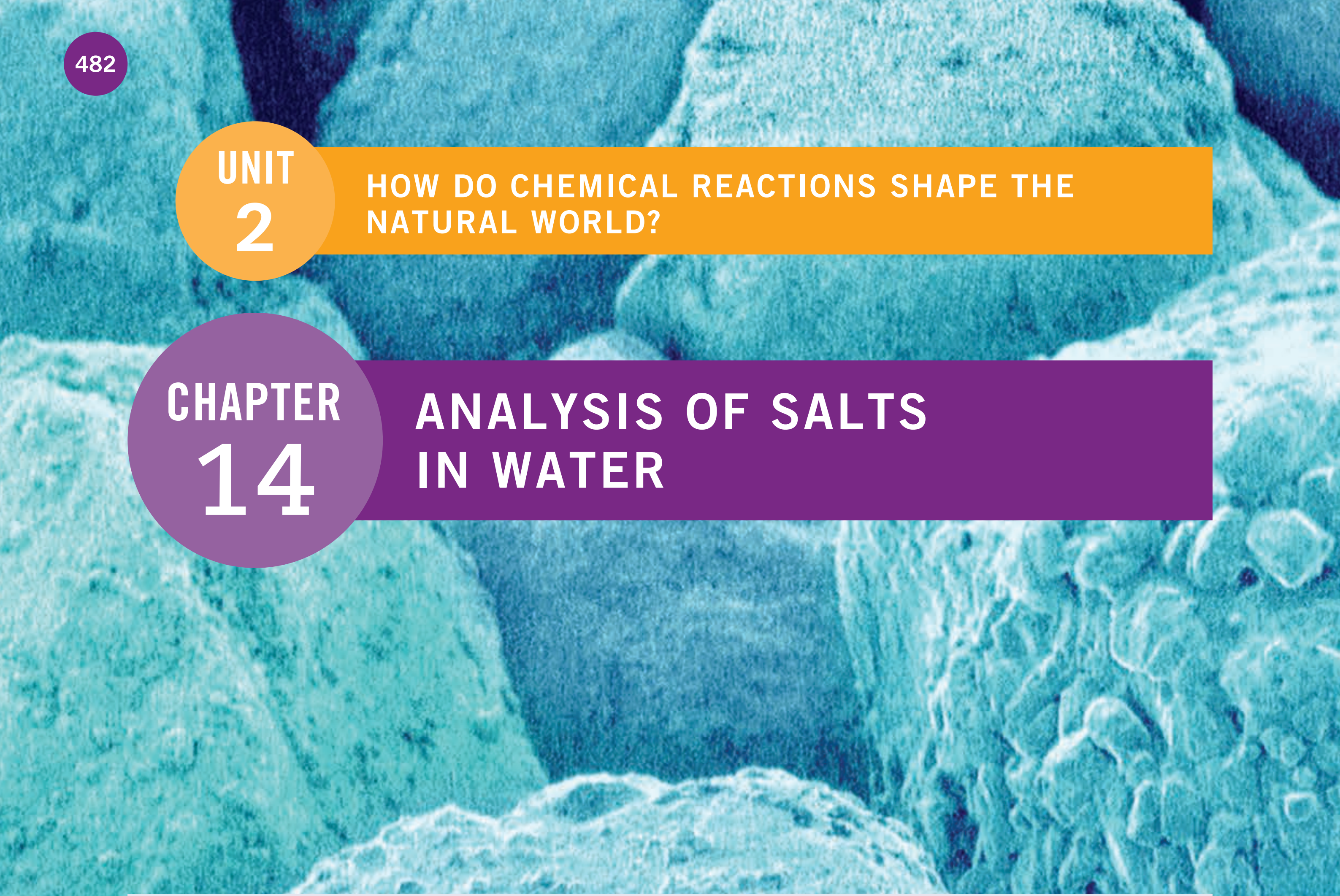

UNIT 2 HOW DO CHEMICAL REACTIONS SHAPE THE NATURAL WORLD?

CHAPTER 14

ANALYSIS OF SALTS IN WATER

Introduction

Drinking a glass of water straight from the tap is something many of us take for granted. However, in different areas of the world, many people do not have such a simple luxury. To make sure everyone has access to safe drinking water, we need to understand what can contaminate it and explore ways to test levels of these substances to treat the water. Drinking water can be contaminated with naturally occurring chemicals, such as arsenic, cadmium, nitrates, uranium and barium. It can also be contaminated via industrial, agricultural and mining waste, and from microbes found in sewage and animal waste. This chapter will look at the different ways and various techniques used to measure the presence of salts in water.

INTRODUCTION VIDEO ANALYSIS OF SALTS IN WATER

Curriculum

Area of Study 2 Outcome 2
How are chemicals measured and analysed?

Study Design:	Learning intentions – at the end of this chapter I will be able to:
• Sources of salts found in water or soil (which may include minerals, heavy metals and organometallic substances) and the use of electrical conductivity to assess the salinity and quality of water or soil samples	**14A Sources of salt in water** **14A.1** Recall the formation of ions and use of solubility rules to predict precipitates **14A.2** Understand the various sources of salts in water samples **14A.3** Measure levels of salinity using a conductivity meter

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Study Design:	Learning intentions – at the end of this chapter I will be able to:
- 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	**14B Gravimetric analysis of salts** **14B.1** Recall how to write a balanced equation for a precipitation reaction **14B.2** Recall how to perform mole calculations **14B.3** Determine the molar ratio of water of hydration for an ionic compound **14B.4** Define ‘gravimetric analysis’ and state what it is used for **14B.5** Explain the steps required in the process of quantitative analysis **14B.6** Identify mole ratios between chemicals **14B.7** Apply (mass–mass) stoichiometry to calculate chemical quantities **14B.8** Conduct a quantitative analysis in the laboratory
- 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	**14C Determining salt concentration using spectroscopy** **14C.1** Understand and apply the principles of colorimetry **14C.2** Understand and apply the principles of UV–visible spectroscopy **14C.3** Recall concentration calculations **14C.4** Convert between units of concentration **14C.5** Create and analyse a calibration curve using primary or secondary data

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Glossary

Bioaccumulation
Biomagnification
Calibration curve
Colorimetry
Conductivity meter
Conductor
Cuvette
Electromagnetic radiation
Electromagnetic spectrum
Full balanced equation
Gravimetric analysis
Heavy metal
Insoluble
Ionic compound
Ionic equation
Organometallic compound
Percentage mass
Quantitative
Quantitative analysis
Saline
Soluble
Spectroscopy
Standard solution
Stoichiometry
Total dissolved solids (TDS)
Ultraviolet–visible (UV–visible) spectrophotometry
Voltmeter

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Concept map

Man made:
- *Mining*
- *Agriculture*
- *Domestic*
- *Sewerage*

Natural:
- *Minerals*
- *Bioaccumulation*

14A Sources of salt in water

Water quality

Measured by:
Conductivity meter

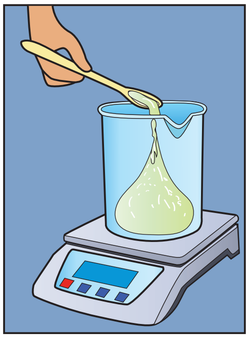
Quantitative analysis

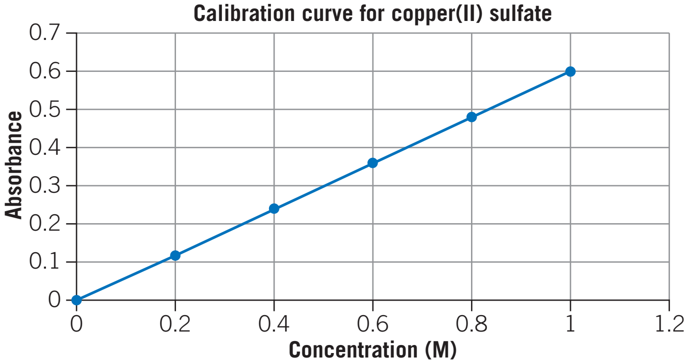
Measured by:
- *Gravimetric analysis*
- *Mass–mass stoichiometry*

Measured by:
- *Colorimetry*
- *UV–visible spectroscopy*
- *Concentration calculations*

14B Gravimetric analysis of salts

14C Determining salt concentration using spectroscopy





See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

14A Sources of salt in water

Study Design:
Sources of salts found in water or soil (which may include minerals, heavy metals and organometallic substances) and the use of electrical conductivity to assess the salinity and quality of water or soil samples

Glossary:

Bioaccumulation
Biomagnification
Conductivity meter
Conductor
Heavy metal
Ionic compound
Organometallic compound
Quantitative
Saline
Total dissolved solids (TDS)
Voltmeter

ENGAGE

Water is the universal solvent

As you learned way back in Chapter 2B, water is a polar molecule. This makes it an excellent solvent of ionic compounds (or salts) and polar molecules, which can either ionise or form hydrogen bonds with water. Water is often referred to as the universal solvent. As a result, all water contains a certain amount of dissolved substances, including salts. Human activity has increased the amount of dissolved salts in many of our water systems, in some cases to dangerously high levels, resulting in detrimental effects on the environment and plant and animal populations.

LINK **2B** MOLECULAR SHAPES AND POLARITY

LINK **13C** PURIFYING WATER

Figure 14A–1 The Great Barrier Reef is one of the seven wonders of the world but is showing the effects of increased dissolved salt levels.

The Great Barrier Reef is recognised as one of the seven natural wonders of the world, being the largest reef in the world. It stretches over 2300 km and is rich in biodiversity. Sadly, it is suffering the effects of increased salt levels due to human activity. Mining, farming, aquaculture and urban development along the Queensland coast have all increased the runoff of various substances into the reef after rainfall. One of the negative effects from this runoff is due to the soluble fertiliser ions, such as nitrates and phosphates. These increased levels can cause eutrophication and algal blooms. Algal blooms block light from reaching organisms such as corals and sea grasses, resulting in the loss of the food sources for other organisms, like the dugong.

LINK **1D** CRITICAL ELEMENTS AND RECYCLING PROCESSES

EXPLAIN

What is a salt?

Did you know that 97% of Earth's water is salt water? And that common salt – sodium chloride – is one of the most widely used substances on Earth? Worldwide, 200 megatonnes of salt is consumed each year.

Figure 14A–2 Raw salt crystals

4A FORMATION AND NAMING OF IONIC COMPOUNDS **LINK**

Ionic compound
a compound or molecule formed when positive cations and negative anions chemically combine

We commonly think of salt as an additive used to flavour our food, but this 'table salt' is called sodium chloride, NaCl, which is an ionic compound. More correctly, the term 'salt' refers to a very large group of **ionic compounds** composed of positive metal cations and negative non-metal anions. Remember in Chapter 4A you learned that ionic compounds are electrically neutral compounds, therefore the total charge of the cation(s) equals the total charge of the anion(s). Many ionic compounds are soluble in water. This solubility can be represented by an equation showing how the solid salt, when placed in water, can ionise into free-moving aqueous ions.

$$\underset{\text{Salt}}{NaCl(s)} \xrightarrow{H_2O(l)} \underset{\text{Cation}}{Na^+(aq)} + \underset{\text{Anion}}{Cl^-(aq)}$$

Check-in questions – Set 1

1 Identify which of the following compounds are salts (ionic compounds).
  - a Cl_2
  - b AgF
  - c $MgSO_4$
  - d $C_6H_{12}O_6$
  - e NH_3
  - f $MgCl_2$

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2 Identify the anion and cation in each of the following salts.

Compound	Cation	Anion
Ag_2O		
$CaCO_3$		
CH_3COONa		

3 Write the formulas for the following salts.

a copper(II) fluoride
b lithium nitrate
c potassium sulfide
d barium iodide

Sources of salts

VIDEO 14A–1 SOURCES OF SALT IN WATER

All naturally occurring water on Earth contains dissolved salts. These salts can be attributed to both natural and artificial sources. Monitoring the levels of these dissolved salts is vital in maintaining safe levels in drinking water and safe environmental levels in waterways.

Salts from minerals

Salts are naturally present in water systems. These salts are generally derived from three sources.

1 Ocean water. There are high levels of salt in ocean water, accounting for much of Earth's salts.
2 Salt is released from rocks during weathering as water runs through soil and rocks. For example, Hepburn Springs is famous for its mineral springs, which contain high levels of dissolved ions due to water passing through rocks that themselves contain millions of years' worth of trapped ions.
3 In some locations, salt remains in sediments left by retreating seas after periods where ocean levels were much higher or the land surface was much lower. The impressive pink lake, Lake MacDonnell in South Australia (Figure 14A-3), contains high levels of salt as a result of this.

Figure 14A–3 Aerial view over pink Lake MacDonnell in South Australia. High in salt concentration, the lake gets its pink colour from a salt-loving bacteria known as halobacteria, which secrete red pigments that give the water its pink colour.

Salts from human activity

1D CRITICAL ELEMENTS AND RECYCLING PROCESSES **LINK**

Humans contribute to increasing salt levels through our everyday domestic activities, including sewerage and use of soaps and detergents that ultimately end up in our waterways. Mining, industry runoff and agriculture, including the use of fertilisers, results in nitrates (NO_3^-), sulfates (SO_4^{2-}) and phosphates (PO^{43-}) being dissolved and transported into waterways. In some countries where snow and ice affects road access, the use of salts to keep roads accessible is another major contributor.

Heavy metal salts

Heavy metal a metal with a relatively high density and a toxic effect on living organisms, such as cadmium, chromium, copper, lead and mercury

Bioaccumulation the build-up of heavy metals in organisms

Biomagnification the build-up (increasing concentration) of heavy metals as we move up the food chain

Heavy metals are metals with relatively high densities and a toxic effect on living organisms, such as cadmium, chromium, copper, lead and mercury. Although heavy metals occur naturally within Earth's crust and can dissolve into rivers and groundwaters, the concentration from these sources is very low. Again, human activity, such as mining, is responsible for increased levels of heavy metals in the environment and in waterways. High concentrations of heavy metals can be detrimental to human health. As such, their acceptable levels are monitored carefully and guidelines are set by the World Health Organization (WHO) and governments.

Heavy metals do not biodegrade over time; once released, they persist in the environment and can actually increase in concentration through the food chain due to **bioaccumulation** and **biomagnification**. For example, small amounts of lead – a toxic heavy metal – released through human activity or pollution, for example from lead-based paints, enter waterways and are absorbed by plankton. These toxins cannot be excreted and don't break down over time. As a result, they are absorbed by the animals that eat the plankton. This repeats up the food chain, resulting in the highest concentration of lead in the body tissues of the species at the top of the food chain (Figure 14A–5, on the following page).

4C PRECIPITATION REACTIONS **LINK**

14B GRAVIMETRIC ANALYSIS OF SALTS **LINK**

Heavy metals can be removed from contaminated water using a precipitation reaction to precipitate out and collect this insoluble compound through filtration. You will learn more about this process when we look at gravimetric analysis later in this chapter.

Organometallic compounds

8B POLYMER RECYCLING **LINK**

Organometallic compound a compound with at least one carbon–metal bond

Organometallic compounds are substances with at least one carbon–metal bond. They are usually synthetic substances that are often used as catalysts for chemical reactions. Methylmercury is an organometallic cation, CH_3Hg^+, used in fluorescent lights, batteries and in the production of polyvinyl chloride (PVC). It can combine with anions to form organometallic salts, such as methylmercury chloride (Figure 14A–4). Such compounds are formed when mercury-containing substances are burned. They are even more toxic than the heavy metal mercury as they are more readily taken up and transported around the body, acting as a neurotoxicant, a substance capable of causing harm to the central nervous system. Like heavy metals, they do not degrade over time, and instead bioaccumulate in the body. Because of this, the Environmental Protection Authority (EPA) has strict measures to control the discharge of these compounds from chemical plants into waterways.

Mines, smelters, industries, landfill and runoff from urban areas are all responsible for adding heavy metals and organometallic compounds into watercourses where they eventually flow down to the sea.

Figure 14A–4 Methylmercury chloride is a neurotoxicant and an environmental hazard.

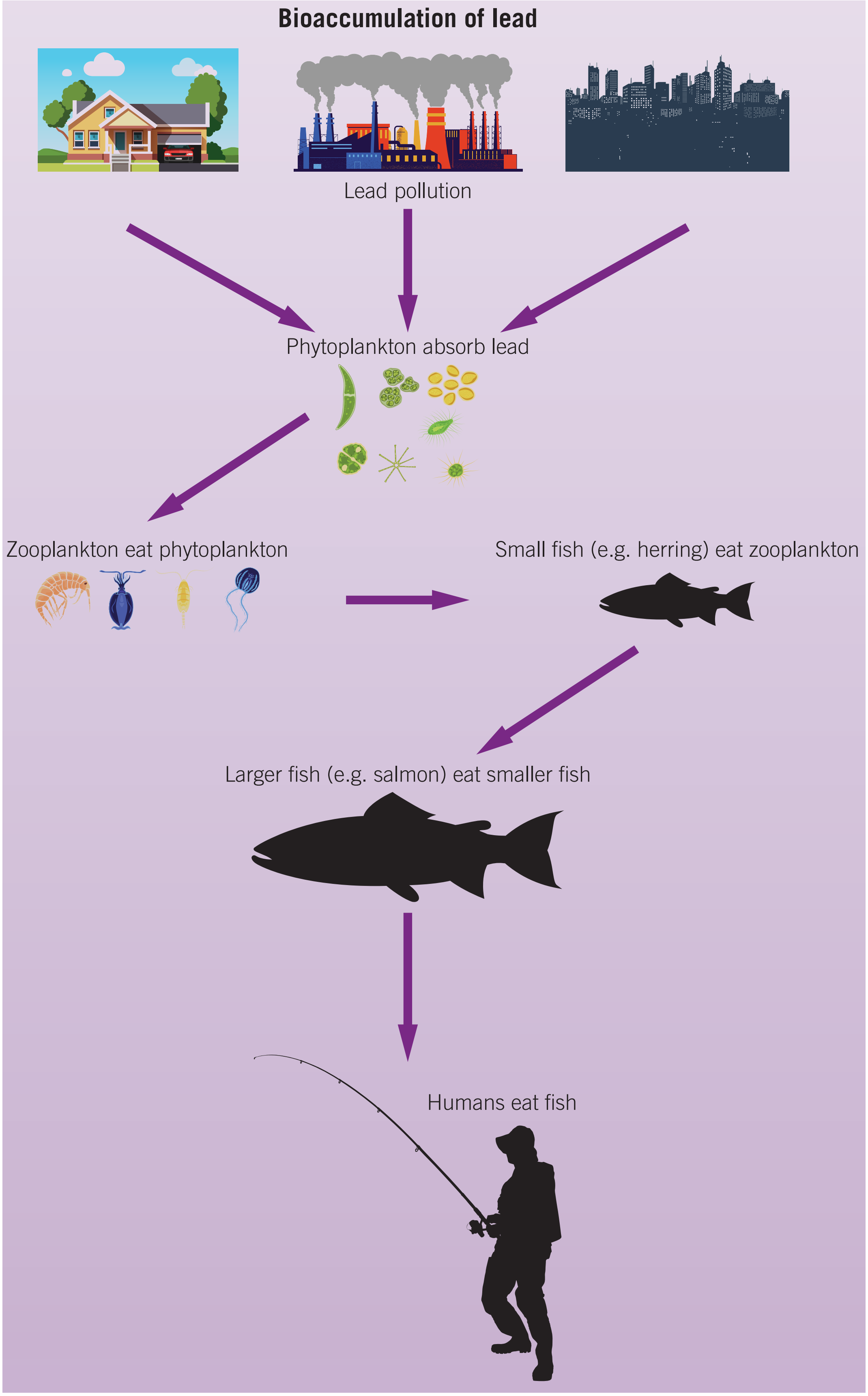

Figure 14A–5 The increasing concentration of lead in smaller species by bioaccumulation impacts the food chain through biomagnification, resulting in high concentrations of the toxin towards the top of the chain.

Check-in questions – Set 2

1 Recall three common sources of salts in Earth's waters.

2 a What is a heavy metal?

  b Heavy metals bioaccumulate in an organism, leading to biomagnification in the food chain. Explain what this statement means.

Measuring salinity in water

Saline water that contains dissolved salt

Approximately 97% of the water on Earth is salt water. Sea water contains a high concentration of dissolved salts and is known as being **saline**. Although sea water does predominately contain dissolved sodium chloride salt ($Na^+(aq) + Cl^-(aq)$), it also contains many other dissolved ions, as shown in Table 14A–1.

Currently, over 700 million people suffer from a shortage of drinkable water. The global water shortage is predicted to increase because of climate change, increasing population and an increase in agricultural demands. Saline water is unsuitable for human consumption or use in agriculture, as it causes dehydration in both animals and plants. The viability and use of desalination plants to increase the quantity of fresh water available for consumption is being investigated in many countries.

Table 14A–1 Some common dissolved ions from water-soluble ionic salts that are found in sea water

Ion	Concentration of ion in sea water mg L^{-1} (or ppm)
Chloride	18 980
Sodium	10 556
Sulfate	2649
Magnesium	1262
Calcium	400
Potassium	380
Bicarbonate	140
Bromide	65
Borate	26
Strontium	13
Fluoride	1
Silicate	1
Iodide	<1
Total dissolved ionic solids	**34 473**

4C PRECIPITATION REACTIONS LINK

You can refer to the solubility rules you learned in Chapter 4 to determine if an ionic compound is soluble or insoluble in water.

Conductivity meters

A simple way we can measure the salt content in water is by evaporating the water and weighing the amount of salt that remains. However, this isn't always a good representation of the salt content of a large area of water and can be time-consuming.

Another way salt levels can be measured in a school laboratory is to set up a simple circuit like the one shown in Figure 14A–6 and measure the electrical conductivity (EC) of a sample.

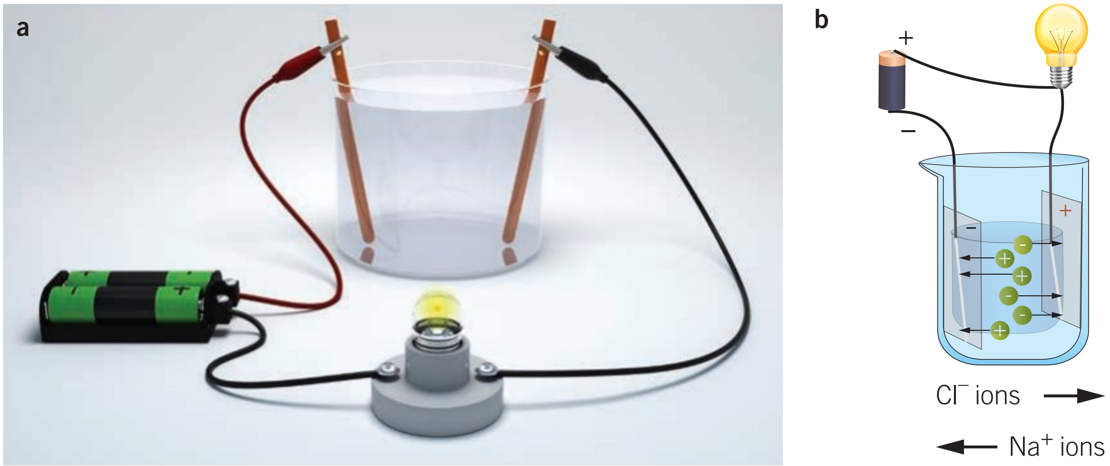

Figure 14A–6 A simple circuit showing how salt water can successfully conduct electricity; **a)** The brighter the bulb, the greater the conductivity of the solution, and hence the higher the level of dissolved salt. **b)** An ionic solution will conduct an electric current as positive ions move towards the negatively charged electrode and negative ions move towards the positively charged electrode.

Conductor a material that allows the flow of charge (electrical current)

It is known that distilled water is a poor **conductor**, while ionic compounds – salts – can conduct electricity when dissolved. The higher the concentration of dissolved ions, the greater the conductivity of the solution. This can be visualised by a bulb glowing brighter or a **voltmeter** giving a higher voltage reading.

LINK **4B** STRUCTURE AND PROPERTIES OF IONIC COMPOUNDS

This method still doesn't give a **quantitative** measure for how many dissolved ions are present in a solution. To do this, we can use a **conductivity meter**. A conductivity meter consists of two probes that emit electrical energy. This small current passes between the two electrodes, working in the same way as our simple circuit does. The higher the concentration of dissolved ions, the higher the conductivity that is measured. The reading obtained gives a correlation between the conductivity of the solution and the amount of **total dissolved solids (TDS)** in ppm or $\text{mg}\,\text{L}^{-1}$. It does not give an actual measurement of the number of ions or the type of ions in the solution. Water containing more than 1000 $\text{mg}\,\text{L}^{-1}$ TDS (Table 14A–2) or having electrical conductivity (EC) greater than 10 $\text{mS}\,\text{cm}^{-1}$ (10 000 $\mu\text{S}\,\text{cm}^{-1}$) is considered unacceptable. Water of this type is also unsuitable for most agricultural purposes due to the harmful effects of high ionic concentrations on plants.

Voltmeter an instrument used for measuring electrical potential

Quantitative relating to measured values

Conductivity meter an instrument that measures the amount of electrical conductivity a solution has; used to indicate levels of dissolved ions

Total dissolved solids (TDS) salinity is often measured using TDS. The total dissolved solids in a water sample is measured in $\text{mg}\,\text{L}^{-1}$

Table 14A–2 Salinity guidelines for water use

Salinity status	Salinity ($\text{mg}\,\text{L}^{-1}$)	Water quality
Fresh	<500	Good for drinking water and all irrigation
Marginal	500–1000	Most irrigation, unpleasant to drink
Brackish	1000–2000	Irrigation for specific crops only
Saline	2000–10 000	Useful for most livestock, can be used for salt-resistant vegetation
Highly saline	10 000–35 000	Limited use except for some livestock
Brine	>35 000	Not suitable for plants or animals

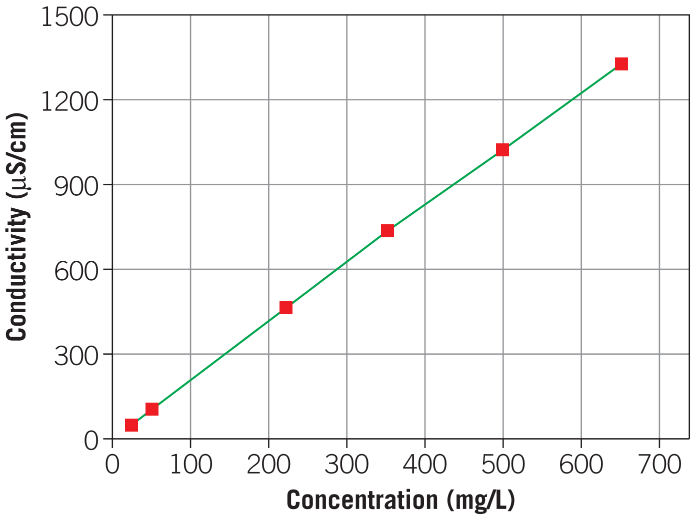

Figure 14A–7 The conductivity of a salt solution increases with concentration.

Conductivity meters are used in wastewater treatment plants, environmental science and agriculture to measure salinity levels of water and soil samples. Often these meters can also be used to measure pH and dissolved oxygen levels.

Another advantage of this method is that the device allows for testing in the field.

WORKSHEET 14A–1 MEASUREMENTS OF SALINITY IN WATER

Figure 14A–8 Testing a water sample with portable digital probe. This will provide a numerical, or quantitative, reading of temperature, pH and electrical conductivity.

Check-in questions – Set 3

1. What happens to electrical conductivity as salinity levels increase?
2. A sample of water has an electrical conductivity of 2700 mS cm^{-1}. Use the salinity guidelines in Table 14A–2 to determine if this sample is safe for consumption.

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14A SKILLS

Errors associated with experimental design

In this section, it was outlined that one way to determine the amount (in grams) of salt present in a sample of water is to boil that sample, thereby evaporating the water and leaving behind only the salt present.

For example, in the high school science experiment method outlined below, two different water samples were analysed – a sample from a marine aquarium and a sample from a planted aquarium.

Method

1. Using scales record the mass of an empty crucible.
2. Collect a sample of water from the marine aquarium and filter this using filter paper to remove any other impurities, such as sand, microorganisms or other debris from the water.
3. Re-zero the digital scales and pour a sample of this water into the crucible until it is approximately half-filled.
4. Record the mass of the water sample added.
5. Place the crucible containing the water sample so it rests on a clay triangle sitting on a tripod and over the top of a Bunsen burner.
6. Turn on the Bunsen burner and with the air hole closed, heat the solution until all the water is evaporated from the crucible.
7. After allowing the crucible to cool, record the mass of the salt formed using the scales.

In performing this method, which sounds simple, you will notice that an error occurs. As the water sample boils, the water splashes over the side of the crucible. As this water dries on the bench around your experiment, you will notice that small white crystals begin to form. This is 'salt' from the water. The result of this is that the mass of the salt recorded will be less than was actually present in the water sample (as some of it is now on the bench and the tripod as opposed to in the crucible, as shown in the image below).

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This type of error occurs due to a flaw in the design or conduction of the experiment and the choice of apparatus. If you were to repeat this experiment, you may find the error occurs again, but to a different extent each time. This error will affect both the precision and accuracy of your collected data. As the experimenter, it would be up to you to redesign your method and/or reconsider your choice of equipment to rectify this in order to improve the accuracy of your collected data.

16B SCIENTIFIC EVIDENCE **LINK**

Different types of errors associated with practical investigations, along with their impact on the accuracy and precision of results, will be explored more in Chapter 16B.

Section 14A questions

1 The three common sources of salts in Earth's waters are minerals, heavy metals and organometallic substances. Explain where each of these salts originates from.

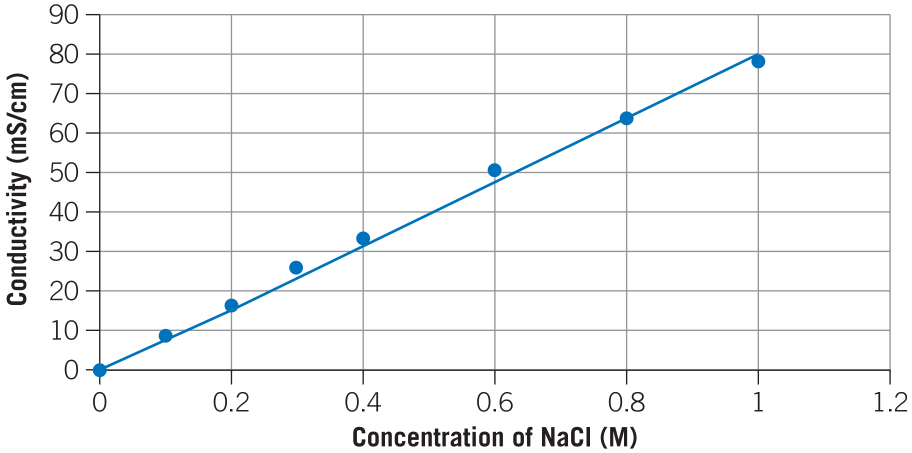
2 Define 'salinity'.

3 Salinity can be measured using a conductivity meter.

   a Explain how the conductivity meter measures salinity, with reference to how it works.

   b Outline two limitations of conductivity meters in terms of what they can record.

Below is a graph representing the conductivity of a salt, sodium chloride, over a range of molar concentrations.



   c Identify the independent and dependent variables from the graph.

   d Using the graph above, determine the conductivity of a 0.5 M NaCl solution.

   e Explain the trend observed from the graphed data.

   f Outline why the dot points representing recorded data do not fall exactly on the line drawn for the graph.

4 Explain 'bioaccumulation' with reference to an example presented in this chapter or from your own knowledge.

5 Recall from Chapter 4C the dissolution of salts in water. Describe, using both an equation and a diagram, how solid sodium chloride (NaCl) can form ions when dissolved in water.

14B Gravimetric analysis of salts

Study Design:

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

Glossary:

Full balanced equation
Gravimetric analysis
Insoluble
Ionic equation
Percentage mass
Quantitative analysis
Soluble
Stoichiometry

ENGAGE

Nutritional analysis

Have you ever wondered how the quantity of nutrients, such as sodium, is determined before they are listed on the nutritional information table on food packaging? One of the ways these nutrients can be measured is by using quantitative analysis. Components in the foodstuff are analysed in a laboratory and listed on the package as a percentage based on their measured mass. Quantitative analysis can also be used to monitor the levels of harmful heavy metals, such as mercury, in drinking water.

Figure 14B–1 A scientist collecting water samples to measure mercury levels. Even low levels of mercury in waterways can be toxic.

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EXPLAIN

Solubility rules and precipitation reactions – review

4C PRECIPITATION REACTIONS LINK

Quantitative analysis any method of analysis that measures the numerical amount of a given substance, including gravimetric analysis and colorimetry

Gravimetric analysis determining the mass of an unknown ion by forming a known mass of precipitate

Soluble when a solute is able to dissolve in a solvent

In Unit 1, you learned that some combinations of solutions result in the formation of an insoluble ionic compound known as a precipitate.

The following two sections examine different forms of quantitative analysis. In chemistry, **quantitative analysis** covers any method of analysis that measures the amount of a given substance. **Gravimetric analysis**, a form of quantitative analysis, uses precipitation reactions, in particular the formation of a known mass of precipitate, to enable us to calculate the amount of a particular ion present in a sample. Unlike a conductivity meter, this technique provides an actual measurement of the amount of a specific ion (or ions) dissolved.

Solubility rules

To successfully utilise a precipitation reaction for the quantitative analytical technique, it is important we understand and remember which ionic compounds (salts) are **soluble** and which are insoluble in water at 25°C.

Worked example 14B–1: Determining the solubility of substances

Determine if the following substances are soluble or insoluble in water at 25°C:

a zinc carbonate

b silver nitrate.

Solution

a

Step 1	Break the ionic compound into its two ions (cation and anion).	Zn^{2+} CO_3^{2-}
Step 2	Look for the ions on the solubility table (Chapter 4, Table 4C–1) to determine if they are soluble or not.	Carbonates (use CHPS) – insoluble
Step 3	INSOLUBLE	

b

Step 1	Silver nitrate Break the ionic compound into its two ions (cation and anion).	Ag^+ NO_3^-
Step 2	Look for the ions on the solubility table to determine if they are soluble or not.	Nitrate (use SNAPE) – soluble
Step 3	SOLUBLE	

Precipitation reactions

Many ionic compounds dissolve in water and form an aqueous solution; for example, sodium chloride or table salt:

$$NaCl(s) \xrightarrow{H_2O\,(l)} Na^+(aq) + Cl^-(aq)$$

When two ionic solutions are mixed, the ions in the solution can sometimes form a new compound that is **insoluble** and separates from solution. This is called a precipitate and therefore why this reaction is called a precipitation reaction.

Insoluble when a solute has limited solubility in a solvent

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Writing a balanced equation

Recall that a precipitation reaction is an example of a double displacement reaction. The cations in the reactants swap their paired anions to form two new products.

There are two ways we can represent a precipitation reaction:

1 Using a **full balanced equation**, which shows all the ions that are present

or

2 Using an **ionic equation**, which shows only the ions that participate in forming the precipitate

Figure 14B–2 Common salt, NaCl, dissolves readily in water.

Full balanced equation a balanced equation that shows all atoms involved in a reaction, including spectator ions

Ionic equation a balanced equation that shows only the chemical species that participate in a reaction

Worked example 14B–2: Writing full equations for precipitation reactions

Write a full balanced equation for the precipitation reaction that results when aqueous solutions of sodium sulfate and barium chloride are mixed at 25°C.

Solution

Step 1	Write out the reactants. **Remember that you will need to use your knowledge of the charges formed by atoms losing or gaining electrons and recall the charges of polyatomic ions.**	Na_2SO_4 + $BaCl_2$ Na^+ SO_4^{2-} Ba^{2+} Cl^-
Step 2	Deduce the products formed by swapping the cation and anion pairings.	Na^+ SO_4^{2-} Ba^{2+} Cl^- Ba^{2+} SO_4^{2-} Na^+ Cl^-
Step 3	Write out the products formed.	$BaSO_4$ NaCl
Step 4	Determine if a precipitate is formed, using the solubility rules.	$BaSO_4$ NaCl Insoluble (precipitate) soluble
Step 5	Assign correct states to all reactants and products.	$Na_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + NaCl(aq)$
Step 6	Balance the equation.	$Na_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

Worked example 14B–3: Writing ionic equations for precipitation reactions

Write a balanced ionic equation for the precipitation reaction that results when aqueous solutions of sodium sulfate and barium chloride are mixed at 25°C.

Solution

Step 1	Using the fully balanced precipitation reaction, write any dissolved ionic compounds (aq) as their separate ions. Remember, (s) substances won't break into ions, so leave these as a full ionic compound.	$2Na^+(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2Cl^-(aq) \rightarrow BaSO_4(s) + 2Na^+(aq) + 2Cl^-(aq)$

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Step 2	Cross out any ions that appear in the same state on both sides of the equation. These ions have not participated in the reaction and are the spectator ions.	~~$2Na^+(aq)$~~ + $SO_4^{2-}(aq) + Ba^{2+}(aq)$ ~~$+ 2Cl^-(aq)$~~ $\rightarrow$ $BaSO_4(s)$ + ~~$2Na^+ (aq)$~~ ~~$+ 2Cl^-(aq)$~~
Step 3	Rewrite the equation, deleting the spectator ions and check states. Ensure the equation is balanced.	$SO_4^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO_4(s)$

Mole calculations – review

6B THE MOLE **LINK**

In Chapter 6, you learned that if you know the mass of a given substance in grams, then you can calculate the amount in moles and vice versa using the relationship

$$n = \frac{m}{M}$$

where:

n = The amount of a substance (in mol)

m = The mass (in grams)

M = The molar mass (in g mol^{-1})

Worked example 14B–4: Calculating amount of moles from mass

Calculate the amount in moles of calcium chloride present in 5.0 g of $CaCl_2$.

Solution

Step 1	Identify the information that is known and unknown from the question.	$m(CaCl_2) = 5.0$ g $n(CaCl_2) = ?$
Step 2	Calculate the molar mass of the substance using relative atomic masses from the periodic table.	$M(CaCl_2) = 40.1 + (2 \times 35.5)$ $= 111.1$
Step 3	Input the data given into the correctly rearranged formula.	$n = \frac{m}{M}$ $n(CaCl_2) = \frac{m}{M} = \frac{5.0}{111.1}$
Step 4	Solve for the amount in moles.	$n(CaCl_2) = 0.045$ mol

Check-in questions – Set 1

1 Write a balanced chemical equation for the following precipitation reactions.
   - a $AgNO_3(aq) + NaCl(aq) \rightarrow$
   - b $CuSO_4(aq) + Na_2CO_3(aq) \rightarrow$
   - c $(NH_4)_2SO_4(aq) + BaCl_2(aq) \rightarrow$

2 Calculate the amount, in mol, of 10.0 g of the following substances.
   - a Na_2CO_3
   - b $K_2Cr_2O_7$

3 Calculate the mass of:
   - a 0.2 mol NaOH
   - b 0.0043 mol H_2SO_4

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Determining the hydration ratio of an ionic compound

Some salts, like the copper sulfate shown in Figure 14B–3, can be found as a hydrated solid, where water molecules are chemically bonded into the crystal lattice. When heated, it is possible to remove the water molecules from the crystal, returning the salt to its anhydrous form (containing no water). The water bound to the crystal is called 'water of hydration'.

Using the mass of the salt before and after heating, the ratio of hydration – how many water molecules are incorporated into the crystal structure – can be calculated.

Copper sulfate is an example of a hydrated salt. Each copper sulfate molecule has a set number of water molecules incorporated into the crystal structure. This can be represented as $CuSO_4.xH_2O$. The actual number of water molecules present can be determined experimentally.

Figure 14B–3 Blue copper sulfate, $CuSO_4.5H_2O$ crystals, are an example of a hydrated salt.

Worked example 14B–5: Calculating hydration ratios

A 250.0 g sample of hydrated copper sulfate ($CuSO_4.xH_2O$) is heated in an evaporating dish until its blue colour is entirely removed. The sample is reweighed and has a mass of 160.0 g. The mass loss is determined to be due to the removal of water molecules.

Calculate the ratio of hydration for the $CuSO_4$ molecule.

Solution

Step 1	List the mass of the salt molecules present. List the mass of the water molecules present.	$CuSO_4.xH_2O(s) \rightarrow CuSO_4(s) + xH_2O(g)$ $m(CuSO_4 \cdot xH_2O) = 250.0$ g $m(H_2O) = 250.0 - 160.0$ $= 90.0$ g
Step 2	Calculate the molar mass of both the copper sulfate and water using relative atomic masses from the periodic table.	$M(CuSO_4) = 63.5 + 32.1 + (4 \times 16)$ $= 159.6\ \text{g mol}^{-1}$ $M(H_2O) = (2 \times 1) + 16$ $= 18\ \text{g mol}^{-1}$
Step 3	Determine the amount in moles of each molecule present using the mole formula.	$n = \frac{m}{M}$ $CuSO_4 : H_2O$ $n(CuSO_4) = \frac{m}{M} : n(H_2O) = \frac{m}{M}$ $n(CuSO_4) = \frac{160.0}{159.6} : n(H_2O) = \frac{90.0}{18}$ $n(CuSO_4) = 1.00 : n(H_2O) = 5.00$

Step 4	Ensure the ratio of salt molecules to water molecules is a whole-number ratio. If it is not, divide both numbers by the smallest to create a whole number.	$n(CuSO_4) = 1\ :\ n(H_2O) = 5$ Ratio = 1 : 5
Step 5	State the formula for the hydrated salt.	$CuSO_4.5H_2O$

Worked example 14B–6: Calculating hydration ratios

A 23.2 g sample of hydrated magnesium nitrate ($Mg(NO_3)_2$) is heated in an evaporating dish. The sample is reweighed and has a mass of 19.6 g. The mass loss is determined to be due to the removal of water molecules.

Calculate the ratio of hydration for the ($Mg(NO_3)_2$) molecule.

Solution

Step 1	List the mass of the salt molecules present. List the mass of the water molecules present.	$Mg(NO_3)_2{\cdot}xH_2O(s) \rightarrow Mg(NO_3)_2(s) + xH_2O(g)$ $m(Mg(NO_3)_2\ {\cdot}xH2O) = 23.2$ g $m(Mg(NO_3)_2) = 19.6$ $m(H_2O) = 23.2 - 19.6$ $= 3.6$ g
Step 2	Determine the number of moles of each molecule present.	$n = \frac{m}{M}$ $(Mg(NO_3)_2)\ :\ H_2O$ $n(Mg(NO_3)_2) = \frac{m}{M}\ :\ n(H_2O) = \frac{m}{M}$ $n(Mg(NO_3)_2) = \frac{19.6}{148.3}\ :\ n(H_2O) = \frac{3.6}{18}$ $n(Mg(NO_3)_2) = 0.132\ :\ n(H_2O) = 0.200$
Step 3	Ensure that the ratio of salt molecules to water molecules is a whole number ratio. If it is not, divide both numbers by the smallest – thus creating a whole number. **Note: If the number remains as a decimal, multiply by an appropriate number to create whole numbers.** **If the number is 0.5, × 2** **0.3, × 3** **0.25, × 4**	$n(Mg(NO_3)_2) = \frac{0.132}{0.132} : n(H_2O) = \frac{0.200}{0.132}$ $n(Mg(NO_3)_2) = 1\ :\ n(H_2O) = 1.52$ $n(Mg(NO_3)_2) = 1\ :\ n(H_2O) = 1.52 \times 2$ $n(Mg(NO_3)_2) = 2\ :\ n(H_2O) = 3$
Step 4	State the formula for the hydrated salt.	$2(Mg(NO_3)_2).3H_2O$

Gravimetric analysis

Gravimetric analysis is an experimental technique used to measure salt concentration using precipitation reactions and **stoichiometry**. The ions of interest are precipitated out of the solution and collected. The mass of these ions allows us to calculate information about the original solution.

VIDEO 14B–1
PERFORMING GRAVIMETRIC ANALYSIS

Stoichiometry the quantitative relationship between two or more substances during a reaction based on the ratio in which they react

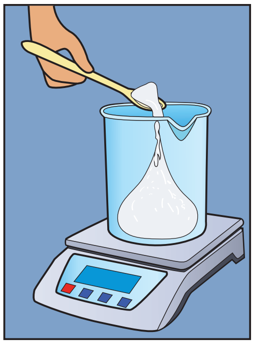
The steps of the process are:

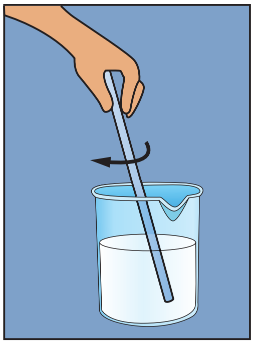
1 Weigh the sample using scales and record the exact mass (in grams).
2 Dissolve the sample in de-ionised water.
3 Add excess of a suitable solution to form a precipitate.

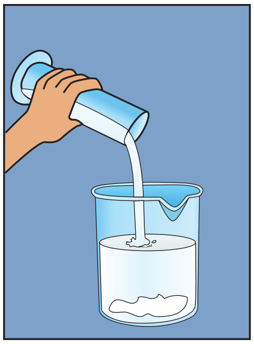
A knowledge of solubility rules is required here. A suitable chemical will contain an ion that will form an insoluble precipitate with the ion of interest.

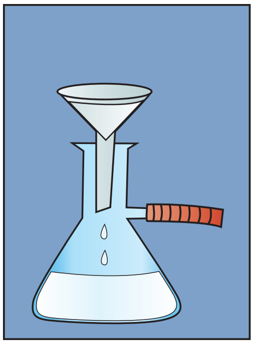
4 Collect the precipitate by filtration. Wash it carefully to ensure that no unwanted soluble species or ions have been trapped by the filter paper.
5 Dry the precipitate carefully by heating, cooling and weighing repeatedly until its mass remains constant, confirming that all water has been removed from the precipitate.

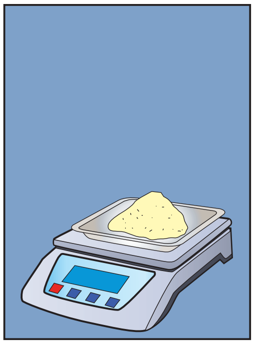
When you have successfully collected the precipitate and recorded its mass, the calculations come next … welcome to STOICHIOMETRY!


(1) Weigh the sample to be analysed


(2) Dissolve the sample in water


(3) Add a suitable chemical to form a precipitate


(4) Filter to collect the precipitate


(5) Repeat drying, cooling and weighing until a constant mass of precipitate is obtained

Check-in questions – Set 2

1 A 3.40 g sample of anhydrous calcium sulfate ($M = 136\ \mathrm{g\,mol^{-1}}$) is formed when 4.30 g of hydrated calcium sulfate is heated to constant mass (and therefore contains no more water). Determine the empirical formula of the hydrated salt.
2 A student wishes to determine the salt content in a packet of potato chips by using gravimetric analysis. List the steps the student would need to undertake.
3 Why must the sample continue to be heated then weighed until a constant mass is achieved?
4 Determine which solution a student could use to precipitate out barium ions from an aqueous sample at 25°C.

Stoichiometry

Stoichiometry allows us to use the amount of one reactant or product to calculate the amount of another reactant or product. This word comes from the Greek words *stoicheion*, meaning 'element' and *metron*, meaning 'measurement'. It is used daily in chemical manufacturing to determine the amount of the desired product made, as well as amounts of potentially damaging by-products. Ultimately, industry wants to maximise the amount of desired product while minimising the amount of waste, as this helps to increase profits and decrease negative impacts on the environment.

Let's look at a reaction to form water:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

The equation tells us that two molecules of hydrogen gas and one molecule of oxygen gas react to produce two molecules of water. However, molecules and atoms aren't useful quantities, so the *mole* was introduced. The equation also tells us that two moles of hydrogen gas and one mole of oxygen gas react to give two moles of water, as shown in Table 14B–1.

Remember, in a chemical reaction, matter cannot be created or destroyed, so:

Total mass of reactants = Total mass of products

Table 14B–1 Stoichiometric calculations for water

		mole × molar mass =	Mass (g)
Reactants	2 H_2	2 × 2	4
	1 O_2	1 × 32	32
Products	2 H_2O	2 × 18	36

Total mass of reactants = 36 g = Total mass of products

Mole ratios

We can define mole ratios between reactants and products. From Table 14B–1, we can now state that:

- for each mole of hydrogen reacted there will be one mole of water produced
- for each mole of oxygen gas reacted there will be two moles of water produced.

Mole ratios form the basis of all stoichiometric calculations.

Mass–mass stoichiometry

The mole ratios, or the coefficients in front of each species, from a balanced chemical equation allow us to calculate the mass of all reactants and products involved in a chemical reaction. Below are the steps involved in mass–mass stoichiometry calculations.

1. Write a balanced equation.
2. Identify the substance that has a known mass and the substance whose mass is to be determined (unknown).
3. Convert *known* quantity into moles using $n = \frac{m}{M}$.
4. Use mole ratio $n(\text{unknown}) = \frac{\text{coefficient of unknown}}{\text{coefficient of known}} \times n(\text{known})$ to calculate amount of *unknown* substance in moles.
5. Calculate the mass of the unknown substance using $m = n \times M$.

Worked example 14B–7: Calculating amount of product from amount of reactant

VIDEO
WORKED
EXAMPLE
14B–7

If 0.010 moles of lead(II) nitrate reacts completely with potassium iodide in aqueous solutions to form a precipitate, determine the amount in moles and grams of the precipitate produced.

Solution

Step 1	Write a balanced equation for the reaction. Identify the KNOWN (*K*) and UNKNOWN (*U*) substances.	$\underset{K}{Pb(NO_3)_2(aq)} + 2KI(aq) \rightarrow \underset{U}{PbI_2(s)} + 2KNO_3(aq)$
Step 2	Calculate the amount (mol) of the KNOWN substance.	$n(\text{known})$ $n(Pb(NO_3)_2) = 0.010 \text{ mol}$
Step 3	Use the mole ratio equation to determine the amount (mol) of the UNKNOWN substance.	$n(\text{unknown}) = \frac{\text{coefficient of unknown}}{\text{coefficient of known}} \times n(\text{known})$ $n(PbI_2) = \frac{1}{1} \times 0.010 = 0.010 \text{ mol}$
Step 4	Calculate the required quantity (mass/vol/conc.) of the UNKNOWN.	$m(PbI_2) = n \times M$ $= 0.010 \times 461$ $= 4.6 \text{ g}$

Worked example 14B–8: Calculating mass of reactant from mass of product

VIDEO
WORKED
EXAMPLE
14B–8

Aqueous solutions of silver nitrate and calcium chloride react to produce 1.2 g of silver chloride precipitate. Calculate the mass of calcium chloride reactant that was present.

Solution

Step 1	Write a balanced equation for the reaction. Identify the KNOWN (*K*) and UNKNOWN (*U*).	$2AgNO_3(aq) + \underset{U}{CaCl_2(aq)} \rightarrow Ca(NO_3)_2(aq) + \underset{K}{2AgCl(s)}$
Step 2	Calculate the amount (mol) of the KNOWN substance. **Remember to only use the chemical formula of the species and not the coefficients, as these are used for the mole ratio steps only (Steps 3–4).**	$n(\text{known})$ $n(AgCl) = \frac{m}{M}$ $= \frac{1.2}{143.4}$ $= 0.00837 \text{ mol}$
Step 3	Use the equation to determine the ratio: $\frac{U\text{ (unknown)}}{K\text{ (known)}}$	$\frac{U}{K} = \frac{CaCl_2}{AgCl} = \frac{1}{2}$
Step 4	From the mole ratio, calculate the amount (mol) of the UNKNOWN substance.	$n(\text{unknown}) = \frac{\text{coefficient of unknown}}{\text{coefficient of known}} \times n(AgCl)$ $= \frac{1}{2} \times 0.00837$ $= 0.00418 \text{ mol}$
Step 5	Calculate the required quantity (mass) of the UNKNOWN.	$m(CaCl_2) = n \times M$ $= 0.00418 \times 111.1$ $= 0.46 \text{ g}$

Percentage mass calculations

Percentage mass the mass of a component represented as a percentage of the total mass of the sample

Once the mass of salt in a sample has been calculated, another way to represent this concentration is as a percentage, known as a **percentage mass**. This is the percentage of the salt in the total mass of a solution that contains the salt. If we were to add up all the components present in the solution, they would add to 100%. It is a useful way to measure and compare levels of salts in water that is easily understood by people who are not chemists.

To calculate the percentage mass of a salt present, we need to know the mass of the salt and the mass of the sample from which it was taken, using the following equation:

$$\% \text{ Mass} = \frac{\text{Mass of salt}}{\text{Total mass of solution}} \times 100$$

Worked example 14B–9: Calculating percentage mass

VIDEO WORKED EXAMPLE 14B–9

A 5.00 g sample of water from the Dead Sea was found through quantitative analysis to contain 1.70 g of sodium chloride. Calculate the percentage mass of sodium chloride in this sample.

Solution

Step 1	Identify the information given.	$m(\text{NaCl}) = 1.70$ g $m(\text{sample}) = 5.00$ g
Step 2	Determine $\% m(\text{salt})$ using the correct formula	$\%\text{mass} = \frac{\text{mass of NaCl}}{\text{total mass of solution}} \times 100$
Step 3	Input the data given into the formula.	$\% m(\text{NaCl}) = \frac{1.70}{5.00} \times 100$
Step 4	Calculate the percentage mass.	$\% m(\text{NaCl}) = 34.0\%$

WORKSHEET 14B–1 MASS–MASS STOICHIOMETRY

Check-in questions – Set 3

1 An aqueous solution of potassium sulfide solution is mixed with an aqueous solution of silver nitrate at 25°C until no further silver sulfide precipitates. The mass of precipitate obtained is 2.5 g.
   - a Write a full equation for the reaction.
   - b Calculate the amount, in mol, of silver sulfide precipitate.
   - c Using your equation, calculate the amount, in mol, of silver nitrate that reacted.
   - d Calculate the mass of silver nitrate that reacted.

14B SKILLS

Memory association

When completing stoichiometric-related problems in chemistry, you saw in this section that a key component of these is using the mole ratio from a balanced chemical equation.

Finding your own ways to remember steps such as this in processes is important. If these are meaningful to you, then it makes this memory even easier. For example, if you've ever been to the United Kingdom, were born or lived there or even have relatives who live there, then the UK may help you to remember the unknown/known step to help you remember your mole ratio!

Think about other aspects of your chemistry studies where this type of memory technique could be applied.

More than just mass–mass stoichiometry

In this section, the focus was solely on mass–mass stoichiometry.

However, recall that in Chapter 13, you learned about volumetric analysis, which involved another version of stoichiometry. Rather than using the mass of a chemical species to determine the mass of another species in the reaction, volumes were used, using the formula $n = c \times V$.

In Chapter 15, you will apply your knowledge of stoichiometry again to solve calculations related to chemical reactions involving gases, whereby the moles, mass and/or volume of gas may need to be determined.

Therefore, at the completion of Unit 2, you may be required (and will be capable with your developed knowledge) to use calculations for moles, mass and volume within one problem. An example of this is presented on the following page.

LINK **13D** VOLUMETRIC ANALYSIS OF ACIDS AND BASES

LINK **15B** GAS STOICHIOMETRY AND THE GREENHOUSE EFFECT

VIDEO 14B–2 SKILLS: MORE THAN JUST MASS–MASS STOICHIOMETRY

Question:

What mass of sodium phosphate, Na_3PO_4, could be obtained by the neutralisation of 0.0500 L of 2.00 M phosphoric acid, H_3PO_4, with sodium hydroxide, NaOH?

Solution:

Step 1	Write a balanced equation for the reaction. Identify the KNOWN (*K*) and UNKNOWN (*U*).	$\underset{K}{H_3PO_4(aq)} + 3NaOH(aq) \rightarrow \underset{U}{Na_3PO_4(aq)} + 3H_2O(l)$
Step 2	Calculate the amount (mol) of the KNOWN substance.	n(known) $n(H_3PO_4) = c \times V$ $n(H_3PO_4) = 2.00 \times 0.0500$ $n(H_3PO_4) = 0.100$ mol
Step 3	Use the equation to determine the ratio: $\frac{U \text{ (unknown)}}{K \text{ (known)}}$	$\frac{U}{K} = \frac{Na_3PO_4}{H_3PO_4} = \frac{1}{1}$
Step 4	From the mole ratio, calculate the amount (mol) of the UNKNOWN substance.	$n(Na_3PO_4) = \frac{1}{1} \times 0.100$ $= 0.100$ mol
Step 5	Calculate the required quantity (mass/vol/conc.) of the UNKNOWN. **Note: In this case, it is the mass.**	$m(Na_3PO_4) = n \times M$ $= 0.100 \times [(3 \times 23.0) + 31.0 + (4 \times 16.0)]$ $= 16.4$ g

As you can see in the example above, a volume and concentration for one of the species is given in the stem information, but it asks you to determine the mass of another species. It is important that you realise from this that the amount in mol, symbol '*n*', is the same across different formula calculations.

Section 14B questions

1 Write balanced equations showing full symbol formulas, with state symbols, for the following precipitation reactions.
  a An aqueous solution of silver nitrate is added to a solution of sodium chloride at 25°C.
  b An aqueous solution of barium hydroxide is added to a solution of potassium sulfate at 25°C.

2 Write balanced ionic equations, with state symbols, for the following precipitation reactions.
  a $AlCl_3(aq) + K_2CO_3(aq)$
  b $NH_4Br(aq) + AgCH_3COO(aq)$

3 An aqueous solution containing 0.2500 mol of cobalt(II) nitrate is added to an aqueous solution of saturated sodium carbonate at 25°C.
  a Write a balanced equation, showing full symbol formulas.
  b Write this equation as an ionic equation.
  c Calculate the mass of precipitate formed in the reaction.

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4 Sodium reacts with water as follows:

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

One mole of sodium is allowed to react completely with excess water.

- a Calculate the number of moles of water used.
- b Calculate the number of moles of sodium hydroxide produced.
- c Calculate the number of moles of hydrogen gas produced.
- d Calculate the mass of hydrogen gas produced.

5 A mass of 3.66 g of iron(III) bromide is dissolved in water at 25°C. To this solution, an aqueous solution containing sodium sulfide is added until no further precipitate forms.

- a Write an ionic equation for the reaction that occurred.
- b Calculate the mass of precipitate formed.

6 A student wishes to determine the percentage of sodium present in a 30.0 g bag of chips. After crushing and dispersing the chips in water to dissolve any soluble matter, they react the solution with silver nitrate, according to the following equation.

$$NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

A10.0 g mass of silver chloride precipitate was formed. Determine the percentage by mass of sodium in the bag of crisps.

7 A 2.60 g mass of anhydrous magnesium sulfate, $MgSO_4$ ($M = 120.4\ g\,mol^{-1}$), is formed when 5.37 g of hydrated magnesium sulfate is heated to constant mass (and therefore contains no more water). Use this data to determine the x value, a coefficient, in the empirical formula of hydrated magnesium sulfate $MgSO_4.xH_2O$.

8 A student completes a gravimetric analysis to determine the percentage of sulfate ions present in a sample of contaminated dirt. They collect the following results.

Mass of dirt sample	10.0 g
Mass of filter paper	0.80 g
Mass of filter paper and precipitate weighing 1	2.3 g
Mass of filter paper and precipitate weighing 2	2.2 g
Mass of filter paper and precipitate weighing 3	2.1 g
Mass of filter paper and precipitate weighing 4	2.1 g
Mass of filter paper and precipitate weighing 5	2.1 g

- a Explain why the student weighed the precipitate and filter paper five times.
- b The student added an excess of barium nitrate to create a barium sulfate precipitate. Write an ionic equation to represent the reaction.
- c Calculate the mass of sulfate ions present in the precipitate.
- d Calculate the percentage by mass of sulfate in the dirt sample.

14C Determining salt concentration using spectroscopy

Study Design:
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

Glossary:
Calibration curve
Colorimetry
Cuvette
Electromagnetic radiation
Electromagnetic spectrum
Spectroscopy
Standard solution
Ultraviolet–visible (UV–visible) spectrophotometry

ENGAGE

The development of instruments from the first colorimeter

There is a strong chance you've never heard of the colorimeter. The earliest true colorimeter was Clerk Maxwell's colour box in 1860, but the colorimeter is considered to have been invented by Louis J. Duboscq, who made his first colorimeter in 1854 and an improved version in 1868. His colorimeter was not the first to be used, but it was the first to permit the simultaneous colour comparison of two liquids. The Duboscq colorimeter gained popularity after biochemists discovered that, when used with appropriate reagents, it provided an efficient and effective method for identifying substances in bodily fluids.

If you've ever had a blood or urine sample taken and sent to a pathologist, you probably wouldn't have thought about the instrument that produced the test results. These instruments are far more modern and sophisticated than the colorimeter that Duboscq invented. However, they are all based on this idea. Like the colorimeter, these modern instruments test the concentration of specific substances in a solution, such as blood or urine, by measuring its absorbance of a specific wavelength of light.

The early forms of these instruments would have required scientists to interpret data and solve mathematical equations specific to the type of molecule being analysed. Nowadays, a computer does these calculations.

Figure 14C–1 Automated blood sample testing for healthcare screening requires instrumentation that analyses the absorbance of different colour intensity solutions.

EXPLAIN

Measuring salt concentration

Gravimetric analysis is not the only form of quantitative analysis available to chemists. In fact, a range of instruments have been developed to determine the concentration of a salt quickly and accurately in a solution without forming a precipitate. These methods are used to determine things like phosphate levels in waterways, as well as non-salt substances, such as cholesterol, in the blood. Two such instruments are **colorimeters** and **ultraviolet–visible (UV–visible) spectrophotometers**, which are examples of the study of **spectroscopy**. Both measure how light interacts with solutions to determine their concentration. We know that the more concentrated a coloured solution becomes, the darker its colour. As such, intensity of the colour allows us to approximate the solution's concentration, as shown in Figure 14C–2. Using a colorimeter or a UV–visible spectrophotometer allows us to measure the intensity of the colour accurately and therefore determine its concentration.

LINK 14B GRAVIMETRIC ANALYSIS OF SALTS

Colorimetry a type of spectroscopy used to determine the concentration of coloured compounds found in solutions. Light is selected using a coloured filter

Ultraviolet-visible (UV–visible) spectrophotometry a type of spectroscopy used to determine the concentration of coloured compounds found in solutions. Light is selected using a monochromator

Spectroscopy a group of techniques that study how matter absorbs and emits light and other forms of radiation

Electromagnetic radiation a form of energy that is all around us and takes many forms, such as radio waves, microwaves, X-rays and gamma rays. Visible light is a form of electromagnetic energy

Figure 14C–2 The concentration of this green-coloured solution increases from left to right, as seen by the increase in darkness of the solution.

Colorimetry

Before we look at how a colorimeter works, we need to understand how the colour we observe in a solution can be used to help us determine what is present and how concentrated it is.

Light is a form of energy; specifically, it is a type of **electromagnetic radiation**. Visible light makes up a small part of the **electromagnetic spectrum** (Figure 14C-3), which shows the types of radiation according to their energy and wavelength, measured in nanometres (nm). The longer the wavelength, the lower the energy.

VIDEO 14C–1 COLORIMETRY

Electromagnetic spectrum the spectrum of electromagnetic radiation, measured in wavelength, from long radio waves to short gamma rays

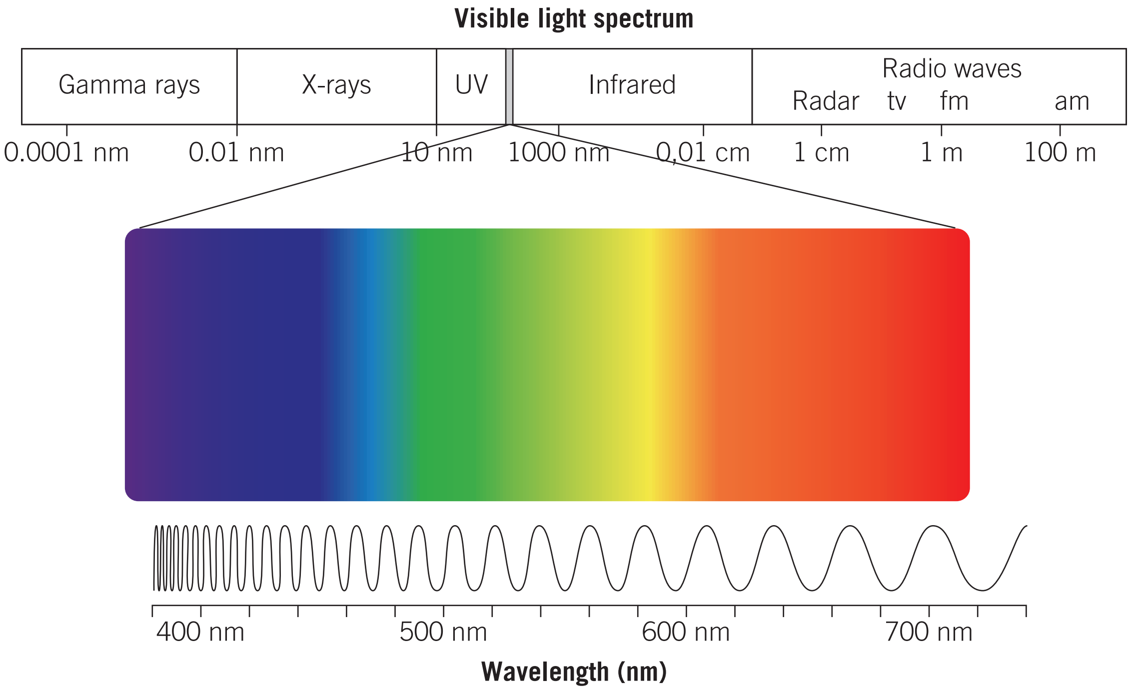


Figure 14C–3 The electromagnetic spectrum. Electromagnetic energy travels in waves, measured in wavelengths, from long radio waves to short gamma rays. We can only see a small portion of the spectrum – visible light.

Energy interacts with atoms. The type of interaction depends on the energy level. When a substance absorbs visible light, it appears coloured. The colour we see in a substance is due to the light that is reflected to our eyes. For example, the leaves of plants are generally green because they contain chlorophyll. Chlorophyll reflects light in the green region of the visible spectrum and absorbs light that is complementary to this – light in the purple and red regions, as shown in Figure 14C–4.

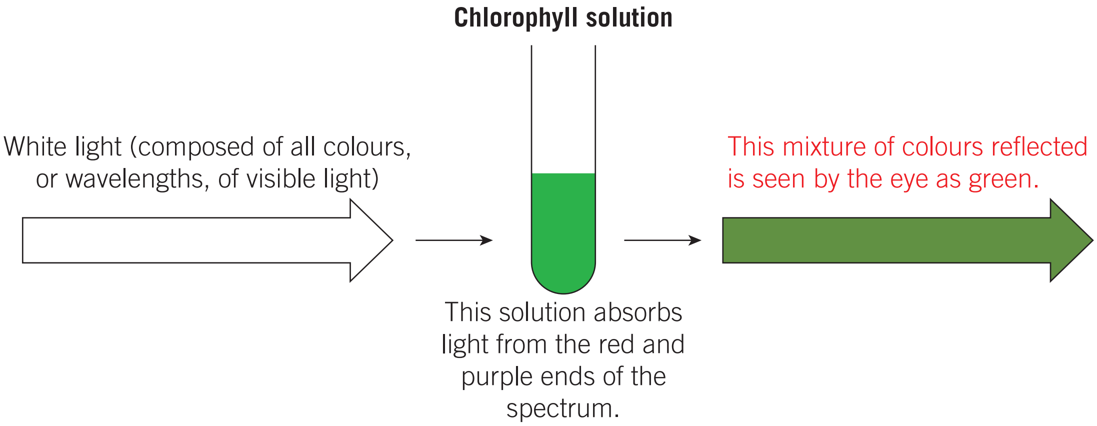
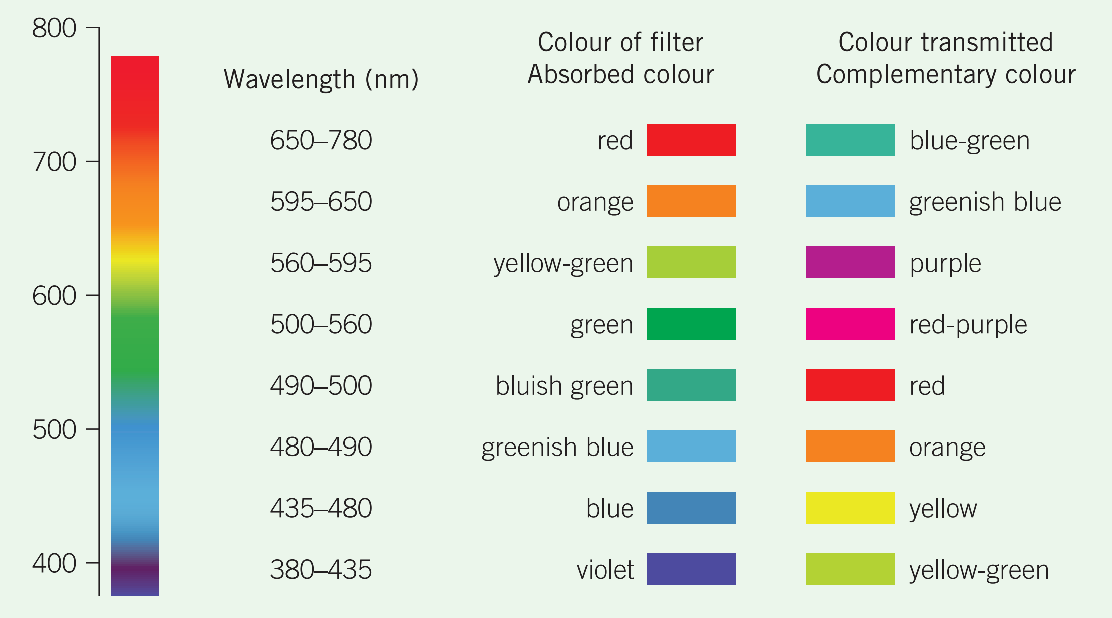


Figure 14C–4 Chlorophyll reflects light in the green region of the visible spectrum and absorbs light that is complementary to this – light in the purple and red regions.

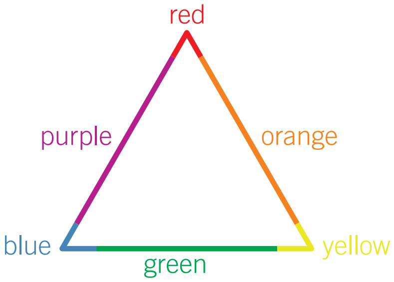
Figure 14C–5 shows some colours transmitted and their complementary colours absorbed at the same time.



Wavelength (nm)	Colour of filter Absorbed colour	Colour transmitted Complementary colour
650–780	red	blue-green
595–650	orange	greenish blue
560–595	yellow-green	purple
500–560	green	red-purple
490–500	bluish green	red
480–490	greenish blue	orange
435–480	blue	yellow
380–435	violet	yellow-green

Figure 14C–5 Visible light has wavelengths between 380 and 780 nm, and different values within this range correspond to different colours of light. The colour absorbed by a substance is complementary to the colour we observe it to be. This knowledge helps us to select an appropriate filter for use in colorimetry.

A colour triangle is helpful to determine the colour absorbed by observing the colour transmitted. For example, if blue is transmitted, using the triangle, the opposite side is the complementary colour – in this case, orange.



How a colorimeter works

Colorimetry is a technique used to determine the concentration of coloured compounds found in solutions according to the Beer–Lambert law, which states that the concentration of a solute is proportional to the amount of light it absorbs. The proportion of light absorbed by a coloured solution will depend on how many coloured solute particles the light interacts with. In a highly concentrated solution, there is an increased number of coloured solute particles present and therefore more particles that interact with the light passing through the sample. In a dilute sample, the absorbance will be low due to the low concentration of coloured solute particles present to interact with the light. A colorimeter gives a reading of the amount of light absorbed by a solution.

A colorimeter only works for coloured compounds. It is possible to treat transparent or non-coloured solutions with another chemical compound to produce a coloured complex that can then be analysed.

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Cuvette a small rectangular tube with two sides that are transparent to UV and visible light

A colorimeter consists of the following components, as shown in Figure 14C–6:

- a light source
- a coloured filter
- a **cuvette** containing the coloured sample
- an electronic detector that measures and records the absorbance of light that passes through the sample contained in the cell.

The colour of the selected filter should be complementary to the colour of the solution. Refer to Figure 14C–5 to determine what colour filter is used for each different coloured solution.

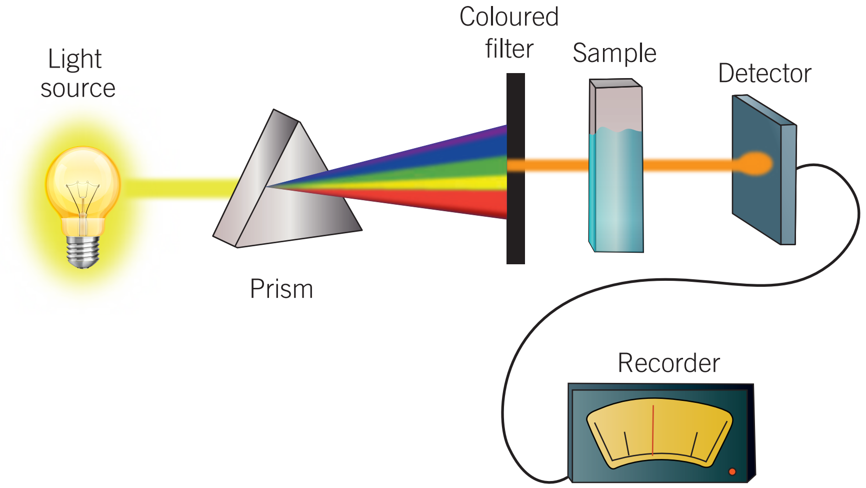

Figure 14C–6 Diagram of the components of a colorimeter. Light is first passed through a prism, which converts light from the source into a spectrum of all wavelengths of light. The light then passes through a coloured filter, selected to be appropriate for the sample, before passing through the sample itself. The detector reads and records the amount of light absorbed by the sample.

Standard solution a solution with a known concentration

The higher the absorbance recorded, the higher the concentration of the solute and therefore the solution. To determine the actual concentration of a coloured solute in a solution, a series of **standard solutions**, solutions with a known concentration, must be prepared and their absorbance measured. Creating and using calibration curves to determine concentration is discussed later in this chapter.

Check-in questions – Set 1

1 Is colorimetry based on the amount of light being reflected or absorbed by the sample?

2 Two samples, Sample A and Sample B, of purple potassium permanganate solution were analysed using a colorimeter.

   a As the solution is purple, which colour filter should be selected to conduct the analysis?

   b Sample A had a concentration of 0.5 M, while sample B had a concentration of 1.5 M. Which sample will absorb the most light?

UV–visible spectroscopy

A UV–visible spectrophotometer is a more sophisticated instrument used to determine the concentration of a solute in a solution by measuring how much light of a particular wavelength is absorbed. It uses a monochromator instead of a filter to select an exact wavelength to be used in the analysis, as shown in Figure 14C–7.

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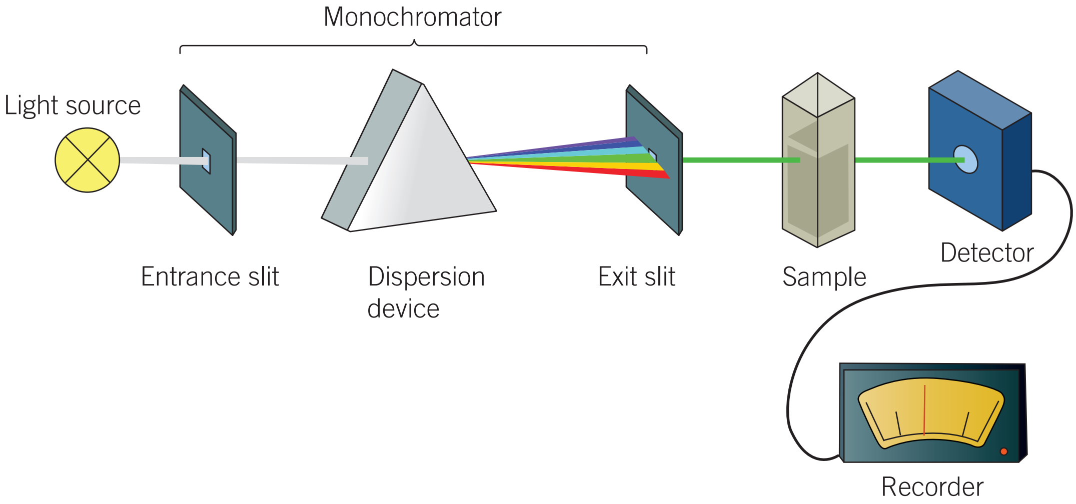


Figure 14C–7 Diagram of the components of a UV–visible spectrophotometer. Light is first passed through a monochromator, which selects an exact wavelength of light to pass through the sample. The detector reads and records the amount of light absorbed by the sample.

When performing UV–visible spectroscopy, the solution to be tested is first scanned across several wavelengths to select the most appropriate wavelength to use when determining its concentration. For example, before testing for the concentration of potassium permanganate, a scan of several wavelengths would be run and wavelengths where absorption is high would then be selected when determining concentration. Figure 14C–8 shows that, for potassium permanganate, a wavelength of 550 nm (the green region from the electromagnetic spectrum in Figure 14C–5) should be used.

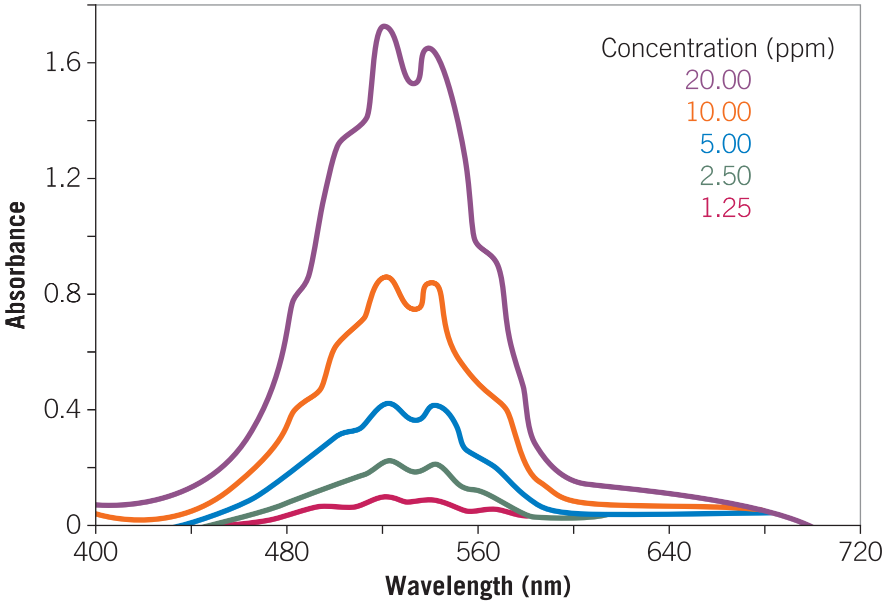


Figure 14C–8 Potassium permanganate is a crystalline salt; it easily dissolves in water and produces a purple coloured solution. The UV–visible spectrum shows that potassium permanganate absorbs best at a wavelength of 550 nm. As the concentration of potassium permanganate increases, so does the absorption of this wavelength.

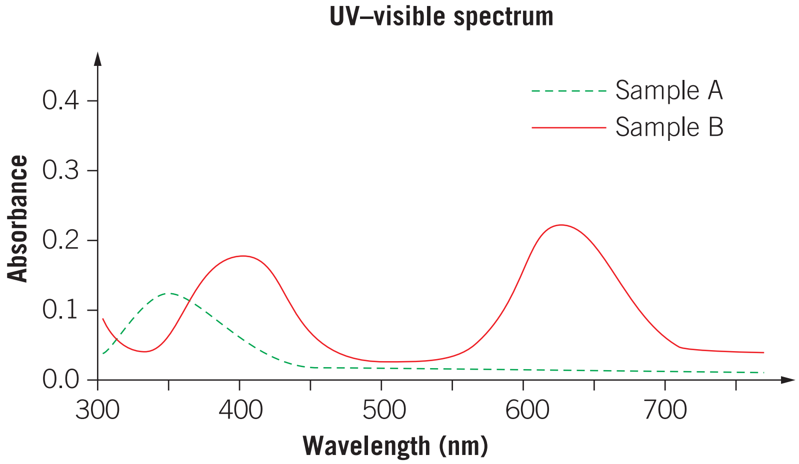
VIDEO 14C–2
UV–VISIBLE SPECTROSCOPY

ISBN 978-1-009-22964-7 Drummond et al

Check-in questions – Set 2

1 What is the advantage of using a UV–visible spectrophotometer rather than a colorimeter?

2 UV–visible spectroscopy was used to measure the spectra of two polluted water samples. The absorption spectra of the two samples are shown below.



  a Using the spectra, determine what wavelength(s) of light is/are absorbed by each solution.
  b Predict what colour solution B would appear.

Creating and using a calibration curve

VIDEO 14C–3 CALIBRATION CURVES

Calibration curve a graph created by plotting the absorbance of a series of standard solutions against their respective concentration

To determine the concentration of a substance in a solution using one of the spectroscopy techniques, a **calibration curve** must be created. A calibration curve plots the absorbance, at a particular wavelength, of a series of standard solutions with increasing concentration. The amount of light the unknown sample absorbs can then be measured and the calibration curve used to determine the solution's unknown concentration. The steps below outline the procedure:

1 Prepare a set of standards of known concentration of the ion of interest.
For example, if you are determining the concentration of a solution containing copper ions, prepare standard solutions of copper(II) sulfate solutions with concentrations of 0.0 M, 0.2 M, 0.4 M, 0.6 M, 0.8 M and 1.0 M.

2 Select an appropriate wavelength for the experiment.

3 Measure the amount of light absorbed by each sample. An example is shown in Table 14C–1.

Table 14C–1 Absorbance of each sample of different copper sulfate concentrations

Concentration of copper(II) sulfate (M)	Absorbance
0.0	0.0
0.2	0.12
0.4	0.24
0.6	0.36
0.8	0.48
1.0	0.60
Sample	0.30

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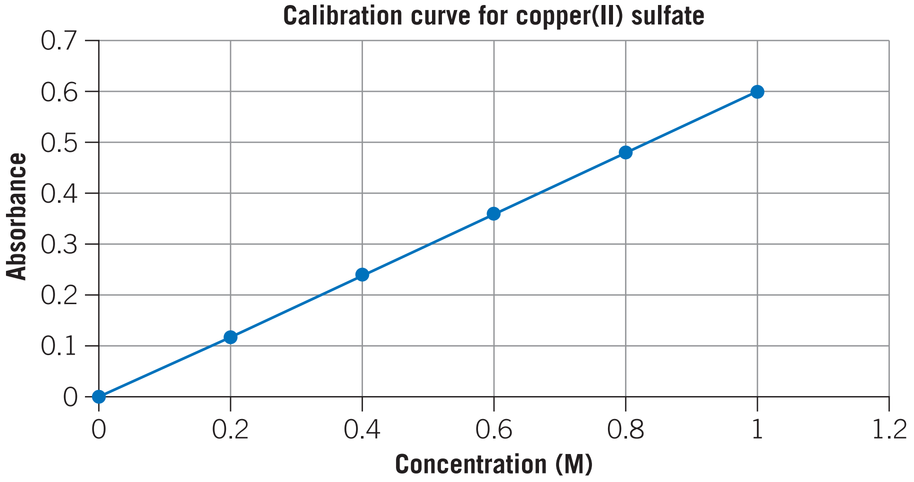
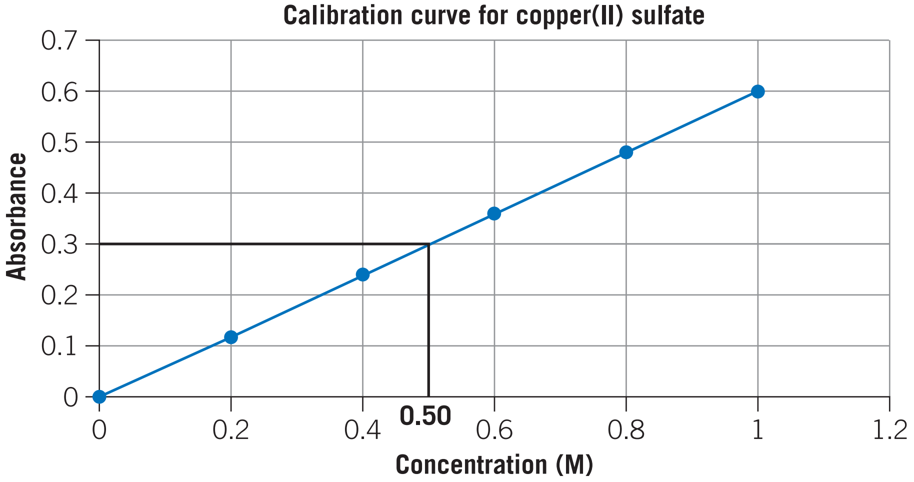
4 Plot a calibration curve showing the absorbance of the different concentrations of copper(II) sulfate solution.



5 Determine the concentration of the sample using its absorbance (0.30), from the calibration curve.



Therefore, the concentration of copper ions in the unknown sample of copper(II) sulfate is 0.50 M.

Units of concentration – review

Concentration is a measure of the quantity of solute present in a solution and there are many different units in which this can be measured. Sometimes we will need to convert these units from one to another to allow an accurate comparison of levels to be made. Worked example 14C–1 is one such case.

LINK **13A** MEASURES OF SOLUBILITY

Worked example 14C–1: Calculating concentration from absorbance

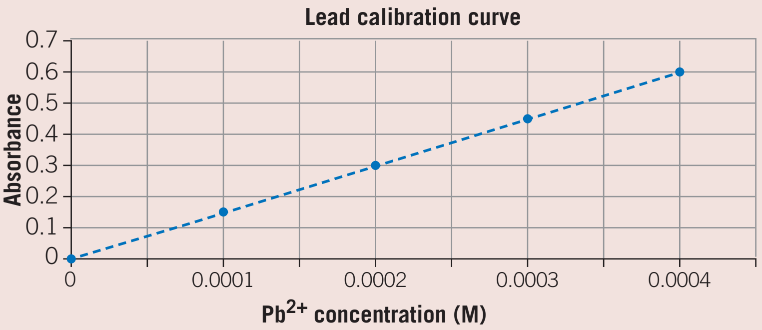
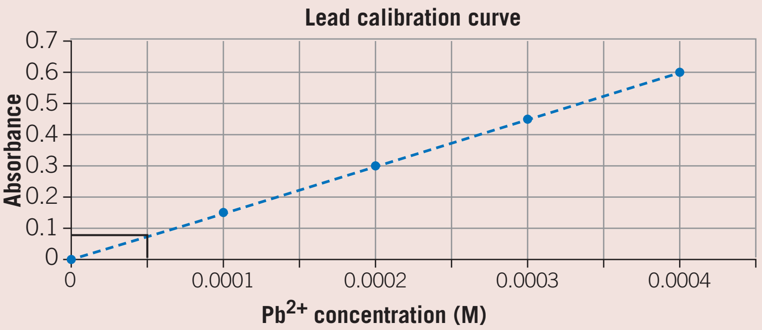
VIDEO
WORKED
EXAMPLE
14C–1

To determine if bore water was safe for drinking, a UV–visible spectrophotometer was used to measure the concentration of lead present. The World Health Organization states that levels of lead must be less than 0.01 ppm.

The absorbance of four standard solutions of lead, in addition to the water sample itself, were measured. Results are shown in the table below. Determine the concentration (ppm) of lead in the bore water.

Concentration (M)	Absorbance
0.00010	0.15
0.00020	0.31
0.00030	0.44
0.00040	0.59
Sample	0.08

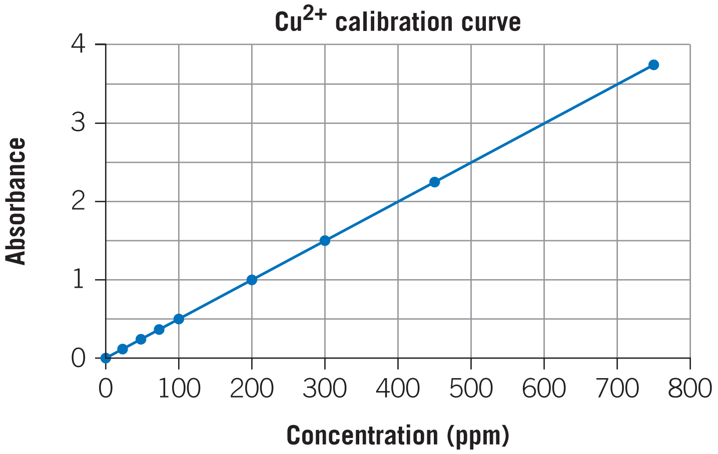
Solution

Step 1	Using the data from the table, plot a calibration curve.	
Step 2	Determine the concentration of the sample, using its absorbance, from the calibration curve.	 $[Pb^{2+}] = 0.00005$ M
Step 3	Convert molar concentration to ppm. Remember ppm = mg L^{-1}	Convert molar concentration to g L^{-1} by multiplying by the molar mass of lead: $0.00005 \text{ M} \times 207.2 \text{ g mol}^{-1} = 0.01036 \text{ g L}^{-1}$ Convert g L^{-1} to ppm: $0.01036 \text{ g L}^{-1} \times 1000 = 10.36 \text{ mg L}^{-1} = 10.36$ ppm The bore water contains dangerously high levels of lead contamination.

Check-in questions – Set 3

WORKSHEET 14C–1
ANALYSING UV–VISIBLE SPECTROSCOPY

1 A calibration curve was constructed using UV–visible spectroscopy of varying concentrations of copper(II) ions. The calibration curve is shown below:



a Use the calibration curve to determine the concentration of copper(II) ions in ppm of an unknown solution with an absorbance of 2.0.

b Use the calibration curve to determine the concentration of copper(II) ions in $mol\,L^{-1}$ of an unknown solution with an absorbance of 0.75.

14C SKILLS

Plotting a line of best fit by hand

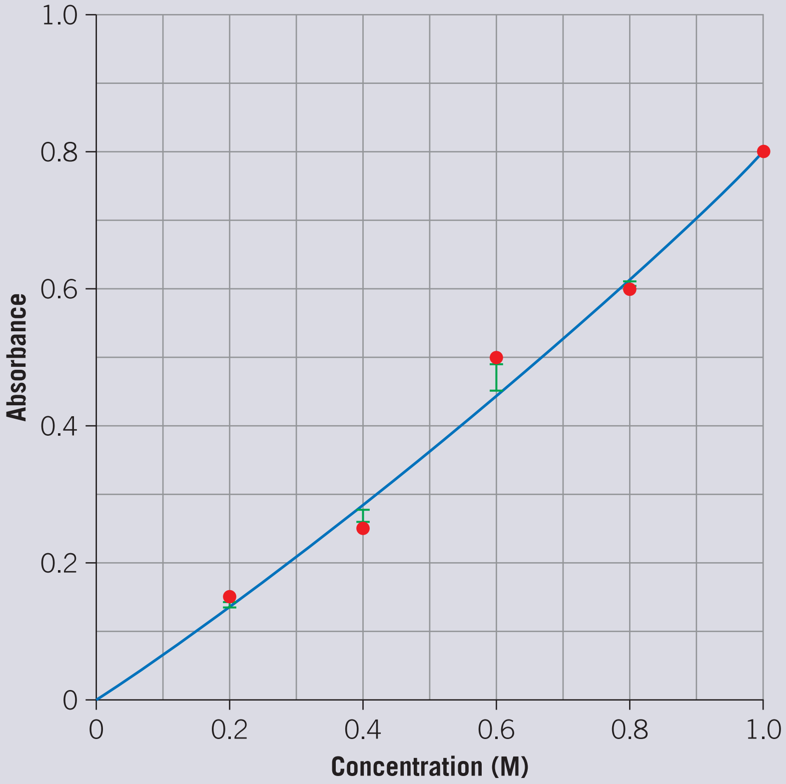
This section focused on the importance of collecting absorbance data for a range of concentrations of a salt, or ion. These are called a set of standard solutions. Graphing this data then allows the chemist to generate a calibration curve. For standard solutions developed for use in both colorimetry and UV–visible spectroscopy, the name is slightly misleading. These graphs are not actually ‘curves’ per se but instead are straight lines, because typically as the concentration of a salt increases, so too does its absorbance. As you saw in this chapter, this is generally a linear relationship.

It is important to note that any time instrumentation such as a colorimeter or UV–visible spectrophotometer is used, there are small opportunities for errors in readings. This means that the absorbance data collected for different concentrations in the standard solutions may not result in a perfect linear relationship.

If this is the case, it is important when drawing your graph to plot a line of best fit as opposed to connecting the data points. If you are constructing your graph digitally, software will have functions that allow you to select ‘line of best fit’ when plotting data. If you are constructing your graph by hand, however, you will need to do it by eye, which can be quite difficult. Your aim here is to ensure that if you were to measure the distance between the data points below the line of best fit and the line itself, they would be equal to the distance between the data points above the line of best fit and the line itself. As you are unlikely to get this perfect on your first attempt, it is crucial that you draw this line of best fit in pencil (as you should for all graphs anyway).

The following data and calibration curve show an example of how to plot a line of best fit by hand.

Concentration (M)	Absorbance
0.0	0
0.2	0.15
0.4	0.25
0.6	0.50
0.8	0.60
1.0	0.80



You can see from the green bars on the calibration curve that the distance of points above the line and below the line from the curve are roughly equal, so this is a good attempt at a calibration curve by hand.

Section 14C questions

1 A solution containing sodium dichromate is orange in colour. Refer to Figure 14C–5 to select an appropriate filter for use in measuring the absorbance of this solution in colorimetry.

2 When analysing a mixture of chemical components, describe the benefit of scanning a wide range of wavelengths in UV–visible spectrophotometry.

3 A sample of lake water was found to contain 1.00×10^{-4} mol L^{-1} of calcium ions. Calculate the calcium ion concentration in the sample in g L^{-1}.

4 A saturated calcium carbonate, or 'limewater', solution contains 1.5 g L^{-1} of calcium carbonate in water. Calculate this concentration of calcium carbonate in percentage by mass – %(m/v).

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5 The World Health Organization (WHO) permissible limit of Al^{3+} in drinking water is 7.4×10^{-6} mol L^{-1}. A sample of Victorian drinking water was analysed for Al^{3+} content by comparing it to a calibration curve produced using UV–visible spectrophotometry.

  a Construct the calibration curve from the data in the table below.

Concentration (ppm)	Absorbance
0.00	0
0.05	120
0.10	245
0.15	380
0.20	505
0.25	630

  b Determine the concentration of Al^{3+} in ppm.

  c Calculate the concentration of Al^{3+} in mol L^{-1} and state whether the concentration is less than the WHO permissible limit.

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Chapter 14 review

Summary

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

Success criteria – I am now able to:		Linked question
14A.1	Recall the formation of ions and use of solubility rules to predict precipitates	1☐, 2☐, 5☐
14A.2	Understand the various sources of salts in water samples	3☐
14A.3	Measure levels of salinity using a conductivity meter	4☐, 11☐
14B.1	Recall how to write a balanced equation for a precipitation reaction	12☐, 13☐
14B.2	Recall how to perform mole calculations	6☐, 15☐
14B.3	Determine the molar ratio of water of hydration for an ionic compound	17☐, 18☐
14B.4	Define ‘gravimetric analysis’ and state what it is used for	7☐, 8☐
14B.5	Explain the steps required in the process of quantitative analysis	7☐, 8☐
14B.6	Identify mole ratios between chemicals	16☐
14B.7	Apply (mass–mass) stoichiometry to calculate chemical quantities	6☐, 16☐
14B.8	Conduct a quantitative analysis in the laboratory	
14C.1	Understand and apply the principles of colorimetry	10☐, 19☐
14C.2	Understand and apply the principles of UV–visible spectroscopy	9☐
14C.3	Recall concentration calculations	14☐, 20☐, 21☐
14C.4	Convert between units of concentration	20☐, 21☐
14C.5	Create and analyse a calibration curve using primary or secondary data	20☐

Multiple-choice questions

1 Which one of the following compounds is insoluble in water?
- **A** calcium bromide
- **B** silver bromide
- **C** potassium bromide
- **D** sodium bromide

2 A precipitate is formed when a concentrated aqueous solution of sodium chloride at 25°C is added to a solution of

A barium hydroxide.
B sodium sulfate.
C lead nitrate.
D potassium nitrate.

3 Which of the following are sources of salts in water samples?

A minerals
B heavy metals
C organometallic substances
D all of the above

4 Which aqueous solution of NaCl would have the greatest electrical conductivity?

A 1 M
B 100 ppm
C $0.1\ \mathrm{g\,L^{-1}}$
D 0.1 M

5 To analyse the salt (NaCl) content of a river water sample using gravimetric analysis, an excess of which reagent could be used to produce a precipitate to collect, dry and weigh?

A ammonium sulfate $(NH_4)_2SO_4$
B silver nitrate ($AgNO_3$)
C potassium hydroxide (KOH)
D sodium carbonate (Na_2CO_3)

6 The lead (Pb^{2+}) content of a river water sample was analysed using gravimetric analysis. Excess potassium hydroxide (KOH) was added and 1.25 g of lead hydroxide ($Pb(OH)_2$) precipitate was collected. The ionic equation below represents the reaction occurring for the formation of the precipitate.

$$Pb^{2+}(aq) + 2OH^-(aq) \rightarrow Pb(OH)_2(s)$$

Calculate the mass of lead in the water sample.

A 0.0052 g
B 2.15 g
C 0.54 g
D 1.07 g

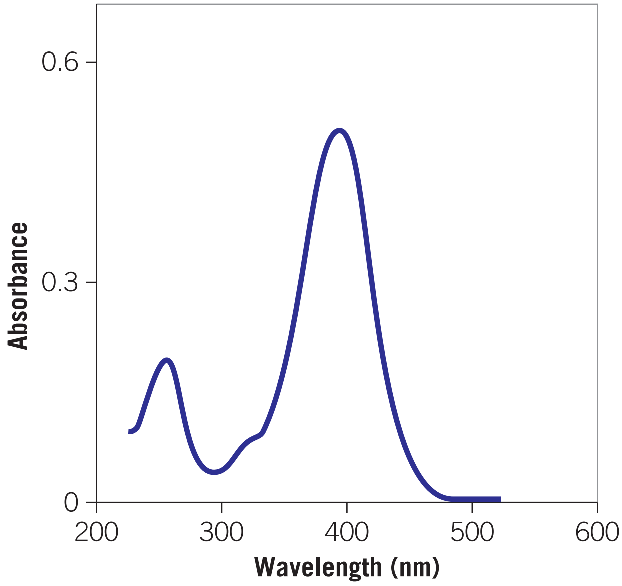
7 During gravimetric analysis, why is it important to reweigh the precipitate until constant mass is achieved?

A to ensure that the scales have equilibrated
B to ensure that the sample is completely dried with all moisture removed
C to eliminate the error associated with air flow
D to ensure that the scales are accurate

8 In gravimetric analysis, if constant mass was not achieved, the percentage mass calculated for the precipitate would

A be higher.
B be lower.
C be unaffected.
D need to be averaged over three concordant (consistent) measurements.

9 The diagram below shows the UV–visible spectrum for a material to be analysed using UV–visible spectroscopy.



The most appropriate wavelength of light to conduct the experiment at would be

A 395 nm.
B 250 nm.
C 500 nm.
D 475 nm.

10 A dilute aqueous solution of potassium dichromate at 25°C has a yellow colour. Which of the following is an incorrect statement?

A It will both absorb and transmit light in the yellow region of the visible spectrum.
B It will absorb light in the blue region of the visible spectrum but remain yellow.
C It will absorb more light as the concentration of the potassium dichromate increases.
D The percentage of light that passes through the sample will increase as the concentration decreases.

Short-answer questions

11 Use a diagram to show the movement of ions in a saltwater solution when a conductivity meter is inserted. (2 marks)

12 Write balanced equations showing full symbol formulas, with state symbols, for the following precipitation reactions.

a A solution of aluminium ethanoate is added to a solution of ammonium phosphate. (2 marks)
b A solution of nickel(II) sulfate is added to a solution of sodium carbonate. (2 marks)

13 Write balanced ionic equations, with state symbols, for the following precipitation reactions.

a An aqueous solution of sodium sulfide is added to an aqueous solution of vanadium(V) iodide at 25°C. (2 marks)
b An aqueous solution of copper(II) nitrate is added to an aqueous solution of potassium hydroxide at 25°C. (2 marks)

14 You have a 200.0 mL sample of river water containing 0.50 mg L^{-1} of NaCl. If you add 500.0 mL of water, what will be the concentration of NaCl in the diluted solution? (2 marks)

15 A sample of sewage water has a high concentration of chloride ions. All of the chloride in the sample was reacted with an aqueous solution containing silver ions to produce 929 mg of silver chloride precipitate. Calculate the moles of chloride ions in the original sewage water sample. (2 marks)

16 A fertiliser is analysed to determine its percentage by mass of phosphate. All of the phosphate in a 10.5 g sample of the fertiliser is dissolved in water and precipitated with barium ions. The precipitate was weighed to be 11.4 g. Determine the percentage by mass of phosphate in the fertiliser. (2 marks)

17 A hydrated magnesium chloride compound has the formula $MgCl_2.xH_2O$. An 80.50 g sample of this compound was heated to remove the water molecules as water vapour, which left 51.53 g of magnesium chloride powder remaining. Determine the value of x and the empirical formula of the hydrated compound. (4 marks)

18 A hydrated zinc(II) carbonate compound has the general formula, $xZnCO_3.yH_2O$. A 5.650 g sample of the compound was heated successively until a weight of 4.648 g of zinc carbonate powder remained constant. Determine the values of x and y and the empirical formula of the hydrated compound. (4 marks)

19 Colorimetry uses the relationship between the concentration of a coloured species and the intensity of the colour of the solution. Explain this statement. (1 mark)

20 A scientist recorded the following absorbance data using nickel(II) ion standard solutions using a UV–visible spectrophotometer.

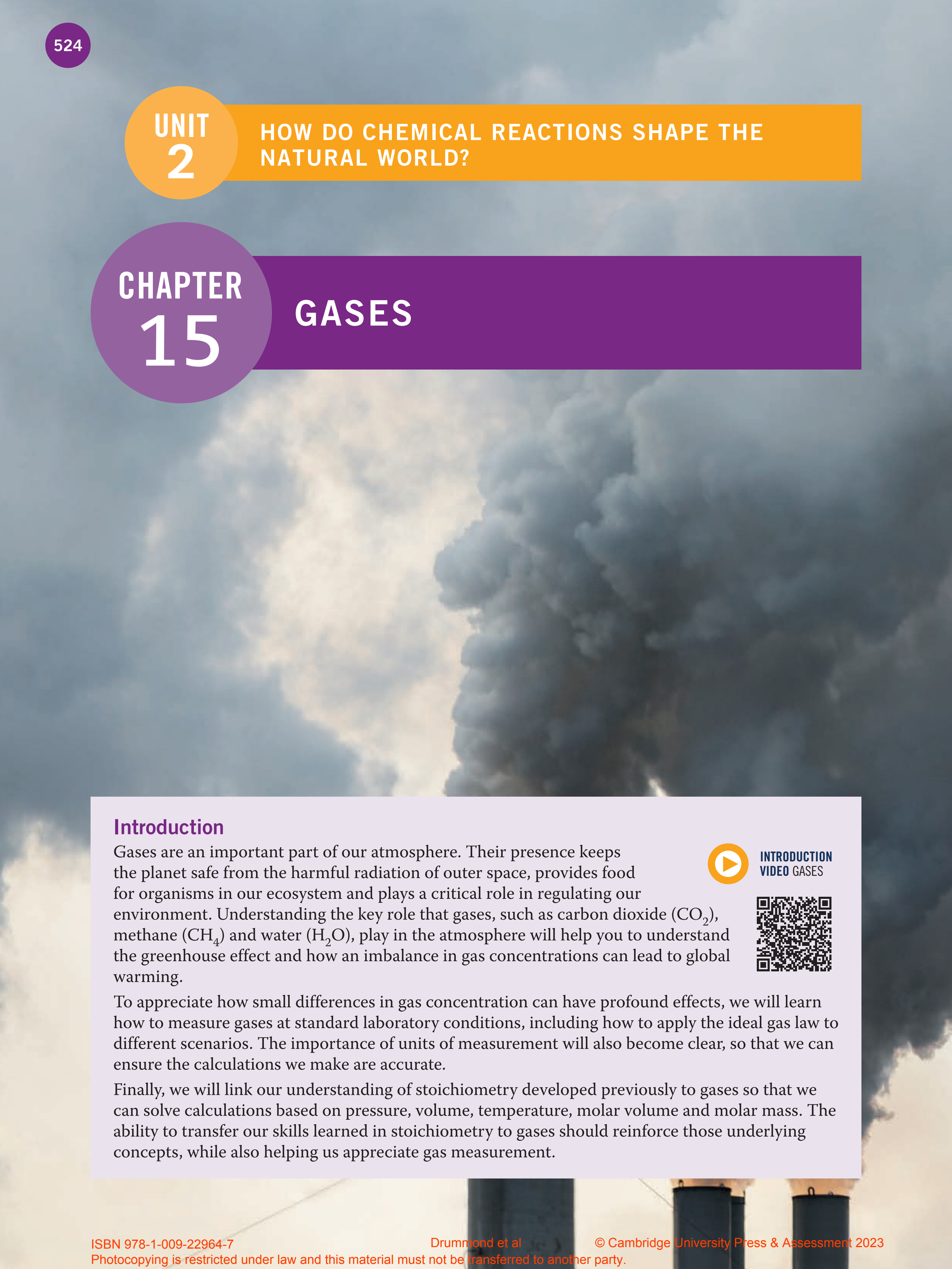
Concentration of Ni^{2+} (ppm)	Absorbance
0	0
1	0.35
2	0.7
5	1.75
9	3.15
11	3.85
13	4.55
15	5.25

a Construct a calibration curve using this data. (4 marks)

b An unknown diluted water sample containing Ni^{2+} was measured to have an absorbance value of 2.8. Using your calibration curve, determine the concentration of Ni^{2+} in this sample in ppm, $g\,L^{-1}$ and $mol\,L^{-1}$. (3 marks)

c This water sample was diluted from the original sample by a factor of 10. Determine the Ni^{2+} concentration in the original water sample, in $mol\,L^{-1}$. (2 marks)

21 A 275 mL sample of groundwater was found to contain lead(II) ions in a concentration of 0.0016 ppm. Determine the concentration of lead(II) ions in $g\,L^{-1}$ and $mol\,L^{-1}$. (2 marks)



UNIT 2

HOW DO CHEMICAL REACTIONS SHAPE THE NATURAL WORLD?

CHAPTER 15

GASES

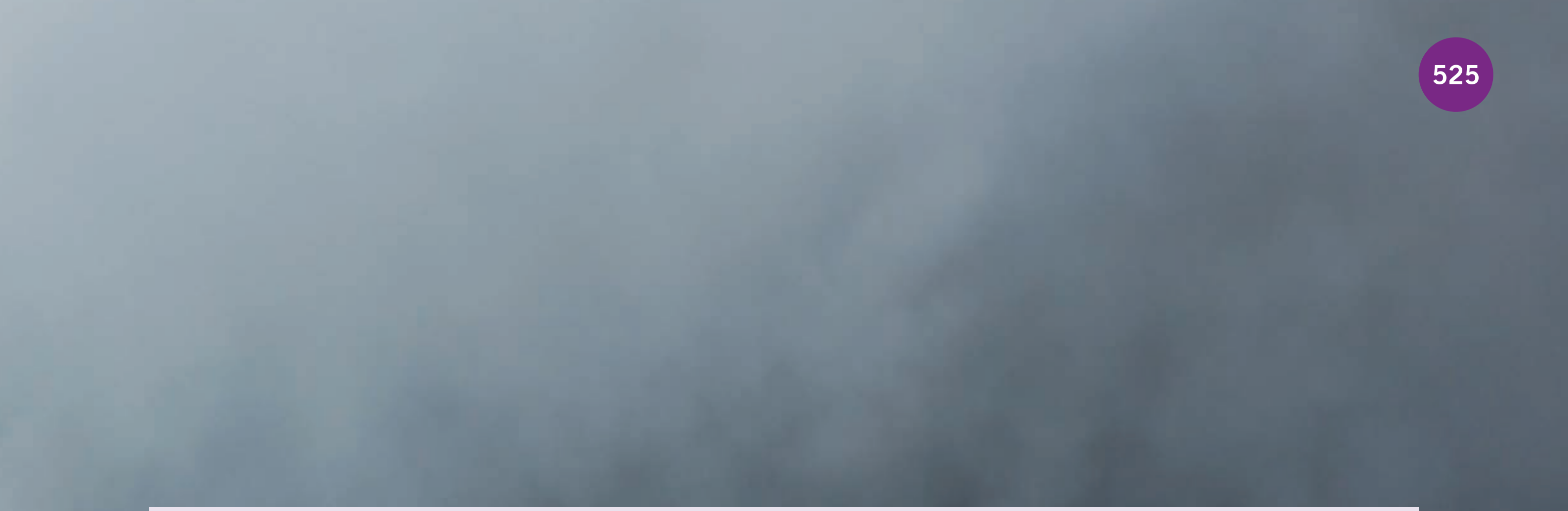
Introduction

INTRODUCTION VIDEO GASES

Gases are an important part of our atmosphere. Their presence keeps the planet safe from the harmful radiation of outer space, provides food for organisms in our ecosystem and plays a critical role in regulating our environment. Understanding the key role that gases, such as carbon dioxide (CO_2), methane (CH_4) and water (H_2O), play in the atmosphere will help you to understand the greenhouse effect and how an imbalance in gas concentrations can lead to global warming.

To appreciate how small differences in gas concentration can have profound effects, we will learn how to measure gases at standard laboratory conditions, including how to apply the ideal gas law to different scenarios. The importance of units of measurement will also become clear, so that we can ensure the calculations we make are accurate.

Finally, we will link our understanding of stoichiometry developed previously to gases so that we can solve calculations based on pressure, volume, temperature, molar volume and molar mass. The ability to transfer our skills learned in stoichiometry to gases should reinforce those underlying concepts, while also helping us appreciate gas measurement.



Curriculum

Area of Study 2 Outcome 2
How are chemicals measured and analysed?

Study Design:	Learning intentions – at the end of this chapter I will be able to:
• 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)	**15A Gases and the ideal gas equation** **15A.1** Recall the properties of gases **15A.2** Define 'gas pressure' and convert between kPa, Pa and atm **15A.3** State the standard laboratory conditions (SLC) **15A.4** Perform conversions between the volume units mL and L **15A.5** Perform conversions between the temperature units °C and K **15A.6** Calculate gas volume or amount in moles using the molar volume of a gas **15A.7** Perform basic calculations using the ideal gas law
• Calculations of the molar volume or molar mass of a gas produced by a chemical reaction • The use of stoichiometry to solve calculations related to chemical reactions involving gases (including moles, mass and volume of gases) • 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	**15B Gas stoichiometry and the greenhouse effect** **15B.1** Calculate the molar volume of a gas produced by a chemical reaction **15B.2** Calculate the molar mass of a gas produced by a chemical reaction **15B.3** Perform calculations on a chemical reaction involving gases using moles, mass and volume **15B.4** Compare and explain the natural and enhanced greenhouse effects **15B.5** Describe the role CO_2, CH_4 and H_2O play in the greenhouse effect with reference to their ability to absorb infrared radiation

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Glossary

Avogadro's law
Chlorofluorocarbon (CFC)
CO_2 equivalent
Enhanced greenhouse effect
Gas pressure
Gas volume
Greenhouse effect
Greenhouse gas
Ideal gas
Ideal gas equation
Infrared radiation (IR)
Molar volume
Photochemical smog
Standard laboratory conditions (SLC)

Concept map

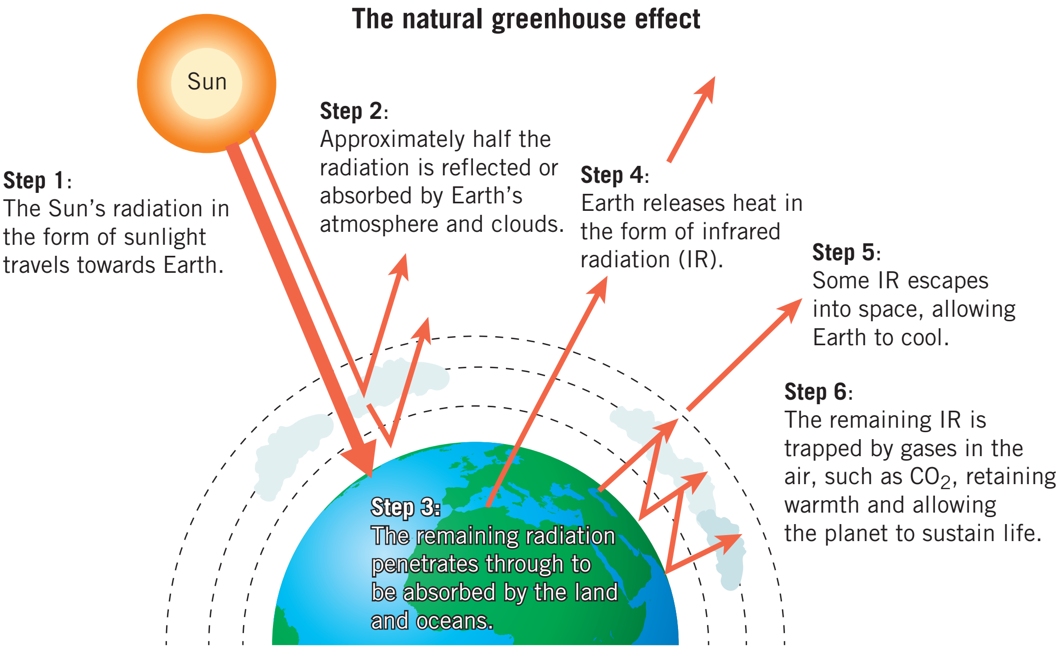
15A Gases and the ideal gas equation

Gases have a huge impact on life on Earth.
To understand the effects of gases we need to be able to:

Have robust methods for measurement

Account for their behaviour

15B Gas stoichiometry and the greenhouse effect



The natural greenhouse effect

Step 1: The Sun's radiation in the form of sunlight travels towards Earth.

Step 2: Approximately half the radiation is reflected or absorbed by Earth's atmosphere and clouds.

Step 3: The remaining radiation penetrates through to be absorbed by the land and oceans.

Step 4: Earth releases heat in the form of infrared radiation (IR).

Step 5: Some IR escapes into space, allowing Earth to cool.

Step 6: The remaining IR is trapped by gases in the air, such as CO_2, retaining warmth and allowing the planet to sustain life.

See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

15A Gases and the ideal gas equation

Study Design:

- 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)

Glossary:

Gas pressure
Gas volume
Ideal gas
Ideal gas equation
Standard laboratory conditions (SLC)

ENGAGE

The seven base units to rule them all!

The International System of Units, commonly referred to as the SI, is the world's most widely used system of measurement. It ensures that no matter what equation we're using or what variable we're solving for, we have a commonly used unit. But who governs and decides this? Well, that responsibility belongs to the General Conference on Weights and Measures, who updated the system in 2019 to include seven defining constants that have exact numerical values when expressed in the SI units. In this chapter, accurate measurement is critical in ensuring that the properties of gases are measured properly.

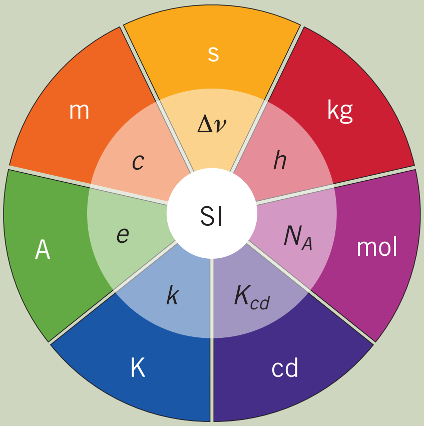


Figure 15A–1 The SI logo showing the seven base units and the defining constants

EXPLAIN

Ideal gas equation

Performing gas calculations requires us to make several assumptions. The first assumption we make is that the gases are behaving 'ideally'. The following must be true for an **ideal gas**:

- The gas takes up the entire volume of its container.
- Collisions between gas particles are elastic and therefore no energy is lost.
- There are no attractive or repulsive forces between the gas particles or between gas particles and the container. (We ignore any forces in the calculations we make.)
- The size of the gas particles is negligible. (The particles themselves have no volume.)

If all these assumptions hold true, then we can consider the gas to be an ideal gas and our calculations, based on information we have about the gas and the system, will be considered to be accurate.

VIDEO 15A–1
THE IDEAL GAS EQUATION

Ideal gas
a gas that obeys the gas laws at all temperatures and pressures

Ideal gas equation
an equation that relates pressure, volume, amount of gas and absolute temperature for an ideal gas

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The calculations are based on the **ideal gas equation**, which is

$$PV = nRT$$

where:

P = The pressure of the system (in kPa)
V = The volume of the system (in litres)
n = The amount of gas (in moles)
T = The temperature of the system (in kelvin)
R = The universal gas constant, which is 8.31 (in $J\ mol^{-1}\ K^{-1}$)

Gas pressure the amount of force exerted by a gas on its container per unit of area

Gas volume the volume of the container occupied by a gas

Gas pressure

The **pressure** of a gas is the force it exerts on its container per unit area. However, as we're usually measuring gases in the atmosphere, it's more appropriate to measure them in terms of their weight/percentage of the total weight of the atmosphere. One atm is equal to the weight of the atmosphere above us. Other common units are Pascals and kilopascals (kPa). Depending on the situation, different units may be used, though it is also easy to go between them using the conversion below:

100 kPa = 100 000 Pa = 0.987 atm

Worked example 15A–1: Converting units of pressure

The pressure of gas found in a container is 132 kPa. Determine the equivalent pressure in Pa and atm.

Solution

Step 1	Determine the conversion factor for each set of units.	100 kPa = 100 000 Pa = 0.987 atm
Step 2	Apply the conversion to 132 kPa.	132 kPa = 132 × 1000 = 132 000 Pa $132\text{ kPa} = \frac{132}{100} \times 0.987$ = 1.30 atm

Gas volume

As gases are defined as filling the entire space of their container, the unit of **volume** often follows the unit of volume most appropriate for a given container. The SI unit for volume is m^3, though an IUPAC-accepted and commonly used alternative is litres (L), where 1 L is equal to one cubic decimetre (dm^3). We will be performing equations using both litres and millilitres, which you can go between using the conversion below:

1 L = 1000 mL

Worked example 15A–2: Converting units of volume

A gas has a volume of 35 mL. Determine the equivalent volume in litres.

Solution

Step 1	Determine the conversion factor for each set of units.	1000 mL = 1 L
Step 2	Apply the conversion to 35 mL.	$35\text{ mL} = \frac{35}{1000} = 0.035\text{ L}$

Gas temperature

The temperature of gases is directly related to their energy, which is something we will learn more about in Unit 3. There are several different measures of temperature, but in this section we are going to focus on the units kelvin (K) and degree Celsius (°C). When converting between units of temperature, the following common conversions, shown in Table 15A–1 and Figure 15A–2, are used.

LINK UNIT 3

Table 15A–1 Converting between degrees Celsius and kelvin

Unit conversion	Formula
Celsius to kelvin °C → K	K = °C + 273
Kelvin to Celsius K → °C	°C = K – 273

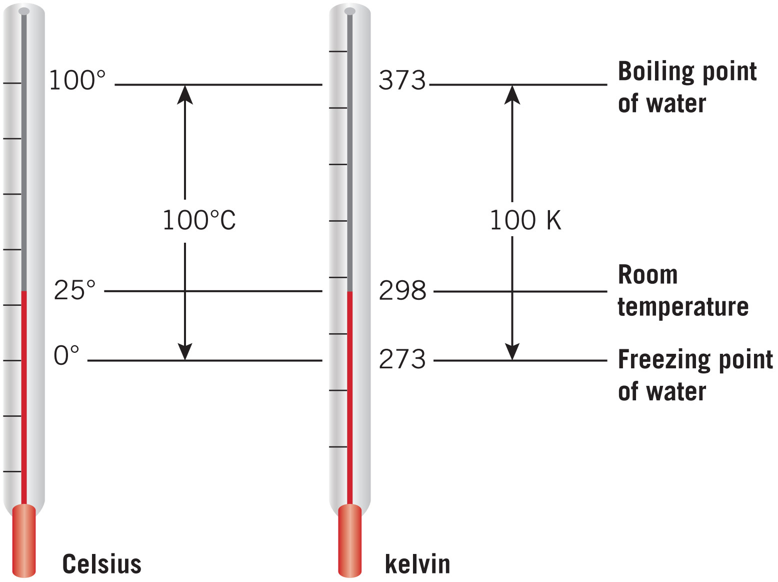

Figure 15A–2 A visual representation of how degrees Celsius and kelvin convert between each other

Worked example 15A–3: Converting units of temperature

A gas has a temperature of 160°C. Determine the equivalent temperature in kelvin (K).

Solution

Step 1	Remember the rule for converting between °C and K.	K = °C + 273
Step 2	Apply the conversion.	K = 160°C + 273 = 433 K

Check-in questions – Set 1

1 Convert the following temperatures to the unit specified.
   - a 25°C to K
   - b 190 K to °C

2 State the ideal gas equation, noting each term and its units.

3 What is the definition of gas pressure?

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4 Convert the following volumes to the unit specified.
- a 0.05 L to mL
- b 2870 mL to L
- c 65 mL to L
- d 1.02 L to mL

Worked example 15A–4: Calculating moles using the ideal gas equation

VIDEO
WORKED
EXAMPLE 15A–4

A 28 L canister of nitrogen gas is being stored at 2 atm in a laboratory cold storage room at a temperature of –23°C. How many moles of nitrogen are in the container?

Solution

Step 1	Extract the information from the question.	$P = 2$ atm = 202.6 kPa $V = 28$ L $T = -23$°C = 250 K $R = 8.31$ $n = ?$
Step 2	Insert values into the ideal gas equation.	$PV = nRT$ $202.6 \times 28 = n(250 \times 8.31)$
Step 3	Solve for *n*.	$n = \frac{PV}{RT}$ $= \frac{202.6 \times 28}{8.31 \times 250}$ $= \frac{5672.8}{2077.5}$ $= 2.73$ mol of N_2 gas

Worked example 15A–5: Calculating volume using the ideal gas equation

VIDEO
WORKED
EXAMPLE 15A–5

A quantity of 42 moles of helium gas was used to fill a balloon at sea level. When the balloon reaches an altitude of 200 m above sea level the air temperature has fallen to 14°C and the air pressure has dropped to 99.99 kPa. Calculate the volume of this balloon at this altitude.

Solution

Step 1	Extract the information from the question.	$P = 99.99$ kPa $V = ?$ $T = 14$°C = 287 K $R = 8.31$ $n = 42$ mol
Step 2	Insert into the ideal gas equation.	$PV = nRT$ $99.99 \times V = 42 \times 8.31 \times 287$
Step 3	Solve for *V*.	$V = \frac{nRT}{P}$ $= \frac{42 \times 8.31 \times 287}{99.99}$ $= \frac{100168}{99.99}$ $= 1000$ L of He gas

WORKSHEET
15A–1
CALCULATIONS
USING THE IDEAL
GAS EQUATION

Common experimental conditions

Most of the experiments involving gases we conduct are occurring in a laboratory or outside at atmospheric pressures. This has led to two common sets of conditions for gas measurement in order to make experiments easily reproducible. These conditions are **standard laboratory conditions (SLC)** and standard temperature and pressure (STP), though in this chapter, we are only going to focus on SLC.

Standard laboratory conditions (SLC) the standard ambient conditions within a laboratory, found to be 25°C and 100 kPa

Table 15A–2 The temperature and pressure at standard laboratory conditions (SLC)

	Temperature (°C)	Pressure (kPa)
SLC	25	100

15A SKILLS

As you will have seen in Chapter 13, and again in this section, unit conversions are something that you will need to understand how to do. While it is possible to try and memorise all of the different conversion factors, this can be a little difficult. One helpful strategy that you can use is to learn the meanings of some of the common prefixes that you encounter in these units.

In this section, we have seen units that contain the prefixes ‘kilo-’ (in kilopascal) and ‘milli-’ (in millilitre). ‘Kilo-’ means 1000 and ‘milli-’ means 1/1000th. This means that to convert a base unit to the same unit with the prefix ‘kilo-’, you must move the decimal point three places to the left. On the other hand, to convert a base unit to the same unit with the prefix ‘milli-’, you must move the decimal point three places to the right.

If we have 350.00 pascals, we could move the decimal place three places to the left and it would give us 0.350 kilopascals.

If we have 350.00 litres, we could move the decimal place three places to the right and it would give us 350 000 millilitres.

You can practise this by creating a table for yourself with the common unit prefixes and the magnitude with which they change the value.

Section 15A questions

1 Convert the following volumes to the units specified.
  - a 2850 mL to L
  - b 190 mL to L

2 Convert each of the following pressures to the units specified.
  - a 160 000 Pa to atm
  - b 8.56 atm to Pa
  - c 120 kPa to atm
  - d 86 000 Pa to kPa
  - e 2300 kPa to Pa

3 Convert the following temperatures to the unit specified.
  - a 25°C to K
  - b 190 K to °C

4 What does SLC stand for? What are the temperature and pressure conditions at SLC?
5 State two requirements for a gas to follow ideal gas behaviour.
6 Calculate the volume of gas, in litres, occupied by the following gases.
  a 0.1 mol of hydrogen at 230 kPa and 40°C
  b 25 mol of neon at 2 atm and –170°C
  c 10 g of carbon dioxide at 100 kPa and 27°C
7 Calculate the volume of the following gases at SLC.
  a 1.4 mol H_2
  b 1.0×10^{-3} mol Cl_2
  c 2.5 mol N_2

15B Gas stoichiometry and the greenhouse effect

Study Design:

- Calculations of the molar volume or molar mass of a gas produced by a chemical reaction
- The use of stoichiometry to solve calculations related to chemical reactions involving gases (including moles, mass and volumes of gases)
- 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

Glossary:

Avogadro's law
Chlorofluorocarbon (CFC)
CO_2 equivalent
Enhanced greenhouse effect
Global warming
Greenhouse effect
Greenhouse gas
Infrared radiation (IR)
Molar volume
Photochemical smog

ENGAGE

Wild weather

If recent years are anything to go by, extreme weather events are starting to become the norm. Extreme weather events increase in likelihood when natural climate change variability is amplified by global warming. For example, the bushfires of 2019 that clouded most of the eastern seaboard of Australia were due to a combination of record-breaking low rainfall coinciding with extreme heat. Extreme heat leads to drying of fire 'fuel' such as grasses, debris and dead wood – the perfect conditions for a bushfire if one were to start (Figure 15B–1). In this section, we are going to learn about the natural and enhanced greenhouse effects, as well as ways that we can calculate the volume of gases in chemical equations.

Figure 15B–1 The 2019 bushfires in eastern Australia were intensified by global warming.

EXPLAIN

Calculating the molar mass and identity of gases from a chemical reaction

The identity of a gas product formed in a chemical reaction can be calculated using a combination of the ideal gas equation, stoichiometry and conversion to molar mass. A key additional piece of information required for this is the mass of the unknown. With the information on mass and the empirical formula of the unknown, we can then use the moles equals mass divided by molar mass equation to solve for molar mass.

Worked example 15B–1: Calculating molar mass and identity of gases

VIDEO
WORKED
EXAMPLE 15B–1

A 7.31 g mass of an unknown straight-chain alkane is held in a 2 L flask at 2.1 atm. The temperature inside the flask is 20°C. Calculate the molar mass of the alkane and determine its identity.

Solution

Step 1	Extract the information from the question.	$m = 7.69$ g $V = 2.0$ L $P = 2.1$ atm $= 2.1 \times 101.3$ kPa $= 212.73$ kPa $R = 8.31$ $T = 20°\text{C} = 20 + 273$ K $= 293$ K $M = ?$ Hydrocarbon is ?
Step 2	Use the ideal gas equation and substitute in the mass–mole equation.	$PV = nRT$ and $n = \frac{m}{M}$ Therefore, $PV = \frac{mRT}{M}$ and furthermore, $M = \frac{mRT}{PV}$.
Step 3	Insert values into the equation.	$M = \frac{mRT}{PV}$ $= \frac{7.69 \times 8.31 \times 293}{212.73 \times 2}$ $= 44 \text{ g mol}^{-1}$
Step 4	Use the equation for alkanes, C_nH_{2n+2}.	Alkanes have the formula C_nH_{2n+2}, so using $n = 3$, the molar mass rounds to 44. This means our unknown hydrocarbon is propane (C_3H_8).

Molar volume of a gas

A special condition of **Avogadro's law** applies when a gas is being measured at SLC. For 1 mole of a gas at SLC, there is a known volume (as temperature and pressure are known at SLC). This means the **molar volume** at SLC is 24.8 L mol^{-1}. Based on this, Avogadro's law for measuring volume change becomes

$$n = \frac{V}{V_m}$$

where:

n = The moles of the gas
V = The volume of gas (in L)
V_m = The molar volume at SLC (in L mol^{-1})

Avogadro's law states that the volume of the gas is proportional to the number of moles of gas when temperature and pressure are constant

Molar volume the volume of one mole of a substance at a specified pressure and temperature; symbol V_m

Worked example 15B–2: Calculating molar volume of gases

Calculate the amount of helium gas in a flask of 12.2 L at SLC.

Solution

Step 1	Extract the information from the question.	$V_m = 24.8$ L mol^{-1} at SLC $V = 12.2$ L $n = ?$ mol
Step 2	Insert into the equation for Avogadro's law for molar volume.	$n = \frac{V}{V_m} = \frac{12.2}{24.8}$
Step 3	Solve for n.	$n = \frac{V}{V_m} = \frac{12.2}{24.8}$ $= 0.492$ mol of He

Check-in questions – Set 1

1 Determine the volume of 14.0 g of nitrogen gas at SLC.
2 A 55 mL fluorescent globe is installed in a laboratory at standard laboratory conditions. How many moles of fluorescent gas fill the globe?

Calculating the amount of gases from a chemical reaction

Ideal gas behaviour is closely linked to stoichiometry. The mole is a part of the ideal gas equation, and thus we can determine not only one, but all gas amounts within a reaction mixture if we have all the information required to use the ideal gas equation.

Chemical reactions tell us the ratio of products formed from reactants used. The mole ratio between reactants and products becomes crucial to solving the amount of gases formed or used in a chemical reaction.

Step 1: Write the balanced reaction down (if not already provided).
Step 2: Note the mole ratios in the reaction.
Step 3: Use the mole ratios to determine the resultant amount of reactant used or product formed.
Step 4: Determine any additional information, such as volume, pressure or temperature, via the moles.

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Worked example 15B–3: Calculating the amount of gases from a reaction

Calcium carbonate decomposes naturally on heating to form calcium oxide and carbon dioxide. If 0.25 moles of calcium carbonate are initially present, how many moles of carbon dioxide are produced?

Solution

Step 1	Write the reaction.	$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
Step 2	Note the mole ratios in the reaction.	The equation is balanced – the mole ratio of $CaCO_3$ to CO_2 is 1 : 1.
Step 3	Use the mole ratios to determine the resultant amount of reactant used or product formed.	0.25 mol of CO_2 will be produced.

Calculating the mass of gases from a chemical reaction

The best way to measure the mass of a gas used or formed in a chemical reaction is from the amount of gas (mole). This is because, in some reactions, the gases produced are so light that it's next to impossible to measure a weight change using scales or other common weighing devices. To do this, we follow a similar process to calculating the amounts of gas in a reaction and use the molar mass of the gas to calculate the mass used or formed.

Worked example 15B–4: Calculating the mass of gases from a reaction

VIDEO
WORKED
EXAMPLE 15B–4

A sample of calcium carbonate is weighed and found to be 1.01 g. It is placed on a heating plate and heated until it has decomposed completely. Calculate the mass of carbon dioxide produced.

Solution

Step 1	Write the reaction.	$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
Step 2	Note the mole ratios in the reaction.	This reaction is balanced and the mole ratio of $CaCO_3$ to CaO is 1 : 1.
Step 3	Use the mole ratios to determine the resultant amount of reactant used or product formed.	If 1.0 g of $CaCO_3$ is decomposed, this will produce $n(CaCO_3) = \frac{m}{M} = \frac{1.01}{100.1} = 0.010\text{ mol}$ If $n(CaCO_3) = 0.01$ mol, then $n(CO_2) = 0.01$ mol.
Step 4	Calculate the mass of CaO produced.	$m(CO_2) = n \times M$ $= 0.010 \times 44.01$ $= 0.44$ g

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Check-in questions – Set 2

1 Standard laboratory conditions (SLC) have a temperature of 25°C and a pressure of 100 kPa.
  - a Avogadro's Law refers to the molar volume of gases. What is this law?
  - b What is the value for the molar volume of gases at SLC?
  - c At SLC, 15 L of hydrogen gas is stored in a container. What amount of hydrogen gas (in mol) is this?
  - d 15 L of oxygen gas is also stored at SLC. What amount of hydrogen gas (in mol) is this?

2 A scientist has been shipped 25 g of a gas that they believe to be methane. However, before using it, they receive a phone call telling them that there may have been a mix-up in the deliveries, and it may actually be propane. To test this, he finds that at 22°C and 100 kPa, the gas takes up a volume of 13.9 L. Which gas have they been shipped?

Calculating the volume of gases from a chemical reaction

The volume of a gas used or formed in a chemical reaction can be calculated by linking our understanding of the ideal gas equation to stoichiometry. To do this, we follow a similar process to calculating the amounts of gases in a reaction and then convert to volume using the ideal gas equation.

Worked example 15B–5: Calculating the volume of gases from a reaction

VIDEO WORKED EXAMPLE 15B–5

Carbon dioxide can be produced spontaneously when carbonic acid breaks down under the right conditions. If this reaction occurs at 23°C and 101.3 kPa for 2.0 L of a 0.050 M solution of carbonic acid, what volume of carbon dioxide will be produced?

Solution

Step 1	Write the reaction.	$H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$
Step 2	Note the mole ratios in the reaction.	This reaction is balanced and the mole ratio of H_2CO_3 to CO_2 is 1:1.
Step 3	Use the mole ratios to determine the resultant amount of reactant used or product formed.	If 2.0 L of 0.05 M H_2CO_3 is decomposed, this will produce $n = c \times V = 0.050 \times 2.0 = 0.10$ mol If $n(H_2CO_3) = 0.10$ mol, then $n(CO_2) = 0.10$ mol.
Step 4	Calculate the volume of CO_2 using the ideal gas equation.	$PV = nRT$ where $P = 101.3$ kPa $V = ?$ $n = 0.10$ mol $T = 23 + 273 = 296$ K $R = 8.31$ J mol^{-1} K^{-1} $PV = nRT$ so $V = \frac{nRT}{P}$ $= \frac{0.10 \times 8.31 \times 296}{101.3}$ $= 2.4$ L

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Volume–volume stoichiometry

In the previous section, we calculated the volume of gas produced from the breakdown of an aqueous solution. In that instance, the ideal gas equation was a perfect tool for making calculations. However, if the reactants and products are gaseous already, then we can use the volumetric stoichiometry to solve without the need for conversion.

For example, in $F_2(g) + H_2(g) \rightarrow 2HF(g)$, one mole of fluorine gas reacts with one mole of hydrogen gas to produce two moles of hydrogen fluoride gas. As equal amounts of gases occupy the same volumes at the same temperature and pressure irrespective of their identity, we can therefore use the ratios in a balanced equation to calculate the volumes of gaseous reactants or products.

Worked example 15B–6: Calculating volume using mole ratios

Methane is burned in a camp stove. If 100 mL of methane is burned in the air at 450°C and 4 atm, calculate the:

a volume of O_2 required for complete combustion of the methane.

b volumes of CO_2 and H_2O vapour produced under the same conditions.

Solution

Step 1	Write the reaction.	$CH_4(g)$ +	$2O_2(g)$ →	$2H_2O(g)$ +	$CO_2(g)$
Step 2	Note the mole ratios in the reaction.	1 mol	2 mol	2 mol	1 mol
Step 3	Use the mole ratios to determine the resultant volumes of reactant used or product formed at constant temperature and pressure.	1 vol 100 mL	2 vol 200 mL	2 vol 200 mL	1 vol 100 mL
		Because we had fixed temperature and pressure and this was an all-gas reaction, we have solved questions **a** and **b** in one step.			

Earth's atmosphere

Gases are an important part of what keeps Earth safe. They are the air we breathe and the food for our forests, and they create clouds that produce rain. They're the atmosphere, the invisible barrier that helps regulate the temperature of our environment, stops water from escaping into outer space and keeps harmful radiation from causing damage to the planet and us. Many gases help form the different layers of the atmosphere. Our air composition is largely constant, and the same proportions occur across the entirety of the atmosphere even though the amount of gas changes as we move further away from Earth's surface.

Table 15B–1 The different gases of the atmosphere and their amount

Gas	Amount (%)
Nitrogen	78.09%
Oxygen	20.95%
Carbon dioxide	0.039%

The natural greenhouse effect

The **greenhouse effect** is a phenomenon that helps regulate Earth's environment. The most critical **greenhouse gases** are carbon dioxide (CO_2), methane (CH_4) and water (H_2O). Each of these gases helps to absorb and reflect **infrared radiation (IR)** from outer space and keep energy within the atmosphere, as shown in Figure 15B–2. This helps regulate the temperature of the planet. Each gas exists at a fixed concentration in the atmosphere and therefore provides a fixed amount of protection to the planet. These gases are needed to ensure the environment on Earth is stable.

VIDEO 15B–1
THE GREENHOUSE EFFECT

Greenhouse effect a natural process in which heat from Earth's surface is re-radiated by gases in the atmosphere, which maintains Earth at a temperature suitable for life

Greenhouse gas a gas that contributes to the greenhouse effect by trapping heat close to Earth's surface; e.g. carbon dioxide

Infrared radiation (IR) the region of the electromagnetic spectrum that lies between the wavelengths 700 nm and 1 mm

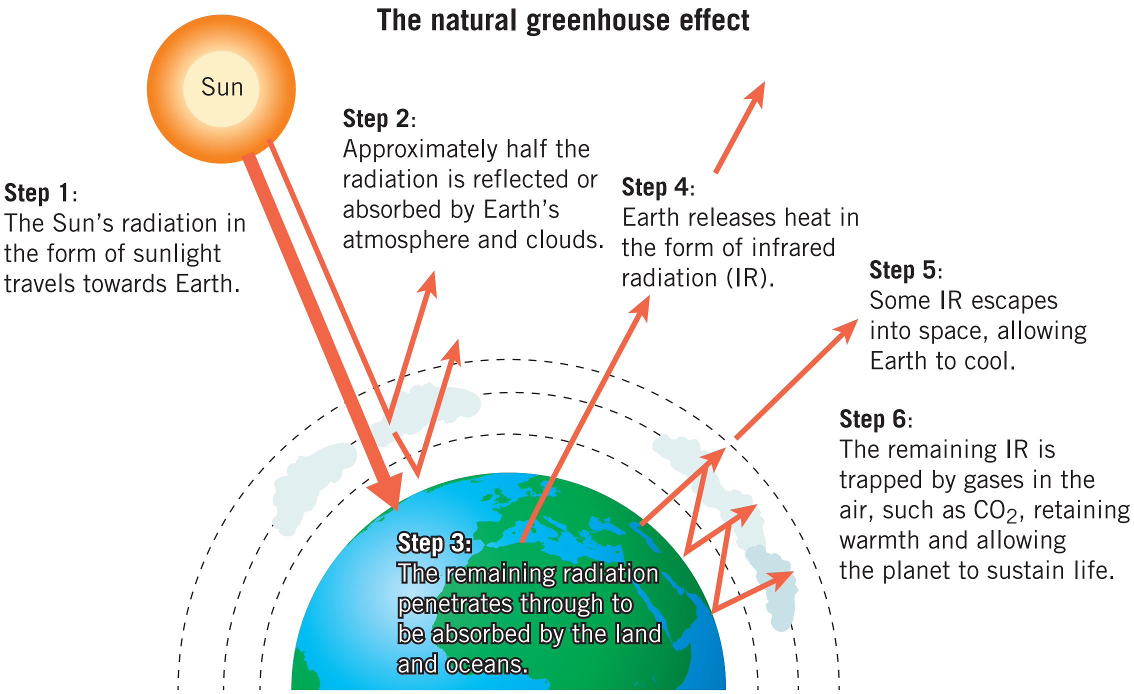

Figure 15B–2 A representation of the natural greenhouse effect

The enhanced greenhouse effect

The **enhanced greenhouse effect**, sometimes referred to as man-made climate change or global warming, occurs when there is a significant change in the levels of greenhouse gases. These additional gases cause more heat emitted from Earth's surface, in the form of infrared radiation, to be reflected back, causing a warming effect. This is further enhanced by pollutants like **photochemical smog** produced by the action of sunlight on vehicle emissions, which cause more heat to become trapped. Australia has warmed on average by 1.44 ± 0.24°C since national records began in 1910 according to the Bureau of Meteorology. The warming leads to extreme weather events, increased average temperatures, increased ocean temperatures, rising sea levels and more, all of which have effects on the environments we live in and the planet as a whole. As Table 15B–2 describes, the causes of the enhanced greenhouse effect are quite complex, with everything from deforestation to industrialisation, dependence on fossil fuels and increased agriculture as contributing factors.

Enhanced greenhouse effect the change of the normal climate equilibrium on Earth, primarily increased global temperatures. Caused by increased emissions of greenhouse gases; also known as global warming.

Photochemical smog a brown haze generated from the reaction between ultraviolet light from the sun and nitrogen oxides in the atmosphere

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Table 15B–2 Causes of increased greenhouse gases and their effects

Cause	Effect
Fires and deforestation	Fires lead to production of smoke and CO_2, which trap infrared radiation. Fires and deforestation also lead to a decrease in tree populations, which decreases the ability of nature to remove CO_2 from the air through photosynthesis.
Industrialisation	Factories and industry burn coal, oil and other fuels that release **CO_2 equivalent** chemicals and photochemical smog into the atmosphere, leading to the trapping of infrared radiation.
Agriculture (practices)	Increased cropping leads to increased use of fuels and pesticides. Fuels release CO_2 when burned, and fertilisers lead to formation of N_2O gas with a greenhouse effect that is approximately 300 times greater than CO_2. Both gases lead to the trapping of infrared radiation.
Agriculture (animal husbandry)	Increased reliance on meat, particularly cattle, leads to increased production of methane from digestion of foodstuffs by cattle. Methane (CH_4) is a CO_2 equivalent with a 21-times-greater effect, leading to the trapping of infrared radiation.

CO_2 equivalent a measure used to compare the global warming potential of different greenhouse gases by converting them to the equivalent amount of carbon dioxide (CO_2)

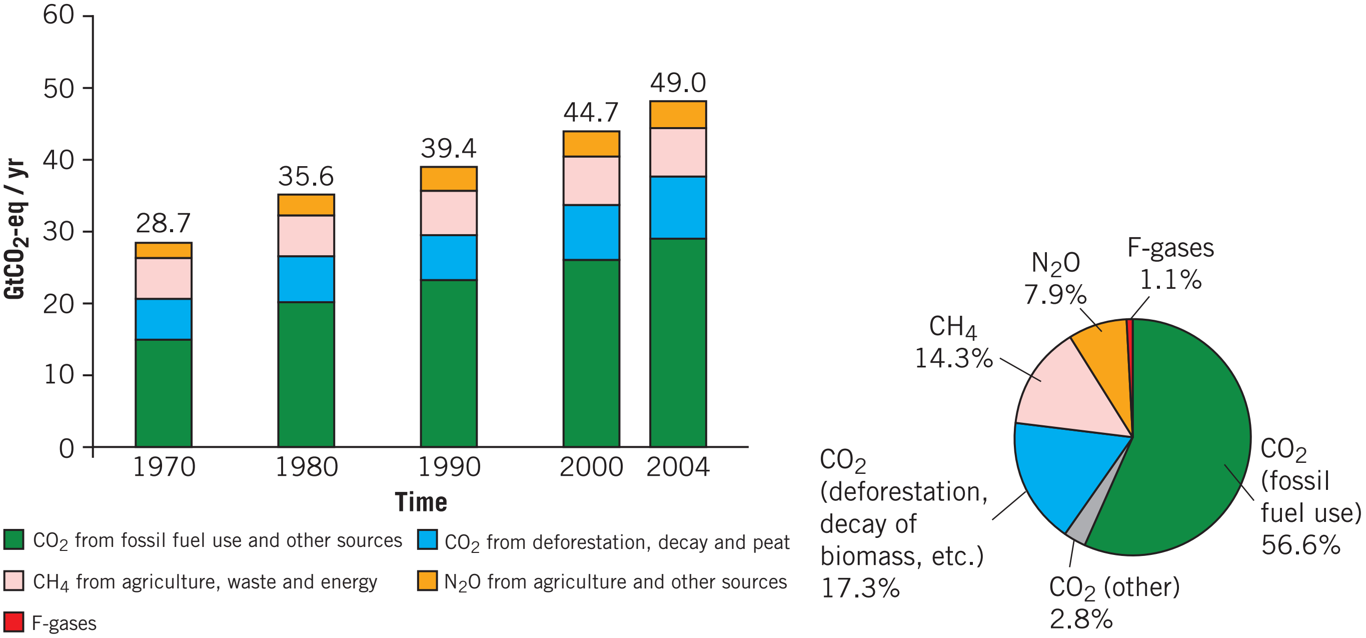

Figure 15B–3 The global levels of CO_2 and other gases per decade 1970–2004 (left). Fossil fuel use is still the primary cause of CO_2 emissions today (right).

The enhanced greenhouse effect correlates strongly with increased global levels of CO_2 emissions, as shown in Figure 15B–3. Other gases are also included in this measurement, as their CO_2 equivalency is factored into their ability to enhance the greenhouse effect (see Table 15B–3). In 2019, the global levels of CO_2 and CO_2 equivalents were 410 parts per million (ppm) and 508 ppm respectively, as shown in Figure 15B–4, which is approximately a 50% increase in levels since the pre-industrial era when the levels were 278 ppm and 310 ppm respectively. Over the last four decades, the rate of CO_2 level increase has also increased, from 14 ppm per decade in the 1980s to 23 ppm per decade in the 2010s. This means that emission levels are increasing, not decreasing, despite many green strategies being implemented to address this.

Table 15B–3 Other gases contributing to the greenhouse effect, expressed as CO_2 equivalence of harm

Greenhouse gas	CO_2 equivalency
Carbon dioxide, CO_2	1
Methane, CH_4	21
Nitrogen oxides, NO_x	310
Chlorofluorocarbons, CFC	4750
Hydrofluorocarbons, HFC	140–11 700
Perfluorochemicals, PFC	6500–9200
Sulfur hexafluoride, SF_6	23 900

Chlorofluorocarbon (CFC) a compound that contains the elements chlorine, fluorine and carbon; long-lived in the atmosphere and damaging to Earth's ozone layer

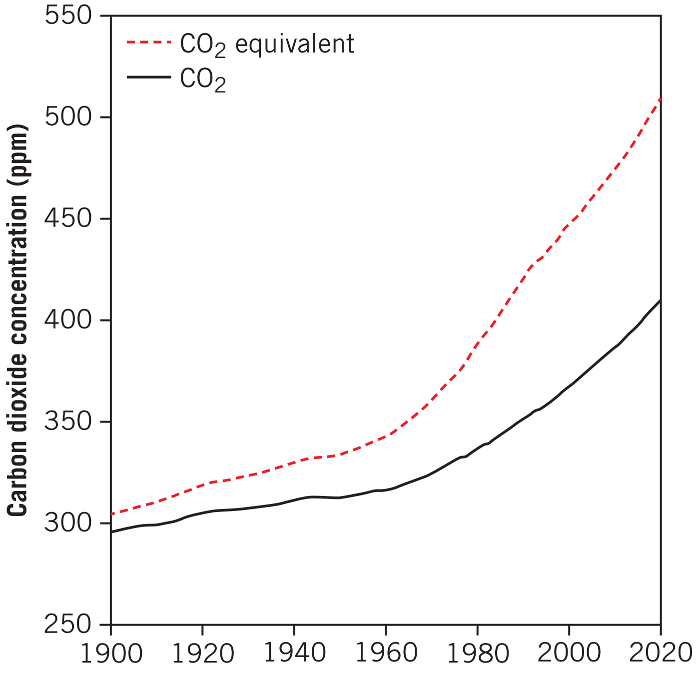

Figure 15B–4 The global levels of CO_2 gas and its equivalents per decade since 1900. Global CO_2 levels are clearly rising.

15B SKILLS

As you have seen in this section, there are lots of different calculations that you will have to perform in VCE Chemistry. For almost any type of calculation, there will be a worked example of a similar question found within this textbook. However, if you go through and attempt one of these worked examples on your own, you may get a bit of a surprise when some of the numbers on your calculator aren't exactly the same as the number that appears in the textbook.

Let's use Worked example 15A–5 to explain what we mean by this.

Worked example 15A–5: Calculating volume using the ideal gas equation

A quantity of 42 moles of helium gas was used to fill a balloon at sea level. When the balloon reaches an altitude of 200 m above sea level the air temperature has fallen to 14°C and the air pressure has dropped to 99.99 kPa. Calculate the volume of this balloon at this altitude.

Solution

Step 1	Extract the information from the question.	$P = 99.99$ kPa $V = ?$ $T = 14°\text{C} = 287$ K $R = 8.31$ $n = 42$ mol
Step 2	Insert into the ideal gas equation.	$PV = nRT$ $99.99 \times V = 42 \times 8.31 \times 287$
Step 3	Solve for V.	$V = \frac{nRT}{P}$ $= \frac{42 \times 8.31 \times 287}{99.99}$ $= \frac{100168}{99.99}$ $= 1000$ L of He gas

In particular, let us look at the line before the final answer. In the worked example, this is written as $= \frac{100\,168}{99.99}$. However, if you were to look at this on your calculator, you would probably find that you have something more along the lines of $= \frac{100\,168.74}{99.99}$.

Why is there a difference, and why is it important that we highlight this?

The difference is due to the fact that certain steps in the working out have been simplified in the textbook. It isn't practical to write down a value that has eight or nine values after the decimal point, for example. However, it is critical that you always continue to use the most exact value that your calculator provides as you progress through the calculation.

In the above example, the only way to get the correct answer is to use the full value (100 168.74). The only step that should involve any rounding is your final answer, as this should always be provided with the appropriate number of significant figures.

Section 15B questions

1 A sample of 0.5 mol of nitrogen is placed in a volumetric flask of 2.5 L and stoppered. A thermometer inside the flask shows the temperature to be 5°C. What is the pressure inside the flask?

2 Determine the volume, in litres, occupied by the following gases:
  - a 0.1 mol of hydrogen at 230 kPa and 45°C
  - b 22 mol of carbon dioxide (CO_2) at 5 atm and 300°C
  - c 17 g of hydrogen sulfide (H_2S) at 101.3 kPa and 27°C

3 At a given temperature, an 8 g sample of helium exerts a pressure of 50.5 kPa on a gas cylinder of volume 5.0 L. Determine the temperature of the gas.

4 Which sample contains a greater amount (in mol) of gas?
  - a 3.2 L of oxygen at 25°C and 120 000 Pa
  - b 2.5 L of hydrogen at 23°C and a pressure of 1.2 atm

5 Calcium reacts with hydrochloric acid according to the equation

$$Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$$

Calculate the volume of hydrogen gas produced at SLC when the following masses of magnesium react with excess hydrochloric acid:
  - a 2.40 g
  - b 1.08 g

6 Octane burns according to the equation

$$2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

What mass of octane must have been used if 50.0 L of carbon dioxide, measured at 150°C and 0.9 atm, was produced?

7 Explain how the natural greenhouse effect works and what it does.

8 How is the enhanced greenhouse effect different from the natural greenhouse effect?

9 List three pieces of evidence for the enhanced greenhouse effect.

10 Outline two causes of the enhanced greenhouse effect.

Chapter 15 review

Summary

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

Success criteria – I am now able to:		Linked question
15A.1	Recall the properties of gases	3☐
15A.2	Define ‘gas pressure’ and convert between kPa, Pa and atm	7☐
15A.3	State the standard laboratory conditions (SLC)	1☐
15A.4	Perform conversions between the volume units mL and L	6☐
15A.5	Perform conversions between the temperature units °C and K	2☐, 16☐
15A.6	Calculate gas volume or amount in moles using the molar volume of a gas	5☐
15A.7	Perform basic calculations using the ideal gas law	9☐, 14☐
15B.1	Calculate the molar volume of a gas produced by a chemical reaction	17☐
15B.2	Calculate the molar mass of a gas produced by a chemical reaction	18☐
15B.3	Perform calculations on a chemical reaction involving gases using moles, mass and volume	15☐, 17☐
15B.4	Compare and explain the natural and enhanced greenhouse effects	12☐
15B.5	Describe the role CO_2, CH_4 and H_2O play in the greenhouse effect with reference to their ability to absorb infrared radiation	4☐ 11☐

Multiple-choice questions

1 Which of the following are the correct temperature and pressure conditions for SLC?

A 298 K, 100 atm
B 25°C, 100 kPa
C 20°C, 100 kPa
D 273 K, 0.987 atm

2 Water freezes at 0°C. This means that it freezes at

A 0 K.
B 100 K.
C 273 K.
D 373 K.

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3 Which of the following is **not** a characteristic property of gases?

A They are easy to compress.
B They expand to fill their container.
C They have no fixed volume or shape.
D They occupy less space than the liquids and solids from which they form.

4 Which of the following is **not** an important greenhouse gas?

A methane
B carbon dioxide
C water
D oxygen

5 A gas occupies a volume of 6.2 L at a temperature of 25°C and a pressure of 100 kPa. The amount in moles of this gas present is

A 0.25.
B 0.50.
C 0.75.
D 1.0.

6 Which of the following is the equivalent of 45 mL?

A 4.5 L
B 0.45 L
C 0.045 L
D 0.0045 L

7 Gas pressure is

A not part of the ideal gas equation.
B different for all gases.
C the volume of the container that the gas is stored in.
D the amount of force exerted by a gas on its container.

8 The combustion of propane with oxygen to make carbon dioxide and water occurs in a 1:5:3:4 molar ratio. If 50 mL of propane is reacted, what volume of oxygen is required for complete combustion?

A 10 mL
B 50 mL
C 250 mL
D 500 mL

9 The correct representation for solving for temperature using the ideal gas equation is

A $T = \frac{nR}{PV}$.
B $T = nRPV$.
C $T = \frac{PV}{nR}$.
D $T = \frac{nRP}{V}$.

10 Which one of the following is **not** an assumption necessary for application of the ideal gas law?

A Gases are infinitely small.
B Gases don't take up any volume.
C Gases have no interactions with the container or each other.
D The energy of gas particle collisions is conserved (that is, the collisions are elastic).

Short-answer questions

11 Draw a diagram outlining the process of the greenhouse effect. How does the enhanced greenhouse effect differ from the natural greenhouse effect? Ensure that your diagram is completely labelled. (6 marks)

12 Describe how greenhouse gases lead to the enhanced greenhouse effect and how different gases have different levels of effect. Your answer should include reference to one cause of increased gas emissions. (6 marks)

13 Biological breakdown of carbon-based compounds often produces CO_2, such as when glucose ($C_6H_{12}O_6$) is fermented by yeast. Adding acid to carbonates such as limestone ($CaCO_3$) also produces CO_2 as a by-product.

a Write balanced equations for the formation of CO_2 via these two methods. (2 marks)

b If you start with 3 g of the glucose and the limestone, which of the two will produce the most CO_2, and by how much at SLC? (2 marks)

14 A weather balloon is a large balloon that carries weather-monitoring equipment to the upper atmosphere. A weather balloon is going to be filled with helium on a day where the temperature is 23.6°C and the pressure is 1.03 atm. How many moles of helium would be required to fill the balloon to its target volume of 2850 L? (6 marks)

15 Butane undergoes complete combustion in excess oxygen according to the following equation:

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$$

At SLC, what volume of carbon dioxide, in litres, would form from the combustion of a 227 g canister of butane? (6 marks)

16 Convert the following temperatures into the units given.

a 120°C into K

b 273 K into °C

c –145°C into K (6 marks)

17 Calculate the following quantities.

a mass of oxygen molecules present in a 100 L container at SLC

b mass of nitrogen gas present in a 20 L container at SLC

c volume of oxygen produced at 27°C and 100 kPa when 20 g of hydrogen peroxide decomposes to give water and oxygen gas (6 marks)

18 A 0.037 g sample of an unknown alkane is held in a 1.1 L flask at 1.5 atm. The temperature inside the flask is 25°C. Calculate the molar mass of the alkane and determine its identity. (3 marks)

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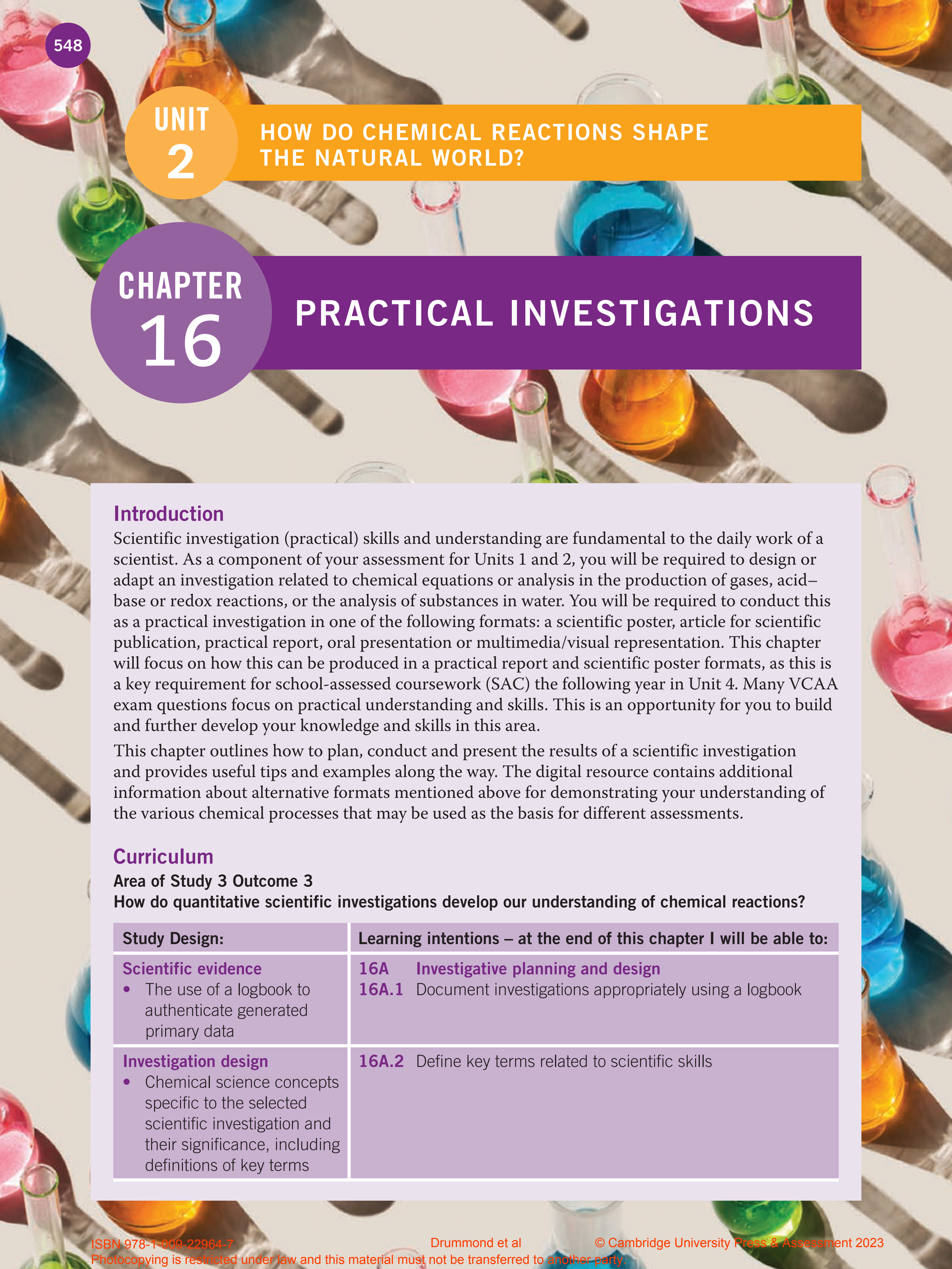

UNIT 2

HOW DO CHEMICAL REACTIONS SHAPE THE NATURAL WORLD?

CHAPTER 16

PRACTICAL INVESTIGATIONS

Introduction

Scientific investigation (practical) skills and understanding are fundamental to the daily work of a scientist. As a component of your assessment for Units 1 and 2, you will be required to design or adapt an investigation related to chemical equations or analysis in the production of gases, acid–base or redox reactions, or the analysis of substances in water. You will be required to conduct this as a practical investigation in one of the following formats: a scientific poster, article for scientific publication, practical report, oral presentation or multimedia/visual representation. This chapter will focus on how this can be produced in a practical report and scientific poster formats, as this is a key requirement for school-assessed coursework (SAC) the following year in Unit 4. Many VCAA exam questions focus on practical understanding and skills. This is an opportunity for you to build and further develop your knowledge and skills in this area.

This chapter outlines how to plan, conduct and present the results of a scientific investigation and provides useful tips and examples along the way. The digital resource contains additional information about alternative formats mentioned above for demonstrating your understanding of the various chemical processes that may be used as the basis for different assessments.

Curriculum

Area of Study 3 Outcome 3
How do quantitative scientific investigations develop our understanding of chemical reactions?

Study Design:	Learning intentions – at the end of this chapter I will be able to:
Scientific evidence • The use of a logbook to authenticate generated primary data	**16A Investigative planning and design** 16A.1 Document investigations appropriately using a logbook
Investigation design • Chemical science concepts specific to the selected scientific investigation and their significance, including definitions of key terms	16A.2 Define key terms related to scientific skills

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Study Design:	Learning intentions – at the end of this chapter I will be able to:
- Scientific methodology relevant to the scientific investigation, selected from the following: classification and identification; controlled experiment; fieldwork; modelling; product, process or system development; and simulation	**16A.3** Define ‘controlled experiment’ **16A.4** List the features that a controlled experiment includes **16A.5** Distinguish between a positive and a negative control **16A.6** Understand the meaning of single variable exploration **16A.7** Define and identify independent, dependent and controlled variables **16A.8** Select and use the procedures and equipment appropriate to an investigation **16A.9** Recall different types of scientific investigations that generate data, including fieldwork, simulations and modelling
Scientific evidence - The distinction between an aim, a hypothesis, a model, a theory and a law	**16A.10** Distinguish between an aim, a hypothesis, a model, a theory and a law **16A.11** Construct aims and questions for investigations **16A.12** Formulate hypotheses and predict possible outcomes
Investigation design - Accuracy, precision, repeatability, reproducibility, resolution and validity of measurements in relation to the investigation - The health, safety and ethical guidelines relevant to the selected scientific investigation	**16A.13** Define reproducibility, repeatability, resolution and validity, and distinguish between these **16A.14** Design an investigation that is valid **16A.15** Design an experiment that is fully reproducible by others **16A.16** Determine potential ethical issues with investigation design **16A.17** Identify how bias can be minimised in an investigation **16A.18** Follow clear guidelines for health and safety when undertaking practical investigations
Investigation design - Techniques of primary qualitative and quantitative data generation relevant to the investigation - Accuracy, precision, repeatability, reproducibility, resolution and validity of measurements in relation to the investigation	**16B** **Scientific evidence** **16B.1** Define qualitative and quantitative **16B.2** Distinguish between qualitative and quantitative data **16B.3** Include appropriate units of measurement for quantitative data **16B.4** Design an investigation that accounts for accuracy and precision

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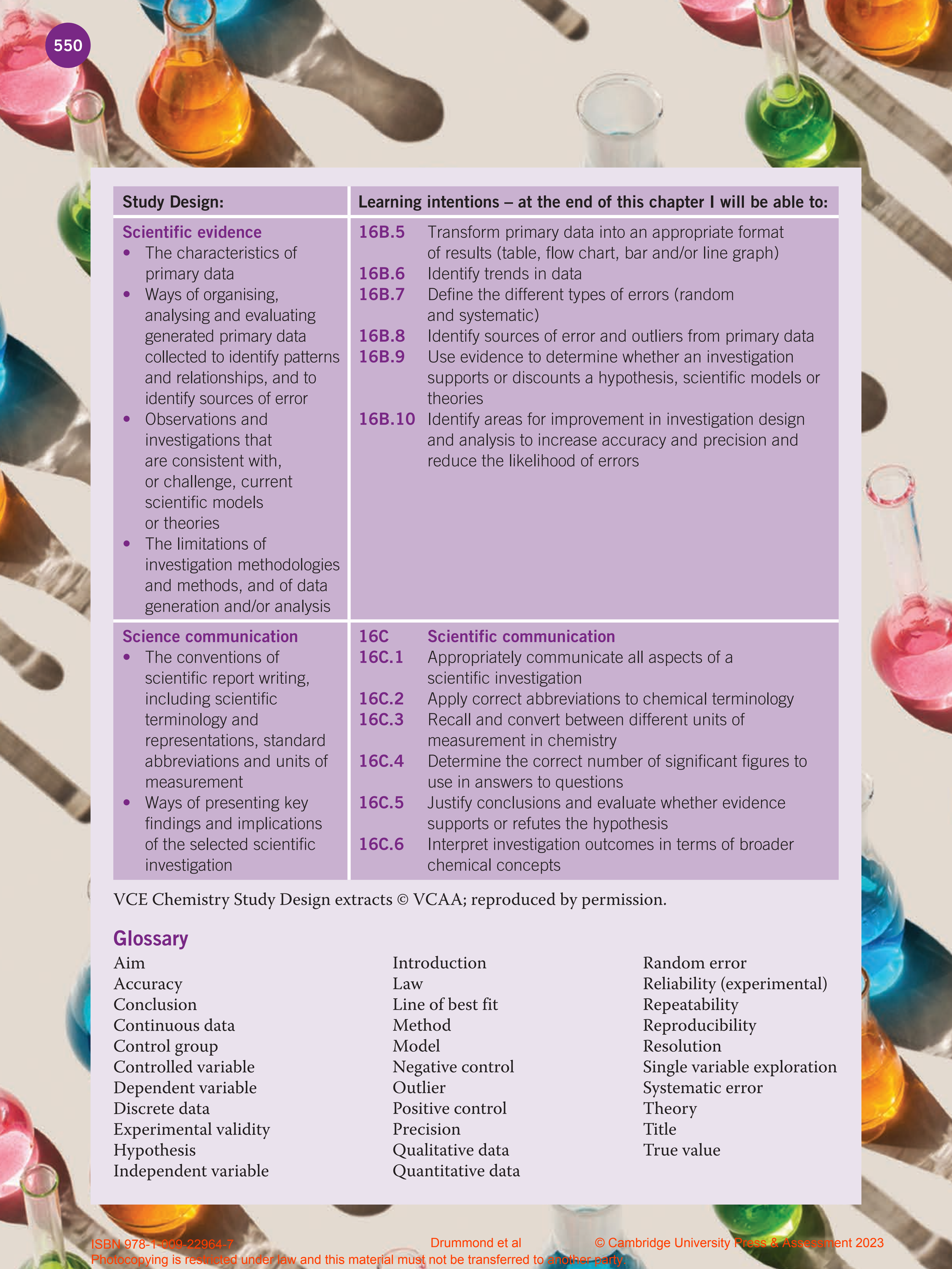

Study Design:	Learning intentions – at the end of this chapter I will be able to:
Scientific evidence • The characteristics of primary data • Ways of organising, analysing and evaluating generated primary data collected to identify patterns and relationships, and to identify sources of error • Observations and investigations that are consistent with, or challenge, current scientific models or theories • The limitations of investigation methodologies and methods, and of data generation and/or analysis	**16B.5** Transform primary data into an appropriate format of results (table, flow chart, bar and/or line graph) **16B.6** Identify trends in data **16B.7** Define the different types of errors (random and systematic) **16B.8** Identify sources of error and outliers from primary data **16B.9** Use evidence to determine whether an investigation supports or discounts a hypothesis, scientific models or theories **16B.10** Identify areas for improvement in investigation design and analysis to increase accuracy and precision and reduce the likelihood of errors
Science communication • The conventions of scientific report writing, including scientific terminology and representations, standard abbreviations and units of measurement • Ways of presenting key findings and implications of the selected scientific investigation	**16C** **Scientific communication** **16C.1** Appropriately communicate all aspects of a scientific investigation **16C.2** Apply correct abbreviations to chemical terminology **16C.3** Recall and convert between different units of measurement in chemistry **16C.4** Determine the correct number of significant figures to use in answers to questions **16C.5** Justify conclusions and evaluate whether evidence supports or refutes the hypothesis **16C.6** Interpret investigation outcomes in terms of broader chemical concepts

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Glossary

Aim
Accuracy
Conclusion
Continuous data
Control group
Controlled variable
Dependent variable
Discrete data
Experimental validity
Hypothesis
Independent variable
Introduction
Law
Line of best fit
Method
Model
Negative control
Outlier
Positive control
Precision
Qualitative data
Quantitative data
Random error
Reliability (experimental)
Repeatability
Reproducibility
Resolution
Single variable exploration
Systematic error
Theory
Title
True value

Concept map

16A Investigative planning and design

Structuring the logbook – investigative design

16B Scientific evidence

Structuring the logbook – scientific evidence

16C Scientific communication

See the Interactive Textbook for an interactive version of this concept map interlinked with all concept maps for the course.

16A Investigative planning and design

Study Design:

Scientific evidence

- The use of a logbook to authenticate generated primary data
- The distinction between an aim, a hypothesis, a model, a theory and a law

Investigation design

- Chemical science concepts specific to the selected scientific investigation and their significance, including definitions of key terms
- Scientific methodology relevant to the scientific investigation, selected from the following: classification and identification; controlled experiment; fieldwork; modelling; product, process or system development; or simulation
- Accuracy, precision, repeatability, reproducibility, resolution and validity of measurements in relation to the investigation
- The health, safety and ethical guidelines relevant to the selected scientific investigation

Glossary:

Aim
Control group
Controlled variable
Dependent variable
Experimental validity
Hypothesis
Independent variable
Introduction
Law
Method
Model
Negative control
Positive control
Reliability (experimental)
Repeatability
Reproducibility
Single variable exploration
Title

ENGAGE

Scientific breakthroughs

Scientists ask questions about things that interest them, trouble them or puzzle them. They plan for new investigations and adjust the practical investigations they are currently undertaking. They work both individually and in teams to share knowledge for the greater good of humans and the existence and survival of many other species. Some of the greatest discoveries in the past two centuries have come from this process, and include:

- the discovery of the structure of DNA by James Watson and Francis Crick in 1953 based on Rosalind Franklin's X-ray crystallography
- Mendeleev’s publication of the first modern periodic table with 66 known elements organised by atomic weights in 1869
- S.P.L. Sørensen's development of the idea of pH and invention of techniques for measuring acidity in 1909.

Figure 16A–1 Our current understanding of pH is indebted to Sørensen’s research from the early 1900s.

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EXPLAIN

Structuring a logbook – investigative design

All these discoveries mentioned in the Engage section on the previous page were possible not only due to the persistence and determination of the scientists but also, no doubt, through careful documentation of their work in a logbook. A logbook allows a scientist to keep track of the specific dates of their investigation ideas, planning, questions, important results, errors and any modifications required, thereby maintaining a record of all work conducted. Information recorded in a logbook is an original source of the information, written, processed or collected by the experimenter or researcher themselves. In doing so, a clear record of all information and data can be authenticated, meaning it is true or valid. A requirement of the VCE is that each student must maintain a logbook of practical activities in Units 1–4 for assessment purposes. The need for a printed logbook, rather than an online or digital logbook, reduces the likelihood of any tampering with results. Of course, this will depend on your school and your teacher's authentication practices, so it is best to work within your school's guidelines.

Primary data refers to any data source that is obtained firsthand by a researcher for their project. Primary data can be collected through many means, including experiments, interviews and field observations. While more time consuming and expensive than collecting second-hand data (already existing collected data), in many instances primary data collection is the best suited for your research.

In this section, you will learn how to develop a logbook for your own practical investigation and the importance of each section in contribution to the end report. This is usually in the form of a practical report or scientific poster (outlined in Chapter 16C).

LINK 16C SCIENTIFIC COMMUNICATION

An important function of an effective logbook is that it clearly demonstrates the development of the ideas that come from your research and investigation, including changes in direction, equipment failures and so on, as well as your collaboration and discussion with peers. This is demonstrated in the example on the following page, which shows two versions of the same investigation:

- Investigation ① is the initial version. The ideas that are highlighted for the investigation are not perfect, but they show the initial planning coming together.
- Investigation ② is the final version. Adjustments have been made to the initial investigation to improve the validity and reliability of the results (terms explored in more detail later).

The investigation on the following page involved determining the volume of carbon dioxide gas produced when varying the mass of a metal carbonate added to the reaction with a strong acid. The transition between Investigation 1 (initial ideas) and Investigation 2 (final product), as well as tips on what to include in each of the key sections, are included to assist you with your own logbook development.

Logbook

Title
the research question under investigation; includes information about what is being tested

Title

① What mass of a metal carbonate reacting with an acid will produce the most carbon dioxide?

② Comparing the theoretical and experimental mass of sodium carbonate required to produce 25 mL of carbon dioxide gas when reacted with hydrochloric acid.

Notes

The title should include reference to the variables being changed (independent variable) and measured (dependent variable), along with enough detail for the reader to decide whether they want to continue reading.

In the introduction (to a poster) it can also be appropriate to present a labelled diagram of the concept/idea being investigated.

Introduction
a detailed but succinct explanation of the reason for undertaking an investigation; includes key chemical concepts, aim and hypothesis

Introduction

① Acids are substances that donate protons. There are many different types of acids, some strong and some weak. Acids have a low pH. Acids react with some substances to produce a gas. The volume of gas produced will vary depending on the amount of reactants used. The theoretical and experimental volumes of gases produced from such reactions can be compared to determine if they are similar or varied.

② Reactions of acid with metal carbonates produce a salt, water and carbon dioxide. The volume of carbon dioxide produced can be determined experimentally using a closed system where the gas is collected by a gas syringe. In this experiment, a 1.00 M solution of hydrochloric acid (HCl) will be reacted with different starting masses of sodium carbonate (Na_2CO_3) in order to determine what mass is required to produce 25 mL of carbon dioxide.

Theoretically, this can be determined by writing out the fully balanced chemical equation for the reaction between HCl and Na_2CO_3, as is shown below:

$$2HCl(aq) + Na_2CO_3(s) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$$

Then, using knowledge of the ideal gas equation, the amount of carbon dioxide gas (in mol) can be determined if it is known that 25.0 mL is to be produced, as shown below:

$$PV = nRT$$

When rearranged, $n = \frac{PV}{RT}$ (assume standard lab conditions for pressure and temperature).

$$n = \frac{(100 \times 0.0250)}{(8.31 \times 298)}$$

$$n = 0.0010 \text{ mol}$$

Using the mole ratios in the equation, sodium carbonate is present in a 1:1 ratio with carbon dioxide. Therefore, the amount in mol of sodium carbonate is also 0.0010 mol.

Then, using the molar mass of sodium carbonate, the theoretical mass of this required to react with hydrochloric acid to produce 25.0 mL of carbon dioxide can be determined, as shown below:

$m = nM$

$m = 0.0010 \times 106$

$m = 0.106$ g

Experimentally, it is very rare to generate 100% yield, so this investigation aims to determine the practical mass required to generate 25 mL of carbon dioxide gas.

Aim
the main purpose of an investigation and what you hope to achieve

Aim

① To determine the mass of metal carbonate that will produce the largest volume of carbon dioxide.

② To confirm if the experimental mass of sodium carbonate predicted from the theoretical calculations will result in the same amount of carbon dioxide gas produced when reacted with hydrochloric acid.

The aim includes explicit reference to the independent and dependent variable. This will be included in your introduction for the final scientific poster presentation.

Hypothesis
an idea or explanation for something that is based on known facts but has not yet been proven

Model
a representation of an idea, object or process that can be used to explain phenomena that cannot be directly experienced

Law
a description of an observed phenomenon

Theory
an explanation based on facts that have been confirmed through both experimentation and observation multiple times

Hypothesis

① 1.00 g of metal carbonate will produce the largest volume of carbon dioxide when reacted with hydrochloric acid.

② If the mass of sodium carbonate is exactly 0.106 g then the volume of carbon dioxide gas collected will be exactly 25.0 mL because this can be determined theoretically using the fully balanced chemical equation for the acid and metal carbonate reaction, combined with the ideal gas and mole equations as represented in the introduction.

The hypothesis is an explanation of what you think will occur. It does not have to be correct, but should be supported by your knowledge of chemistry, which is the difference between Investigations 1 and 2.

It is important to use appropriate language and not confuse an aim with a hypothesis. This is also the case with many other scientific terms that you may use in your investigation, such as **model**, **law** and **theory**.

Independent variable
the variable for which quantities are changed by the experimenter

Independent variable

① The mass of metal carbonate

② The starting mass of sodium carbonate weighed in grams using an electronic balance (0.086 g, 0.096 g, 0.106 g, 0.116 g, 0.126 g)

The independent variable should be specific and list all quantities or changes being investigated. It should also list the control group, if one is included in the investigation.

Dependent variable the variable that changes in response to changes in the independent variable; the experimenter measures these changes

Dependent variable

① Volume of carbon dioxide produced

② The volume of carbon dioxide gas produced, measured using a gas syringe in millilitres.

State how the dependent variable will be measured in your experiment. This could be using an instrument and units of measurement, or even a simple scale of effects. The meaning of each number in the scale (e.g. 0 = no symptoms, 5 = extreme symptoms) must always be given.

Controlled variable anything kept constant, or monitored, so it does not affect the independent and dependent variables and therefore the validity of experimental results

WORKSHEET 16A–1
IDENTIFYING VARIABLES

Controlled variables

① Environmental conditions
Equipment
Acid
Metal carbonate

② Same concentration of hydrochloric acid (1.00 M)
Same volume of hydrochloric acid (10.0 mL)
Same purity of sodium carbonate
Same electronic balance used to record mass of sodium carbonate
Same gas syringe used in each trial
Same temperature (even though this may not be at standard laboratory conditions, or SLC, it should be consistent throughout all trials)

There will always be more than one controlled variable in any investigation. List as many as you can for your investigation. (When answering questions on this, usually only two are required.)

The word 'same' is used for each variable to show that the variables have been kept constant and so the results are valid.

Control group the set-up or group in an experiment that does not receive treatment; it is used as the 'standard of comparison'

Positive control a control group that receives a treatment with a known response that can then be compared to the experimental group(s)

Negative control a control group that isn't expected to produce a result

Control group

① The control group is the experiment set-up at standard laboratory conditions.

② As this experiment is determining if the theoretical mass of 0.106 g of sodium carbonate will result in exactly 25 mL of carbon dioxide gas produced, there is not really a control group required. A negative control could be included, which would be a reaction whereby no sodium carbonate is added to the reaction to prove that the production of carbon dioxide gas is caused by the addition of the sodium carbonate.

If possible, a control group should be included in an investigation.

Including a control group gives a known result for comparison with the experimental groups. You may need to account for the difference between a **positive control** and a **negative control**.

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Method
a series of numbered steps describing the procedure

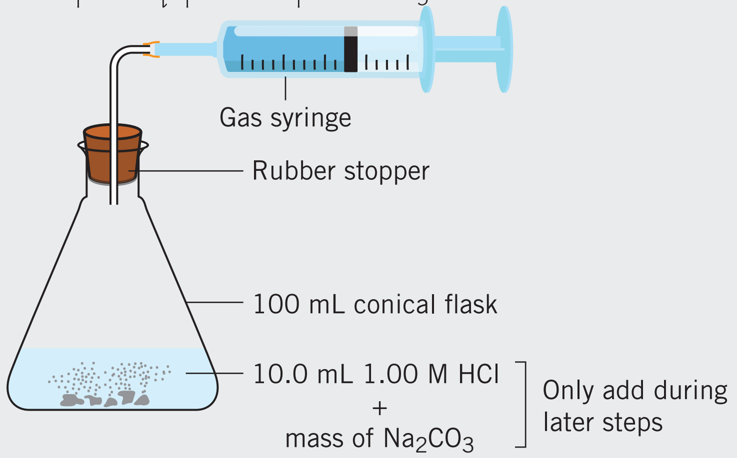
Method

①

1 Place metal carbonate into test tube.
2 Add enough acid to cover the metal carbonate in the test tube.
3 Place stopper in top of test tube with the syringe connected.
4 Measure the volume of gas produced.

②

1 Set up the equipment as per the diagram below:



2 Using the electronic balance, weigh 0.086 g of sodium carbonate into a weigh boat. Then add this to the 100 mL conical flask.
3 Measure out 10.0 mL of 1.00 M HCl using a 10 mL measuring cylinder.
4 Add the HCl to the conical flask and immediately place the stopper containing the gas syringe to seal the conical flask. Note: make sure you do not prevent the movement of the gas syringe plunger in any way as the gas is being produced.

The method must contain enough information for the investigation to be repeated by yourself (repeatability) and others (reproducibility). Therefore, it must include specific quantities of any substances and specific equipment used. This helps to establish validity.

The investigation also needs to be repeated, ensuring there is a large sample of results – the larger the sample size, the more reliable the results.

The method should also include the set-up for the control group (if there is a control group).

5 Once the reaction has gone to completion (as indicated by all sodium carbonate disappearing or no more bubbles present), record the volume of carbon dioxide produced in the gas syringe by reading the graduated markings.
6 Repeat steps 1–5 two more times for the same mass of sodium carbonate and record results.
7 Repeat steps 1–6 for the other masses of sodium carbonate to be tested (0.096 g, 0.106 g, 0.116 g and 0.126 g).

Health, safety and ethical considerations

① Wearing of lab coat, safety glasses and gloves.
There are no ethical issues to consider.

② For this experiment, wearing of personal protective equipment (lab coat, safety glasses and gloves) is required, especially when handling chemicals such as hydrochloric acid, and particularly if diluting this from a higher concentration stock solution. It is important to ensure that hands are washed thoroughly after the investigation.
(An example of an ethical issue for a different experiment using lamb liver is the following:
The lamb liver was obtained from a verified butcher that only obtains meat from farms that practise correct handling of animals and processing of different body parts following death. When alive, these animals were treated well and given access to sufficient food and water and suitable environments for exercise.)

Many schools have access to an online risk assessment program that allows you to input equipment and any chemicals specific to your own investigation and generates safety requirements to be followed. With the use of living organisms, ethical integrity when using data and reporting on outcomes should be clearly considered and outlined. Special consideration should be given to the ethics of any investigation (such as Investigation topic 3 in the VCE Chemistry Study Design) involving Aboriginal and Torres Strait Islander knowledge and culture. Refer to the online guidelines listed in the overview at the front of this textbook for further information.

Repeatability recording of results produced when the experiment is repeated in one lab by one operator under the same conditions

Reliability (experimental) the extent to which an experiment always yields the same results under the same conditions

Repeatability

① Repeat experiment three times to determine the maximum amount of metal carbonate required to produce the most amount of gas.

② This experiment should be conducted three times for each of the different masses of sodium carbonate added to the 1.00 M hydrochloric acid in the conical flask. This is to increase the sample size and ensure repeatability (or reliability) of results.

Without repeating your experiment, your results are potentially valid but not **reliable** (representative of normal conditions). It is important in any investigation to have a large sample size.

Reproducibility
when the same results are obtained for the same experiment by different operators using different equipment

Reproducibility

① Repeat experiment three times to determine the maximum amount of metal carbonate required to produce the most amount of gas.

② The experiment could be repeated by other investigators using a different concentration of acid or type of strong acid, or with different masses of sodium carbonate (or another metal carbonate) to achieve the same (or different) volume of carbon dioxide produced. Results could then be compared, based on similarity of experimental set-up and aim of the investigation.

Other experimenters achieving the same result as you after conducting the same investigation further strengthens the validity and reliability of your results. It is worth researching if others have done your experiment previously, or getting your peers to perform your experiment using your method to see if they achieve the same results. It is important not to confuse reproducibility and repeatability.

Experimental validity
the extent to which all variables in the experiment have been controlled, so that the independent variable is the only factor that changes

Single-variable exploration
an investigation that contains only one independent and one dependent variable

Experimental validity

① All variables are controlled in the experiment. Therefore, it is a valid test.

② This experiment contains one independent variable (mass of sodium carbonate added to the hydrochloric acid) and one dependent variable (volume of carbon dioxide gas produced), where all other factors in the experimental design have been controlled (same concentration and volume of hydrochloric acid, same purity of sodium carbonate and same temperature). Hence, it is a single variable exploration.

It is crucial to ensure that the independent variable is the only aspect changing in your investigation (**single-variable exploration**) and that all other variables are controlled. If this is not done or noted, then your work is unable to be accepted by your peers, teacher or the broader community.

16A SKILLS

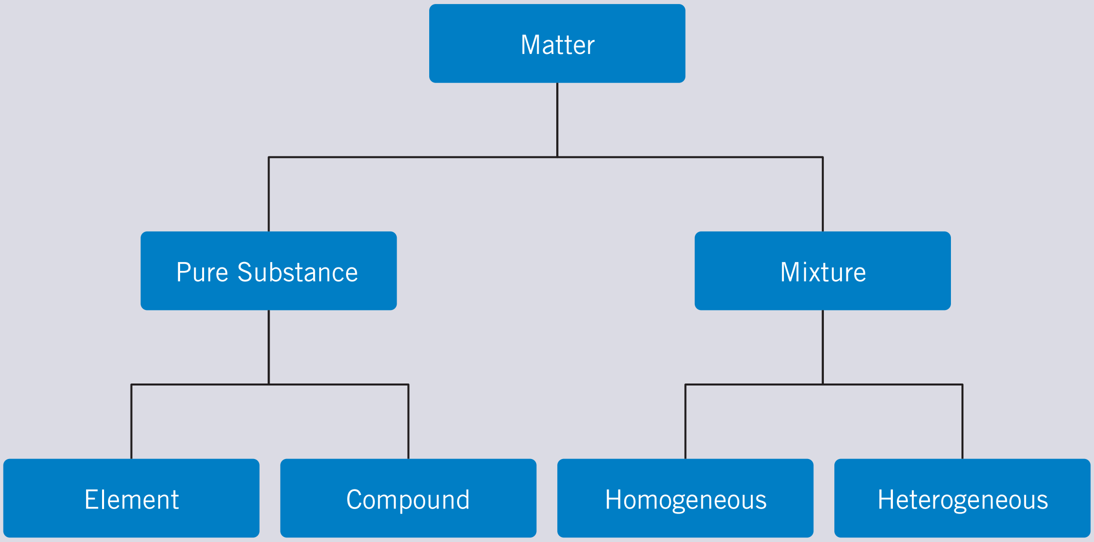
Scientific methodologies

In this chapter, you have seen how the scientific investigation process is used to establish facts or acquire knowledge. It involves formulating a hypothesis, collecting data, analysing the data and drawing conclusions. These principle steps of the scientific process can be applied across all disciplines of science; however, the specifics of each step may change depending on the aim of the investigation and what is being examined. In this 'scientific skills' section, you will learn about different scientific methodologies that may be applied to investigations.

Classification and identification

Classification is the arrangement of phenomena, objects or events into groups or categories based on their relationships. Identification involves allocating unidentified objects to the correct category once a classification has been established.

Classification can involve a simple set of categories (for example, acids and bases) or can be performed using a tree structure to show how some classes can be further divided into more specific groups. For example, matter may be classified as follows:



Observations from scientific investigations can be used to identify an unknown object or event. Using the classifications examples above, we could make the following identifications based on observations:

Observations	Identification
• Does not separate into components • Made up of more than one element	Compound
• Liquid appears uniform • Separates into multiple components when analysed by TLC	Homogeneous mixture

Think about some of the other classification systems in chemistry that you have come across in Units 1 and 2.

Fieldwork

Fieldwork involves collecting data at a specific location, rather than in the semi-controlled environment of a lab or classroom. Environmental chemists often perform fieldwork, travelling to locations to collect and analyse samples to investigate the effects that chemicals have on the air, water, and soil.

The generation of site-specific data should be recorded in your logbook. An example of a portion of a logbook entry for a fieldwork investigation is shown below. Section 16B further explores how to present results of your investigations in your logbook.

LINK 16B SCIENTIFIC EVIDENCE

Aim: To investigate the water quality of Lake Daylesford

Location: End of Lake Daylesford jetty

Date: 9/10/2022

Results:

Test	Result	Units
1. Temperature	12	°C
2. pH	6.8	pH units
3. Dissolved oxygen	7.1	mg L^{-1}
4. Total phosphates	0.4	mg L^{-1}
5. Nitrates	8.9	mg L^{-1}
6. Total solids	5.5	mg L^{-1}

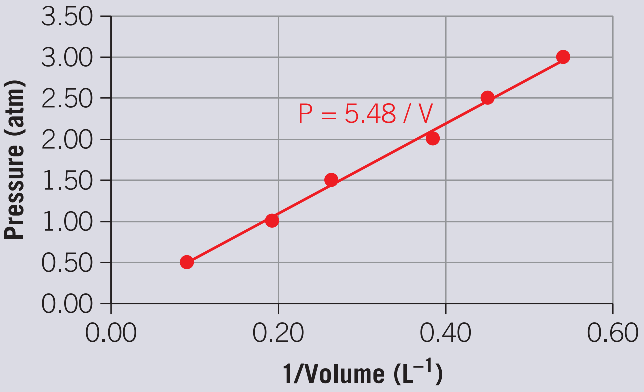
Modelling and simulation

Scientific modelling involves the construction of a physical, conceptual or mathematical model, which represents a real phenomenon. Models provide a visual way of linking theory with experiment and provide a way to explain or predict behaviour of objects or systems. Simulation involves using a model to study the behaviour of a real or theoretical system. Performing simulations can overcome experimental limitations; for example, a system may be too complex, too large or small, too fast or slow, inaccessible or too dangerous.

One model that you will have already come across is Bohr's model of the structure of the atom, which was primarily based on the emission spectrum of hydrogen.

An example of a simple mathematical model is shown below. In this investigation, a student examined the relationship between volume and pressure using a gas syringe and a pressure gauge. An inverse relationship between volume and pressure was observed. Further analysis involved mathematical modelling of the data by fitting to a linear function.

Volume (L)	1/Volume (L^{-1})	Pressure (atm)
11.0	0.09	0.50
5.20	0.19	1.00
3.80	0.26	1.50
2.60	0.38	2.00
2.22	0.45	2.50
1.85	0.54	3.00



Based on the mathematical model determined, can you predict what the volume of the system would be if the pressure is 0.75 atm?

Far more complicated models have been developed by scientists to identify and understand patterns in our world. For example, mathematical models have been used to simulate future temperature increases due to climate change.

Product, process or system development

Some scientific investigations involve the design of a product, process or system to meet a particular need. This may involve technological applications in addition to scientific knowledge and procedures.

You will have encountered many examples where scientific knowledge has been used to generate a range of everyday products, such as:

- knowledge of measurement of pH to design kits for testing the acidity of swimming pools and aquariums
- knowledge of removal of impurities from water to build industrial water treatment and sewerage plants
- knowledge of redox reactions to generate batteries and power sources that produce greener, cleaner energy.

Section 16A questions

1 What features does a controlled experiment include?

2 Define the following terms: independent variable, dependent variable, controlled variable.

3 What is the difference between a positive control group and a negative control group?

4 Compare validity and reliability.

5 Outline the difference between repeatability and reproducibility.

6 Explain what single-variable exploration means.

7 Two scientists are investigating the role of UV light on the substitution reaction between hexane and bromine gas. Michael says that a suitable control group would be placing the reaction vessel in question in a dark cupboard away from any light. However, Simar says that the control group should be the volume of hexane used in the experiment.

   a Explain why Michael is correct.

   b What was the mistake Simar made in her comment?

   c Design an experiment to test the question: 'Does UV light promote substitution reactions of alkanes with halogens?' In your answer include the following: independent variable, dependent variable, controlled variables, control group and method.

8 A chemistry student performs an acid–base titration to determine the concentration of citric acid in **three** different brands of orange juice. The student dilutes 5 mL of each orange juice to 100 mL with distilled water. The student transfers 25 mL of a diluted juice solution to a conical flask and adds three drops of phenolphthalein indicator. The student then performs a titration with a standardised NaOH solution. The student performs **three** repeated titrations for each brand of orange juice.

   Using what you have learned in this chapter, write a method for this experiment. Ensure you mention all relevant variables in your response.

9 A group of chemistry students have developed a polymer for 3D-printing applications that they believe thermally degrades into a series of small organic molecules and oxygen gas at moderate temperatures. Design an experiment that could be used to test the temperature dependence of the polymer degradation. Include all necessary information.

10 Many pharmaceuticals are formulated as ionic compounds in order to increase their water solubility and therefore reach the desired concentration in systemic circulation for achieving the required pharmacological response. Three common anions used in the formulation of ionic pharmaceuticals are chloride, sulfate and acetate. Design an experiment to test the solubility of a novel drug formulated with these anions. In your experimental design, be sure to include:

- the independent variable
- the dependent variable
- at least two controlled variables
- the expected outcomes.

16B Scientific evidence

Study Design:

Investigation design

- Techniques of primary qualitative and quantitative data generation relevant to the investigation
- Accuracy, precision, repeatability, reproducibility, resolution, and validity of measurements in relation to the investigation

Scientific evidence

- Ways of organising, analysing and evaluating generated primary data collected to identify patterns and relationships, and to identify sources of error
- Observations and investigations that are consistent with, or challenge, current scientific models or theories
- The characteristics of primary data
- The limitations of investigation methodologies and methods, and of data generation and/or analysis

Glossary:

Accuracy
Conclusion
Continuous data
Discrete data
Line of best fit
Outlier
Precision
Qualitative data
Quantitative data
Random error
Resolution
Systematic error
True value

Structuring the logbook – scientific evidence

This section discusses how to represent the results of your investigation in your logbook, including: drawing up tables and graphs, analysing the data collected, commenting on any errors, ensuring the precision and accuracy of the data, and noting areas for improvement.

Qualitative data data that is descriptive (not numeric)

Quantitative data data that is measured and represented numerically

Resolution the smallest change in the quantity being measured that results in a noticeable change in the value shown on the piece of measuring equipment

Before recording your results, it is important to understand the difference between **qualitative data** and **quantitative data**. Qualitative data is *descriptive* – this means it is in the form of words, not numbers. For example, it could be the appearance of something (e.g. 'cloudy' or 'clear') or colour (e.g. 'red' or 'yellow'). Quantitative data is *numerical* – this means it is in the form of numbers, based on counting or measuring. For example, it could be volume (e.g. 25 mL) or pH recorded on a scale of 0 to 14.

Knowing this difference is important when determining what type of graph should be used to represent the data you have generated.

When collecting quantitative data using measuring equipment or devices, it is important to also consider the **resolution** of this data. Resolution refers to the smallest change in the quantity being recorded that results in a noticeable change in the value shown on the measuring equipment. For example, if the gas syringe, shown in Figure 16B–1, used to record the volume of carbon dioxide gas produced has a resolution of 1.0 mL, then the experimenter can only estimate the volume between the two marked intervals on the syringe. As such, the data recorded will only be a whole number. This means that a data reading of 11 or 12 mL is possible, but a reading of 11.5 mL would not be possible. Alternatively, if you were using a burette, shown in Figure 16B–1, where the resolution is 0.1 mL, then the experimenter can estimate the volume to one decimal place. For example, a volume recorded could be 40.5 mL or 40.6 mL. However, it again cannot be between these intervals. The meniscus of the liquid in the burette would be either sitting on the interval line, so it would be 40.6 mL, or somewhere between the two interval lines of 40.5 mL and 40.6 mL, and therefore estimated to be 40.55 mL.

Consequently, when recording primary data – that is, data collected firsthand by the experimenter in their results table – the number of decimal places recorded is dependent on the instrumentation used. This should therefore be consistent for all data collected.

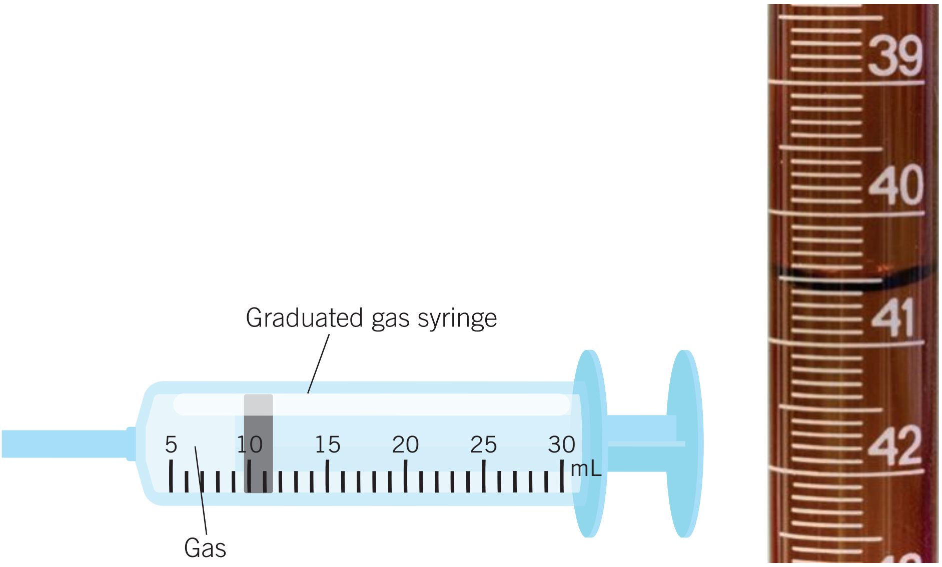


Figure 16B–1 Different pieces of equipment will have different resolutions, depending on their graduated markings. The gas syringe (left) has markings every 1 mL and therefore has a resolution of 1.0 mL, whereas a section of the burette (right) has markings every 0.1 mL and therefore has a resolution of 0.10 mL. Instruments, such as these affect the number of decimal places data can be recorded to.

Results – table

①

Mass of metal carbonate	0.086	0.096	0.106	0.116	0.126
Trial 1	20	22.5	25	26	26

A results table should include a title with both variables (independent and dependent) mentioned.

②

Title: Changes in initial mass of sodium carbonate and volume of carbon dioxide gas produced experimentally

Results – table

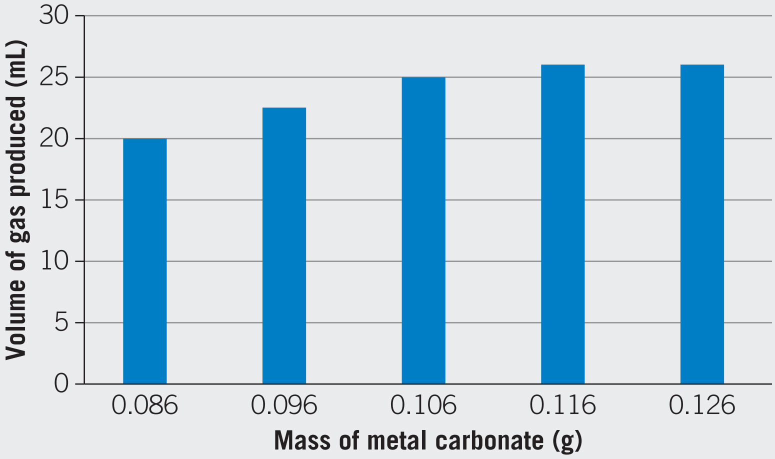
Initial mass of Na_2CO_3 (g)	Volume of CO_2 produced (mL)			
	Trial 1	Trial 2	Trial 3	Average
0.086	19.4	19.1	18.9	19.1
0.096	20.6	20.7	20.9	20.7
0.106	22.7	22.3	23.0	22.7
0.116	24.1	23.7	23.9	23.9
0.126	24.8	25.2	25.0	25.0

All columns/rows must be labelled with an appropriate heading and relevant units. It is good to include an average, or a percentage, from multiple repeats. The averages should be rounded to the same number of decimal places as the data, one decimal place in this case. The table can be drawn by hand or generated digitally.

WORKSHEET 16B–1
REPRESENTING DATA

①

Results graph – Investigation 1



Graphs should include a main title and have both *x*- and *y*-axes labelled with units (if required). The *x*-axis is for the independent variable and the *y*-axis is for the dependent variable.

The correct type of graph must be used. **Discrete data** (based on counting), should be represented as a bar graph. If data is **continuous** (based on measuring, e.g. changing time, temperature, volume, concentration, pH), the data should be represented as a line graph.

The graph can be hand-drawn, but for reports it is probably best if it is digital.

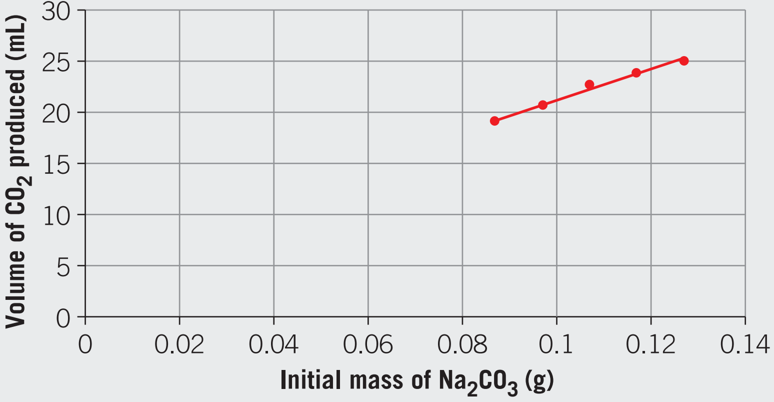
Discrete data
data that is countable and in discrete categories; contains distinct or separate values; best represented by a bar graph

Continuous data
data that is measurable and continuous, with infinite possible values; best represented by a line graph

②

Results graph – Investigation 2

Changes in initial mass of sodium carbonate and volume of carbon dioxide gas produced experimentally



Accuracy

① This experiment is accurate, as data was generated using correct experimental procedure.

② To fully determine the **accuracy** of this experiment, the experimentally achieved results for the volume of carbon dioxide gas produced would need to be compared to the theoretical volume of carbon dioxide gas produced for each initial mass of sodium carbonate weighed and tested. As such, there is a definitive **'true' value** for the volume of gas produced.

This experiment was not accurate as less carbon dioxide gas was produced for each initial mass of sodium carbonate than was predicted theoretically from the chemical equation and calculations.

The accuracy of an experiment is not always known. It depends on how unique your investigation is. Most experiments you perform will have been completed by others previously, so you are just reproducing results. In those cases, a true value is known and you can compare your own results to them.

Accuracy
how close the measurements are to the 'true' value of the quantity being measured

True value
the value or range of values that would be obtained if the quantity could be measured perfectly

Systematic error
when the readings obtained from measurements differ from the 'true' value consistently in one direction every time

Systematic errors

① Each time the mass of metal carbonate was weighed using the scales, the measurements fluctuated, which could have affected the results.

② The way the experiment was set up meant that the carbon dioxide produced from the reaction needed to travel from the conical flask through the tubing and into the gas syringe. This would not be an accurate representation of all the carbon dioxide collected in the gas syringe as some remained in the conical flask, thereby affecting the results of the investigation for each mass of sodium carbonate tested.

As all the results in this experiment were lower than the predicted theoretical calculation for the volume of carbon dioxide, this could be in part attributed to the fact that the investigation was not conducted at true standard laboratory conditions (SLC). This was supposed to be 25°C; however, the room temperature was always lower than this.

A systematic error cannot be improved by repeating measurements, having a larger sample size or taking a mean. As the results are always out by a consistent value, they will always be inaccurate. Therefore, a change in the method or equipment must occur.

Random error
an unpredictable variation in the readings obtained due to variables not all being controlled (extraneous variables), resulting in the readings being higher or lower than expected

Outlier
a reading that is very different from other results obtained for the same measurement

Random errors

① There were no random errors because the experiment was only conducted once for each mass of metal carbonate tested.

② Random errors were present in this experiment, as it is clear that there are differences in measurements between the three trials for each mass of sodium carbonate tested. This could be due to variations in measuring the precise volume of 10 mL of 1 M hydrochloric acid. Depending on how the experimenter reads the bottom of the meniscus in the measuring cylinder, this could be equally likely to be higher or lower if not measured correctly.

Each time the 10 mL of 1 M HCl was added to the initial measured mass of sodium carbonate in the 100 mL conical flask, there was a short period of delay before the stopper connected to the gas syringe could be placed on the conical flask to seal the carbon dioxide gas in. This resulted in a small amount of gas escaping into the atmosphere and therefore not being measured by the change in position of the gas syringe plunger.

Other examples of errors in data were discussed in Chapter 14A Skills.

14A SOURCES OF SALT IN WATER LINK

The effect of a random error can be reduced by repeating measurements, having a larger sample size and/or finding the mean. As results are equally likely to be high or low, averaging can improve the precision of results. For most experiments in a laboratory where multiple trials are conducted, you will need to account for the effect of these errors on your investigation. You may need to also account for **outliers** by repeating measurements.

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Precision
how close all the measurements are to each other

Precision

❶ The results were not precise as the volume of gas produced reached a maximum at a mass of metal carbonate of 0.12 g but then did not further increase at 0.14 g.

❷ The three trials for all masses of sodium carbonate tested were similar but not always the most precise. However, an average of the three trials was taken to improve the precision of the results as much as possible.

Precision can be achieved in your own investigations by ensuring that a large sample size (repeats) is used and a mean is calculated. Precision is more easily identifiable with continuous data measurements.

Discussion

The discussion is the most crucial aspect of the investigation. Use your logbook to write down any notes about what you plan to address in this section. The discussion should include the following key content:

- *Interpret and evaluate the trends and patterns in your data* – use a graph if possible, as this will clearly show the relationship between the independent and dependent variables. It is important here to also quote relevant data from the table/graph when referring to trends in the data.
- *Acknowledge any deviations (outliers) in the data from the results that were expected* – relate the results to the relevant chemistry or scientific theory and key terms. This is very important, as this outlines how well you understand your results and can interpret them based on your knowledge of the theory. Therefore, use the knowledge in previous chapters of this text to help with the respective topic you are investigating.
- State whether the data you collected supports or doesn't support the hypothesis.
- Identify *any limitations in the data or the method* – refer to random and systematic errors, and accuracy and precision, as discussed earlier in the chapter.
- Suggest *future improvements to the investigation* if it were to be performed again – this means referring to experimental errors, not human errors (for example, incorrect measuring of volume or not calibrating electronic balances). If your experiment is performed correctly, there should be no human errors.

NOTE

Avoid terms such as 'proved', 'disproved', 'correct' and 'incorrect' in relation to your hypothesis, as it is unlikely that you can be this certain from a single investigation. Instead use terms such as 'supported', 'not supported', 'indicated' and 'suggested'. If your results contradict your predictions, this would warrant repeating the experiment, if you have time. If you cannot repeat the experiment, then your report should include a discussion of flaws in the design or method and suggestions for how the investigation could be altered to minimise or eliminate these.

Conclusion
a summary of what you can deduce from the results of the investigation, including whether the tested hypothesis was supported

Conclusion

The main purpose of the **conclusion** is to briefly summarise the position of the experiment in the wider understanding of the chemistry topic(s). You need to state the important overall trend of the data (referring specifically to data from your results) and whether or not the results support the tested hypothesis. The conclusion should assess whether the results of the experiment have contributed new information to what is known about the topic and any further investigations that need to be undertaken. The conclusion should not introduce any information that has not already been discussed in the results and discussions sections.

For example, in the experiment described in this chapter, the conclusion might read like this:

> In conclusion, the results indicate that the experimental volume of carbon dioxide produced from the initial starting mass of sodium carbonate did not match that of the predicted theoretical calculations. At a mass of sodium carbonate of 0.106 g, the average volume of carbon dioxide gas produced was only 22.7 mL, as opposed to the predicted 25.0 mL, which actually required 0.126 g of initial sodium carbonate experimentally. As such, the results do not support the hypothesis. However, whether this is completely true should be further explored by conducting the investigation at true SLC conditions.

Check-in questions – Set 1

1 What is the difference between qualitative data and quantitative data?
2 What is the appropriate method of representing continuous data?
3 Compare random and systematic errors, including what results from each type of error.
4 What is the difference between accuracy and precision?

16B SKILLS

When completing work within this 'scientific skills' section, you will draw upon some very important skills. Many of these you will have learned in previous years of studying science, but it is particularly important to highlight some here, to ensure that you maximise your performance on any given assessment.

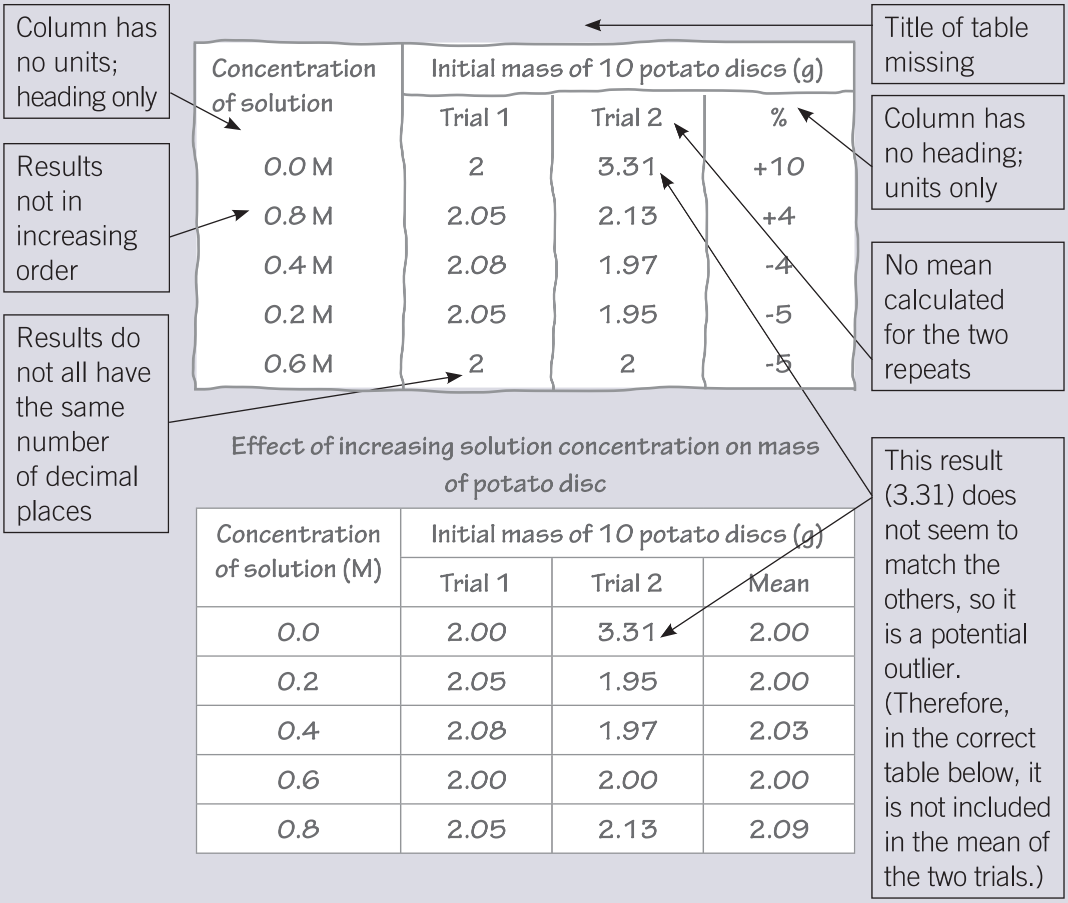
Recording results in a table

VIDEO 16B–1
RECORDING RESULTS IN A TABLE

When constructing and recording your results in a table:

- Rule the table in pencil, so any amendments can be made easily.
- Give each column a clear heading, including both the quantity and the unit it is measured in. Do not enter the units in the table along with each numerical value – the units go in the heading only.
- The independent variable is usually placed in the first column, with the dependent variable to follow in the other columns.
- Organise the results appropriately. For example, if your experiment involved testing an increasing concentration of a solute solution, your results should start with the lowest concentration and continue to the highest concentration.
- If recording quantitative results, all values should have the same number of decimal points.

- Include results for all repeats in the table and the mean (average) calculated for these.
- Any results that are outliers should be recorded again (repeat the measurement). If there is no time to repeat the experiment, include the outliers but ignore them when calculating the mean.
- Give the table an overall title. This should include mention of both the independent and dependent variables.
- In most cases, data from a table also need to be displayed as a graph: a line graph for continuous data, a bar graph for discrete data.



Column has no units; heading only

Results not in increasing order

Results do not all have the same number of decimal places

Title of table missing

Column has no heading; units only

No mean calculated for the two repeats

Concentration of solution	Initial mass of 10 potato discs (g)		
	Trial 1	Trial 2	%
0.0 M	2	3.31	+10
0.8 M	2.05	2.13	+4
0.4 M	2.08	1.97	-4
0.2 M	2.05	1.95	-5
0.6 M	2	2	-5

Effect of increasing solution concentration on mass of potato disc

Concentration of solution (M)	Initial mass of 10 potato discs (g)		
	Trial 1	Trial 2	Mean
0.0	2.00	3.31	2.00
0.2	2.05	1.95	2.00
0.4	2.08	1.97	2.03
0.6	2.00	2.00	2.00
0.8	2.05	2.13	2.09

This result (3.31) does not seem to match the others, so it is a potential outlier. (Therefore, in the correct table below, it is not included in the mean of the two trials.)

The top table lacks ruled lines, and there are errors in the representation of data. A corrected version is shown in the table below it.

Drawing graphs

When constructing a graph:

- Use pencil, as this will allow you to make any amendments easily.
- Put the independent variable on the x-axis (horizontal axis) and the dependent variable on the y-axis (vertical axis).
- Fully label both axes and include units (units should be the same as the results table if headings are correct there).
- The scale on the axes should have increasing values spaced at equal intervals, and it should be easy to read values between these intervals. Do not extend the scales too far beyond the recorded data values. Note: *You do not have to begin your scale at 0.*
- Make the graph as large as possible, so it is easy to read precise values.

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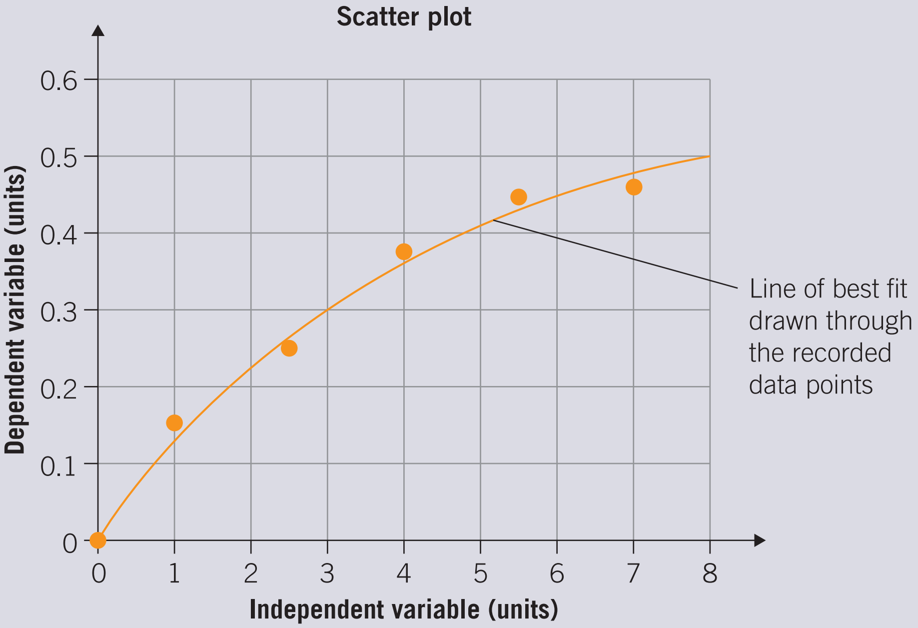
When drawing *scatter plots*:

- Use a scatter plot to represent continuous data.
- Plot data points as crosses (×) or dots (•). If using dots, be sure to draw them large enough so that they are not covered by the line drawn in the next step.
- Draw a **line of best fit**. This does not need to go through the first and last point, nor does it need to be a straight line (both common mistakes made by students).
- If constructing your line of best fit digitally, software will have functions that allow you to select 'line of best fit' when plotting data. If constructing this by hand, you will need to do this by eye, which can be quite difficult. The line can be linear or non-linear. Your aim here is to ensure that if you were to measure the distance between the data points below the line of best fit and the line itself, they would be equal to the distance between the data points above the line of best fit and the line itself. As you are unlikely to get this perfect on your first attempt, it is crucial you draw this line of best fit in pencil (as you should for all graphs anyway!).

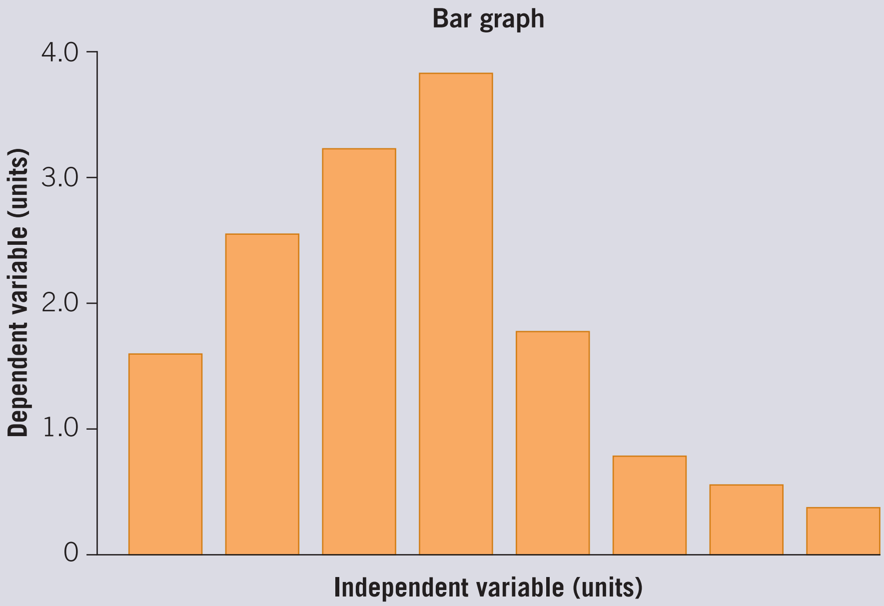
Line of best fit
a line on a graph that shows the general trend of the data points; the distance to the points above the line should equal the distance to the points below the line



VIDEO 16B–2
DRAWING GRAPHS

When drawing *bar graphs*:

- Use a bar graph to represent discrete data, and draw the bars with gaps between them.
- However for a histogram, a special type of bar graph showing the distribution of numerical data, the rules are different. It displays the frequency of values on the vertical axis that fall into defined ranges called 'bins' marked on the horizontal axis.



Ordering your discussion

When choosing an order for all the points to talk about in your discussion, follow this sequence:

1 Describe the overall trend (overall relationship between the independent and dependent variable).
2 Describe any changes in the gradient of the graph, particularly focusing on sharp changes and where they occur.
3 Quote data from the graph.
4 Explain the results using your scientific knowledge of the key concepts studied in your investigation.
5 Identify sources of errors in the experiment. Begin this section with any *systematic errors* and how these affected the accuracy of your measurements due to limitations in the apparatus, experimental technique or experimental design. Do not include human errors (e.g. incorrectly measuring an exact volume of solution). Then identify any *random errors* and how these affected the precision of your measurements.
6 Identify areas for improvement. Focus on how to reduce the errors mentioned in step 5. This could include things such as:
    - using better techniques for measuring the dependent variable
    - using equipment that is more likely to keep controlled variables constant and therefore make your data valid, as well as more precise and accurate
    - repeating the investigation to increase reliability.

Section 16B questions

1 A student is investigating the effect of pH on the rate of hydrolysis of a condensation polymer. The products of hydrolysis are water-soluble. They cut a piece of the polymer into five, roughly equal, pieces. They prepare five solutions of varying pH by performing dilutions of a 0.1 M HCl solution with distilled water to a total volume of 10 mL (pH = 1, 2, 3, 4 and 5). They place a piece of polymer into each of the tubes. After one hour, the solution is filtered and the remaining solid polymer is rinsed, dried and weighed, and compared to the initial mass (in grams). The student calculates the percentage mass loss of the polymer. The following results are produced.

pH	Initial mass of polymer (g)	Final mass of polymer (g)	Change in mass (final – initial) (g)	Percentage mass loss (%)
1	2.05	0.75	–1.30	63
2	2.10	1.50	–0.60	29
3	2.06	1.75	–0.31	15
4	2.05	1.86	–0.19	9
5	2.09	1.95	–0.14	7

- a Identify the independent variable and the dependent variable in this investigation.
- b What would be two controlled variables required in this experiment? Explain by stating:
  - i how you would control these variables.
  - ii the effect that these variables would have on the rate of hydrolysis if they were not controlled.
- c What is the purpose of a control group?
- d Does this experiment have a control group? If yes, identify which set-up it is. If no, identify and explain what would be an appropriate control group.
- e Are the results of this investigation qualitative or quantitative? Explain.
- f Describe whether the results of this investigation are reliable (use information in the question and results table to assist you).
- g Use the information from the table to construct a graph of the results. Be sure to draw the appropriate type of graph.
- h By looking at the table of results above, what could be done to improve the precision of these results?
- i Using the graph you drew for part **g**, in which pH solution, 2 or 4, is the rate of hydrolysis faster?

2 Ella was performing an experiment based on one completed previously by other researchers. Their experiment demonstrated that when exposed to air and light, olive oils undergo oxidation to form conjugated dienes and trienes (compounds containing two or three alternating double bonds). These compounds absorb UV light in the 200 to 300 nm region.

Like the previous researchers, Ella prepared six flasks containing the same extra virgin olive oil. She left three on a bench and placed the other three in a dark cupboard for one month. Ella prepared dilutions of the oil samples in octane and recorded the UV–visible spectra.

- a If Ella managed to get the same results as those of the previous researchers, does this refer to 'reproducibility' or 'repeatability'?

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- b List two controlled variables that Ella would include in her experiment.
- c State the independent variable in Ella's investigation.
- d State a possible dependent variable in Ella's investigation and how this could be measured.
- e Is the measurement of the dependent variable you identified in part **d** classified as qualitative or quantitative?
- f What results from this investigation would support the results of the previous researchers?
- g Ella runs a sample of octane on the UV–visible spectrometer, as well as the oil samples diluted in octane. What purpose does this serve?

3 A scientist wishes to analyse an unknown metal sulfate salt via precipitation reactions, using gravimetric analysis. This technique involves collecting and weighing the precipitate formed to calculate the percentage by mass of sulfate in the unknown sulfate salt and then using this percentage to identify the metal. The scientist has two potential reactants available, barium chloride and magnesium chloride. Recalling the solubility rules, the scientist hypothesises that only one of these reactants would form a precipitate with the metal sulfate. A method for an investigation is described in detail below.

Method

Part A: Precipitate formation

1. Label one clean dry beaker 'Reaction 1: $BaCl_2$' and another 'Reaction 2: $MgCl_2$.'
2. Add 0.50 g of the unknown metal sulfate salt to each beaker.
3. With a 100 mL graduated measuring cylinder, add 50 mL of distilled water to each beaker, followed by 20 drops of 5 M HCl.
4. Stir each beaker while heating over a Bunsen burner until the salt has completely dissolved.
5. Measure out 50 mL of 0.1 M $BaCl_2$ in a 100 mL measuring cylinder.
6. While the metal sulfate salt solution remains heating, slowly add the $BaCl_2$ solution, while stirring.
7. Repeat steps 5–7 using 0.1 M $MgCl_2$. Record your observations of both beakers in a table, as shown on the following page.

Part B: Gravimetric analysis

8. If a precipitate forms in Part A, allow the beaker to cool and any precipitate formed to settle in the beaker for approximately 30 minutes.
9. Sit a funnel in the mouth of a 500 mL Erlenmeyer flask. Take a piece of filter paper folded into a cone and place into the funnel.
10. Slowly pour the reaction mixture containing the precipitate into the funnel. Rinse any remaining precipitate from the beaker into the funnel using distilled water.
11. Carefully remove the filter paper from the funnel and transfer the precipitate to a pre-weighed glass crucible.
12. Heat the crucible over a Bunsen burner for five minutes.
13. Record the mass of the crucible precipitate.
14. Repeat steps 12 and 13 until a constant mass is achieved.

The results of the investigation are displayed in the table below.

		Reagent	
		Reaction 1: $BaCl_2(aq)$	**Reaction 2: $MgCl_2(aq)$**
Part A	Observation	White precipitate forms	No precipitate forms
Part B	Mass of unknown metal sulfate (g)	0.50 g	
	Mass of empty crucible (g)	50.02 g	
	Constant mass of crucible and precipitate (g)	50.81 g	

a Write a suitable hypothesis for Part A of this investigation.

b Write a balanced chemical equation for Reaction 1 (unknown metal salt and barium chloride). Hint: The formula of the metal sulfate can be written $M_2(SO_4)_x$ where M = metal and x = charge of the metal ion. Ensure that you include the states.

c Write a balanced chemical equation for Reaction 2 (unknown metal salt and magnesium chloride). Why was a precipitate not formed?

d Why was it necessary to heat the precipitate and weigh until a constant mass was achieved?

e In this investigation, an excess of the $BaCl_2$ solution was used. How would the results be affected if an excess was not used?

f Is this investigation reliable? Explain.

g Are the results of this experiment qualitative or quantitative? Explain.

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16C Scientific communication

Study Design:

Science communication

- The conventions of scientific report writing, including scientific terminology and representations, standard abbreviations and units of measurement
- Ways of presenting key findings and implications of the selected scientific investigation

ENGAGE

Communicating your work

Think about how you've probably presented your experimental work in science in previous years. You would have written up your planning stages and presented your results in a typical scientific report. While this is still done even at university, would you want to read a scientist's research presented to you in this way? Probably not. So scientists often communicate their findings in the form of a scientific poster. In the past few decades, posters have become a popular way to showcase new investigations and findings to the scientific community, nationally and internationally.

Producing a scientific poster involves developing skills and is not easy. It often requires a lot of thought and preparation, as well as the ability to keep your ideas concise.

EXPLAIN

Scientific poster template

A key part of the School Assessed Coursework (SACs) in VCE Chemistry Units 3 & 4 is a self-designed practical investigation presented in the format of a poster, using the following template:

<table>
<tr><td colspan="3">Title
Student name</td></tr>
<tr><td>Introduction
Methodology and methods
Results</td><td>Communication statement reporting the key finding of the investigation as a one-sentence summary</td><td>Discussion
Conclusion</td></tr>
<tr><td colspan="3">References and acknowledgements</td></tr>
</table>

This template is the same as the one provided in the VCE Chemistry Study Design. You'll notice from the template that the heading of each section correlates closely with the information that is recorded in your logbook from Sections 16A and 16B. This should therefore highlight the importance of the progression of information that is recorded there, as it will need to be transferred in a more succinct fashion onto your poster. The VCAA has also stipulated that this should not exceed 600 words. The centre of the poster occupies approximately one-quarter of the page and needs to be one sentence summarising the main outcomes of your investigation, thereby answering your investigation question (title). This should be as engaging as possible to encourage people to read the other sections of your poster.

16A INVESTIGATIVE PLANNING AND DESIGN LINK

16B SCIENTIFIC EVIDENCE LINK

You might be wondering where the aim and hypothesis from your logbook should be included on the poster. They are included in the introduction, following a brief explanation of the reasons for conducting your investigation, and link to the relevant chemical concepts.

The poster can be produced either electronically or in hard-copy format. It will of course depend on what your school chooses as the preferred format, and you should therefore follow this.

Purpose of the scientific poster

The scientific poster's main goal is to get your message, the findings of your investigation, across to everyone. It is designed not just for those with a scientific understanding of the investigation conducted, but those with a non-scientific background as well. The reason for setting it out in the template proposed on the previous page is for maximum impact and visual appeal. There are no large blocks of text that are not inviting to read. It is organised into columns, to help your readers follow the information. Including a table of results and a graph breaks up the text and avoids large blank areas. Lastly, the limitation on word count is so that the text is clear and concise. Your investigation should therefore be explicit and not too broad or overly complicated, so the results can be explained concisely.

VIDEO 16C–1 SCIENTIFIC POSTER EXAMPLE

In the poster, you should also include the sources of information (references) that you used when planning and conducting the investigation. You can also acknowledge those who supported your investigation, such as peers, your teacher or others. How to correctly complete references is explored in Chapter 9.

LINK **9A** PREPARING A RESEARCH INVESTIGATION

Other methods for presenting an investigation

In addition to presenting your results in the form of a poster, you should be able to discuss your investigation with your peers, teacher, family or anyone from a non-scientific background. Most of those in the scientific community prefer to have individual discussions with other experimenters about their work, rather than reading it from a poster. Talking about their work with others also gives the experimenter a chance to discuss aspects of their investigation that have not been included in the final presentation and to answer any follow-up questions.

Therefore, other methods for presenting the purpose and results of your investigation include: oral presentation, multimedia presentation, visual representation or even an article for a scientific publication. Tips and suggestions for these are included in the digital resource, in the form of downloadable documents on preparing an article for scientific publication and preparing an oral, multimedia or visual presentation.

DOCUMENT 16C–1 PREPARING AN ARTICLE FOR SCIENTIFIC PUBLICATION

DOCUMENT 16C–2 PREPARING AN ORAL, MULTIMEDIA OR VISUAL PRESENTATION

Section 16C questions

1. Where should you draw information from to complete your scientific poster?
2. What is included in the introduction section of the scientific poster?
3. What are the key sections of a scientific poster?
4. What is the purpose of the communicating statement in the centre of the poster?
5. What are some other ways of presenting the results of an investigation in chemistry, and in general, to the scientific community?

Chapter 16 review

Summary

Create your own set of summary notes for this chapter on paper or in a digital document. A model summary is provided in the Teacher Resources, which can be used to compare with yours.

Checklist

In the Interactive Textbook, the success criteria are linked from the review questions and will be automatically ticked when answers are correct. Alternatively, print or photocopy this page and tick the boxes when you have answered the corresponding questions correctly.

Success criteria – I am now able to:		Linked question
16A.1	Document investigations appropriately using a logbook	8☐
16A.2	Define key terms related to scientific skills	13☐, 16☐
16A.3	Define 'controlled experiment'	15☐
16A.4	List the features that a controlled experiment includes	2☐, 6☐, 13☐
16A.5	Distinguish between a positive and a negative control	16☐
16A.6	Understand the meaning of single variable exploration	15☐
16A.7	Define and identify independent, dependent and controlled variables	1☐, 11☐, 13☐, 15☐
16A.8	Select and use the procedures and equipment appropriate to an investigation	13☐
16A.9	Recall different types of scientific investigations that generate data, including fieldwork, simulations and modelling	16☐
16A.10	Distinguish between an aim, a hypothesis, a model, a theory and a law	3☐, 14☐
16A.11	Construct aims and questions for investigations	11☐
16A.12	Formulate hypotheses and predict possible outcomes	3☐
16A.13	Define reproducibility, repeatability, resolution and validity, and distinguish between these	13☐
16A.14	Design an investigation that is valid	4☐, 12☐, 15☐
16A.15	Design an experiment that is fully reproducible by others	5☐, 8☐
16A.16	Determine potential ethical issues with investigation design	ITB
16A.17	Identify how bias can be minimised in an investigation	16☐
16A.18	Follow clear guidelines for health and safety when undertaking practical investigations	16☐
16B.1	Define qualitative and quantitative	15☐
16B.2	Distinguish between qualitative and quantitative data	15☐, 16☐
16B.3	Include appropriate units of measurement for quantitative data	13☐, 15☐
16B.4	Design an investigation that accounts for accuracy and precision	4☐, 12☐, 15☐

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Success criteria – I am now able to:		Linked question
16B.5	Transform primary data into an appropriate format of results (table, flow chart, bar and/or line graph)	13☐, 15☐
16B.6	Identify trends in data	7☐, 15☐
16B.7	Define the different types of errors (random and systematic)	9☐
16B.8	Identify sources of error and outliers from primary data	13☐
16B.9	Use evidence to determine whether an investigation supports or discounts a hypothesis, scientific models or theories	16☐
16B.10	Identify areas for improvement in investigation design and analysis to increase accuracy and precision and reduce the likelihood of errors	16☐
16C.1	Appropriately communicate all aspects of a scientific investigation	10☐, 16☐
16C.2	Apply correct abbreviations to chemical terminology	15☐, 16☐
16C.3	Recall and convert between different units of measurement in chemistry	15☐
16C.4	Determine the correct number of significant figures to use in answers to questions	13☐, 15☐
16C.5	Justify conclusions and evaluate whether evidence supports or refutes the hypothesis	7☐, 10☐, 15☐, 16☐
16C.6	Interpret investigation outcomes in terms of broader chemical concepts	7☐, 15☐, 16☐

Multiple-choice questions

1 An investigation is performed to observe the reactivity of a series of metals with concentrated hydrochloric acid. What could be the dependent variable?

- **A** the volume of hydrogen gas produced
- **B** the change in mass of the reaction flask
- **C** the time taken for the production of bubbles to cease
- **D** all of the above

2 Aside from the independent and dependent variables in an investigation, what other type of variable is included to ensure it is valid?

- **A** changed
- **B** controlled
- **C** control group
- **D** precise

3 A hypothesis is best described as

- **A** a statement describing what the investigation hopes to determine.
- **B** a series of steps involved in planning an investigation.
- **C** clear ethical guidelines for how any living organisms should be handled in the experiment.
- **D** a prediction of what will be observed.

The following information relates to Questions 4–6.
A sphygmomanometer is an instrument used to measure blood pressure.

The table below shows measurements from four sphygmomanometers, W, X, Y and Z. Sphygmomanometer W is known to be accurate.

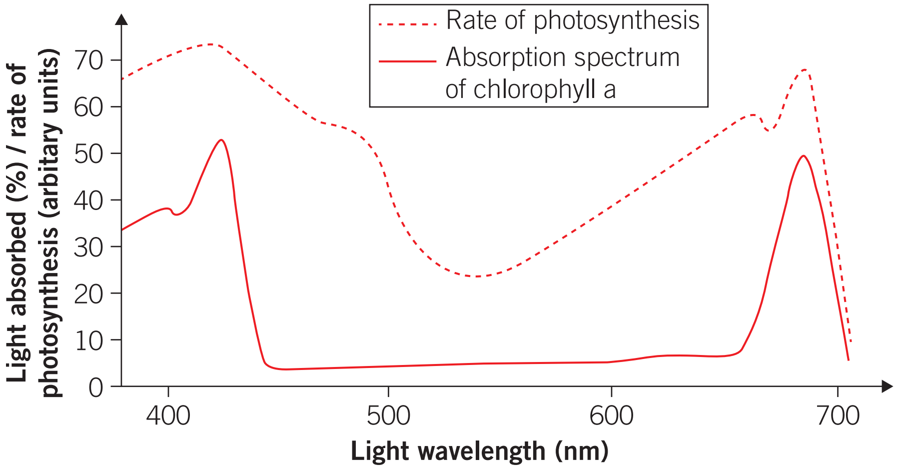
Sphygmomanometer	Pressure (mmHg)				
W	20.00	20.00	20.00	20.00	20.01
X	19.91	20.23	20.13	19.42	19.59
Y	19.00	20.40	19.50	20.10	21.00
Z	19.00	19.00	19.00	19.00	19.00

4 Which statement is correct?
- **A** The measurements from sphygmomanometer Z are more precise but less accurate than the measurements from sphygmomanometer Y.
- **B** The measurements from sphygmomanometer Z indicate a random error and are more accurate than the measurements from sphygmomanometer X.
- **C** The measurements from sphygmomanometer X indicate a systematic error.
- **D** The measurements from sphygmomanometer X have the same degree of precision as the measurements from sphygmomanometer Y.

5 Which of the following is **least** likely to improve the quality of the data in this investigation?
- **A** more blood pressure measurements recorded for each sphygmomanometer
- **B** more sphygmomanometers being tested
- **C** all groups using the same sphygmomanometer, W
- **D** taking an average of the combined results of all groups

6 The reason that sphygmomanometer W was included as a reference for accuracy in the investigation was to
- **A** remove any possible random errors from the results.
- **B** allow the pressure to be measured accurately.
- **C** ensure the other instruments were precise.
- **D** ensure the experiment generated qualitative data.

7 The graph below compares the absorption spectrum of chlorophyll a with the rate of photosynthesis of a plant upon exposure to a range of light wavelengths.



From this graph, it can be concluded that chlorophyll a is not the only pigment involved in photosynthesis. Evidence that supports this conclusion includes the fact that

A the rate of photosynthesis remains high when the plant is exposed to light wavelength between 450 nm and 650 nm. The percentage of light absorbed by chlorophyll a over these wavelengths is low.

B the rate of photosynthesis is low when the plant is exposed to light wavelengths such as 450 nm. Absorption of light by chlorophyll a is higher at these wavelengths.

C the rate of photosynthesis and the percentage of light absorbed by chlorophyll a are equal at 650 nm.

D between wavelengths 450 nm and 700 nm, the percentage of light absorbed by chlorophyll a is constant.

8 Which of the following statements would be appropriate in the 'method' section of a logbook?

A Exactly 50 mL of solution was measured using a measuring cylinder.

B Exactly 50 mL of solution was measured using a beaker.

C Solution was measured and placed in a tube.

D Add solution to a measuring cylinder.

9 A systematic error

A can be minimised by increasing the sample size and taking a mean.

B is equally likely to be quantitatively higher or lower.

C affects the precision of results.

D occurs consistently in one direction.

10 A conclusion should **not**

A summarise the key findings of an investigation.

B state key data from results.

C introduce new information relevant to the investigation.

D state whether or not the hypothesis was supported.

Short-answer questions

11 **a** Copy and complete the table below. (4 marks)

Hypothesis	Independent variable	Dependent variable	Controlled variables (×2)
Magnesium reacts more readily than iron with hydrochloric acid	Type of metal (magnesium or copper)	Time for bubbles to appear, time for metal to disappear	Same mass and shape of metal, same concentration and volume of HCl, same temperature
The solubility of potassium nitrate increases with increasing temperature			
The rusting of iron in water requires oxygen gas			
A solution of NaCl is a better conductor of electricity than pure water			

b For each of the hypotheses in the table, write an aim that would be appropriate for that investigation. (4 marks)

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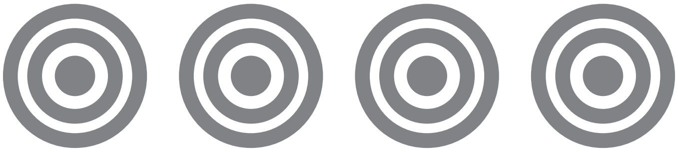
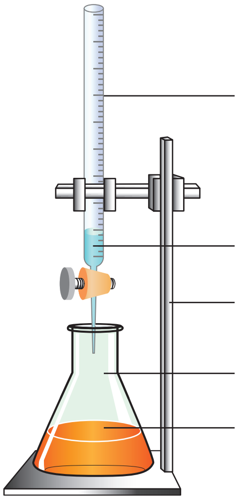
12 Redraw the standard bullseye/dartboard-style shown. Add three dots to each board to show each of the following types of data:



a accurate and precise
b accurate but not precise
c precise but not accurate
d neither accurate nor precise (4 marks)

13 A scientist conducted volumetric analysis on a sample of ethanoic acid that had absorbed some water from the air, to determine the percentage purity. The scientist weighed 1.00 g of the acid and diluted to 100 mL with distilled water. A 20 mL aliquot of this solution, containing three drops of phenolphthalein indicator, was titrated against 0.14 M sodium hydroxide. The scientist repeated the titration five times. The experiment set-up and results are shown below.



	Experiment				
	1	2	3	4	5
Initial burette reading (mL)	0.00	22.35	0.00	21.95	0.00
Final burette reading (mL)	22.35	44.50	21.95	44.00	22.21
Volume of NaOH (mL)					

a Label each of the pieces of equipment shown above and which solution is in which piece of glassware. (5 marks)
b Complete the final row in the table, indicating the volume of NaOH added in each titration. Highlight the values that are concordant. (2 marks)
c Draw a bar graph for the data shown in the final row of the table. Include axis titles and units. Indicate the concordant results on your graph with an asterisk (*). (4 marks)
d Explain why the scientist performed the experiment five times. (2 marks)

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e Calculate the average volume of NaOH added that you would use in your calculation to determine the ethanoic acid purity. Indicate this value with a line on your graph from part **c**. (2 marks)

f The scientist conducting the analysis averaged the results from all five repeated measurements. What effect would this have on the percentage purity determined? (2 marks)

g The scientist rinsed the burette with distilled water prior to the first titration. Explain why this could have caused the observed result determined from part **f**. (2 marks)

Another scientist repeated the investigation on the same ethanoic acid, following the same experimental method, however, obtained different results.

h Does this refer to the 'reproducibility' or 'repeatability'? In your answer, explain the difference between the two terms. (2 marks)

i Suggest one possible reason for the differences in the results obtained by the two scientists. (1 mark)

j Identify three controlled variables in this experiment. (3 marks)

14 Outline the difference between the following terms: model, law, theory. (3 marks)

15 A year 12 student designed an experiment to investigate the effect of light intensity on the rate of photosynthesis in *Elodea*, a freshwater pond weed. The experiment involved exposing sections of *Elodea* to varying light intensities for one hour. Light intensity was measured in arbitrary units, and the amount of oxygen gas produced was recorded and used to indicate the rate of photosynthesis. The student's results are shown in the table.

Light intensity (arbitrary units)	Oxygen production (μL)
0 (no light)	1
10	6
20	11
30	37
40	51
50	49
60	7
70	No data
80	1
90	1
100	1

a Is the data represented here qualitative or quantitative? Explain the difference between these two types of data. (2 marks)

b From the table of results provided, plot a graph of oxygen production versus light intensity. Then, using your graph, explain the trend observed in the data. (6 marks)

c **i** Before starting their experiment, the student wrote a hypothesis. Give an example of the hypothesis they may have written. (1 mark)

ii Identify the independent and dependent variables of the experiment. (2 marks)

iii To ensure the experiment was valid, the student made a list of factors that needed to be controlled. State two factors that the student would have included in their list. (2 marks)

iv Explain how this experiment supports a single variable exploration. (2 marks)

d **i** The student was unable to collect data for a light intensity of 70 arbitrary units. They used the graph to estimate the amount of oxygen produced at this intensity. Write their estimate. (1 mark)

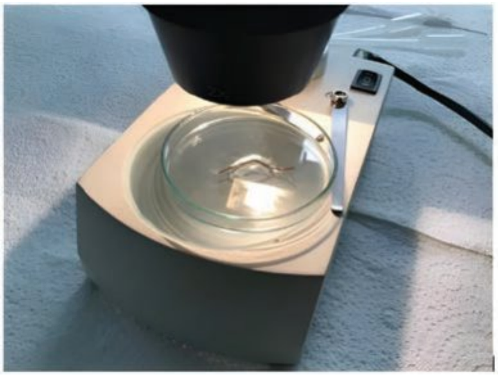
ii Consider your answer to part **d i**. Is this likely to be accurate? Give reason(s) for your answer. (1 mark)

e Explain why the student used the amount of oxygen gas produced by *Elodea* as an indicator of the rate of photosynthesis. (2 marks)

f Write a conclusion based on what the student's results suggest about the relationship between light intensity and the rate of photosynthesis. (2 marks)

g The student then decided to repeat the experiment using different sections of *Elodea*. Why was this a good idea? (1 mark)

16 Fatima investigated whether a displacement reaction occurred using combinations of different metals and ionic compound (salt) solutions. She used a stereomicroscope to observe the outcomes of the different reactions as shown below.



The method for the experiment is written below.

1. After your teacher has demonstrated how to set up and focus the stereomicroscope, place a Petri dish on the platform and use the forceps to place a piece of copper wire in the centre of dish. Focus the stereomicroscope on the wire. Then place two or three drops of the $AgNO_3$ solution over the centre of the wire and observe the changes. You will notice them immediately.
2. After a few minutes, carefully remove the Petri dish from the platform so that the crystals that have formed are not disturbed and place it next to the stereomicroscope for further observations. Record your observations in the table below.
3. Place a fresh clean Petri dish on the platform. Carefully place a small iron nail in the centre of the dish. Focus the stereomicroscope on the nail. Then place two or three drops of the $CuCl_2$ solution over the centre of the nail and observe the changes. You will notice them immediately.
4. Continue observing both reactions, swapping them out of the stereomicroscope carefully so that the crystals and solutions are not disturbed. If practicable, observe the crystals again on the following day. Record your observations in the table below.

Fatima recorded her results in a table, as shown below.

Mixture	Initial appearance of metal	Initial colour of salt solution	Changes observed
$Cu + AgNO_3$	Shiny bronze colour	Clear transparent solution	Brown substance was deposited on the metal. Solution turned slightly blue in colour
$Fe + CuCl_2$	Shiny silver colour	Light blue coloured transparent solution	Solution became clear over time and the metal had growth of black crystals on the outside

- **a** What type of data did Fatima collect? Explain. (1 mark)
- **b** What is missing from Fatima's results table? (1 mark)

Answer the following questions in relation to the chemical concepts covered by Fatima in this experiment.

- **c** Give the formulas of the species that were present before each reaction started (this should include any charged ions). (4 marks)
- **d** Based on Fatima's observations and given that in a chemical reaction, no new elements are created and none are destroyed, what must be the products of both reactions? State their formulas. (2 marks)
- **e** How can these changes be explained? What must have moved between the reactants? (2 marks)
- **f** Why do metals react in this way? (2 marks)
- **g** From the reactions that took place in this experiment, what can be determined about the reactivity of copper, iron and silver? Explain your reasoning. (3 marks)
- **h** From your answer to part **g** above, predict what would happen if you were to add some silver nitrate solution to an iron nail. State your reasoning. (1 mark)
- **i** Was a control used in this experiment? If yes, explain what it was. If no, explain what negative and positive controls could be included. (2 marks)
- **j** What safety requirements would Fatima have needed to implement in this experiment? Explain why in relation to the reactants that were used in the displacement reactions. (4 marks)
- **k** Identify an aspect of Fatima's method that could be adjusted to increase both the accuracy and precision of her results. Explain how making such an adjustment would improve the results by reducing the chance of bias or errors in Fatima's measurements. (4 marks)

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Unit 2 Revision exercise

Multiple-choice questions

1 Most water on Earth is located in

A lakes.
B oceans.
C glaciers
D the atmosphere.

2 The latent heat of vaporisation of water is best defined as

A the temperature at which water boils.
B the energy required to raise the temperature of one gram of water by one degree.
C the energy required to convert one mole of solid water to one mole of liquid water.
D the energy required to convert one mole of liquid water to one mole of gaseous water.

3 Consider the following statements.

I Water molecules consist of two hydrogen atoms and one oxygen atom.
II Hydrogen bonding exists between water molecules.
III There are covalent bonds between the hydrogen and oxygen atoms in a water molecule.

Which statements are true for gaseous water?

A I only
B I and III
C II and III
D I, II and III

4 In which of the following reactions is H_2O acting as a Brønsted–Lowry acid?

A $H_2SO_4(aq) + H_2O(l) \rightarrow HSO_4^-(aq) + H_3O^+(aq)$
B $Na(s) + H_2O(l) \rightarrow NaOH(aq) + H_2(g)$
C $NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$
D $HF(aq) + H_2O(l) \rightarrow F^-(aq) + H_3O^+(aq)$

5 What is the conjugate base of HPO_4^{2-}?

A PO_4^{3-}
B HPO_4^{2-}
C $H_2PO_4^-$
D H_3PO_4

6 Which of the following is an example of a concentrated solution of a weak acid?

A 0.01 M HNO_3
B 0.01 M CH_3COOH
C 10 M HNO_3
D 10 M CH_3COOH

7 What is the pH of a 0.005 M solution of $Mg(OH)_2$, assuming complete dissociation?

A 2.0
B 2.3
C 11.7
D 12.0

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8 Which of the following is **not** a redox reaction?

A $2KClO_3\ (s) \rightarrow 2KCl\ (s) + 3O_2\ (g)$

B $Cl_2\ (g) + 2KBr\ (aq) \rightarrow 2KCl\ (aq) + Br_2\ (l)$

C $Pb(NO_3)_2\ (aq) + Na_2SO_4\ (aq) \rightarrow 2NaNO_3\ (aq) + PbSO\ (s)$

D $2C_8H_{18}\ (g) + 25O_2\ (g) \rightarrow 16CO_2\ (g) + 18H_2O\ (l)$

9 In which of the following compounds is sulfur's oxidation state +4?

A S_8

B SO_2

C SO_3

D H_2SO_4

10 The oxidation state of chromium in $Cr_2O_7^{2-}$ is

A −2.

B +3.

C +6.

D +12.

11 In the redox reaction $2Ag^+(aq) + Sn(s) \rightarrow 2Ag(s) + Sn^{2+}(aq)$

A $Ag^+(aq)$ is the oxidising agent and gains electrons.

B $Ag^+(aq)$ is the oxidising agent and loses electrons.

C Sn(s) is the oxidising agent and gains electrons.

D Sn(s) is the oxidising agent and loses electrons.

12 Which of the following correctly describes the preparation of a solution with a 1 in 50 dilution?

A 5 μL of sample mixed with 250 μL of water

B 1 mL of sample mixed with 49 L of water

C 4 mL of sample mixed with 0.196 L of water

D 98 L of sample mixed with 2 L of water

13 The solubility of potassium nitrate in water at 25°C is 38 g per 100 g water. A solution of potassium nitrate is prepared at 25°C by adding 45 g to 200 g of water. This solution could be classified as

A unsaturated.

B saturated.

C supersaturated.

D none of the above.

14 Which of the following lists of steps involved in precipitation reactions is in the correct order?

A Add precipitation reagent, pour through filter paper, allow precipitate to settle.

B Pour through filter paper, add precipitation reagent, allow precipitate to settle.

C Allow precipitate to settle, pour through filter paper, add precipitation reagent.

D Add precipitation reagent, allow precipitate to settle, pour through filter paper.

15 The formula of hydrated copper(II) sulfate is $CuSO_4.5H_2O$. This tells us that in a sample of hydrated copper(II) sulfate there is

A 5 mol of $CuSO_4$ for every 1 mol of H_2O.

B 1 mol of $CuSO_4$ for every 5 mol of H_2O.

C 5 g of $CuSO_4$ for every 1 g of H_2O.

D 1 g of $CuSO_4$ for every 5 g of H_2O.

16 Aluminium reacts with oxygen to form aluminium oxide according to the following equation: $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$.

Which of the following statements is true for this reaction?

A 6 mol of Al will react with 6 mol of O_2 to form 4 mol of Al_2O_3.
B 3 mol of Al will react with 1 mol of O_2 to form 2 mol of Al_2O_3.
C 2 mol of Al will react with 1.5 mol of O_2 to form 1 mol of Al_2O_3.
D 4 mol of Al will react with 3 mol of O_2 to form 4 mol of Al_2O_3.

17 A sodium dichromate solution is orange in colour. Which of the following would be the appropriate colour filter to use to analyse a sodium dichromate solution in a colorimetry experiment?

A yellow
B orange
C red-purple
D greenish-blue

18 Which line in the table has gas pressures correctly converted across all three units?

	kPa	Pa	atm
A	50.00	5000	0.4935
B	1.00	1000	98.7
C	10 000	10	9.87
D	175	175 000	1.73

19 The enhanced greenhouse effect differs from the natural greenhouse effect in that

A it is driven by the release of greenhouse gases through human activities.
B it involves greenhouse gases that are better at absorbing infrared radiation.
C it results in an overall cooling of Earth's surface.
D it doesn't involve methane or water vapour.

20 Which of the following statements about 1.00 mol of hydrogen gas (H_2) and 1.00 mol of oxygen gas (O_2) at standard laboratory conditions (SLC) is correct?

A The mass and volume of the oxygen sample will be greater.
B The mass and volume of the two gas samples will the same.
C The mass of the oxygen sample will be greater, and the volumes are the same.
D The mass of the two samples is the same and the volume of oxygen is greater.

Short-answer questions

1 Calculate the pH of the following solutions.

a 0.35 M HCl(aq) (1 mark)
b 0.80 M KOH(aq) (2 marks)

2 Methanoic acid (H_3COOH) is a weak acid.

a Define the term 'weak acid'. (1 mark)
b Write the equation for the ionisation of methanoic acid in water to show it acting as a weak acid. (1 mark)
c A student measures 100 mL of 1.0 M methanoic acid into a beaker. In a second beaker, they measure 100 mL of 1.0 M hydrochloric acid. However, they forget to label the beakers and can't remember which solution is contained in each beaker. Explain how the following pieces of equipment could be used to distinguish between the two acids.

  i pH meter (2 marks)
  ii conductivity meter (2 marks)

3 An unknown ideal gas stored at a temperature of 4.0°C and a pressure of 2.47 atm occupies a volume of 750 mL.

a Perform the following unit conversions:

i Pressure from atm to kPa (1 mark)

ii Temperature from °C to K (1 mark)

iii Volume from mL to L (1 mark)

b Calculate the amount, in mol, of the unknown gas. (1 mark)

c It is thought that the sample in the container is either methane (CH_4) or carbon dioxide (CO_2), both of which are greenhouse gases. Given the mass of the unknown gas is 1.30 g, determine its likely identity. (2 marks)

d Explain how some gases, such as methane or carbon dioxide, cause the greenhouse effect. (2 marks)

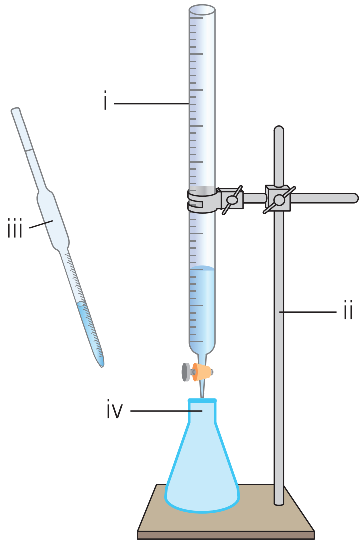
4 You are conducting a volumetric analysis to determine the concentration of hydrochloric acid in a brick cleaner. To do this, you will prepare a standard solution of sodium carbonate to titrate with a 1 in 50 dilution of the brick cleaner. The equation for this reaction is

$$HCl(aq) + Na_2CO_3(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

a Calculate the mass of Na_2CO_3 required to prepare 250.0 mL of a 0.100 M standard solution. (2 marks)

b State two properties of sodium carbonate that make it suitable for use as a primary standard. (2 marks)

c You set up your volumetric analysis experiment as shown in the diagram below. Name the pieces of equipment labelled **i–iv**. (2 marks)



d The titration was performed with the 0.100 M sodium carbonate standard in the burette and a 20.00 mL aliquot of the diluted brick cleaner solution in the conical flask with a few drops of a pH indicator. Titres of 19.45, 18.95, 19.05 and 19.00 mL were required to reach the end point.

i Calculate the average volume of the concordant titres. (1 mark)

ii Calculate the amount, in mol, of sodium carbonate in the average titre. (1 mark)

iii Calculate the amount, in mol, of hydrochloric acid in the conical flask. (1 mark)

iv Calculate the concentration of hydrochloric acid in the diluted brick cleaner. (1 mark)

v Calculate the concentration of hydrochloric acid in the original, undiluted, brick cleaner. (1 mark)

5 You have conducted a series of experiments to determine the reactivity series of a number of metals. Based on the results of your experiments, you are able to construct the following equations, which show whether a displacement reaction has occurred:

$Ca(s) + CuSO_4(aq) \rightarrow CaSO_4(aq) + Cu(s)$

$Al(s) + CaCO_3(aq)$ (no reaction)

$2Al(s) + 3CuSO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3Cu(s)$

a What is a displacement reaction? (1 mark)

b Using the information above, what is the reactivity series of the metals tested? (2 marks)

c You have been told that nickel is more reactive than aluminium. Write the equation for a reaction that would support this claim. (1 mark)

6 Redox reactions involve the transfer of electrons. They can be represented by the oxidation and reduction half-equations and these half-equations can be combined to give the overall equation for the reaction.

a Chromium is a metal that can exist in different oxidation states and many of the chromium-containing ions form brightly coloured solutions. A reaction occurs where the orange $Cr_2O_7^{2-}$ ion forms the green Cr^{3+} ion in acidic solution. Determine the balanced half-equation for this reaction and indicate whether it is an oxidation or a reduction half-reaction. (2 marks)

b Hydrogen sulfide, $H_2S(aq)$, reacts with silver ions, $Ag^+(aq)$, in acidic conditions to form sulfur, S(s), and silver, Ag(s).

i Write the two balanced half-equations for this reaction. (2 marks)

ii Hence, write the overall ionic equation for this redox reaction. (1 mark)

7 Water in a new housing development is pale blue in colour and suspected of being contaminated with Cu^{2+} ions leaching from the copper pipes. Analysis by UV–visible spectroscopy is conducted to determine the concentration of Cu^{2+} ions in the water.

a Explain why light in the blue region of the spectrum would not be suitable to choose for this analysis. (1 mark)

b A series of standard solutions of $CuSO_4(aq)$ are prepared and their absorbance measured. Use the data in the table below to prepare a calibration curve. (2 marks)

Cu^{2+} concentration (mg L^{-1})	Absorbance
0.00	0.000
50.0	0.251
100	0.505
150	0.754
200	1.097

c Use your calibration curve to determine the concentration, in mg L^{-1}, of Cu^{2+} ions in a sample of water from the housing development that has an absorbance of 0.321. (1 mark)

d A sample from another location in the housing development is found to have an absorbance so high that the calibration curve cannot be used to determine its concentration. A 1 in 4 dilution of this sample is prepared.

i A volume of 10.0 mL of the diluted sample is required. Determine the volumes of the sample and water that are needed to prepare this dilution. (1 mark)

ii The diluted sample is found to have an absorbance of 0.604. Determine the concentration, in mg L^{-1}, of Cu^{2+} ions in the undiluted sample. (2 marks)

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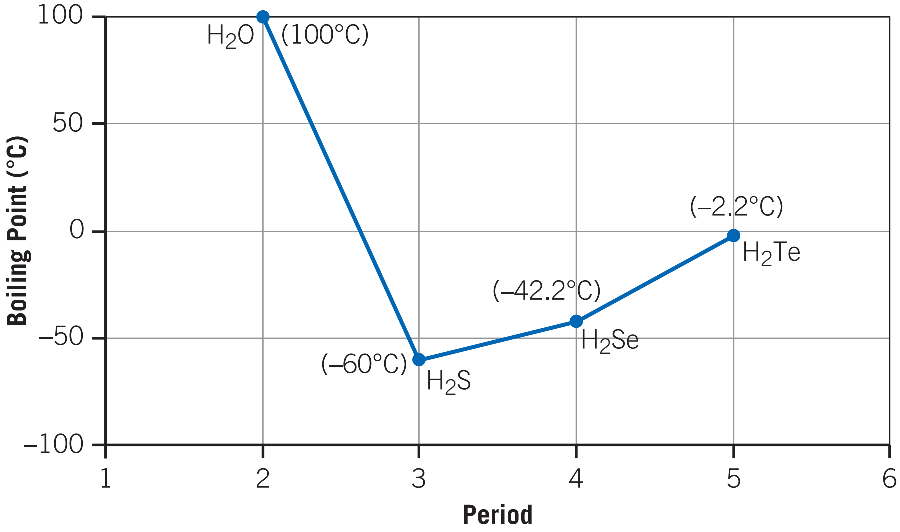
8 Below is a graph of the boiling points for Group 16 hydrides.



Use this information to explain the following:

a the difference in boiling points observed for H_2S and H_2Se (3 marks)

b the difference in boiling points observed for H_2O and H_2Te (3 marks)

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Glossary

Accuracy
how close the measurements are to the ‘true’ value of the quantity being measured

Acid
an acidic substance that is a proton, H^+, donor. It has a pH value less than 7 in aqueous solution at 25°C

Acid rain
precipitation with a pH less than 5

Addition reaction
a reaction where two or more reactants combine to form a larger product. Addition reactions require an unsaturated reactant

Adsorption
the process in which a compound is interacting with the stationary phase during a chromatography experiment

Aim
the main purpose of an investigation and what you hope to achieve

Alcohol
a family of organic compounds characterised by a hydroxyl functional group

Alkaline
a term for a water-soluble base. A substance that will gain an H^+ ion and has a pH greater than 7 in water at 25°C

Alkanol
an alkane in which a hydroxyl group (–OH) has replaced a hydrogen atom

Alkyl group
a group that is part of a larger molecule and is represented as an alkane minus a hydrogen atom, e.g. methyl ($-CH_3$)

Allotrope
a different structural form of the same chemical element

Alloy
a metallic substance composed of two or more elements

Amphiprotic
a substance that is able to donate or accept a proton, H^+ ion

Anion
an atom that has gained electron(s), filling its valence shell and resulting in an overall negative charge

Anode
the electrode at which oxidation occurs

Aragonite
a source of calcium carbonate

Atom
the smallest piece of an element that retains the properties of that element

Atomic number
the number of protons present in the nucleus of an atom

Atomic radius
(plural: radii) a measure of the size of an atom from the centre of its nucleus to its valence shell

Aufbau’s principle
describes the order of filling subshells with electrons from their lowest possible energy levels first

Authority
the level of expertise and relevance of an author in relation to the information they provide

Avogadro’s constant
the number of particles in a mole: $6.02 \times 10^{23}\ \text{mol}^{-1}$

Avogadro’s law
states that the volume of the gas is proportional to the number of moles of gas when temperature and pressure are constant

Base
an alkaline substance that is a proton, H^+, acceptor. It has a pH value between 8 and 14 in aqueous solution at 25°C

Beneficence
the commitment to maximising benefits and minimising risks and harms in a particular position or action

Bibliography
a list of referred to resources, listed in alphabetical order according to the author’s surname or organisation, appearing at the end of your report

Bioaccumulation
the build-up of heavy metals in organisms

Biodegradable
a substance or object capable of being decomposed by bacteria or other living organisms and thereby avoiding pollution

Biomagnification
the build-up (increasing concentration) of heavy metals as we move up the food chain

Biomass
renewable organic matter that comes from plants, animals and other living organisms

Bioplastic
plastic produced from renewable plant sources

Bohr model
also called the shell model. Shows the position of electrons in shells around the nucleus

Boiling point
the temperature at which the vapour pressure above the liquid equals the pressure exerted by the surroundings on the liquid; temperature at which liquid boils

Bond length
the optimal distance between the nuclei of two atoms within a molecule, allowing them to not repel each other but still share bonding electrons

Bond strength
the amount of energy required to break apart one mole of covalently bonded molecules in the gas state, measured in kJ mol^{-1}

Branched alkane
an alkane where the carbon atoms form a branched structure

Brittle
hard and rigid but easily broken. A result of the same charged ions in ionic compounds coming together, causing them to repel and the crystal lattice to shatter/break

Bromo-
a functional group characterised by a bromine atom bonded to a carbon

Calibrated
any instrument for quantitative measurement that has been checked, verified and rectified if required

Calibration curve
a graph created by plotting the absorbance of a series of standard solutions against their respective concentration

Capillary action
the movement of a substance due to adhesion (attractive forces between the solvent and stationary phase) and cohesion (attractive forces between molecules of the solvent)

Carbon dioxide
a gas, CO_2, that is an important part of the carbon cycle

Carbon neutral
a replenishing process where there is an overall balance between releasing and absorbing CO_2 from the atmosphere

Carboxyl group
a functional group consisting of –COOH

Carboxylic acid
a family of organic molecules characterised by the presence of a carboxyl group

Catalysis
a process where a substance known as a catalyst is used during a reaction, speeding up the reaction rate

Cathode
the electrode at which reduction occurs

Cation
an atom that has lost electron(s) from its valence shell, resulting in an overall positive charge

Chemical property
the behaviour of an element or substance when it reacts with another element or substance

Chloro-
a functional group characterised by a chlorine atom bonded to a carbon

Chlorofluorocarbon (CFC)
a compound that contains the elements chlorine, fluorine and carbon; long-lived in the atmosphere and damaging to Earth's ozone layer

Circular economy
an economy that functions on a continuous cycle with a focus on optimising the re-use of resources and materials

CO_2 equivalent
a measure used to compare the global warming potential of different greenhouse gases by converting them to the equivalent amount of carbon dioxide (CO_2)

Colorimetry
a type of spectroscopy used to determine the concentration of coloured compounds found in solutions. Light is selected using a coloured filter

Combustion reaction
a reaction involving oxygen that produces carbon dioxide (or carbon monoxide or carbon), water and heat

Complex half-equation
either the reduction or oxidation part of a redox reaction involving more than one element

Compound
a substance formed from two or more different types of atoms in a fixed ratio

Concentrated
a solution with a large number of molecules of the solute dissolved in a given volume

Conclusion
a summary of what you can deduce from the results of the investigation, including whether the tested hypothesis was supported

Concordant
in titration, the volume of two or more titres that are within 0.1 mL of each other

Conductivity meter
an instrument that measures the amount of electrical conductivity a solution has; used to indicate levels of dissolved ions

Conductor
a material that allows the flow of charge (electrical current)

Conjugate acid–base pair
represents the reactant (acid)/product (base) pairs that differ by the loss or gain of a hydrogen ion, H^+, e.g. HNO_3/NO_3^-

Continuous data
data that is measurable and continuous, with infinite possible values; best represented by a line graph

Control group
The set-up or group in an experiment that does not receive treatment; it is used as the ‘standard of comparison’

Controlled variable
Anything kept constant, or monitored, so it does not affect the independent and dependent variables, and therefore the validity of experimental results

Core charge
the electrostatic force of attraction between valence electrons and the nucleus. It can be calculated by subtracting the number of non-valence electrons from the number of protons

Corrosion
an irreversible reaction at the surface of a material with its environment, which results in deterioration of the material

Covalent bond
a shared pair of electrons between two atoms within a molecule; *co* means ‘together’; *valent* means ‘valence electrons’

CRAP detection (or analysis)
a method of determining whether a source is reliable, based on the currency, reliability, authority and purpose of the information

Critical element
a chemical element essential to the economic or national security of a nation and vulnerable to disruption of supply

Crude oil
a naturally occurring mixture of liquid hydrocarbons that can be used as a fuel source

Cryogenics
the production and effects of elements at very low temperatures

Currency
timeliness; how up to date information is

Cuvette
a small rectangular tube with two sides that are transparent to UV and visible light

Degradation
the process of a substance breaking down into a less complex state

Delocalised electron
an electron that is not associated with any particular atom and is free to move

Dependent variable
the variable that changes in response to changes in the independent variable; the experimenter measures these changes

Desorption
the process in which a compound interacts with the mobile phase during a chromatography experiment

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Diatomic
an element that consists of two atoms; the prefix *di* means ‘two’

Dilute
a solution with a small number of molecules of the solute dissolved in a given volume

Dilution
the process of adding a solvent to a solution to reduce its concentration

Dimer
a complex where two molecules are linked together

Dipole
a bond or molecule where the ends have opposite partial charges, one positive and one negative

Dipole–dipole attraction
an intermolecular force present between polar molecules that have permanent dipoles

Diprotic acid
a substance that can act as a Brønsted–Lowry acid twice, as it has two protons, H^+ ions, to donate to a Brønsted–Lowry base

Discrete data
data that is countable and in discrete categories; contains distinct or separate values; best represented by a bar graph

Dispersion force
a weak intermolecular force present in all molecules created between instantaneous dipoles. Also known as London dispersion forces or van der Waals forces

Displacement reaction
when a more reactive element replaces another. In a metal displacement reaction, the ions of the less reactive element are displaced from the solution and form a solid

Dissolution
the process in which molecules or ions become separated and mobile when surrounded by solvent molecules. For ionic compounds, they are split into ions that then form an electrostatic force of attraction with the solvent

Double displacement reaction
a chemical reaction in which ionic compounds swap their cations and anions

Ductility
the ability of a material to be drawn out into a long wire without breaking

Electrode
a connecting terminal that allows current to enter or leave

Electrolysis
the process of forcing a redox reaction to occur by passing an electric current through an electrolyte

Electrolyte
a non-reactive solution or paste that enables the flow of charged particles and helps to facilitate the redox reactions

Electromagnetic radiation
a form of energy that is all around us and takes many forms, such as radio waves, microwaves, X-rays and gamma rays. Visible light is a form of electromagnetic energy

Electromagnetic spectrum
the spectrum of electromagnetic radiation, measured in wavelength, from long radio waves to short gamma rays

Electron
a negatively charged subatomic particle that moves around the nucleus of an atom

Electron domain
the negatively charged area around an atom

Electron transfer
the exchange of electrons that occurs during a redox reaction. Electrons are transferred from the reducing agent to the oxidising agent

Electronegativity
the ability of an atom to attract electrons towards itself. This is affected by an atom’s core charge and the distance electrons are from the nucleus (atomic radius)

Electrostatic force of attraction
the attraction between positive and negatively charged particles that are not directly in contact

Element
a pure substance made of only one type of atom

Emission spectrum
(plural: spectra) the electromagnetic radiation emitted by a source, such as atoms

Empirical formula
the simplest whole-number ratio of elements within a compound

End point
the point in titration when the indicator changes colour

Enhanced greenhouse effect
the change of the normal climate equilibrium on Earth, primarily an increase in temperatures globally. Caused by increased emissions of greenhouse gases; also known as global warming

Equivalence point
the point in a titration when the number of moles of the acid and the base are present in the stoichiometric ratio that achieves neutralisation

Ethics
moral principles that guide our beliefs about what is right or wrong conduct

Eutrophication
a dense growth of plant life as a result of high concentrations of nutrients in water, commonly due to runoff from land

Excited state
when electron(s) in an atom have gained energy and are no longer in their lowest possible energy state closest to the nucleus

Experimental validity
the extent to which all variables in the experiment have been controlled, so that the independent variable is the only factor that changes

Feedstock
raw material that is used to supply an industrial process

Filtration
the process by which precipitate is removed from a solution by passing it through a medium that only allows fluid to pass through, such as filter paper

First ionisation energy
the energy required to remove the first, or outermost, electron from a gaseous atom

Fluoro-
a functional group characterised by a fluorine atom bonded to a carbon

Fossil fuel
a source of fuel made from decomposing plants and animals over hundreds of millions of years. It consists mainly of carbon and hydrogen atoms

Fossil fuel-based plastic
plastic in which the monomers are derived from either petroleum or natural gas

Fractional distillation
separation of a mixture of liquids into separate fractions based on their boiling points

Full balanced equation
a balanced equation that shows all atoms involved in a reaction, including spectator ions

Functional group
an atom or a group of atoms in an organic molecule that largely determines the molecule's properties and reactions

Gas pressure
the amount of force exerted by a gas on its container per unit of area

Gas volume
the volume of the container occupied by a gas

Gravimetric analysis
determining the mass of an unknown ion by forming a known mass of precipitate

Green chemistry
the design of new chemical products and manufacturing processes that are safer and more sustainable than traditionally used products and processes

Greenhouse effect
a natural process in which heat from Earth's surface is re-radiated by gases in the atmosphere, which maintains Earth at a temperature suitable for life

Greenhouse gas
a gas that contributes to the greenhouse effect by absorbing infrared radiation (heat) close to the Earth's surface – e.g. carbon dioxide

Ground state
when all electrons in an atom are in their lowest possible energy state closest to the nucleus

Half-equation
either the reduction or oxidation part of a redox reaction

Haloalkane
an alkane in which one or more halogen atoms has replaced a hydrogen atom or atoms

Hardness
a measure of the resistance of a material to abrasion or surface indentation

Heavy metal
a metal with a relatively high density and a toxic effect on living organisms, such as cadmium, chromium, copper, lead and mercury

Homologous series
a sequence of compounds that have identical functional groups and similar chemical properties but differ by a CH_2 group in their structure

Hund's rule
describes the situation where every orbital in a subshell contains only one electron before any orbital is fully occupied by two electrons

Hydrocarbon
a compound made up of carbon and hydrogen atoms only

Hydrogen bonding
a stronger type of dipole–dipole attraction. Occurs between molecules that contain either an N–H, O–H or F–H covalent bond

Hydroxide
a chemical compound containing a hydroxide group (an oxygen and hydrogen covalently bonded together) combined with another element

Hydroxyl group
a functional group consisting of –OH

Hygroscopicity
the ability to attract and hold water from the environment

Hypothesis
an idea or explanation for something that is based on known facts but has not yet been proven

Ideal gas
a gas that obeys the gas laws at all temperatures and pressures

Ideal gas equation
an equation that relates pressure, volume, amount of gas and absolute temperature for an ideal gas

Independent variable
the variable for which quantities are changed by the experimenter

Indicator
a substance that changes colour in the presence of an acid or a base

Infrared radiation (IR)
the region of the electromagnetic spectrum that lies between the wavelengths 700 nm and 1 mm

Insoluble
when a solute has limited solubility in a solvent

Instantaneous dipole
when a dipole or opposite charges are created randomly and very quickly within a molecule

Integrity
the commitment to searching for knowledge and understanding, and the honest reporting of sources and results

Intermolecular force
an attraction or force between separate, or discrete, molecules

Introduction
a detailed but succinct explanation of the reason for undertaking an investigation; includes key chemical concepts, aim and hypothesis

Iodo-
a functional group characterised by an iodine atom bonded to a carbon

Ion
a positively or negatively charged atom that has either lost or gained electron(s)

Ion–dipole attraction
the bond formed between an ion and the oppositely charged end of a polar dipole molecule, such as water

Ionic bond
the electrostatic force of attraction between cations and anions in a lattice

Ionic compound
a compound or molecule formed when positive cations and negative anions chemically combine

Ionic equation
a balanced equation that shows only the atoms and ions that participate in a reaction

Ionic product of water (K_w)
the equilibrium constant K_w, where $K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$ M^2 at 25°C. It represents the ionisation reaction of water with itself. K_w is the product of the concentration of ions in solution, so it is referred to as the ionic product of water

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Ionisation
when an atom or a compound becomes charged via the gain or loss of electrons

Ionisation energy
the energy required to remove an electron from a gaseous atom or ion

Isotope
an atom of the same element with the same number of protons (same atomic number) but a different number of neutrons (different mass number)

Justice
ensuring fair consideration of competing claims, with no unfair burden on any particular group, and fair distribution of benefits

Kinetic energy
the energy an object has due to its motion

KOHES
acronym for remembering how to write and balance complex half-equations

Latent heat of vaporisation
the amount of energy required to convert one mole of liquid water into one mole of gaseous water (water vapour)

Lattice
an organised arrangement of atoms or ions into fixed positions in a 3D structure

Law
a description of an observed phenomenon

Lewis structure
a representation of the electron arrangement in a molecule in which the valence shell electrons are represented by dots or crosses

Line of best fit
a line on a graph that shows the general trend of the data points; the distance to the points above the line should equal the distance to the points below the line

Linear economy
an economy that functions on a model whereby products are made and then simply discarded after use

Lustre
the appearance of a surface in reflected light. Lustrous metals appear shiny

Malleability
the ability of a material to be distorted by an applied force without breaking

Mass number
the total number of protons and neutrons present in the nucleus of an atom

Melting point
the temperature at which a solid melts

Metallic character
the ability of a metal to lose its valence electron and display common metallic properties

Method
a series of numbered steps describing the procedure

Mobile phase
the liquid or gas that flows during a chromatography experiment, which transports the compounds to be separated at different rates over the stationary phase

Model
a representation of an idea, object or process that can be used to explain phenomena that cannot be directly experienced

Molar mass
the mass of one mole of a substance measured in g mol^{-1}; symbol M

Molar volume
the volume of one mole of a substance at a specified pressure and temperature; symbol V_m

Mole
one mole is exactly $6.02214076 \times 10^{23}$ particles (6.02×10^{23} to three significant figures)

Molecular formula
the chemical formula that gives the total number of atoms of each element within a compound

Molecular weight
the sum of the atomic masses of the atoms as given in the molecular formula

Monoatomic
an element that consists of only one atom; the prefix *mono* means ‘one’

Monomer
a small molecule that reacts with other monomer molecules to form long chains (polymers)

Monoprotic acid
a substance that can act as a Brønsted–Lowry acid once, as it has one proton, H^+ ion, to donate to a Brønsted–Lowry base

Natural indicator
a substance that occurs in nature and changes to different colours when in contact with acids or bases

Negative control
a control group that isn't expected to produce a result

Neutralisation reaction
a type of reaction involving an acid reacting with a base. The products formed are neutral water and salt

Neutron
an uncharged subatomic particle present within the nucleus of an atom

Non-maleficence
avoiding harm; in scientific research the harm resulting from a position or course of action should not be disproportionate to the benefits

Non-polar
a molecule with no permanent net dipole

Non-renewable resource
a natural resource that cannot be readily replaced by natural means at a pace quick enough to keep up with consumption

Octet rule
a rule to explain that atoms bond to other atoms in a way that results in each atom containing eight electrons in its valence shell

OILRIG
a mnemonic to recall redox – **O**xidation **I**s **L**oss (of electrons), **R**eduction **I**s **G**ain (of electrons)

Opinion
a view or judgement that may be based on personal feelings, not necessarily on facts or research

Orbital
a three-dimensional description of the most likely location for an electron within a region of an atom. Each can hold zero, one or two electrons

Ore
a deposit within Earth's crust that contains one or more valuable minerals, typically with a high concentration of a certain element (particularly metals)

Organometallic compound
a compound with at least one carbon–metal bond

Outlier
a reading that is very different from other results obtained for the same measurement

Oxidant
a reagent that causes an oxidation reaction to occur

Oxidation
a loss of electrons

Oxidation reaction
a reaction that occurs when there is an increase in the number of C–O bonds or a decrease in the number of C–H bonds

Oxidation state
a measure of the degree of oxidation of an atom in a substance. It is defined as the charge an atom might have when electrons are counted to help identify redox reactions, oxidising agents and reducing agents

Oxide
a chemical compound with one or more oxygen atoms combined with another element

Oxidising agent
the reactant in a redox reaction that causes oxidation. It is itself reduced and therefore will gain electrons – its oxidation state will decrease

Parent hydrocarbon
the longest continuous carbon chain in a molecule

Percentage abundance
the relative proportions of the different isotopes for an element. Also called relative abundance

Percentage mass
the mass of a component represented as a percentage of the total mass of the sample

Peroxide
a class of chemical compounds in which two oxygen atoms are linked together by a single covalent bond

pH
the quantitative value of how acidic or basic a substance is

Photochemical smog
a brown haze generated from the reaction between ultraviolet light from the sun and nitrogen oxides in the atmosphere

Photosynthesis
a chemical reaction that plants utilise to create glucose from carbon dioxide

Physical property
the features of an element or substance that can be measured without altering the chemical composition of that substance

pOH
the quantitative value of how acidic or basic a substance is on a scale of 1–14. It is a measure of the concentration of hydroxide ions, $[OH^-]$

Polar
a molecule with two distinct poles of partial electric charge, a partial positive charge on one side of the molecule and a partial negative charge on the other side

Polarity
the distribution of electric charge between two atoms joined by a bond

Polyatomic ion
an ion that consists of more than one atom covalently bonded together to form a charged compound; *poly* means 'many'

Polymer
a long-chain molecule formed by large numbers of repeating units (monomers)

Polymerisation
a reaction that covalently links repeated units of smaller molecules to make a larger molecule

Polyprotic
a substance that can act as a Brønsted–Lowry acid more than once, as it has several protons, H^+ ions, to donate. Diprotic and triprotic acids are examples

Positional isomer
a molecule with the same molecular formula and parent hydrocarbon but different positions of functional groups

Positive control
a control group that receives a treatment with a known response that can then be compared to the experimental group(s)

Precipitate
the solid formed in a reaction in which two or more solutions of soluble salts are mixed

Precipitation reaction
a chemical reaction in which two or more solutions of soluble salts are mixed and one or more of the products formed is an insoluble solid – known as a precipitate

Precision
how close all the measurements are to each other

Primary cell
a device that converts chemical energy into electrical energy while discharging but cannot be recharged

Proton
a positively charged subatomic particle present within the nucleus of an atom

Qualitative data
data that is descriptive (not numeric)

Quantitative
relating to measured values

Quantitative analysis
any method of analysis that measures the numerical amount of a given substance, including gravimetric analysis and colorimetry

Quantitative data
data that is measured and represented numerically

Random error
an unpredictable variation in the readings obtained due to variables not all being controlled (extraneous variables), resulting in the readings being higher or lower than expected

Reactivity
how likely a chemical is to take part in a chemical reaction

Redox couple
a reducing agent, which donates electrons, and its corresponding oxidising agent, which accepts electrons

Redox reaction
a reaction involving a transfer of electrons. Both reduction and oxidation occur simultaneously

Reducing agent
the reactant in a redox reaction that causes reduction. It is itself oxidised and therefore will lose electrons – its oxidation state will increase

Reduction
a gain of electrons

Refining
the elimination of impurities from metals to obtain highly pure versions of the metal

Relative atomic mass
the average of the relative masses of all atoms for an element based on the isotopes for an element and their natural percentage abundances. Also known as relative atomic weight or standard atomic weight

Relative isotopic mass
the mass of an atom of an isotope relative to the mass of a carbon-12 atom, which is set at 12 exactly

Reliability
if research findings are repeated or supported by evidence more than once

Reliability (experimental)
the extent to which an experiment always yields the same results under the same conditions

Renewable
a resource that can be replenished at a higher rate than it is consumed

Repeatability
recording of results produced when the experiment is repeated in one lab by one operator under the same conditions

Reproducibility
when the same results are obtained for the same experiment by different operators using different equipment

Resolution
the smallest change in the quantity being measured that results in a noticeable change in the value shown on the piece of measuring equipment

Respect
giving consideration to the extent to which living things have an intrinsic value; giving due regard to the welfare, liberty, autonomy, beliefs, perceptions, customs and cultural heritage of individual and collective

Retention factor (R_f)
the distance between the origin and the spot of interest relative to the distance between the origin and the solvent front

Rubric
a scoring guide; a set of criteria used to evaluate students’ responses

Rust
the product of corrosion, which involves iron reacting with oxygen and moisture in the air, converting the iron into iron oxides

Rusting
an example of corrosion, where the process specifically involves iron reacting with oxygen and moisture in the air, converting the iron into iron oxides and iron hydroxides

Saline
water that contains dissolved salt

Saturated
an organic molecule where all C–C bonds are single bonds

Saturated solution
a solution that cannot dissolve any more solute in the volume of solvent at a given temperature

Semi-structural formula
a representation of organic molecules that shows the order in which atoms or groups of atoms are connected

Simple half-equation
either the reduction or oxidation part of a redox reaction involving only one element

Single variable exploration
an investigation that contains only one independent and one dependent variable

Slag
a silicon dioxide and metal oxide mixture that is left over as a by-product from smelting to extract metal from its ore

Smelting
the process of heating and melting an ore to extract a metal

Social norms
standards of what is considered socially acceptable within a community

Solubility
the ability of a solute to dissolve in a solvent

Solubility curve
a graph of solubility versus temperature for a certain solute dissolved in a certain solvent

Solubility table
a reference table used to predict the solubility of ionic compounds

Soluble
when a solute is able to dissolve in a solvent

Solute
the solid, liquid or gas that dissolves in a solvent

Solution
a mixture of a solute dissolved in a solvent

Solvent
the substance that the solute dissolves in

Solvent front
the distance the solvent has travelled up the TLC plate/stationary phase

Species
any reactant or product in a chemical reaction

Specific heat capacity
the amount of energy required to raise the temperature of one gram of a substance by one degree Celsius

Spectator ion
an ion that does not take part in the precipitation of the solid but remains soluble in solution

Spectroscopy
a group of techniques that study how matter absorbs and emits light and other forms of radiation

Stakeholder
the person or organisation that will be affected by the factor under consideration

Standard laboratory conditions (SLC)
the standard ambient conditions within a laboratory, found to be 25°C and 100 kPa

Standard solution
a solution with a known concentration

Stationary phase
a solid, liquid or gel that remains static during a chromatography experiment

Stoichiometry
the quantitative relationship between two or more substances during a reaction based on the ratio in which they react

Straight-chain alkane
an alkane consisting of a continuous chain of carbon atoms

Strong acid
an acid that will ionise completely in solution – i.e. every acid molecule will donate its H^+ ion to form a conjugate base

Strong base
a base where every base molecule accepts an H^+ ion from an acid

Structural formula
a representation of how the individual atoms of the compound may be arranged in three-dimensional space

Structural isomer
a molecule with the same molecular formula but different connectivity of atoms

Subatomic particle
a particle that is present within an atom; includes protons, neutrons and electrons

Subshell
a region within a shell composed of certain numbers of orbitals. Subshells are labelled as s, p, d and f

Substitution reaction
a reaction where an atom or a group of atoms on a molecule is replaced with another atom or group of atoms

Supersaturated solution
a solution that contains more solute than a saturated solution in a certain volume at a given temperature

Sustainable
a process or practice that meets the needs of the current generation without compromising the needs of future generations

Sustainable development
an approach that meets the needs of the present without compromising the ability of future generations to meet their own needs

SWOT analysis
a method of assessing an issue by considering strengths, weaknesses, opportunities and threats associated with the issue

Systematic error
when the readings obtained from measurements differ from the 'true' value consistently in one direction every time

Systematic nomenclature
the name given to an organic molecule that follows a set of systematic rules based on the structure of the molecule

Tetrahedral structure
a structure where the central atom is bonded to four other atoms that are located at the corners of a tetrahedron

Theory
an explanation based on facts that have been confirmed through both experimentation and observation multiple times

Thin-layer chromatography
an analytical technique for separating and identifying mixtures in which a silica or polar aluminium oxide gel acts as the stationary phase

Title
the research question under investigation; includes information about what is being tested

Titration
a type of volumetric analysis whereby a measured volume of a standard solution is added to a measured volume of a solution whose concentration is unknown

Titre volume
the volume of solution required to reach the end point in a titration

Total dissolved solids (TDS)
salinity is often measured using TDS. The total dissolved solids in a water sample is measured in $mg\,L^{-1}$

Triprotic acid
a substance that can act as a Brønsted–Lowry acid three times, as it has three protons, H^+ ions, to donate to a Brønsted–Lowry base

True value
the value or range of values that would be obtained if the quantity could be measured perfectly

Ultraviolet–visible (UV–visible) spectrophotometry
a type of spectroscopy used to determine the concentration of coloured compounds found in solutions. Light is selected using a monochromator

Unsaturated
a molecule that has one or more $C{=}C$ double bonds or $C{\equiv}C$ triple bonds

Unsaturated solution
a solution that contains less solute than a saturated solution in the volume of solvent at a given temperature

Valence electron
an electron found in the outermost shell of an atom

Valence shell
the outermost shell of an atom

Valence shell electron pair repulsion (VSEPR)
a model used to predict the geometric shape of molecules based on the number of valence shell electron pairs surrounding the central atoms in molecules

Validity
the extent to which evidence supporting an argument is true or accurate

Voltmeter
an instrument used for measuring electrical potential

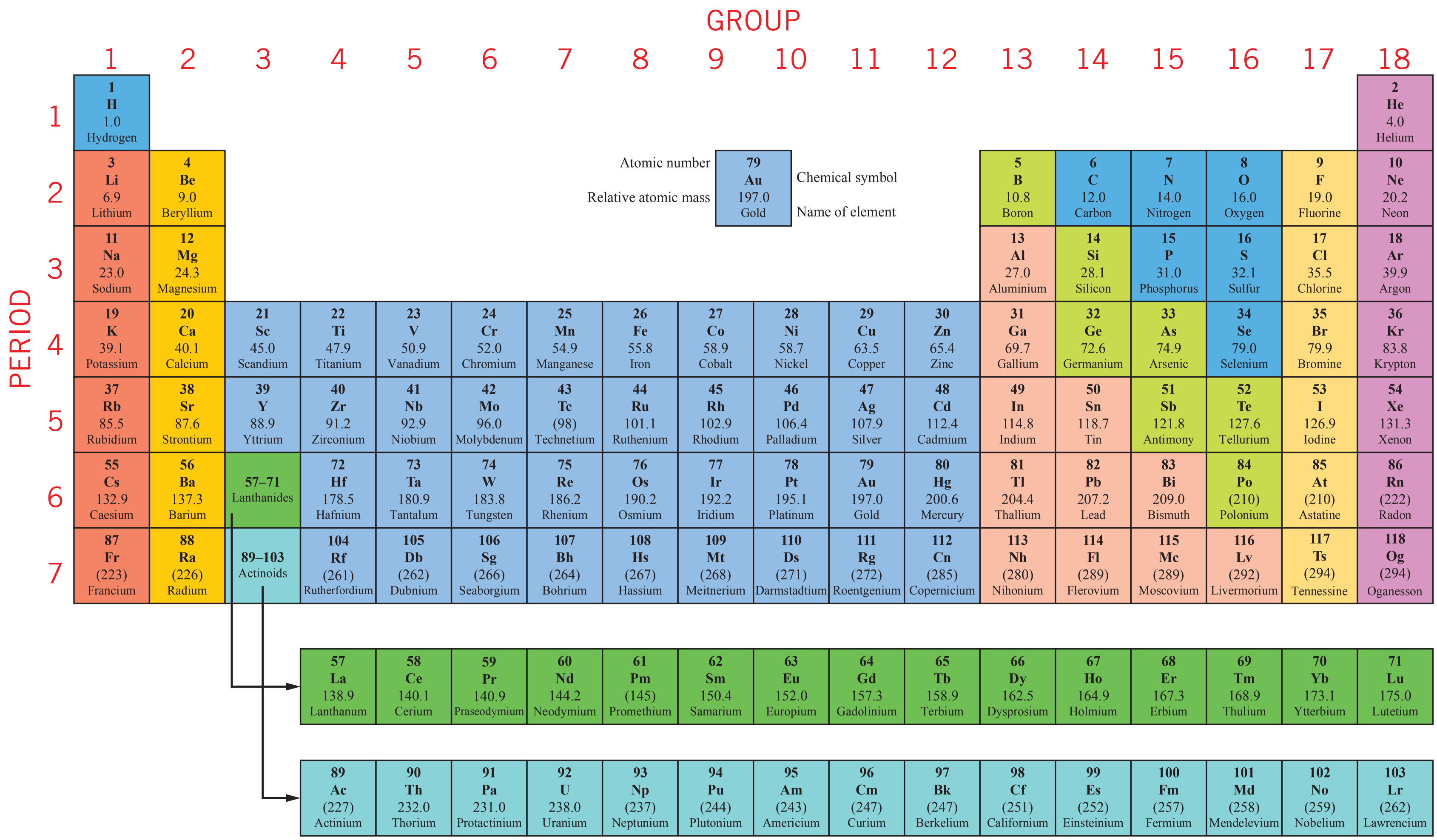
Volumetric analysis
analysis of solutions by using measurements of volumes

Weak acid
an acid that will only partially ionise in solution – i.e. not every acid molecule will donate an H^+ ion to form a conjugate base, resulting in a solution that contains both undissociated acid molecules and conjugate base molecules

Weak base
a base where not every base molecule accepts an H^+ ion from an acid

Weighted average
a calculation of the average that takes into account the frequency of a number within the data set

Appendix: Periodic table



Atomic number — 79 — Chemical symbol — Au
Relative atomic mass — 197.0 — Name of element — Gold

GROUP

PERIOD	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1.0 Hydrogen																	2 He 4.0 Helium
2	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
3	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
4	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
5	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
6	55 Cs 132.9 Caesium	56 Ba 137.3 Barium	57–71 Lanthanides	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
7	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

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