CHEMISTRY

VCE UNITS 1 AND 2

1

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Contents

UNIT 1

AREA OF STUDY 1

CHAPTER 1

Atomic structure 2

CHAPTER 2

The periodic table 25

CHAPTER 3

Ionic bonding 42

CHAPTER 4

Metallic bonding 65

CHAPTER 5

AREA OF STUDY 2

CHAPTER 6

CHAPTER 7

CHAPTER 8

CHAPTER 9

AREA OF STUDY 3

CHAPTER 10

Investigating and reporting 194

UNIT 2

AREA OF STUDY 1

CHAPTER 11

Connecting chemical reactions and equations 206

CHAPTER 12

Properties of water 219

CHAPTER 13

Reactions in water 237

AREA OF STUDY 2

CHAPTER 14

CHAPTER 15

CHAPTER 16

CHAPTER 17

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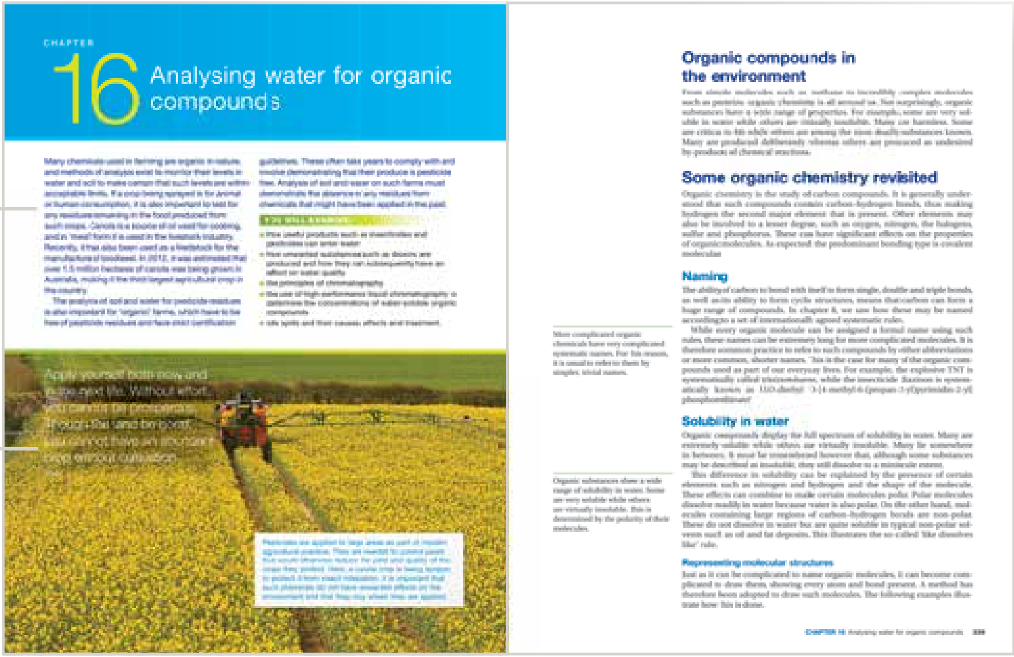

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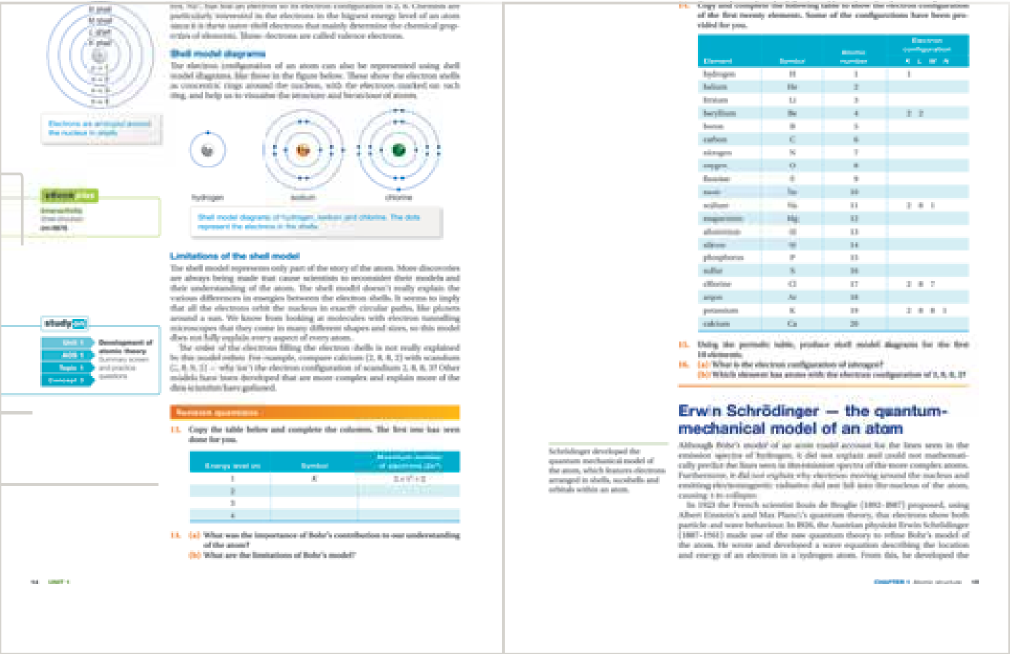

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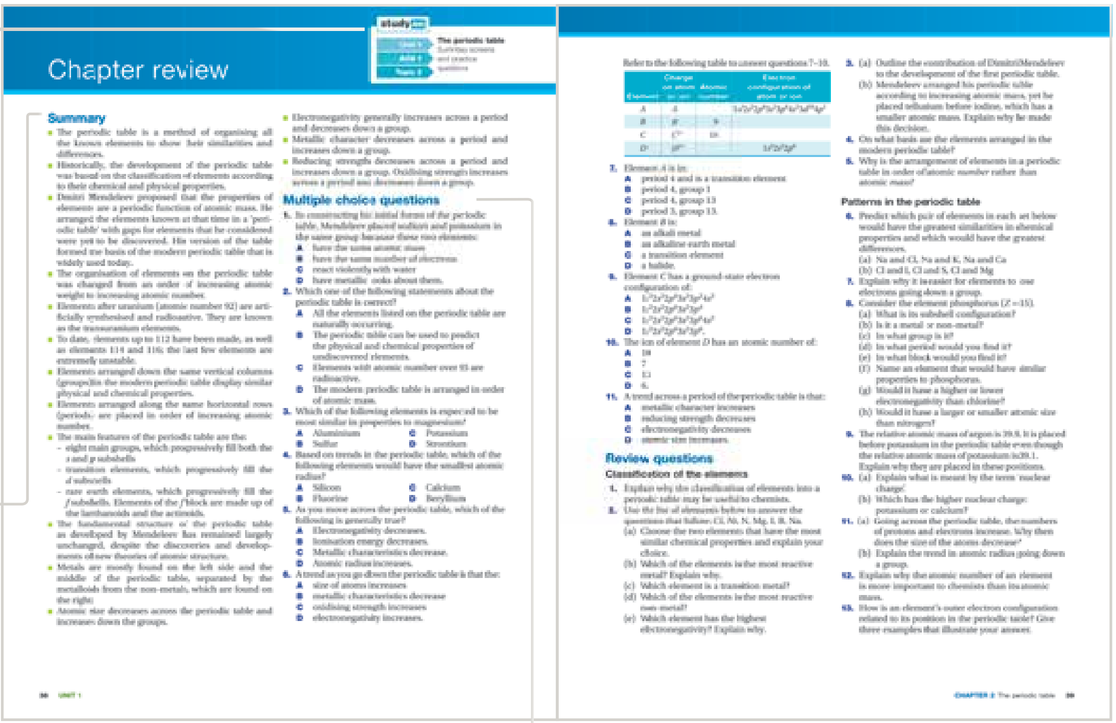

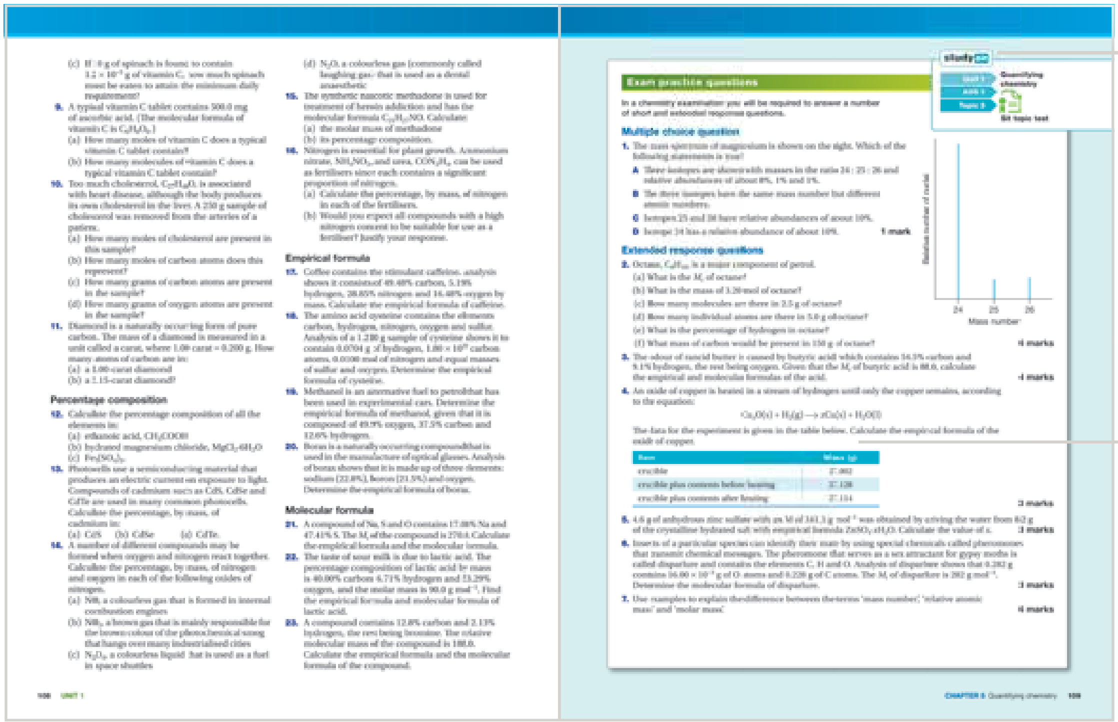

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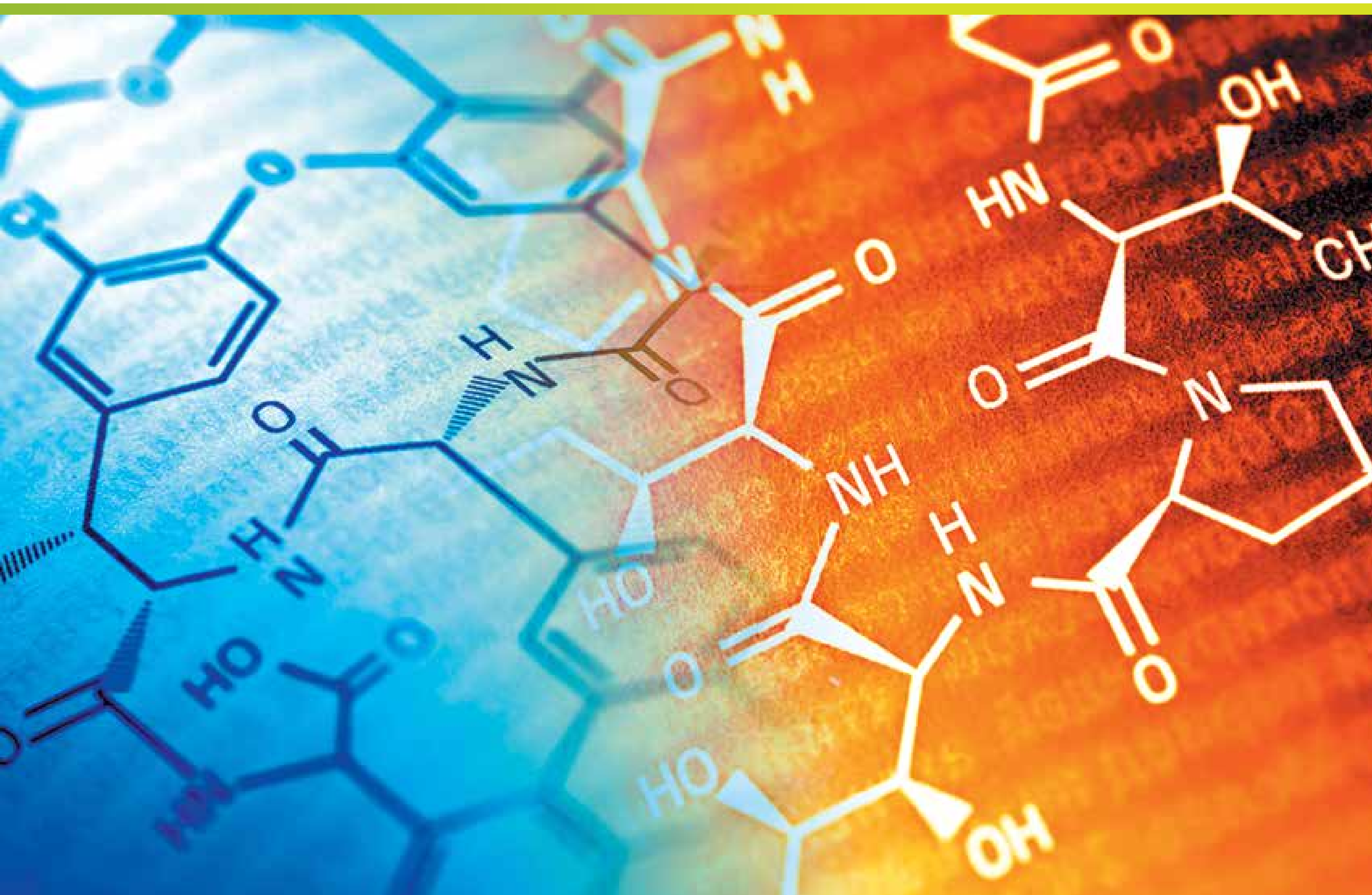
UNIT 1

AREA OF STUDY 1

AREA OF STUDY 2

AREA OF STUDY 3



CHAPTER

1

Atomic structure

YOU WILL EXAMINE:

- the historical development of atomic theory
- the structure of atoms
- the comparative sizes of particles
- nanotechnology
- isotopes and isotopic symbols
- arrangement of electrons in atoms
- limitations of the shell model
- the quantum mechanical model.

> Nothing exists except atoms and empty space; everything else is opinion.
>
> *Democritus*

Small amounts of compounds containing elements with characteristic flame colours are added to fireworks. When the fireworks are ignited, beautiful colours are produced. Strontium produces red light, copper produces green light and potassium produces lilac light.

Atoms

Life is a mystery. Where did we come from? What are we made of? Scientists tell us that we are made of very small particles called atoms and that these atoms have their origin in stars. What a remarkable journey these atoms must have undertaken while being recycled over the millions of years since the origin of the universe. Matter is made up of atoms. Every material thing that you can see, smell and touch, that occupies space and has mass, is a form of matter. Studying the structure and behaviour of matter — of which life, the Earth and the universe are composed — has been an ongoing human preoccupation.

Early Greek philosophers and alchemists, working from their basic observations of matter, believed that all materials were made up of different proportions of four basic elements: earth, air, fire and water. We have come a long way since then. Basing our approach to the study of matter on experimental measurement, we have been able to modify such old theories and develop new ones that more clearly explain the structure of matter. This improved understanding of matter has enabled us to further investigate the different properties of materials that make them useful to our society.

The structure of atoms

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Rutherford's gold foil experiment
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Atomic theory attempts to explain the structure of materials. According to this theory, all matter is made of **atoms**.

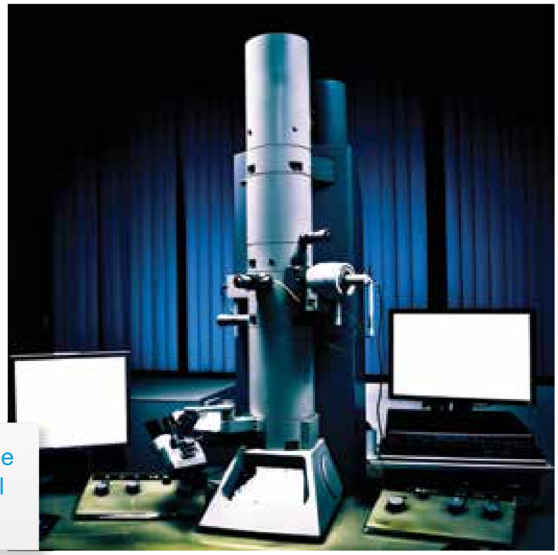
Take, for example, a strip of tin foil. Imagine cutting the foil into ever smaller pieces. Imagine that you could continue doing this, even after the pieces were too small to be seen under a powerful microscope. In theory, you could continue until you had a piece of tin the size of just one atom! This atom would be the smallest particle of foil that could exist and still be tin. If the tin atom were further divided, it would no longer be a tin atom.

Atoms are so small that it was not until 1981 that their images could finally be seen using the newly invented scanning tunnelling microscope. Ideas about the atom had been refined over many centuries. These new images were very exciting to the scientific community at the time they were developed.

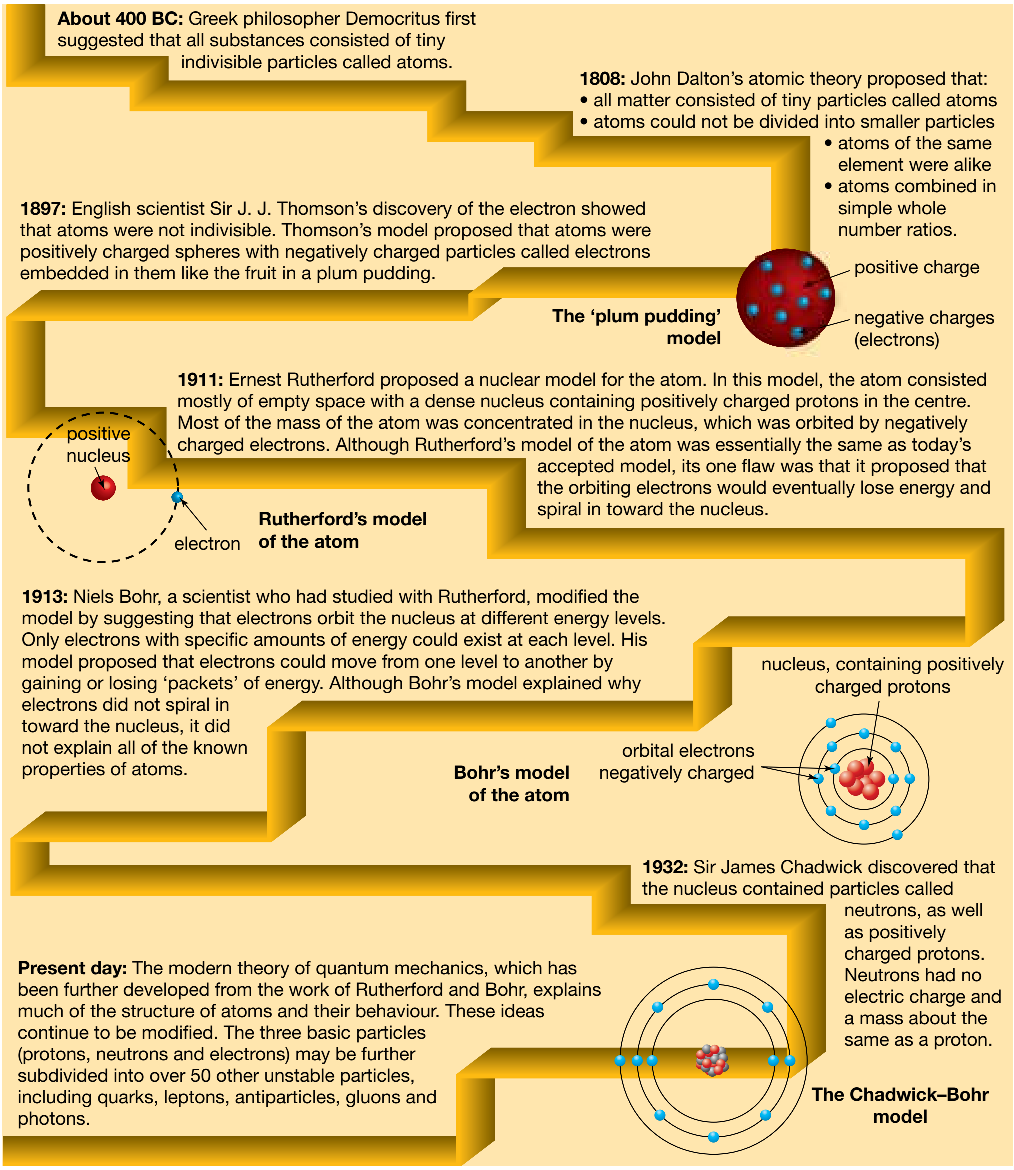
Rutherford's nuclear model of the atom contained a nucleus with protons and neutrons and orbiting electrons.

The most powerful microscopes cannot yet give detailed pictures of the internal structure of atoms, so chemists use models to represent them. One very useful model is the nuclear model of the atom proposed by Rutherford in 1911. This describes an atom as being mostly empty space, with a dense central structure called a **nucleus**. The nucleus, though its volume is very small relative to the atom as a whole, contains most of the mass of the atom. The nucleus is made up of positively charged particles called **protons**. The simplest nucleus is that of the hydrogen atom, which contains just one proton. All other atoms have nuclei that also contain **neutrons**. A neutron has no charge but has virtually the same mass as a proton.

The empty space around the nucleus contains negatively charged particles called **electrons**. Electrons take up most of the space occupied by the atoms and move very rapidly around the nucleus in orbits. Each electron has a definite energy and moves in a specific energy level. The mass of an electron is very much less than that of a proton or a neutron.



An electron microscope used to view individual atoms



The development of the model of the atom

About 400 BC: Greek philosopher Democritus first suggested that all substances consisted of tiny indivisible particles called atoms.

1808: John Dalton's atomic theory proposed that:
- all matter consisted of tiny particles called atoms
- atoms could not be divided into smaller particles
- atoms of the same element were alike
- atoms combined in simple whole number ratios.

1897: English scientist Sir J. J. Thomson's discovery of the electron showed that atoms were not indivisible. Thomson's model proposed that atoms were positively charged spheres with negatively charged particles called electrons embedded in them like the fruit in a plum pudding.

1911: Ernest Rutherford proposed a nuclear model for the atom. In this model, the atom consisted mostly of empty space with a dense nucleus containing positively charged protons in the centre. Most of the mass of the atom was concentrated in the nucleus, which was orbited by negatively charged electrons. Although Rutherford's model of the atom was essentially the same as today's accepted model, its one flaw was that it proposed that the orbiting electrons would eventually lose energy and spiral in toward the nucleus.

1913: Niels Bohr, a scientist who had studied with Rutherford, modified the model by suggesting that electrons orbit the nucleus at different energy levels. Only electrons with specific amounts of energy could exist at each level. His model proposed that electrons could move from one level to another by gaining or losing 'packets' of energy. Although Bohr's model explained why electrons did not spiral in toward the nucleus, it did not explain all of the known properties of atoms.

1932: Sir James Chadwick discovered that the nucleus contained particles called neutrons, as well as positively charged protons. Neutrons had no electric charge and a mass about the same as a proton.

Present day: The modern theory of quantum mechanics, which has been further developed from the work of Rutherford and Bohr, explains much of the structure of atoms and their behaviour. These ideas continue to be modified. The three basic particles (protons, neutrons and electrons) may be further subdivided into over 50 other unstable particles, including quarks, leptons, antiparticles, gluons and photons.

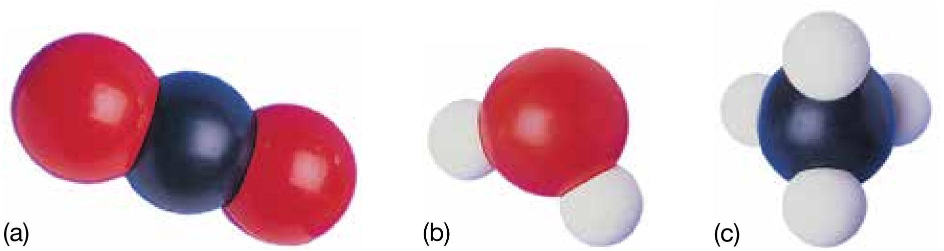
Protons, electrons and neutrons are called **subatomic particles**. Table 1.1 summarises the properties of these particles. Atoms that are neutrally charged have the same number of electrons and protons. If an atom gains or loses an electron it becomes charged and is then called an **ion**. You will study ions and their behaviour in more detail later.

TABLE 1.1 Particles in an atom and their properties

Subatomic particle	Relative mass	Relative charge
electron	$\frac{1}{1837} = 0.0005$	–1
proton	1	+1
neutron	1	0

Atoms are not all the same. To date, chemists have identified about 118 different types of atom. Some substances contain only one type of atom. These substances are called **elements**. For example, the element oxygen contains only oxygen atoms and pure lead contains only lead atoms. The atoms of each element have special characteristics, and these are used to classify them.

Very few elements exist as individual atoms; examples are helium and neon. Many more substances consist of two or more atoms that are chemically combined. These are called **molecules**. The atoms in molecules can be the same, such as nitrogen, N_2, and oxygen, O_2, which are molecules in the air that we breathe. Most molecules contain different atoms; for example, water, H_2O, contains hydrogen and oxygen atoms, and carbon dioxide, CO_2, contains carbon and oxygen atoms. These are quite small molecules, but some molecules can consist of hundreds or even thousands of atoms. Synthetic polymers, proteins and DNA are described as macromolecules. The types of atoms present and the sizes, arrangements and shapes of molecules significantly affect the properties of materials, and this will be a focus of your study for this unit of chemistry.



Molecules of (a) carbon dioxide, CO_2, (b) water, H_2O, and (c) methane, CH_4. These models are called space-filling models; they show the relative sizes of the centre-to-centre distances between the atoms. Such models are useful in showing the dimensions of a molecule but don't show the bonds involved.

A matter of size

Nanotechnology works at the scale of single atoms or small molecules.

A nanometre is one billionth of a metre. 1 nm = 10^{-9} m = 10^{-7} cm

Understanding the dimensions of atoms can be quite a challenge. If the nucleus in the atom on the next page was 5 millimetres in diameter then the outer edge of the atom would be 50 000 millimetres or 50 metres away. When discussing the sizes of atoms, a more useful unit is the nanometre.

Nano- is a prefix like *milli-* and *micro-*:

- A millimetre is one thousandth of a metre: mm = 10^{-3} m.
- A micrometre is one millionth of a metre: µm = 10^{-6} m.
- A nanometre is one billionth of a metre: nm = 10^{-9} m.

We are aware that humans worked for tens of thousands of years with tools made of stone, wood, bone and, much later, bronze and iron. These tools were made on a scale that humans could handle, from a few centimetres to a few metres. Later technologies enabled humans to create things on a larger scale, such as large construction machines and ocean-going ships. The development of the microscope (c. 1600) was the first step in developing technologies that could work at a scale invisible to the unaided eye. Most human eyes cannot see objects smaller than the width of a human hair, or about 0.05 mm. The spectacular developments in electronics since 1970 have been possible due to the increasingly sophisticated control of objects at the microscopic level. Technologies on this scale, such as microelectronics and microsurgery, are often referred to as microtechnologies.

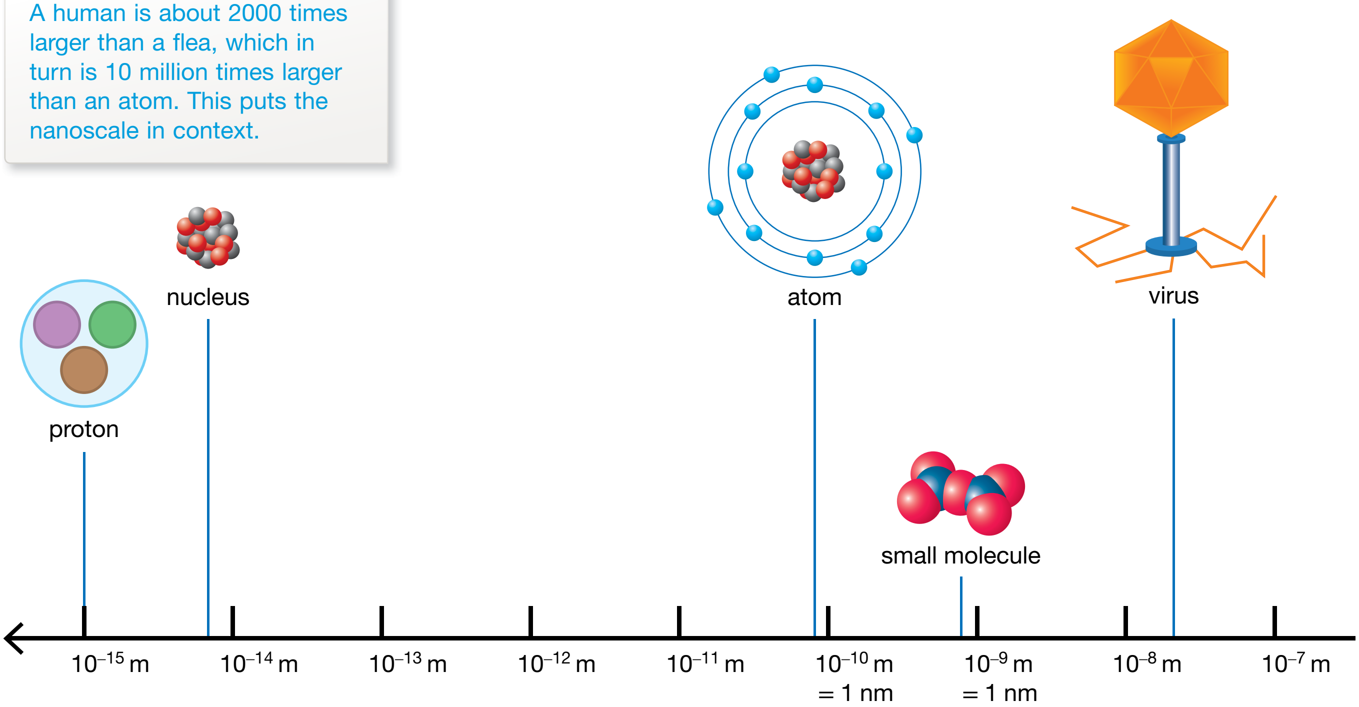
Nanotechnology takes our ability to control things one step beyond the microworld, to the nanoworld. Special tools are needed to work at the nanoscale. Scientists use scanning probe microscopes (SPM) to make images of individual atoms that are much too small to see. SPMs work by 'feeling' the bumps caused by atoms on a surface. Any technology that involves manipulation, construction or control of objects whose size is at the nanometre scale is a branch of nanotechnology. A large atom is about 0.1 nm in diameter. Molecules can range from about 0.15 nm (H_2) to a visible size (macromolecules). Bacteria range in size from about 500 nm to about 5000 nm, and viruses from about 10 to 50 nm.

studyon

Unit 1
AOS 1
Topic 1
Concept 1

The sizes of particles
Summary screen and practice questions

A human is about 2000 times larger than a flea, which in turn is 10 million times larger than an atom. This puts the nanoscale in context.



To compare two objects, their measurements must be converted to the same unit.

$$1\ \text{m} = 10^2\ \text{cm} = 10^3\ \text{mm} = 10^6\ \mu\text{m} = 10^9\ \text{nm}$$

Revision questions

1. **A human hair is about 50 μm thick. A pet flea is about 1 mm in size. How many times smaller is the thickness of a human hair than a flea?**
2. **A red blood cell is about 7 μm in diameter. A virus particle is 50 nm wide. How many times smaller is the virus particle than a red blood cell?**

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Weblink
Scale of the universe

Nanoscience is an interdisciplinary field that involves combinations of knowledge from biology, chemistry, physics, engineering and computer science. It deals directly with atoms, molecules or macromolecules. The extremely small size of nanoparticles gives them unique properties that are best described using quantum rather than classical physics. According to the quantum mechanical model, energy is not continuous but comes in small ‘packets’ known as quanta (singular = quantum) and particles have wave properties. Calculations based on quantum theory can accurately predict the properties of nanoparticles.

The study of nanoparticles is dominated by surface effects such as electronic charge, dispersion forces, and hydrogen, ionic, metallic and covalent bonding. For chemists, nanotechnology will enable materials to be assembled literally atom by atom. This idea is referred to as ‘molecular manufacturing’. Molecular manufacturing has as its goal the ability to control assembly of individual atoms into specific molecules. This could mean production of materials that have no waste products, no defects in materials, no impurities, and no requirement for machining or assembly as the atoms are placed directly where required to make the final object. Nanotechnology has the potential to build cleaner, safer, longer lasting and smarter products for communication, medicine, transportation, industries, agriculture and the home.

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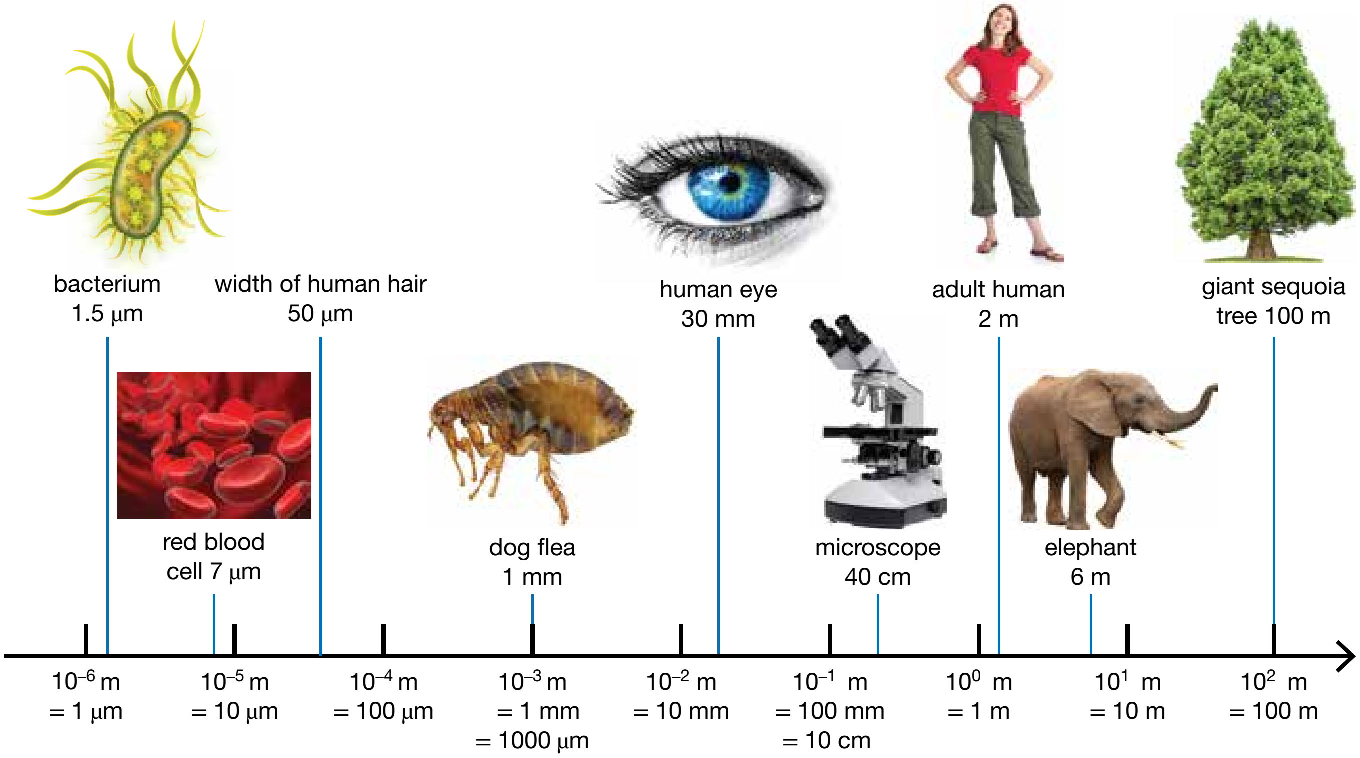
Unit 1
AOS 1
Topic 1
Concept 2

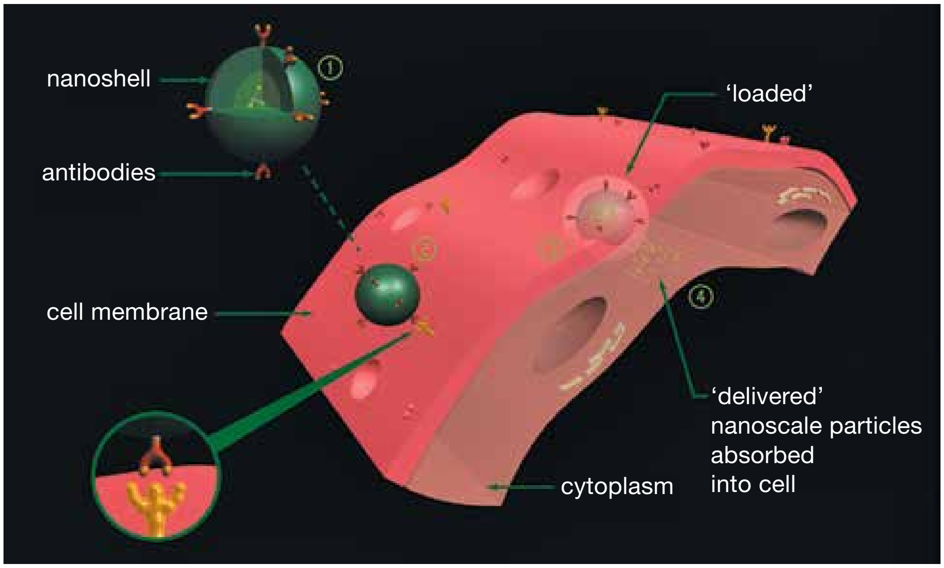
Nanoparticles and nanotechnology
Summary screen and practice questions



Biomedicine

Combined with biotechnological knowledge, nanotechnology has immense potential in saving, improving and extending lives. Nanoscale technologies are helping medicine with devices in areas of drug discovery, diagnostics and therapeutics. For example, trials are being undertaken in which micro-capsules with nanoscale pores containing pancreatic cells are used so that diabetic patients can obtain insulin without having to inject it. The pores in the microcapsules are big enough to allow insulin to flow out of the capsules and be released into the body but small enough to prevent the entry of harmful antibodies. Other research is developing cancer-fighting drugs that can be adsorbed onto nanoparticles so they can bypass the body's natural defence mechanisms and precisely target cancer cells.

Medicines and pharmaceuticals are being designed on a nanoscale to allow more effective and accurate treatments.



Nanoshells with antibodies attached can be delivered through the surface of a cell membrane. The antibodies are attracted to the protein receptor and are loaded into the cell membrane.

Microarrays and labs-on-chips (also known as microfluidic devices) are tools at the microscale for analysing large numbers of DNA or protein samples at a single time. These are used for diagnosis and drug discovery. Microarrays perform one type of analysis thousands of times, and lab-on-a-chip devices can perform multiple types of analysis at once. Research is now being undertaken into producing nanoarrays, in which thousands of binding sites can be printed onto the area of a single conventional microarray spot. Coupled with computer technology, this will potentially enable the detection of diseases and discovery of useful drugs to proceed much more quickly and at lower cost.

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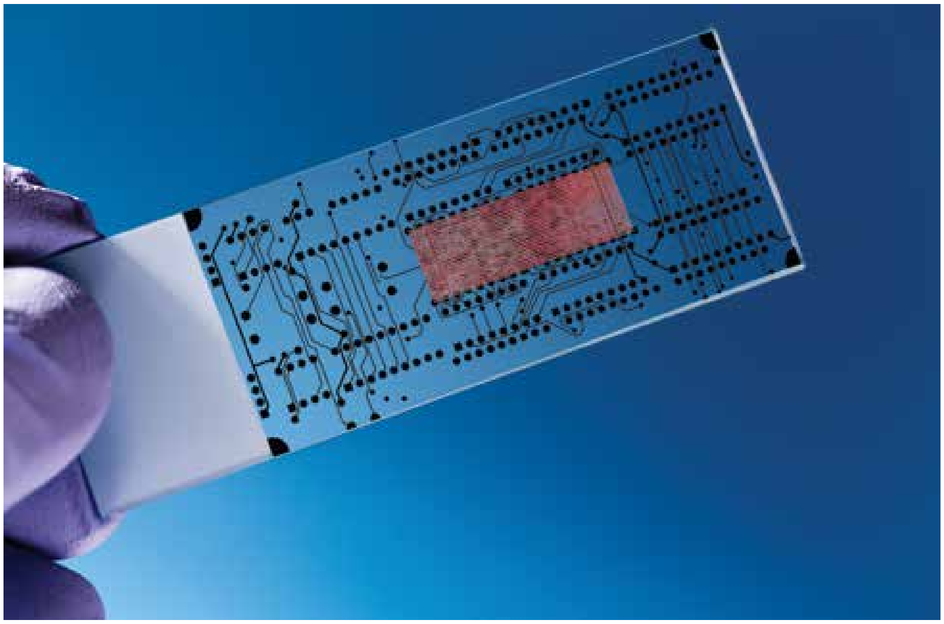
Lab-on-a-chip



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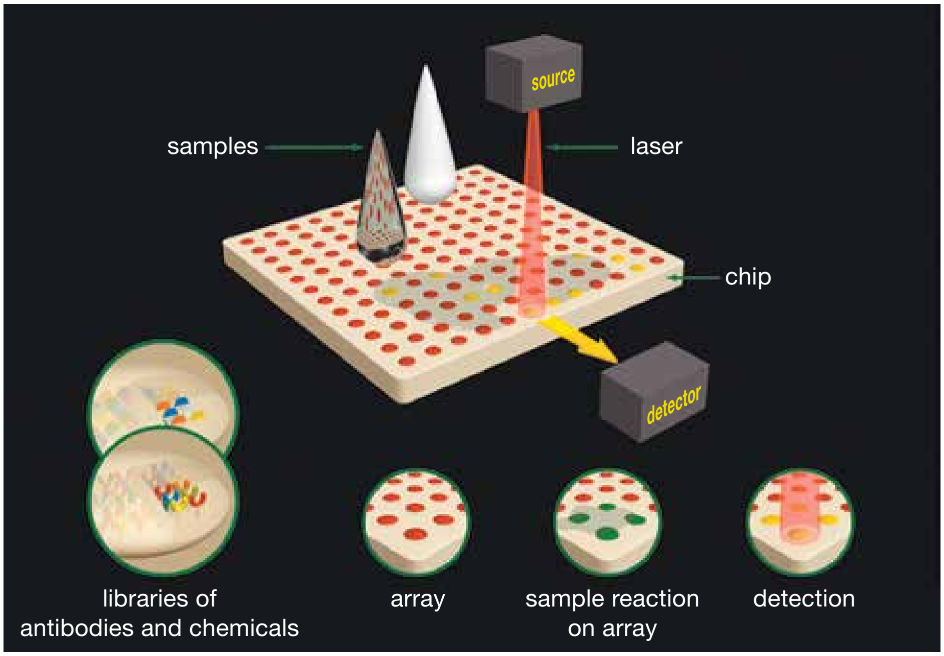

Samples are placed onto the microarray. These may react with the antibodies or chemicals that are placed already onto the microarray surface; the detector then determines if the reaction has occurred. This can be repeated many times with multiple samples.

Responsible use of nanotechnology

There are potential risks with nanotechnology such as military use, privacy issues and accidental release of nanomaterials.

Nanotechnology is likely to benefit many areas of our lives in the future, particularly in the area of improving and extending life and ensuring a cleaner and safer environment to live in. However, there are also potential risks with nanotechnology, such as military use in creating miniaturised weapons and explosives, or in designing biological organisms at molecular levels. Who will own the technology? What risks may be generated as a result of the technology? These and other ethical and social questions have to be raised and discussed by scientists, policymakers and the general public so that we can responsibly develop and use nanotechnology.

Representing atoms

Atomic number (symbol *Z*)

The atomic number (symbol Z) is the number of protons in the nucleus of an atom.

Each of the 118 elements known to chemists has its own **atomic number**. The atomic number (symbol Z) of an element is defined as the number of protons in the nucleus of an atom of that element. When an atom is neutrally charged, the atomic number of the atom corresponds to the number of electrons, because the number of positive charges must be the same as the number of negative charges. For example, oxygen has an atomic number of 8 and therefore has 8 protons and 8 electrons.

Mass number (symbol *A*)

The mass number (symbol A) is the number of protons and neutrons added together.

Protons have approximately the same mass as neutrons. The electron's mass is negligible. Therefore, the mass of an atom depends on the number of particles in the nucleus. The **mass number** (symbol A) is defined as the total number of protons and neutrons in an atom of an element.

We can use the periodic table to identify each element by its atomic number. The relative atomic mass of each element is also shown on the table (this is discussed in chapter 5).

An element is commonly represented as follows:

$$\begin{array}{rl} \text{mass number} \longrightarrow & A \\ \text{atomic number} \longrightarrow & Z \end{array} E \longleftarrow \text{symbol for element}$$

This is known as the **isotopic symbol** of an element. We can determine the number of neutrons in an atom by subtracting the atomic number, Z, from the mass number, A. For example, sodium, Na, has atomic number 11 and mass number 23, and can be represented as ${}^{23}_{11}\text{Na}$. An atom of sodium, therefore, has 11 protons and 12 neutrons.

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Unit 1
AOS 1
Topic 1
Concept 4

Atomic number, mass number and isotopes
Summary screen and practice questions

Revision questions

3. **Look up your periodic table to find the atomic number of each of the following elements.**
- **(a)** **H**
- **(b)** **Ne**
- **(c)** **Ag**
- **(d)** **Au**

4. **An atom has 13 protons and 13 neutrons. What is its:**
- **(a)** **atomic number**
- **(b)** **mass number**
- **(c)** **name?**

5. **Find the symbols for elements with atomic numbers:**
- **(a)** **5**
- **(b)** **12**
- **(c)** **18**
- **(d)** **20.**

Isotopes

Isotopes of an element have the same number of protons but different numbers of neutrons.

All atoms of a particular element contain the same number of protons. However, atoms of the same element can contain different numbers of neutrons, and these atoms are called **isotopes**. Isotopes have similar chemical properties because their electron structure is the same. They do, however, have different physical properties due to their different masses.

Naturally occurring oxygen consists of three isotopes: ${}^{16}_{8}O$, ${}^{17}_{8}O$ and ${}^{18}_{8}O$. Isotopes are named by their element name followed by their mass number to distinguish them; for example, the isotopes of oxygen are oxygen-16, oxygen-17 and oxygen-18. Aluminium has only one isotope, aluminium-27, ${}^{27}_{13}Al$.

Isotopes have various applications in society. One of these is radiocarbon dating. Pure carbon has three isotopes: ${}^{12}_{6}C$, ${}^{13}_{6}C$ and ${}^{14}_{6}C$. The radioactive isotope ${}^{14}_{6}C$, with a half-life of 5700 years, is present in small amounts in living things. The amount of ${}^{14}_{6}C$ left in the remains of an organism can be measured, so the age of the remains can be determined. In the past, Aboriginal Australians sometimes used human blood in the red pigment of their rock painting. This enables chemists to determine the ages of such paintings. Tiny fragments of the pigmented layers that flake off the rock walls are collected, and carbon-dating experiments are performed on the organic matter in the blood residues. Some motifs are as ancient as 20 000 years old.

Sample problem 1.1

1. The nucleus of an atom contains 19 protons and 20 neutrons.
   (a) Write the isotopic symbol for the element.
   (b) What is the value of *Z*?
   (c) What is the value of *A*?
   (d) How many electrons does this atom have?
2. Find the number of electrons, protons and neutrons in ${}^{35}_{17}Cl$.

Solution: 1. (a) The atom with 19 protons is potassium, which can be written as ${}^{39}_{19}K$.
(b) The number of protons, *Z*, is 19.

(c) The mass number is equal to the number of protons plus the number of neutrons; therefore, $A = 39$.

(d) The number of electrons is equal to the number of protons; therefore, there are 19 electrons in the potassium atom.

2. If $Z = 17$, then the number of protons is 17.
   The number of electrons is 17 since atoms are neutral.

   If $A = 35$, then the number of neutrons $= 35 - 17$
   $= 18$.

Revision questions

6. **In the element argon, $Z = 18$ and $A = 40$. For argon, state:**
   - (a) **the number of neutrons**
   - (b) **the number of electrons**
   - (c) **the isotopic symbol for this element.**
7. **How many protons and neutrons are in ${}^{79}_{35}\text{Br}$?**
8. **An atomic nucleus consists of one proton and one neutron. What is its symbol?**
9. **Copy the following table into your book and complete it.**

Element	Number of protons	Number of electrons	Number of neutrons
${}^{12}_{6}\text{C}$			
${}^{56}_{26}\text{Fe}$			
${}^{40}_{18}\text{Ar}$			
${}^{235}_{92}\text{U}$			
${}^{238}_{92}\text{U}$			
${}^{19}_{9}\text{F}$			

   - (a) **Identify any isotopes in the table.**
   - (b) **Explain the difference between the isotopes.**
10. (a) **Write the isotopic symbols for the atoms nitrogen-14 and nitrogen-15.**
    - (b) **How many protons does each atom have?**
    - (c) **How many neutrons does each atom have?**
    - (d) **What are these atoms called?**
11. **Why do we identify an element by its atomic number rather than its mass number?**

Exciting electrons

What causes rainbows? Why does the blue gas flame turn yellow when salt spills on it? Why is it that when you look into a fire you see different coloured flames? The answers lie in the way the electrons are arranged around the nucleus of the atom. This arrangement largely determines the properties and the behaviour of elements and the materials made from them.

When white light is separated by a prism, a *continuous spectrum* of colour is observed. Every element emits light if it is heated by passing an

Emission spectrum lines correspond to energy differences between electron shell levels.

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Digital document
Experiment 1.1 Flame colours
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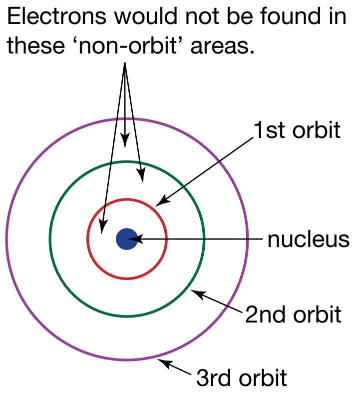
electric discharge through its gas or vapour. This happens because the atoms of the element absorb energy, and then lose it, emitting it as light. Passing the light emitted by an element through a prism produces an **atomic emission spectrum** for that element.

The emission spectra of elements are quite different from the spectrum of white light. White light gives a continuous spectrum, whereas atomic emission spectra consist of separate lines of coloured light. Each line in an emission spectrum corresponds to one particular frequency of light being given off by the atom; therefore, each line corresponds to one exact amount of energy being emitted.

Bohr's energy levels

In 1913, Niels Bohr suggested an explanation for the emission spectrum by proposing a model for the hydrogen atom. He proposed that electrons of specific energy move around the central nucleus in circular orbits or energy levels. Electrons cannot exist between these orbits. Although an electron cannot lose energy while orbiting a nucleus, it could be given excess energy (by a flame or electric current) and then move to a higher orbit. If this happens, the electron has moved from the **ground state** (lowest energy level) to an **excited** state. When it drops back down to a lower, more stable orbit, this excess energy is given out as a **photon** or quantum of light. This is seen as a line of a particular colour on the visible spectrum.



The Bohr model of an atom

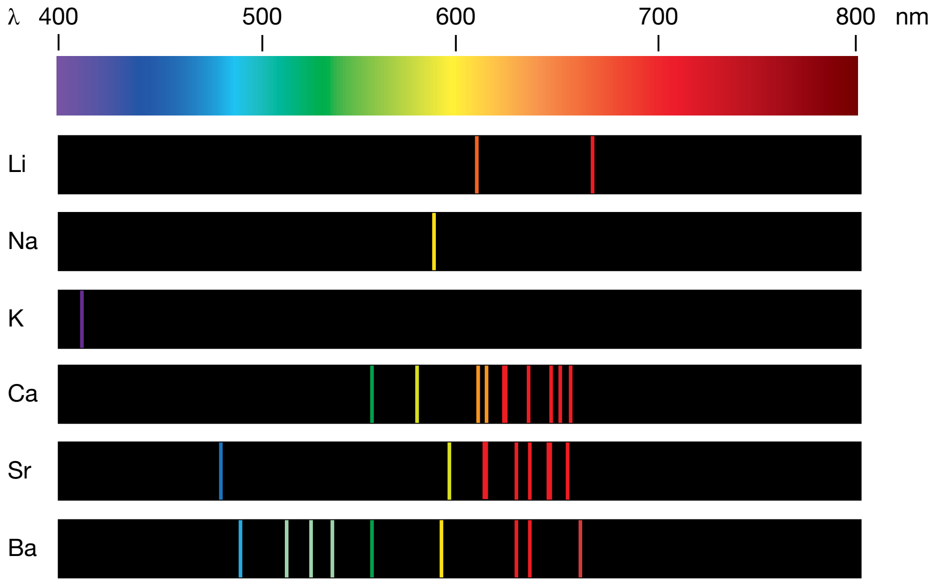
The energy given out is the difference in energy between the two energy levels. Since there are only certain allowed energy levels, the energy released has specific allowed values, each corresponding to a line in the emission spectrum. This spectrum is different for each element, so it is often called the 'fingerprint' by which an element may be identified.

Some metallic elements can be identified simply by their characteristic flame colours when heated. Copper burns with a blue-green flame, and sodium burns with a yellow flame.

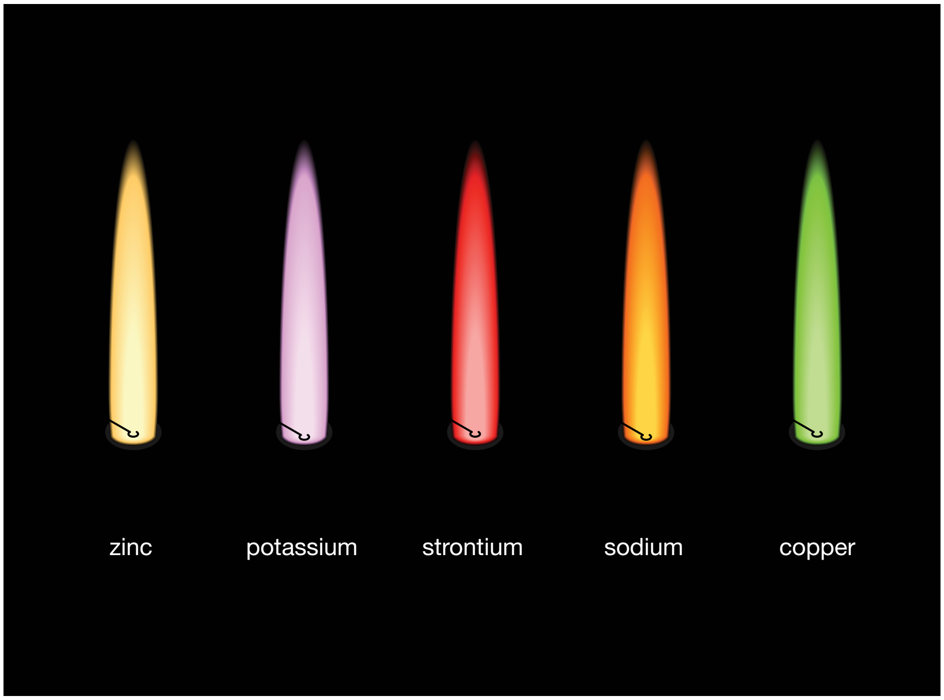
study on

Unit 1
AOS 1
Topic 1
Concept 5

Spectroscopy and spectral evidence for Bohr's model of the atom
Summary screen and practice questions



White light is a continuous spectrum (top). The emission spectra of various atomic elements consist of distinct lines that correspond to differences in energy levels.



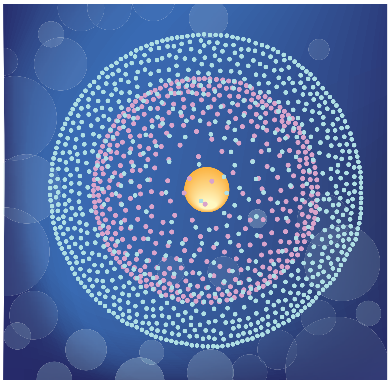
When wires with small amounts of different metal salts are placed in a flame, the electrons are excited and emit characteristic coloured light.

If an atom gains or loses an electron it becomes charged and is called an ion.

Electron shells

Electrons may be visualised as moving within a region of space surrounding the nucleus. The regions are called electron shells, labelled *K*, *L*, *M* and *N*, and numbered 1, 2, 3 and 4. A definite energy level is associated with each shell; the one closest to the nucleus, *K*, has the lowest energy level. An electron has to gain energy to move further away from the nucleus. If it gains enough energy to completely leave the atom, the particle that is left is no longer neutral and is called a positive ion. Potassium, *K*, has 19 protons and 19 electrons. If it loses an electron, it becomes the positive ion K^+ as it now has 19 protons and only 18 electrons.

Further studies of line spectra in the 1910s and 1920s led to the prediction that there was a maximum number of electrons that could be present in a given energy level.



Beryllium has two electrons in the inner shell, *K*, and two electrons in the next shell, *L*.

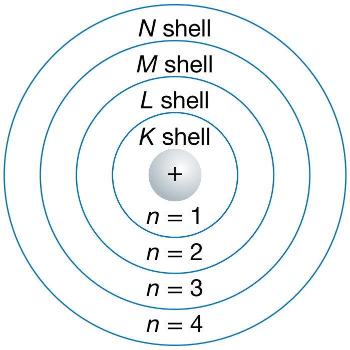
The maximum number of electrons a shell can hold is $2n^2$. For the first 20 elements, the maximum number of electrons in an outer shell is 8.

Electron configuration

The arrangement of electrons in the shells is called the atom's **electron configuration**. The electron capacity of each shell is limited. The maximum number of electrons that each shell can hold is $2n^2$ where n is the shell number or energy level.

Electron shells are filled in order from the nucleus, starting with the *K* shell, so that the electrons are in their lowest possible energy levels (or ground state). The one electron of a hydrogen atom would be in the first shell, and the electron configuration is written as 1. Sodium has 11 electrons, so two go into the first shell, eight in the second and the last electron in the third shell. The electron configuration of sodium, therefore, is written as 2, 8, 1. Chlorine has 17 electrons and an electron configuration of 2, 8, 7. Note that for the first 20 elements *the outer shell never has more than eight electrons*. Potassium, for example, has 19 electrons and an electron configuration of 2, 8, 8, 1 rather than 2, 8, 9. This means that the fourth shell is the outer shell for potassium electrons, rather than the third.

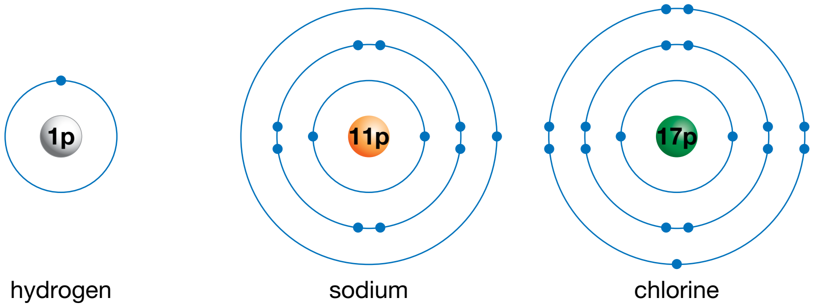
Ions are atoms that have lost or gained one or more electrons. For example, a sodium atom has 11 electrons, so its electron configuration is 2, 8, 1. A sodium ion, Na^+, has lost an electron so its electron configuration is 2, 8. Chemists are particularly interested in the electrons in the highest energy level of an atom since it is these outer-shell electrons that mainly determine the chemical properties of elements. These electrons are called valence electrons.



Electrons are arranged around the nucleus in shells.

Shell model diagrams

The electron configuration of an atom can also be represented using shell model diagrams, like those in the figure below. These show the electron shells as concentric rings around the nucleus, with the electrons marked on each ring, and help us to visualise the structure and behaviour of atoms.



Shell model diagrams of hydrogen, sodium and chlorine. The dots represent the electrons in the shells.

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Limitations of the shell model

The shell model represents only part of the story of the atom. More discoveries are always being made that cause scientists to reconsider their models and their understanding of the atom. The shell model doesn't really explain the various differences in energies between the electron shells. It seems to imply that all the electrons orbit the nucleus in exactly circular paths, like planets around a sun. We know from looking at molecules with electron tunnelling microscopes that they come in many different shapes and sizes, so this model does not fully explain every aspect of every atom.

The order of the electrons filling the electron shells is not really explained by this model either. For example, compare calcium (2, 8, 8, 2) with scandium (2, 8, 9, 2) — why isn't the electron configuration of scandium 2, 8, 8, 3? Other models have been developed that are more complex and explain more of the data scientists have gathered.

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Unit 1	**Development of atomic theory**
AOS 1	Summary screen and practice questions
Topic 1	
Concept 3	

Revision questions

12. **Copy the table below and complete the columns. The first one has been done for you.**

Energy level (*n*)	Symbol	Maximum number of electrons ($2n^2$)
1	*K*	$2 \times 1^2 = 2$
2		
3		
4		

13. (a) **What was the importance of Bohr's contribution to our understanding of the atom?**
    (b) **What are the limitations of Bohr's model?**

14. **Copy and complete the following table to show the electron configuration of the first twenty elements. Some of the configurations have been provided for you.**

			Electron configuration			
Element	**Symbol**	**Atomic number**	***K***	***L***	***M***	***N***
hydrogen	H	1	1			
helium	He	2				
lithium	Li	3				
beryllium	Be	4	2	2		
boron	B	5				
carbon	C	6				
nitrogen	N	7				
oxygen	O	8				
fluorine	F	9				
neon	Ne	10				
sodium	Na	11	2	8	1	
magnesium	Mg	12				
aluminium	Al	13				
silicon	Si	14				
phosphorus	P	15				
sulfur	S	16				
chlorine	Cl	17	2	8	7	
argon	Ar	18				
potassium	K	19	2	8	8	1
calcium	Ca	20				

15. **Using the periodic table, produce shell model diagrams for the first 10 elements.**

16. **(a)** **What is the electron configuration of nitrogen?**
(b) **Which element has atoms with the electron configuration of 2, 8, 8, 2?**

Erwin Schrödinger — the quantum-mechanical model of an atom

Schrödinger developed the quantum mechanical model of the atom, which features electrons arranged in shells, subshells and orbitals within an atom.

Although Bohr's model of an atom could account for the lines seen in the emission spectra of hydrogen, it did not explain and could not mathematically predict the lines seen in the emission spectra of the more complex atoms. Furthermore, it did not explain why electrons moving around the nucleus and emitting electromagnetic radiation did not fall into the nucleus of the atom, causing it to collapse.

In 1923 the French scientist Louis de Broglie (1892–1987) proposed, using Albert Einstein's and Max Planck's quantum theory, that electrons show both particle and wave behaviour. In 1926, the Austrian physicist Erwin Schrödinger (1887–1961) made use of the new quantum theory to refine Bohr's model of the atom. He wrote and developed a wave equation describing the location and energy of an electron in a hydrogen atom. From this, he developed the

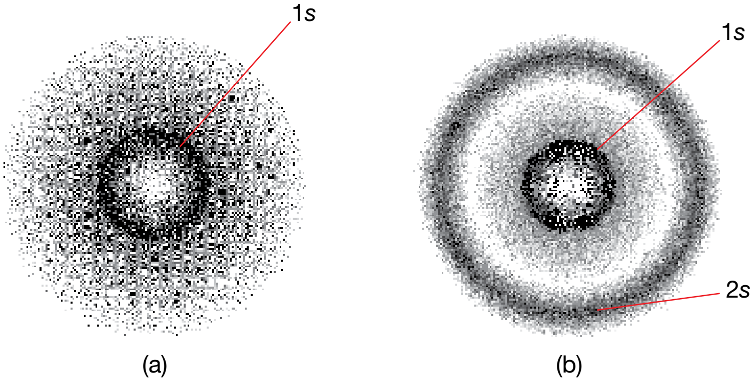
Erwin Schrödinger proposed a charge cloud model of the atom.

studyon

Unit 1	**Electron configuration and the Schrödinger model**
AOS 1	Summary screen and practice questions
Topic 1	
Concept 6	

currently accepted quantum-mechanical model of the atom — a complex mathematical model based on particles such as electrons showing wave-like behaviour.

According to quantum mechanics, the electron is not considered as moving along a definite path. Instead, the electron is found in a region of space around a nucleus called an **orbital**. An orbital may be visualised as a blurry cloud of negative charge; the cloud is most dense where the probability of finding the electron is large and less dense where the probability of finding the electron is small.



In the quantum-mechanical model, an electron around a nucleus may be visualised as a cloud of negative charge: (a) the charge cloud for the 1*s* electron in hydrogen (b) charge clouds for the electrons in the 1*s* and 2*s* subshells. As in the Bohr model, the electron is attracted to the nucleus by electrostatic forces and moves in such a way that its total energy has a specific value.

In the quantum-mechanical model of the atom:

1. The energy levels of electrons are designated by **principal quantum numbers**, ***n***, and are assigned specific values: $n = 1, 2, 3, 4, 5$ and so forth. These principal quantum numbers may be referred to as shells and are also called the *K*, *L*, *M* and *N* shells respectively.
2. Within each shell, several different energy levels called subshells may be found. The number of subshells equals the shell number; for example, if the shell number is 2, there are two subshells at that energy level. Each subshell corresponds to a different electron cloud shape. Subshells are represented by the letters *s*, *p*, *d*, *f* and so on.

TABLE 1.2 Energy levels within shells of an atom

Shell number (*n*)	Shell symbol	Number of subshells	Subshell symbol	Maximum number of electrons in subshell
1	*K*	1	*s*	2
2	*L*	2	*s*	2
			p	6
3	*M*	3	*s*	2
			p	6
			d	10
4	*N*	4	*s*	2
			p	6
			d	10
			f	14

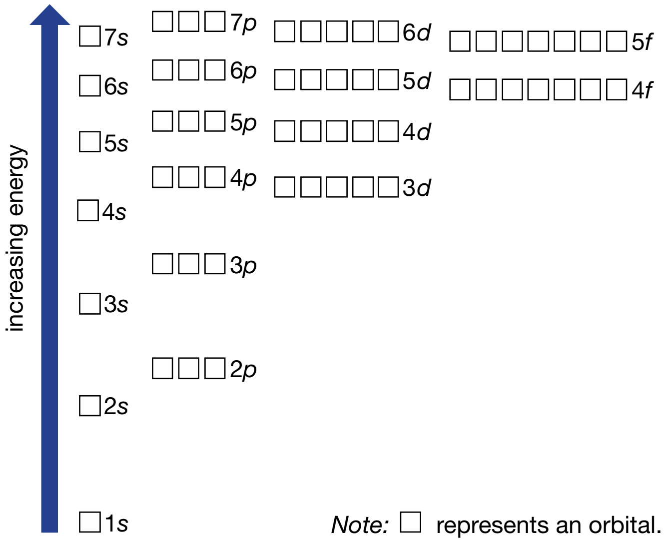
Electron configuration

The order of filling subshells is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$.

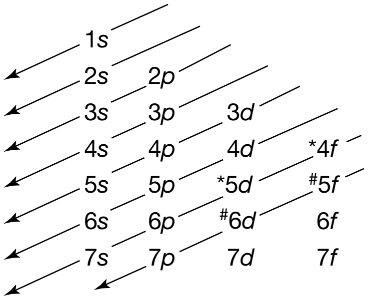
The way in which electrons are arranged around the nucleus of an atom is called the electron configuration of the atom.

Generally, the order of subshell filling is from lowest energy first:

$$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p \ldots$$



The energy levels of atomic orbitals



* and #: In both these cases it is found that the two subshells fill at the same time.

The order of filling of subshells may be found by following the arrows from tail to head, starting with the top arrow.

Notice that the 4*s* subshell is filled before the 3*d* subshell, which is of a higher energy than the 4*s* subshell. Likewise, the 4*d* subshell is higher in energy than the 5*s* subshell, and so on.

Excited states

When an atom moves to a higher energy level than the ground state by absorbing energy, its electron configuration changes. The outermost electron moves to a higher energy level subshell. For example, neon has two electrons in the first shell and eight in the second shell; when the outermost electron gains energy, $1s^22s^22p^6$ becomes $1s^22s^22p^53s^1$. Note that, once the order of filling subshells has been determined, the subshells are written in increasing numerical order, rather than the order of increasing energy. So $1s^22s^22p^63s^23p^64s^23d^2$ becomes $1s^22s^22p^63s^23p^63d^24s^2$.

The electron configuration for the elements hydrogen, sodium and scandium, for example, can be written as:

H $(Z = 1)$	$1s^1$
Na $(Z = 11)$	$1s^22s^22p^63s^1$
Sc $(Z = 21)$	$1s^22s^22p^63s^23p^64s^23d^1$

The electron configuration for scandium can also be written as $1s^22s^22p^63s^23p^63d^14s^2$ to show that the last occupied shell is the fourth one; nevertheless, the fourth shell starts to fill before the third shell is complete.

The group and period in which an element is found is easily read from the electron configuration, as shown in the example below:

group 2 (is the number of valence electrons in the last shell being filled)

Be $(Z = 4)$ $1s^22s^2$

period 2 (the highest shell number being filled)

group 18 (the outer shell is filled; argon is a noble gas)

Ar $(Z = 18)$ $1s^22s^22p^63s^23p^6$

period 3

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Interactive periodic table

The group number is found by using the total number of electrons in the last shell to count across the periodic table. The period is the number of the last shell occupied by electrons.

Sample problem 1.2

Find the ground state electron configuration of a fluorine atom.

Solution: Fluorine has nine electrons. According to the order of subshell filling, the 1*s* subshell in a fluorine atom fills first, and contains two of fluorine's nine electrons. The next energy level is the 2*s* subshell — this holds another two electrons. The 2*p* subshell can hold six electrons, but, since only five electrons remain to be placed, one of the 2*p* subshell orbitals is incomplete.

Fluorine's electron configuration may be written as follows:

$1s^2 2s^2 2p^5$

Sample problem 1.3

Find the ground state electron configuration of a potassium atom.

Solution: Potassium has 19 electrons. Two electrons fill the 1*s* subshell, two electrons fill the 2*s* subshell, six electrons fill the 2*p* subshell, two electrons fill the 3*s* subshell and six electrons fill the 3*p* subshell. This means that one electron remains to be placed. The next subshell, 3*d*, is of a higher energy level than the 4*s* subshell. The remaining electron is placed in the 4*s* subshell. The electron configuration of potassium is:

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

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Unit 1
AOS 1
Topic 1
Concept 7

Writing electron configurations using subshell notation
Summary screen and practice questions

Revision questions

17. **Write the full electron configuration of the following elements: lithium, potassium, carbon, chlorine, argon and nitrogen.**
18. **The isotope $^{32}_{15}P$ is used in the treatment of leukaemia.**
    (a) **Write the full electron configuration of this isotope.**
    (b) **How does it differ from $^{31}_{15}P$?**
19. **An element *X* has configuration $1s^2 2s^2 2p^6 3s^2 3p^4$.**
    (a) **What group is it in?**
    (b) **What period is it in?**
    (c) **Give its name and symbol.**
20. **Name the elements with the following configurations, and state the group number and period number in each case.**
    (a) $1s^2 2s^1$
    (b) $1s^2 2s^2 2p^6$
    (c) $1s^2 2s^2 2p^6 3s^1$
    (d) $1s^2 2s^2 2p^6 3s^2 3p^1$
    (e) $1s^2 2s^2 2p^6 3s^2 3p^6$
    (f) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
21. **For each of the elements in the following two sets of atomic numbers, write the electron configuration and state which elements belong to the same group.**
    (a) $Z = 20, 12, 4, 9$
    (b) $Z = 5, 6, 8, 16$

22. **Write the electron configurations and element symbols of:**
  - **(a)** **the first three elements in group 18**
  - **(b)** **the element in group 13, period 3**
  - **(c)** **the element in group 15, period 2**
  - **(d)** **the element in group 17, period 3.**

Sample problem 1.4

Find the ground state electron configuration of an aluminium ion, Al^{3+}.

Solution: An aluminium ion, Al^{3+}, is an aluminium atom ($Z = 13$) that has lost three electrons. Therefore, its electron configuration is as follows:

$1s^22s^22p^6$

Sample problem 1.5

The electron configuration below represents an atom or ion. Is it in an excited state?

$1s^22s^22p^63p^1$

Solution: The species contains 11 electrons. It represents an atom or ion in an excited state, because the final electron would, in the ground state, be placed in a 3*s* orbital.

Revision questions

23. **Distinguish between the terms 'shell' and 'subshell'.**

24. **Write the ground state electron configurations of the first 30 elements of the periodic table.**

25. **Which of the following electron configurations are ground state configurations and which are excited state configurations?**
  - **(a)** $1s^22s^22p^63s^1$
  - **(b)** $1s^22s^22p^63s^23p^63d^1$
  - **(c)** $1s^22s^22p^63s^23p^63d^54s^2$

26. **A neutral magnesium atom has an electron configuration of** $1s^22s^22p^63s^16s^1$**.**
  - **(a)** **How can you tell that the atom is in an excited state?**
  - **(b)** **Describe what would happen if the atom changed its electron configuration to the ground state.**
  - **(c)** **Write the ground state electron configuration of the magnesium ion** Mg^{2+}**.**

27. **Identify the following elements with ground state electron configurations of:**
  - **(a)** $1s^22s^22p^63s^23p^63d^54s^2$
  - **(b)** $1s^22s^22p^63s^23p^64s^2$.

Chromium and copper — atypical electron configurations

Chromium has a half-filled *d* subshell. Copper has a filled *d* subshell. These unusual electron configurations make the elements more stable.

A few elements have electron configurations that do not follow the usual pattern. Chromium (atomic number 23) and copper (atomic number 29), for example, may be expected to be written as follows:

chromium $1s^22s^22p^63s^23p^63d^44s^2$

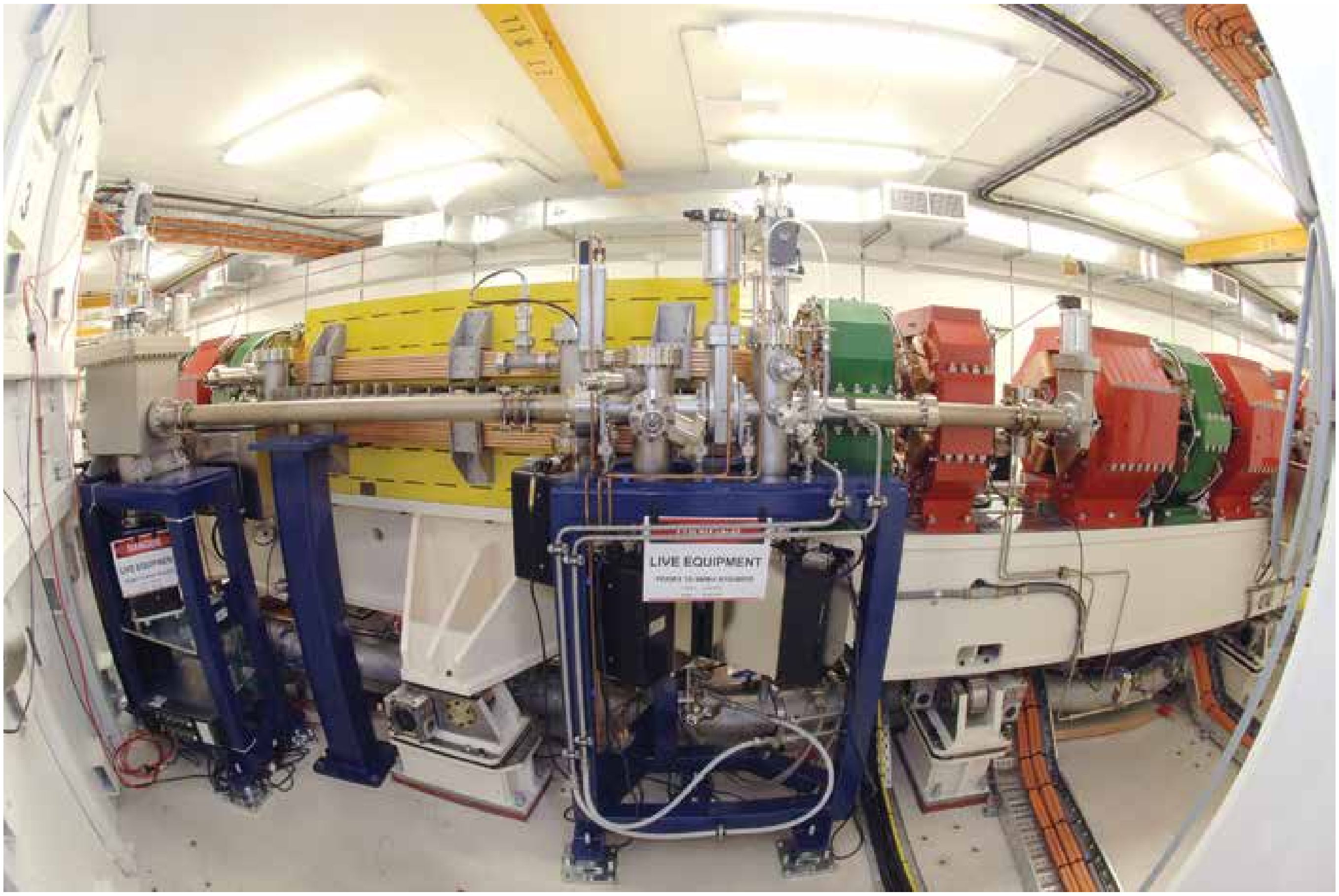
copper $1s^22s^22p^63s^23p^63d^94s^2$.

The correct electron configurations are shown below:

chromium $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$

copper $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$.

These arrangements give chromium a half-filled *d* subshell and copper a filled *d* subshell. Filled subshells are more stable than half-filled subshells. However, half-filled subshells are more stable than other partly filled subshells.



The Australian Synchrotron is a particle accelerator in Victoria that is used to discover more about the structure of matter.

From the outside, the Australian Synchrotron resembles a football stadium. Inside, however, instead of footballs going in different directions, electrons are accelerated around a large loop at almost the speed of light. These high-energy electrons are deflected by magnets and produce a very bright light that is considerably brighter than the sun. This light is directed to a number of workstations where many different experiments take place. Research areas include agricultural science, environmental science, minerals analysis, medical investigations, materials science, nanotechnology and forensics.

Chapter review

studyon

Unit 1 **Atomic structure**
AOS 1 Summary screens and practice questions
Topic 1

Summary

- The nuclear model of the atom proposes that atoms are composed of the following subatomic particles:
  - protons, which are positively charged particles found in the nucleus
  - neutrons, which are neutrally charged particles found in the nucleus
  - electrons, which are negatively charged particles found in energy levels around the nucleus.
- Each element is represented by a specific chemical symbol and is composed of only one kind of atom.
- The atoms of an element:
  - all contain the same number of protons
  - are neutral, as the number of electrons is equal to the number of protons.
- The atom is mainly empty space.
- A molecule is a group of atoms bonded together.
- Particles at the atomic level are measured using the nanometre (nm); 1 nm = 10^{-9} m.
- Nanotechnology works at the scale of atoms and molecules.
- The high surface area to volume ratio of nanoparticles and extremely small size gives them unique properties that are best described using quantum rather than classical physics.
- Nanotechnology has important applications in items such as sunscreen, fabric protection, self-cleaning surfaces and surface adhesion. It has many potential benefits in biotechnology and drug delivery.
- An isotope is an atom of an element with a different number of neutrons and hence a different mass number.
- The isotopic symbol of an element E may be written as ${}^{A}_{Z}E$ where:
  - atomic number (Z) = number of protons
  - mass number (A) = number of protons + number of neutrons.
- Niels Bohr's model suggested that electrons orbited around the nucleus in definite paths of energy called quantum levels or shells. A quantum of energy or photon is emitted when an electron that has gained enough energy to move to a higher energy level then falls back to its ground state position. Bohr determined the different energy levels by using mathematical formulas that measured the wavelengths of the different energy levels.
- Erwin Schrödinger's most important contribution to modern atomic theory was his development of the mathematical description of the paths that electrons would most likely follow in their orbits around the nucleus. The formulas he developed formed the basis of the quantum-mechanical model of the atom. In it he proposed that, instead of Bohr's idea of electrons following predetermined paths, they move around in regions of space called orbitals. The quantum-mechanical model has the following features:
  - Orbitals are of various shapes and are found within subshells, which in turn are found in shells.
  - Each shell has a different energy level, with the shell furthest away from the nucleus having the highest energy level and the one closest to the nucleus having the lowest energy level.
  - Each energy shell has a principal quantum number, n. The first four shells have $n = 1, 2, 3$ and 4; they are also called the K, L, M and N shells, respectively.
  - Subshells are energy levels found within shells. There are four different types of subshells — the s, p, d and f subshells.
  - The maximum number of electrons the different subshells can hold are:
    s subshell: 2 electrons
    p subshell: 6 electrons
    d subshell: 10 electrons
    f subshell: 14 electrons.
  - Electrons fill shells and subshells of lowest energy first. The order of filling is: $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d$.
- The ground state electron configuration of an atom refers to electrons in their lowest energy level. Any other configurations represent the atom in an excited state and in a higher energy level.

Multiple choice questions

1. An atomic particle has a net charge of zero and is found in the nucleus. Which of the following particles is it?
A Proton
B Electron
C Neutron
D Positron

2. Atoms:
A have a nucleus that occupies most of the atom's volume
B are all the same size
C contain only protons in their nucleus
D have most of their mass in the nucleus.

3. Which of the following particles have approximately the same mass?
A A proton and an electron
B A proton and a neutron
C A neutron and an electron
D An electron and a hydrogen atom

4. A nanometre is:

A one billionth of a metre
B one hundredth of a metre
C one thousandth of a metre
D one hundred thousandth of a metre.

5. What is the nanometre scale in centimetres?

A 10^{-5}
B 10^{-9}
C 10^{-7}
D 10^{-12}

6. Which of the following is the correct ranking of prefixes from largest to smallest when they refer to particle size?

A Nano, micro, macro
B Nano, macro, micro
C Micro, macro, nano
D Macro, micro, nano

7. What is the size of the smallest thing that the human eye can see?

A 2000 nm
B 5000 nm
C 10 000 nm
D 50 000 nm

8. In the nucleus of an atom there are 11 protons and 12 neutrons. Its mass number is:

A 11
B 12
C 1
D 23.

9. The element $^{16}_{8}\mathrm{O}$ has the atomic number:

A 24
B 16
C 8
D 2.

10. Which of the following statements regarding subatomic particles is correct?

A Protons are positively charged particles and neutrons are negatively charged.
B The relative masses of an electron, a proton and a neutron are all about 1 unit.
C In a neutral atom, the number of neutrons is equal to the number of protons.
D Isotopes of an element have the same number of protons but different numbers of neutrons.

11. Bohr's theory of the atom proposed that:

A electrons orbit the nucleus like planets move around the sun
B no more than two electrons are allowed in any energy level
C electrons move around the nucleus in fixed orbits, each of which has a different energy level
D energy shells have subshells that contain regions of space called orbitals.

12. The maximum number of electrons that can be placed in the shell $n = 3$ is:

A 18
B 8
C 2
D 32.

13. Which of these electron configurations represents an atom in an excited energy state?

A $1s^2 2s^2 2p^4 3s^2$
B $1\mathrm{s}^2 2s^2 2p^6 3s^2 3p^4$
C $1s^2 2s^2 2p^6$
D $1s^2 2s^2 2p^6 3s^2 3p^6 3d^3 4s^2$

14. The electron configuration of $^{35}_{17}\mathrm{Cl}^-$ is:

A $1s^2 2s^2 2p^6 3s^2 3p^5 3d^1$
B $1s^2 2s^2 2p^6 3s^2 3p^4 4s^2$
C $1s^2 2s^2 2p^6 3s^2 3p^6$
D $1s^2 2s^2 2p^6 3s^2 3p^4 3d^1 4s^1$.

15. The electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6$ represents which of the following ions?

A O^{2-}
B S^{2-}
C Al^{3+}
D Na^+

16. The electron configuration of an atom X is $1s^2 2s^2 2p^6 3s^2 3p^1$. Which of the following formulas is most likely to be a compound formed with X?

A $X\mathrm{F}_2$
B $\mathrm{Ca}X$
C $X\mathrm{Cl}_3$
D $\mathrm{Mg}X_2$

Review questions

The structure of atoms

1. Briefly discuss the similarities and differences between protons, neutrons and electrons.

A matter of size

2. Match each object with its correct size.

Object	Size
red blood cell	0.15 nm
virus	7000 nm
hydrogen molecule	30 nm

3. Which of these is at the nanoscale level (between 1 and 100 nanometres)?

(a) A water molecule
(b) Width of a DNA molecule
(c) A white blood cell
(d) A crystal of salt
(e) A snowflake

4. Reorder the following from smallest to largest: bacteria, flea, width of a human hair, DNA molecule, helium atom.

Representing atoms

5. For each of the atoms below state the:

(i) atomic number
(ii) mass number
(iii) number of protons
(iv) number of neutrons
(v) number of electrons
(vi) name of element.

(a) $^{23}_{11}\mathrm{Na}$
(b) $^{19}_{9}\mathrm{F}$
(c) $^{28}_{14}\mathrm{Si}$
(d) $^{56}_{26}\mathrm{Fe}$
(e) $^{197}_{79}\mathrm{Au}$
(f) $^{235}_{92}\mathrm{U}$

6. Copy and complete the following table.

Name of atom	Atomic number	Mass number	Protons	Neutrons	Electrons
	18	36			
		34		18	
	18			20	
		31			15
lead				126	82
			19	20	
				20	16

7. What is the general name for the group of atoms that includes carbon-12, carbon-13 and carbon-14?

8. For each of the atoms in question 7, state:
(a) the atomic number
(b) the mass number
(c) the number of neutrons
(d) the number of electrons.

Isotopes

9. From the information in the table below:
(a) identify those species that are isotopes of the same element
(b) write the isotopic symbol for each atom.

Atom	Neutrons	Protons	Electrons
A	20	20	20
B	20	17	17
C	22	20	20
D	26	20	20

10. The isotopes below belong to three elements. Identify the elements and list the isotopes next to the name of each element.

${}^{37}_{17}A$ ${}^{35}_{17}D$ ${}^{60}_{27}F$
${}^{26}_{12}B$ ${}^{25}_{12}E$ ${}^{24}_{12}G$
${}^{59}_{27}C$

Electron arrangement in atoms

11. (a) What is required for an electron in an atom to move from a lower energy level to a higher energy level?
(b) What happens when an electron returns to a lower level?

12. If the flame spectra of calcium chloride, calcium carbonate and calcium nitrate were observed, would you observe one main colour or three?

13. Sodium vapour street lights emit a characteristic yellow light. Explain this.

14. An atom has several excited states but only one ground state. What does this mean?

15. How does an atom become an ion?

16. What is the maximum number of electrons that can be accommodated in the first four energy levels?

17. Name the elements that have the following electron configurations.
(a) 2, 6
(b) 2, 8, 3
(c) 2, 8, 7
(d) 2, 8, 8
(e) 2, 8, 8, 1

18. Write the electron configurations of:
(a) fluorine
(b) potassium
(c) oxygen
(d) an atom for which $Z = 12$
(e) an atom containing 20 electrons.

19. Explain the difference between:
(a) a 'shell' and a 'subshell'
(b) an atomic 'orbit' and an 'orbital'.

20. What is the maximum number of electrons that may be found in:
(a) a $3p$ subshell
(b) a $2s$ subshell
(c) a $4d$ subshell
(d) the third shell?

21. How many valence electrons are there in an atom with the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^3$?

22. Chromium (atomic number 23) and copper (atomic number 29) do not follow the usual pattern of subshell electron configuration.
(a) Why does this occur?
(b) Write the subshell electron configuration for chromium.
(c) Write the subshell electron configuration for copper.

Exam practice questions

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Unit 1 | AOS 1 | Topic 1

Atomic structure

Sit topic test

In a chemistry examination, you will be required to answer a number of short and extended response questions.

Multiple choice questions

1. The number of valence electrons in the element calcium is:

A 1
B 2
C 20
D 40. **1 mark**

2. All isotopes of a given element have the same:

A atomic number
B number of neutrons in their nuclei
C mass number
D atomic mass. **1 mark**

3. A neutral atom of the isotope ${}^{13}_{6}C$ consists of:

A 6 protons, 13 neutrons and 13 electrons
B 0 protons, 13 neutrons and 13 electrons
C 13 protons, 7 neutrons and 13 electrons
D 6 protons, 7 neutrons and 6 electrons. **1 mark**

4. The ground state electron configuration of a neutral atom with an atomic number of 19 is:

A $1s^22s^22p^63s^23p^63d^1$
B $1s^22s^22p^63s^23p^7$
C $1s^22s^22p^63s^23p^64s^1$
D $1s^22s^22p^63s^23p^53d^14s^1$. **1 mark**

Extended response questions

5. Explain what nanotechnology is. Discuss two advantages and two disadvantages of nanotechnology. **6 marks**

6. (a) Write the electron configuration of each of the following in their ground states.

(i) Argon atom, $Z = 18$
(ii) Calcium ion, Ca^{2+}, $Z = 20$

(b) By referring to the electron configuration of the argon atom, explain how its group and period on the periodic table can be determined. **4 marks**

CHAPTER

2

The periodic table

YOU WILL EXAMINE:

- the importance of the periodic table as a critical tool for chemists and researchers
- the main features of the modern periodic table, including 118 naturally occurring and artificially synthesised elements, the metals and non-metals, the main-group elements and the transition elements
- groups and periods and the *s*, *p*, *d* and *f* blocks
- increasing atomic number and the physical and chemical properties of groups of elements
- patterns in electron configurations
- trends in properties of elements across and down the periodic table including electronegativity, first ionisation energy, metallic/non-metallic character and reactivity.

> Man is a microcosm, or a little world, because he is an extract from all the stars and planets of the whole firmament, from the earth and the elements; and so he is their quintessence.
>
> *Paracelsus*

Stars are formed in the pillars of creation in the Eagle nebula. This is also where some of the elements found in the periodic table are created. The pillar on the left is about three light-years long and contains massive clouds of dense, cool gas and dust. Young stars often emit huge jets of gas. The Hubble telescope has been able to show the existence of these jets in unprecedented detail. With a new infra-red sensor, scientists hope to see inside such clouds of gas and probe even deeper into the universe.

Classification of the elements

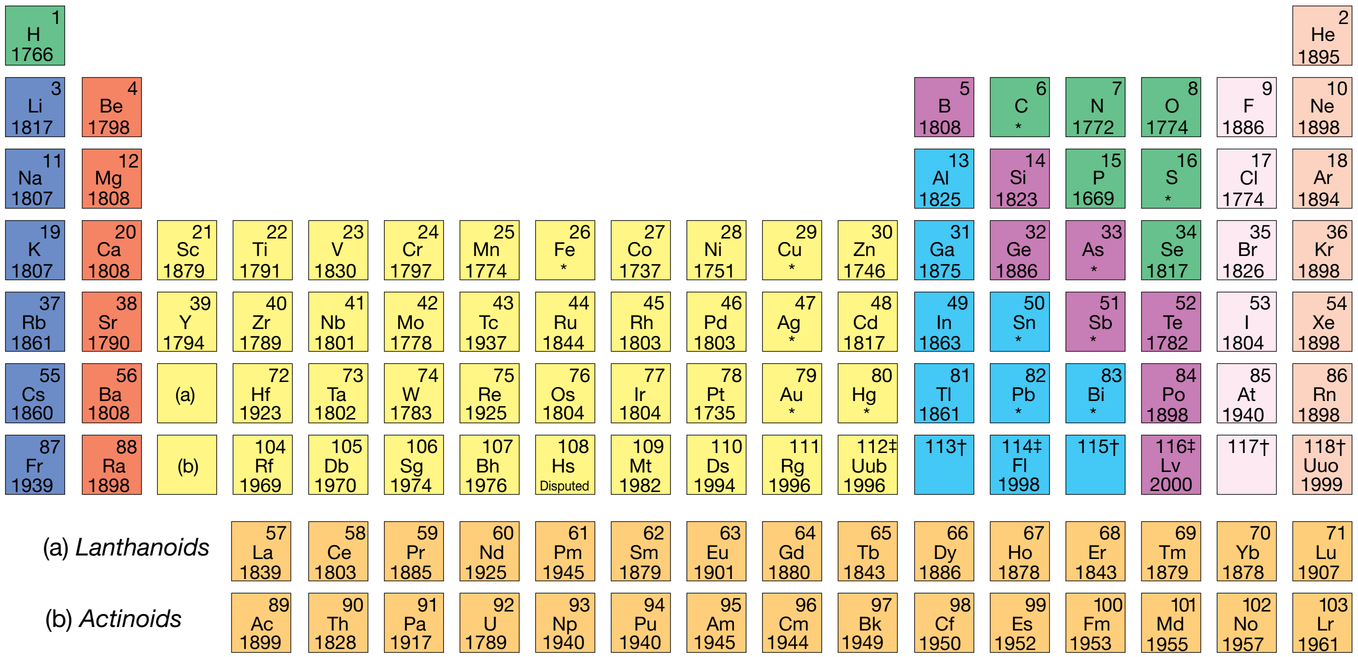
Chemists, like anyone else, would find it difficult to remember details of a particular element, given that there are over 100 known elements. Fortunately, a system of grouping or classifying these elements has evolved over the last 200 years. This system of classification is the arrangement of elements, based on similar chemical properties, in a table known as the periodic table of the elements. The term ‘periodic’ suggests that the elements show regular patterns in their chemical properties. Chemists need to remember the properties of only a handful of typical elements, and the rest fall into groups or families with similar properties.

The periodic table of the elements has become one of the most important icons in science today. A chart of this table hangs on the wall of almost every classroom and chemical laboratory in schools, universities and research institutions around the world. It is a single document that consolidates much of our knowledge of chemistry and is a vital tool for modern chemists. The modern periodic table now consists of 118 known elements. Elements up to and including uranium are naturally occurring. All the elements beyond uranium have been synthesised by chemists since 1940. They are all radioactive elements and are called the transuranium elements. These elements do not occur in nature but have been synthesised in nuclear reactors, which accelerate bombarding particles to very high speeds and shoot them at very heavy target nuclei.

The periodic table shows the patterns and properties of the 118 elements that have been discovered and synthesised. Element 117, temporarily named ununseptium, has been synthesised by scientists but not officially confirmed by the International Union of Pure and Applied Chemistry (IUPAC). IUPAC is an international scientific organisation that aims to promote the benefits of chemistry and set standards for units, names and formulas.

Revision question

1. **Find out why element 117 is called unununseptium and how it was synthesised.**



1 H 1766																	2 He 1895
3 Li 1817	4 Be 1798											5 B 1808	6 C *	7 N 1772	8 O 1774	9 F 1886	10 Ne 1898
11 Na 1807	12 Mg 1808											13 Al 1825	14 Si 1823	15 P 1669	16 S *	17 Cl 1774	18 Ar 1894
19 K 1807	20 Ca 1808	21 Sc 1879	22 Ti 1791	23 V 1830	24 Cr 1797	25 Mn 1774	26 Fe *	27 Co 1737	28 Ni 1751	29 Cu *	30 Zn 1746	31 Ga 1875	32 Ge 1886	33 As *	34 Se 1817	35 Br 1826	36 Kr 1898
37 Rb 1861	38 Sr 1790	39 Y 1794	40 Zr 1789	41 Nb 1801	42 Mo 1778	43 Tc 1937	44 Ru 1844	45 Rh 1803	46 Pd 1803	47 Ag *	48 Cd 1817	49 In 1863	50 Sn *	51 Sb *	52 Te 1782	53 I 1804	54 Xe 1898
55 Cs 1860	56 Ba 1808	(a)	72 Hf 1923	73 Ta 1802	74 W 1783	75 Re 1925	76 Os 1804	77 Ir 1804	78 Pt 1735	79 Au *	80 Hg *	81 Tl 1861	82 Pb *	83 Bi *	84 Po 1898	85 At 1940	86 Rn 1898
87 Fr 1939	88 Ra 1898	(b)	104 Rf 1969	105 Db 1970	106 Sg 1974	107 Bh 1976	108 Hs Disputed	109 Mt 1982	110 Ds 1994	111 Rg 1996	112‡ Uub 1996	113†	114‡ Fl 1998	115†	116‡ Lv 2000	117†	118† Uuo 1999
(a) *Lanthanoids*			57 La 1839	58 Ce 1803	59 Pr 1885	60 Nd 1925	61 Pm 1945	62 Sm 1879	63 Eu 1901	64 Gd 1880	65 Tb 1843	66 Dy 1886	67 Ho 1878	68 Er 1843	69 Tm 1879	70 Yb 1878	71 Lu 1907
(b) *Actinoids*			89 Ac 1899	90 Th 1828	91 Pa 1917	92 U 1789	93 Np 1940	94 Pu 1940	95 Am 1945	96 Cm 1944	97 Bk 1949	98 Cf 1950	99 Es 1952	100 Fm 1953	101 Md 1955	102 No 1957	103 Lr 1961

Periodic table showing date of discovery of elements

Element groups (families)		
Alkali earth	Alkaline earth	Transition metals
Rare earth	Other metals	Metalloids
Non-metals	Halogens	Noble gases

* Element known in ancient times

† Undiscovered element

‡ Element not yet confirmed by IUPAC

In 1869, after painstakingly collecting and collating many chemical facts, Dmitri Mendeleev (1834–1907) noticed the existence of 'groups' of different elements with similar chemical properties. He then produced a periodic table upon which the modern classification of elements is based.

Dmitri Mendeleev

Most of the credit for arranging the elements in a periodic table is given to a Russian chemist, Dmitri Mendeleev (1834–1907). Mendeleev spent many years collecting and sorting information about each of the 63 elements known at the time and constructed a set of data cards (one data card for each element). On each card he noted the atomic mass and other properties of the element and its compounds.

Mendeleev noticed that there were groups of different elements that had similar chemical properties. He was able to arrange the elements into a periodic table according to increasing order of relative atomic mass and the periodicity of their properties. He even left gaps for elements that, he said, had not yet been discovered; and he listed separately some 'odd' elements (for example, cobalt and nickel) whose properties did not fit in with those of the main group in which they were located. He then proposed a periodic law that stated that: The properties of the elements are periodic functions of their relative atomic masses. This means that, if the elements are arranged in order of increasing atomic mass, similar physical and chemical properties occur at regular intervals.

TABLE 2.1 A form of Mendeleev's table published in 1871*

Row	Group I	Group II	Group III	Group IV	Group V	Group VI	Group VII	Group VIII
1	H = 1							
2	Li = 7	Be = 9.4	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27.3	Si = 28	P = 31	S = 32	Cl = 35.5	
4	K = 39	Ca = 40	= 44	Ti = 48	V = 51	Cr = 52	Mn = 55	Fe = 56, Co = 59, Ni = 59, Cu = 63
5		Zn = 65	= 68	= 72	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	Yt = 88	Zr = 90	Nb = 94	Mo = 96	= 100	Ru = 104, Rh = 104, Pd = 106, Ag = 108
7		Cd = 112	In = 113	Sn = 118	Sb = 122	Te = 125	I = 127	
8	Cs = 133	Ba = 137	Di = 138	Ce = 140				
9								
10			Er = 178	La = 180	Ta = 182	W = 184		Os = 195, Ir = 197, Pt = 198, Au = 199
11		Hg = 200	Tl = 204	Pb = 207	Bi = 208			
12				Th = 231		U = 240		

*In particular, Mendeleev used the formulas of compounds to classify the elements. For example, he saw that the group 1 metals have chlorides with the general formula MCl and oxides with the general formula M_2O. Note the spaces left for elements with atomic weights of 44, 68, 72 and 100.

Mendeleev's table is organised in a similar way to the one used today, showing vertical columns (called groups) of elements with similar physical and chemical properties. Elements in horizontal rows (called periods) are arranged in order of increasing atomic mass.

Mendeleev arranged his periodic table in order of atomic mass. He organised the known elements into groups and periods.

If a theory is to be useful, it should not only explain the known facts but also enable new predictions to be made. The modern periodic table is largely attributed to Mendeleev. He went a step further than simply arranging the known elements of the time into a systematic order. The table enabled predictions to be made about the properties of other, as yet undiscovered, elements. The accuracy of Mendeleev's periodic table was borne out by the later discovery of elements to fill the gaps that had originally been left in the table. For example, the existence of an element with the properties of the element now known as gallium was predicted by Mendeleev in 1871. He called this as-yet undiscovered element 'eka-aluminium' (meaning 'first after aluminium', *eka* being a Sanskrit word meaning 'one'). Table 2.2 shows the properties that Mendeleev had predicted for gallium and the actual properties of gallium discovered four years after his prediction of its existence.

TABLE 2.2 Comparison of predicted and actual properties of gallium

Property	Properties of gallium predicted by Mendeleev in 1871	Actual properties of gallium, discovered in 1875
relative atomic mass	68	69.9
density (g cm^{-3})	5.9	5.94
melting point	low	30 °C
solubility in acids and bases	dissolves slowly	dissolves slowly

Revision questions

2. **Explain how Mendeleev was able to predict the properties of the element gallium even though gallium had not been discovered.**
3. **What are the similarities and differences between Mendeleev's periodic table and the current periodic table?**

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Unit 1
AOS 1
Topic 2
Concept 1

Organisation of the periodic table by electron configuration
Summary screen and practice questions

Organisation of the periodic table

In the modern periodic table, all the chemical elements are arranged in order of increasing atomic number (the number of protons in a nucleus of an atom of that element). The elements are arranged in rows and columns in relation to their electronic structures and also their chemical properties. Modern understanding of the periodic table arose from the recognition of four principles:

1. Atomic number, rather than atomic mass, was the basic property that determined the order of the elements in the periodic table.
2. Repeating patterns of electron configuration were observed when the electrons around the nucleus of an atom were arranged in order of increasing energy level.
3. The arrangement of the outer-shell electrons was most important in determining the chemical properties of an element.
4. The periodic recurrence of similar properties was seen to result from the periodic change in the electronic structure.

Periods and groups in the periodic table

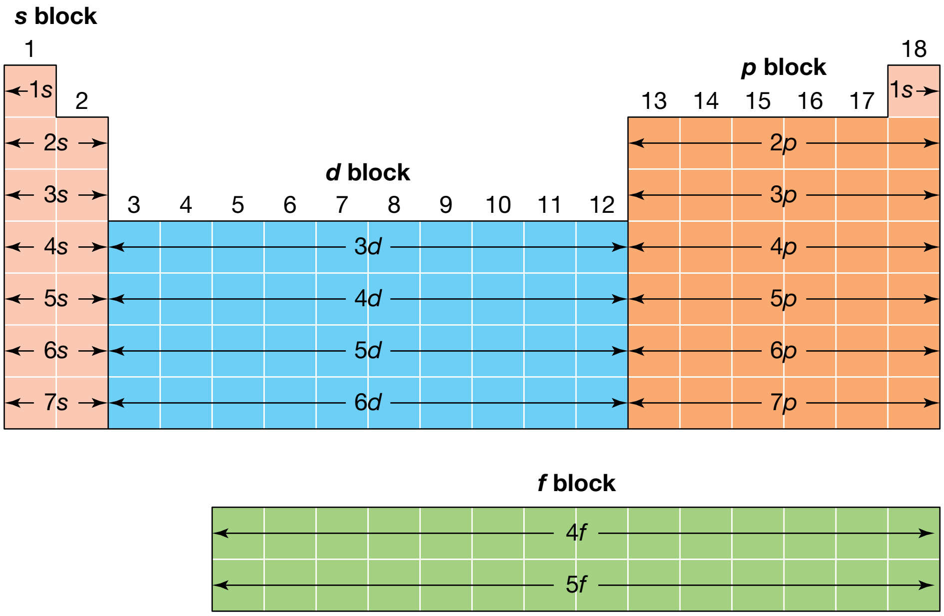
The seven horizontal rows in the periodic table are called periods. Each period corresponds to the filling of a shell. The location of an element in a period tells you the number of shells each atom of that element has. Elements in the third period, for example, have three shells.

Vertical columns of elements are called groups. For example, all atoms of group 2 elements have 2 electrons in their outer shell. Groups 1, 2 and 13–18 are also referred to as main-group elements. These include not only the most abundant elements on Earth but also the most abundant elements in the universe.

Modern periodic tables have groups numbered 1 to 18; older versions are often numbered with roman numerals I to VIII.

Electron configuration and blocks of elements in the periodic table

Elements in the periodic table can be divided into four main blocks according to their electron configurations.



Each block of the periodic table corresponds to a particular subshell.

The elements in group 1 and group 2 form a block of reactive metals and are known as the *s*-block elements. These elements have their outermost electrons in the *s* subshell. Group 1 elements have outer shells of s^1 and group 2 elements have outer shells of s^2. Helium is a group 2 element with a filled *s* subshell of the innermost *K* shell of the atom, rendering it unreactive. It is often grouped with the group 18 noble elements with similar properties.

The *s*, *p*, *d* and *f* blocks in the periodic table correspond to that subshell being filled. The group and period of an element can be found from its electron configuration.

The elements in groups 13 to 18 form the *p* block, in which elements have their outermost electrons in the *p* subshells. These elements have outershell electron configurations of s^2p^1 to s^2p^6.

The *d*-block elements, from group 3 to group 12 are the transition metals or transition elements. These elements have their *d* subshells progressively filled only after their next *s* subshell has been filled. Their outershell electron configurations are d^1s^2 to $d^{10}s^2$.

The lanthanoids and actinoids form a block of elements within the transition metals and are sometimes known as the inner transition elements. These elements form the *f* block of the periodic table and have their *f* subshells progressively filled.

	Period 1		
		1 Hydrogen **H** **1.01** 1	2 Helium **He** **4.00** 2

Key
Atomic number
Name
Symbol
Relative atomic mass
Electron configuration

Alkali metals → Group 1

Alkaline earth metals → Group 2

Transition metals → Group 3 – Group 9

	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9
Period 2	3 Lithium **Li** **6.94** 2,1	4 Beryllium **Be** **9.01** 2,2							
Period 3	11 Sodium **Na** **22.99** 2,8,1	12 Magnesium **Mg** **24.31** 2,8,2							
Period 4	19 Potassium **K** **39.10** 2,8,8,1	20 Calcium **Ca** **40.08** 2,8,8,2	21 Scandium **Sc** **44.96** 2,8,9,2	22 Titanium **Ti** **47.90** 2,8,10,2	23 Vanadium **V** **50.94** 2,8,11,2	24 Chromium **Cr** **52.00** 2,8,13,1	25 Manganese **Mn** **54.94** 2,8,13,2	26 Iron **Fe** **55.85** 2,8,14,2	27 Cobalt **Co** **58.93** 2,8,15,2
Period 5	37 Rubidium **Rb** **85.47** 2,8,18,8,1	38 Strontium **Sr** **87.62** 2,8,18,8,2	39 Yttrium **Y** **88.91** 2,8,18,9,2	40 Zirconium **Zr** **91.22** 2,8,18,10,2	41 Niobium **Nb** **92.91** 2,8,18,12,1	42 Molybdenum **Mo** **95.94** 2,8,18,13,1	43 Technetium **Tc** **98.91** 2,8,18,13,2	44 Ruthenium **Ru** **101.07** 2,8,18,15,1	45 Rhodium **Rh** **102.91** 2,8,18,18,1
Period 6	55 Caesium **Cs** **132.91** 2,8,18,18,8,1	56 Barium **Ba** **137.34** 2,8,18,18,8,2	57–71 Lanthanoids	72 Hafnium **Hf** **178.49** 2,8,18,32, 10,2	73 Tantalum **Ta** **180.95** 2,8,18,32, 11,2	74 Tungsten **W** **183.85** 2,8,18,32, 12,2	75 Rhenium **Re** **186.2** 2,8,18,32, 13,2	76 Osmium **Os** **190.2** 2,8,18,32, 14,2	77 Iridium **Ir** **192.22** 2,8,18,32,17
Period 7	87 Francium **Fr** **(223)** 2,8,18,32,18, 8,1	88 Radium **Ra** **(226)** 2,8,18,32,18, 8,2	89–103 Actinoids	104 Rutherfordium **Rf** **(261)** 2,8,18,32,32, 10,2	105 Dubnium **Db** **(262)** 2,8,18,32,32, 11,2	106 Seaborgium **Sg** **(266)** 2,8,18,32,32, 12,2	107 Bohrium **Bh** **(264)** 2,8,18,32,32, 13,2	108 Hassium **Hs** **(269)** 2,8,18,32,32, 14,2	109 Meitnerium **Mt** **(268)** 2,8,18,32,32, 15,2

The period number refers to the number of the outermost shell containing electrons.

Note that, although elements 113, 115 and 117 are not known, they have been included in the periodic table in their expected positions.

Lanthanoids

57 Lanthanum **La** **138.91** 2,8,18,18,9,2	58 Cerium **Ce** **140.12** 2,8,18,20,8,2	59 Praseo-dymium **Pr** **140.91** 2,8,18,21,8,2	60 Neodymium **Nd** **144.24** 2,8,18,22,8,2	61 Promethium **Pm** **(145)** 2,8,18,23,8,2	62 Samarium **Sm** **150.4** 2,8,18,24,8,2

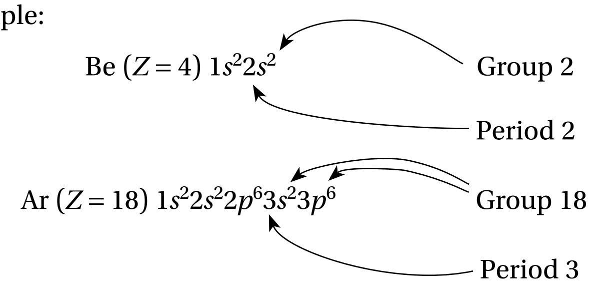
Actinoids

89 Actinium **Ac** **(227)** 2,8,18,32,18, 9,2	90 Thorium **Th** **232.04** 2,8,18,32,18, 10,2	91 Protactinium **Pa** **231.04** 2,8,18,32,20, 9,2	92 Uranium **U** **238.03** 2,8,18,32,21, 9,2	93 Neptunium **Np** **237.05** 2,8,18,32,22, 9,2	94 Plutonium **Pu** **(244)** 2,8,18,32,23, 9,2

This periodic table is based on the IUPAC periodic table and shows relative atomic masses to 2 decimal places. It is common to use only 1 decimal place in calculations (see chapter 5). Relative atomic mass (A_r) is based on the carbon-12 atom, the most common isotope of carbon. This isotope is assigned a mass of exactly 12. On this scale, 1 is therefore equal to 1/12 of the mass of a carbon-12 atom. Values in brackets are for the most stable or best-known isotopes.

The period corresponds to the *number of the outer shell* of the atom. To find the group number, first add the total number of electrons to find the atomic number, and then look it up the periodic table.

For example:



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Electron configuration

							Halogens	Noble gases
			Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
			5 Boron **B** **10.81** 2,3	6 Carbon **C** **12.01** 2,4	7 Nitrogen **N** **14.01** 2,5	8 Oxygen **O** **16.00** 2,6	9 Fluorine **F** **19.00** 2,7	10 Neon **Ne** **20.18** 2,8
Group 10	Group 11	Group 12	13 Aluminium **Al** **26.98** 2,8,3	14 Silicon **Si** **28.09** 2,8,4	15 Phosphorus **P** **30.97** 2,8,5	16 Sulfur **S** **32.06** 2,8,6	17 Chlorine **Cl** **35.45** 2,8,7	18 Argon **Ar** **39.95** 2,8,8
28 Nickel **Ni** **58.71** 2,8,16,2	29 Copper **Cu** **63.55** 2,8,18,1	30 Zinc **Zn** **65.38** 2,8,18,2	31 Gallium **Ga** **69.72** 2,8,18,3	32 Germanium **Ge** **72.59** 2,8,18,4	33 Arsenic **As** **74.92** 2,8,18,5	34 Selenium **Se** **78.96** 2,8,18,6	35 Bromine **Br** **79.90** 2,8,18,7	36 Krypton **Kr** **83.80** 2,8,18,8
46 Palladium **Pd** **106.4** 2,8,18,18	47 Silver **Ag** **107.87** 2,8,18,18,1	48 Cadmium **Cd** **112.40** 2,8,18,18,2	49 Indium **In** **114.82** 2,8,18,18,3	50 Tin **Sn** **118.69** 2,8,18,18,4	51 Antimony **Sb** **121.75** 2,8,18,18,5	52 Tellurium **Te** **127.60** 2,8,18,18,6	53 Iodine **I** **126.90** 2,8,18,18,7	54 Xenon **Xe** **131.30** 2,8,18,18,8
78 Platinum **Pt** **195.09** 2,8,18,32,17,1	79 Gold **Au** **196.97** 2,8,18,32,18,1	80 Mercury **Hg** **200.59** 2,8,18,32,18,2	81 Thallium **Tl** **204.37** 2,8,18,32,18,3	82 Lead **Pb** **207.2** 2,8,18,32,18,4	83 Bismuth **Bi** **208.98** 2,8,18,32,18,5	84 Polonium **Po** **(209)** 2,8,18,32,18,6	85 Astatine **At** **(210)** 2,8,18,32,18,7	86 Radon **Rn** **(222)** 2,8,18,32,18,8
110 Darmstadtium **Ds** **(271)** 2,8,18,32,32,17,1	111 Roentgenium **Rg** **(272)** 2,8,18,32,32,18,1	112 Ununbium **Uub** **(272)** 2,8,18,32,32,18,2	113 **Uut**	114 Flerovium **Fl** **(285)** 2,8,18,32,32,18,4	115 **Uup**	116 Livermorium **Lv** **(289)** 2,8,18,32,32,18,6	117 **Uus**	118 Ununoctium **Uuo** **(293)** 2,8,18,32,32,18,8

Metals ← → **Non-metals**

63 Europium **Eu** **151.96** 2,8,18,25,8,2	64 Gadolinium **Gd** **157.25** 2,8,18,25,9,2	65 Terbium **Tb** **158.93** 2,8,18,27,8,2	66 Dysprosium **Dy** **162.50** 2,8,18,28,8,2	67 Holmium **Ho** **164.93** 2,8,18,29,8,2	68 Erbium **Er** **167.26** 2,8,18,30,8,2	69 Thulium **Tm** **168.93** 2,8,18,31,8,2	70 Ytterbium **Yb** **173.04** 2,8,18,32,8,2	71 Lutetium **Lu** **174.97** 2,8,18,32,9,2

95 Americium **Am** **(243)** 2,8,18,32,24,9,2	96 Curium **Cm** **(247)** 2,8,18,32,25,9,2	97 Berkelium **Bk** **(247)** 2,8,18,32,26,9,2	98 Californium **Cf** **(251)** 2,8,18,32,27,9,2	99 Einsteinium **Es** **(254)** 2,8,18,32,28,9,2	100 Fermium **Fm** **(257)** 2,8,18,32,29,9,2	101 Mendelevium **Md** **(258)** 2,8,18,32,30,9,2	102 Nobelium **No** **(255)** 2,8,18,32,31,9,2	103 Lawrencium **Lr** **(256)** 2,8,18,32,32,9,2

Revision questions

4. **Identify the period and group to which each of the following elements belongs.**
- **(a)** $1s^22s^22p^4$
- **(b)** $1s^22s^1$
- **(c)** $1s^22s^22p^63s^23p^63d^64s^2$
- **(d)** $K\,L\,M\,4s^24p^6$
- **(e)** $K\,L\,3s^23p^63d^14s^2$
- **(f)** $1s^22s^22p^63s^23p^63d^{10}4s^2$

5. **Name the elements that you have identified in question 4.**

6. **Write the ground-state electron configuration for elements with the following atomic numbers, and state the group and period of each element.**

(a) **8**

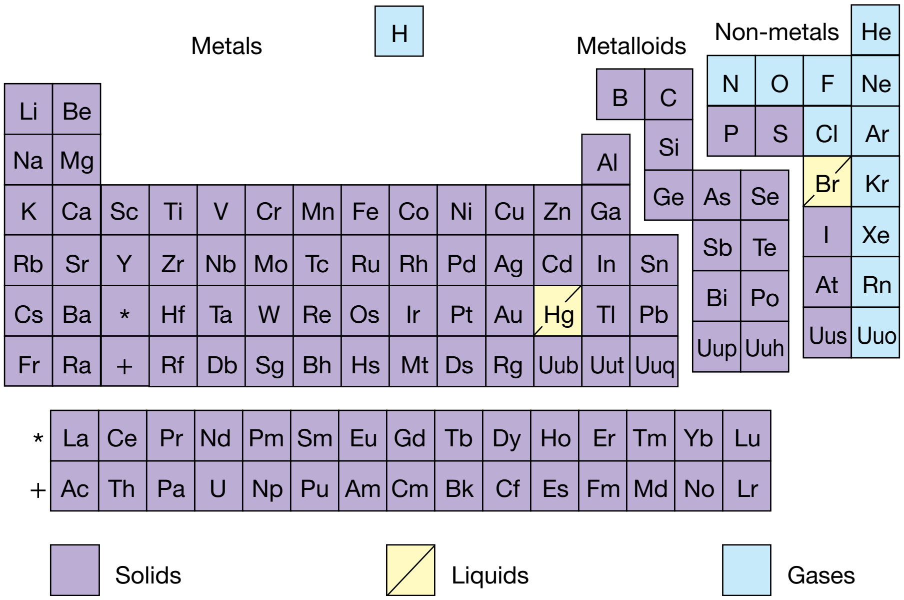
(b) **17**

(c) **20**

(d) **35**

Metals and non-metals in the periodic table

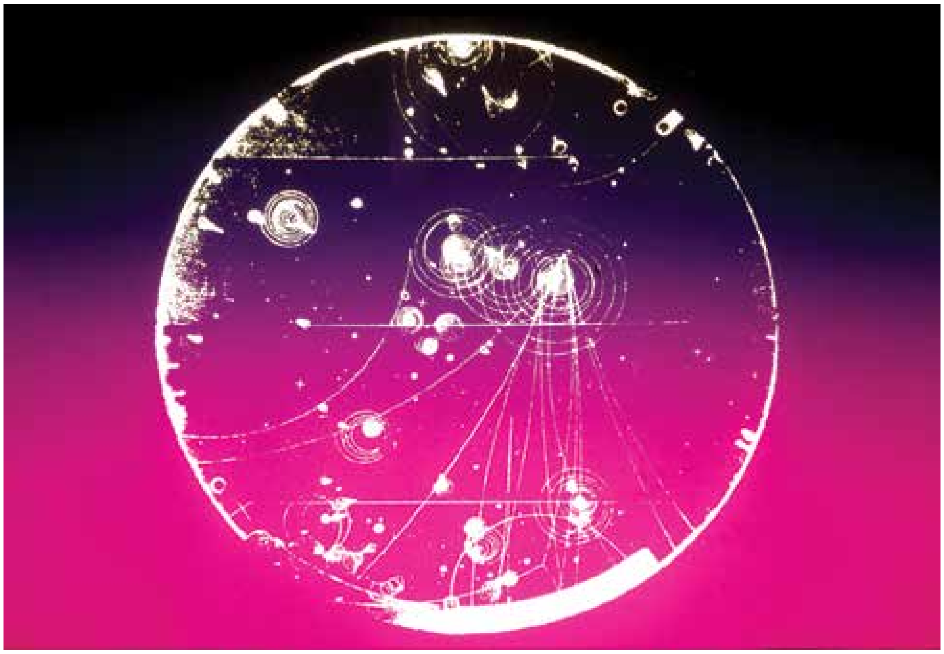
Elements may be classified as **metals** or **non-metals**. In the periodic table, the metals are found towards the left side and the non-metals are found towards the right side, as shown in the figure below.



Metals, non-metals and metalloids in the periodic table

Metals are mostly found on the left side of the periodic table, and non-metals on the right, with the metalloids in between.

Most elements are metals, although some elements show both metallic and non-metallic characteristics. These elements are known as *metalloids*.



Particle tracks like this one are part of the evidence to show that a new particle has been produced. Darmstadtium, Ds, element 110, and roentgenium (pronounced rent-ghen-i-em), Rg, element 111, were first discovered in 1994. Both were produced in a heavy ion accelerator from the fusion of lead and other elements. Only a few atoms were produced of these new species before they radioactively decayed; they are expected to be metallic in character.

TABLE 2.3 General properties of metals and non-metals

Metals	Non-metals
high melting and boiling points	low melting and boiling points
good conductors of heat and electricity	poor conductors of heat and electricity
opaque	transparent in a thin sheet
shiny appearance	dull colour
ductile and malleable	brittle when solid
strong	weak
form positive ions	form negative ions

The elements in groups 3 to 12 include some familiar metals, such as iron, which is commonly used in construction, copper, which is particularly useful for electrical wiring, and gold, which is used in jewellery. These groups are known as the transition elements. Transition elements contain atoms with filling *d* subshells. These elements are metals and have the properties listed in table 2.3 but are less reactive than *s*-block metals, and most form coloured compounds. In addition, most can form ions with different charges; for example copper can form Cu^+ ions and Cu^{2+} ions.

Revision question

7. (a) Zinc is an element that forms mostly colourless compounds. Write the subshell configuration for zinc ($Z = 30$).

(b) State two reasons why zinc does not fit into the usual pattern for transition elements.

Patterns in the periodic table

Periodic trends in atomic size

Atomic size decreases across the periodic table and increases down the groups.

Since an atom does not have a sharply defined boundary to set the limit of its size, the radius of an atom cannot be measured directly. However, several methods are available to gain an estimate of the relative sizes of atoms.

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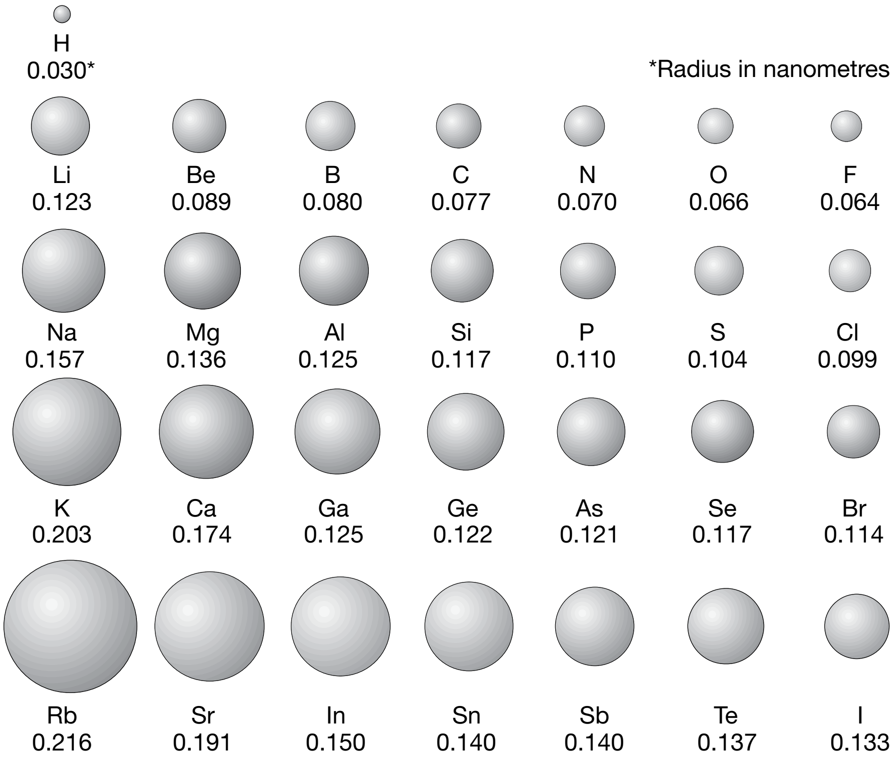
Interactivity
Periodic table trends
int-6352



Atomic radii (in nanometres) of selected elements. Atomic radii decrease across a period but increase down the group.

Atomic size generally increases down a group of the periodic table. Going down a group, electrons are added to successively higher energy levels, or main shells, further out from the nucleus. As the number of positive charges in the nucleus also increases as you go down a group, the nuclear charge (attraction of positive charges in the nucleus to the electrons) increases. The inner electrons, however, create a 'shielding' effect, thereby decreasing the pull of the nucleus on the outermost electrons.

Atomic size generally decreases from left to right across a period. Across a period, each atom maintains the same number of main shells. Each element has one proton and one electron more than the preceding element. The electrons are added to the same main shell so the effect of the increasing nuclear charge on the outermost electrons is to pull them closer to the nucleus. Atomic size therefore decreases.

Periodic trends in ionisation energy

When an atom gains or loses an electron it forms an ion. The energy required to remove an electron from a gaseous atom is known as the *ionisation energy*. Since the amount of energy required to do this is very small, it is more realistic to compare the amount of energy required to ionise one mole of atoms simultaneously. Therefore, the unit used is kilojoules per mole (kJ mol^{-1}).

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Unit 1
AOS 1
Topic 2
Concept 2

Trends in attraction between the nucleus and valence electrons
Summary screen and practice questions

TABLE 2.4 Ionisation energies of the first 20 elements (kJ mol^{-1}). The red letters and numbers indicate the elements and first ionisation energies for group 1.

Symbol of element	Ionisation energy (kJ mol^{-1}) First	Second	Third
H	1 312		
He	2 371	5 247	
Li	520	7 297	11 810
Be	900	1 757	14 840
B	800	2 430	3 659
C	1 086	2 352	4 619
N	1 402	2 857	4 577
O	1 314	3 391	5 301
F	1 681	3 375	6 045
Ne	2 080	3 963	6 276
Na	495.8	4 565	6 912
Mg	737.6	1 450	7 732
Al	577.4	1 816	2 744
Si	786.2	1 577	3 229
P	1 012	1 896	2 910
S	999.6	2 260	3 380
Cl	1 255	2 297	3 850
Ar	1 520	2 665	3 947
K	418.8	3 069	4 600
Ca	589.5	1 146	4 941

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Periodic table patterns

Removal of one electron results in the formation of a positive ion with a 1+ charge:

$$A(\mathrm{g}) \longrightarrow A^{+}(\mathrm{g}) + \mathrm{e}^{-}$$

The energy required to remove this first (outermost) electron is called the first ionisation energy. To remove the outermost electron from the gaseous 1+ ion:

$$A^{+}(\mathrm{g}) \longrightarrow A^{2+}(\mathrm{g}) + \mathrm{e}^{-}$$

an amount of energy called the second ionisation energy is required, and so on. Table 2.4 shows the first three ionisation energies of the first 20 elements in the periodic table.

In general, the first ionisation energy decreases moving down a group of the periodic table. Since the size of the atoms is increasing when moving down a group, the outermost electrons are further from the nucleus. The nucleus therefore does not hold these electrons as strongly, so they are more easily removed. The atom therefore has a lower ionisation energy.

The first ionisation energy generally increases as we move from left to right across a period. The nuclear charge is increasing, whereas the shielding effect is relatively constant. A greater attraction of the nucleus for the electron therefore leads to an increase in ionisation energy.

Periodic trends in electronegativity

Electronegativity generally increases across a period and decreases down a group.

The electronegativity of an element is a measure of the degree to which an atom can attract an electron to itself. This is most evident when it is chemically combined with another element. The extent of attraction is expressed in arbitrary units on the Pauling electronegativity scale.

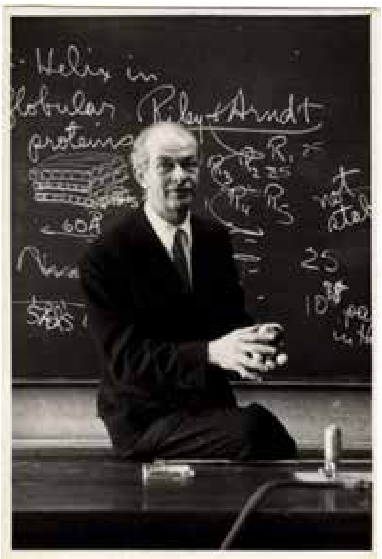
Each element except the noble gases, which do not readily form compounds, is assigned an electronegativity number. Caesium and francium, the least electronegative elements, have a value of 0.7, whereas fluorine, the most electronegative element, has a value of 4.0.

Increasing electronegativity →

Increasing electronegativity ↑

Group 1	2	13	14	15	16	17	18
H 2.1							He —
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne —
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar —
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr —
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe —
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn —
Fr 0.7	Ra 0.9						

The Pauling scale of electronegativities



Linus Carl Pauling (1901–1994) received the Nobel Prize for chemistry in 1954, and the Nobel Peace Prize in 1962.

Going across a period from left to right, the electronegativity of the main-group elements increases. This is because, as you move from one element to the next across a period, the nuclear charge increases by one unit, as one electron is added to the outer shell. As the positive charge in the nucleus increases, the atom has an increasing electron-attracting power and therefore an increasing electronegativity. Moving down a group, the electronegativity decreases because the outer electrons are further away from the nucleus and the shielding effect of the inner electrons decreases the electron-attracting power of the atom.

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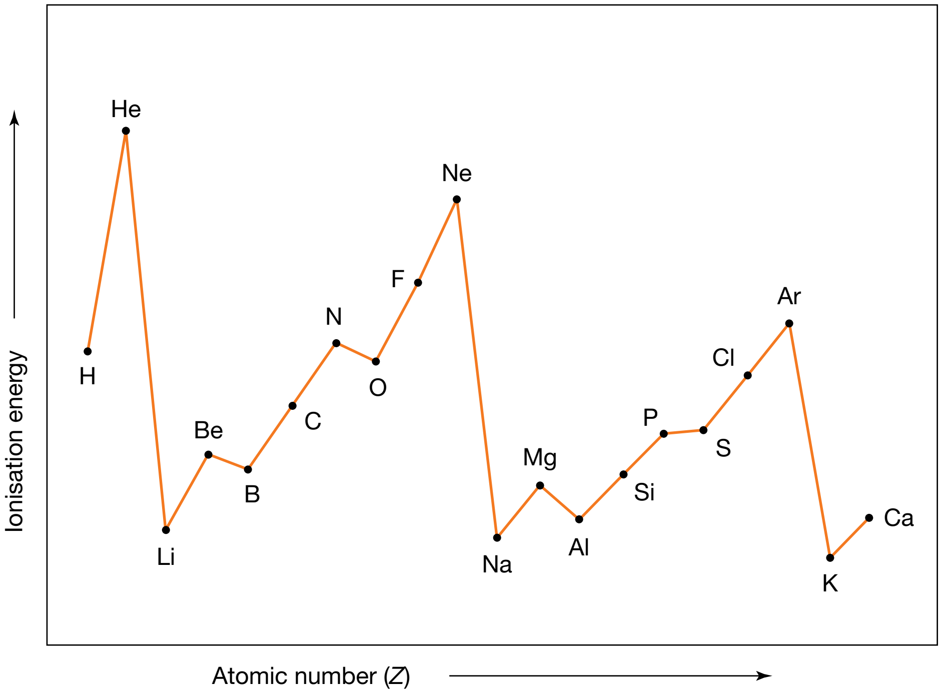
Unit 1
AOS 1
Topic 2
Concept 3

Electronegativity, ionisation energy and metal/non-metal character
Summary screen and practice questions



Trends in ionisation energy of the first 20 elements. Ionisation energy is the energy required to form an ion. It is high for noble gases such as He, Ne and Ar. Table 2.4 shows the change in ionisation energies for Li, Na and K.

Metallic character decreases across a period and increases down a group.

Periodic trends in metallic characteristics

In terms of electronic structure, the metallic characteristics of an element are determined by its ease in losing electrons. As elements move across a period, they lose their metallic characteristics. This is because, as the number of electrons in the same shell increases across a period and the nuclear charge also increases, the electrons become less easily lost to form positive ions. As elements move down a group, they become more metallic because the outer-shell electrons are further away from the nucleus (due to an increased number of shells) and are less strongly attracted. Hence, the elements lose their outer-shell electrons more easily.

TABLE 2.5 The melting points (t_m) and boiling points (t_b) of selected elements

Group 1			Group 2			Group 13			Group 14			Group 15			Group 16			Group 17		
	t_m (°C)	t_b (°C)		t_m (°C)	t_b (°C)		t_m (°C)	t_b (°C)		t_m (°C)	t_b (°C)		t_m (°C)	t_b (°C)		t_m (°C)	t_b (°C)		t_m (°C)	t_b (°C)
Li	180	1320	**Be**	1283	3000	**B**	2030	2550*	**C**	3600	4800	**N**	−210	−196	**O**	−218	−183	**F**	−220	−188
Na	98	890	**Mg**	650	1100	**Al**	660	2500	**Si**	1400	2400	**P**	44	280	**S**	113	444	**Cl**	−101	−35
K	63	770	**Ca**	850	1500	**Ga**	30	2400	**Ge**	940	2800	**As**	820	613*	**Se**	220	685	**Br**	−7	59
Rb	39	690	**Sr**	770	1400	**In**	157	2000	**Sn**	232	2300	**Sb**	630	1380	**Te**	450	1390	**I**	113	184*
Cs	29	690	**Ba**	710	1140	**Tl**	304	1460	**Pb**	327	1750	**Bi**	271	1560						
Low melting points that decrease down the group. Group 1 metals are soft and have low densities.			Much higher melting points of the first elements due to some covalent character in the bonds between atoms. Boron is a metalloid made up of a giant ring structure with covalent bonds.						High melting points that decrease down the group. C, Si and Ge exist as giant covalent networks of great hardness and high melting point.			Elements at the top of groups 15 and 16 and all the elements of group 17 exist as small covalent molecules. The melting points and boiling points increase down the groups as metallic bonding takes over in groups 15 and 16 elements.								

*sublimes

Periodic trends in oxidising and reducing strength

The oxidising strength of an element can be defined as how readily an element gains electrons. Elements that gain electrons easily are strong oxidants and are themselves reduced. Likewise, the reducing strength of an element is defined as how readily an element loses electrons.

The more readily an element gives up its electrons, the more easily it is oxidised, making it a stronger reductant (it has more reducing strength). As elements move across a period, the reducing strength decreases as the atoms give up their outershell electrons less readily, and the oxidising strength of these elements increases as elements gain electrons more readily. The extreme in oxidising/reducing behaviour of elements across the periods can be seen in examples such as sodium and potassium metals, which give up their electrons very readily, and the non-metals fluorine and chlorine, which prefer to hold on to their electrons. Hence, sodium and potassium are strong reductants while fluorine and chlorine are strong oxidants. Going down a group, the elements release their electrons more readily, making them stronger reductants (the reducing strength increases). For example, potassium is a stronger reductant than sodium and is more reactive.

An oxidant causes oxidation by gaining electrons but is itself reduced. An oxidant is also called an oxidising agent.

A reductant causes reduction by losing electrons but is itself oxidised. A reductant is also called a reducing agent.

Reducing strength decreases across a period and increases down a group. Oxidating strength increases across a period and decreases down a group.

TABLE 2.6 Trends in the periodic table

	Across a period		Down a group	
Metallic character	↓	The attraction force increases with the increasing number of protons.	↑	Electrons are less attracted to the nucleus because they are shielded by the increasing number of shells.
Atomic size	↓	The attraction force increases with the increasing number of protons and so pulls the electrons closer.	↑	Increasing number of electron shells
Reactivity	↓	Reactivity is high at the start of the period, less in the middle and more reactive at the end.	↑	Outer electrons are less strongly attracted.
Electronegativity	↑	Fluorine is the most electronegative element. The further away an element is from fluorine in the periodic table, the less electronegative it is.	↓	Increasing numbers of inner shells shield outershell electrons from the nucleus.

study on

Unit 1
AOS 1
Topic 2
Concept 4

Reactivity and interactions between elements on Earth
Summary screen and practice questions

Revision questions

8. **Account for and explain the general trends in:**
  (a) **electronegativity across the periodic table**
  (b) **the atomic radii of elements down a group**
  (c) **the reducing strength of elements across a period.**

9. **Explain why, in the periodic table, there are:**
  (a) **two elements in the first period**
  (b) **eight elements in the second period**
  (c) **no transition elements in the first three periods.**

10. **For each of the following pairs of elements, state which element is the more electronegative.**
  (a) **K, Ca**
  (b) **Be, Ca**
  (c) **Cl, Br**

Chapter review

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Unit 1 **AOS 1** **Topic 2**

The periodic table Summary screens and practice questions

Summary

- The periodic table is a method of organising all the known elements to show their similarities and differences.
- Historically, the development of the periodic table was based on the classification of elements according to their chemical and physical properties.
- Dmitri Mendeleev proposed that the properties of elements are a periodic function of atomic mass. He arranged the elements known at that time in a 'periodic table' with gaps for elements that he considered were yet to be discovered. His version of the table formed the basis of the modern periodic table that is widely used today.
- The organisation of elements on the periodic table was changed from an order of increasing atomic weight to increasing atomic number.
- Elements after uranium (atomic number 92) are artificially synthesised and radioactive. They are known as the transuranium elements.
- To date, elements up to 112 have been made, as well as elements 114 and 116; the last few elements are extremely unstable.
- Elements arranged down the same vertical columns (groups) in the modern periodic table display similar physical and chemical properties.
- Elements arranged along the same horizontal rows (periods) are placed in order of increasing atomic number.
- The main features of the periodic table are the:
  - eight main groups, which progressively fill both the *s* and *p* subshells
  - transition elements, which progressively fill the *d* subshells
  - rare earth elements, which progressively fill the *f* subshells. Elements of the *f* block are made up of the lanthanoids and the actinoids.
- The fundamental structure of the periodic table as developed by Mendeleev has remained largely unchanged, despite the discoveries and developments of new theories of atomic structure.
- Metals are mostly found on the left side and the middle of the periodic table, separated by the metalloids from the non-metals, which are found on the right.
- Atomic size decreases across the periodic table and increases down the groups.
- Electronegativity generally increases across a period and decreases down a group.
- Metallic character decreases across a period and increases down a group.
- Reducing strength decreases across a period and increases down a group. Oxidising strength increases across a period and decreases down a group.

Multiple choice questions

1. In constructing his initial forms of the periodic table, Mendeleev placed sodium and potassium in the same group because these two elements:
   A have the same atomic mass
   B have the same number of electrons
   C react violently with water
   D have metallic looks about them.
2. Which one of the following statements about the periodic table is correct?
   A All the elements listed on the periodic table are naturally occurring.
   B The periodic table can be used to predict the physical and chemical properties of undiscovered elements.
   C Elements with atomic number over 95 are radioactive.
   D The modern periodic table is arranged in order of atomic mass.
3. Which of the following elements is expected to be most similar in properties to magnesium?
   A Aluminium **C** Potassium
   B Sulfur **D** Strontium
4. Based on trends in the periodic table, which of the following elements would have the smallest atomic radius?
   A Silicon **C** Calcium
   B Fluorine **D** Beryllium
5. As you move across the periodic table, which of the following is generally true?
   A Electronegativity decreases.
   B Ionisation energy decreases.
   C Metallic characteristics decrease.
   D Atomic radius increases.
6. A trend as you go down the periodic table is that the:
   A size of atoms increases
   B metallic characteristics decrease
   C oxidising strength increases
   D electronegativity increases.

Refer to the following table to answer questions 7–10.

Element	Charge on atom or ion	Atomic number	Electron configuration of atom or ion
A	*A*		$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$
B	B^-	9	
C	C^{2-}	16	
D	D^{3+}		$1s^2 2s^2 2p^6$

7. Element *A* is in:
- **A** period 4 and is a transition element
- **B** period 4, group 1
- **C** period 4, group 13
- **D** period 3, group 13.

8. Element *B* is:
- **A** an alkali metal
- **B** an alkaline earth metal
- **C** a transition element
- **D** a halide.

9. Element *C* has a ground-state electron configuration of:
- **A** $1s^2 2s^2 2p^6 3s^2 3p^2 4s^2$
- **B** $1s^2 2s^2 2p^6 3s^2 3p^4$
- **C** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
- **D** $1s^2 2s^2 2p^6 3s^2 3p^6$.

10. The ion of element *D* has an atomic number of:
- **A** 10
- **B** 7
- **C** 13
- **D** 6.

11. A trend across a period of the periodic table is that:
- **A** metallic character increases
- **B** reducing strength decreases
- **C** electronegativity decreases
- **D** atomic size increases.

Review questions

Classification of the elements

1. Explain why the classification of elements into a periodic table may be useful to chemists.

2. Use the list of elements below to answer the questions that follow: Cl, Ni, N, Mg, I, B, Na.
- (a) Choose the two elements that have the most similar chemical properties and explain your choice.
- (b) Which of the elements is the most reactive metal? Explain why.
- (c) Which element is a transition metal?
- (d) Which of the elements is the most reactive non-metal?
- (e) Which element has the highest electronegativity? Explain why.

3. (a) Outline the contribution of Dimitri Mendeleev to the development of the first periodic table.
- (b) Mendeleev arranged his periodic table according to increasing atomic mass, yet he placed tellurium before iodine, which has a smaller atomic mass. Explain why he made this decision.

4. On what basis are the elements arranged in the modern periodic table?

5. Why is the arrangement of elements in a periodic table in order of atomic *number* rather than atomic *mass?*

Patterns in the periodic table

6. Predict which pair of elements in each set below would have the greatest similarities in chemical properties and which would have the greatest differences.
- (a) Na and Cl, Na and K, Na and Ca
- (b) Cl and I, Cl and S, Cl and Mg

7. Explain why it is easier for elements to lose electrons going down a group.

8. Consider the element phosphorus ($Z = 15$).
- (a) What is its subshell configuration?
- (b) Is it a metal or non-metal?
- (c) In what group is it?
- (d) In what period would you find it?
- (e) In what block would you find it?
- (f) Name an element that would have similar properties to phosphorus.
- (g) Would it have a higher or lower electronegativity than chlorine?
- (h) Would it have a larger or smaller atomic size than nitrogen?

9. The relative atomic mass of argon is 39.9. It is placed before potassium in the periodic table even though the relative atomic mass of potassium is 39.1. Explain why they are placed in these positions.

10. (a) Explain what is meant by the term ‘nuclear charge’.
- (b) Which has the higher nuclear charge: potassium or calcium?

11. (a) Going across the periodic table, the numbers of protons and electrons increase. Why then does the size of the atoms decrease?
- (b) Explain the trend in atomic radius going down a group.

12. Explain why the atomic number of an element is more important to chemists than its atomic mass.

13. How is an element’s outer electron configuration related to its position in the periodic table? Give three examples that illustrate your answer.

14. Table 2.7 shows the approximate date on which each of the first nineteen elements in the periodic table was isolated.

Plot a graph of the date of isolation (vertical axis) against increasing atomic number, then answer the following questions:

(a) Discuss any periodic trends evident in your graph.

(b) Predict the date of isolation of calcium (atomic number 20).

(c) Explain why four of these elements were not isolated until the end of the nineteenth century.

TABLE 2.7 Dates of isolation of elements

Element	Atomic number	Date of isolation
H	1	1766
He	2	1895
Li	3	1817
Be	4	1828
B	5	1808
C	6	known to ancients
N	7	1775
O	8	1775
F	9	1886
Ne	10	1900
Na	11	1807
Mg	12	1808
Al	13	1824
Si	14	1809
P	15	1669
S	16	known to ancients
Cl	17	1774
Ar	18	1894
K	19	1808

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Unit 1
AOS 1
Topic 2

The periodic table

Sit topic test

Exam practice questions

In a chemistry examination you will be required to answer a number of short and extended response questions.

Extended response questions

1. Why do metals generally have low electronegativities, whereas non-metals have high electronegativities? **3 marks**

2. (a) What is meant by the term 'electronegativity'?

(b) Explain how electronegativity is related to ionisation energy. **3 marks**

3. The table below represents part of the periodic table. The letters shown represent some of the elements but they may not be the symbols of those elements.

																	E
A	*T*											*M*	*X*	*L*		*G*	
												Q	*D*		*R*	*J*	
				B													

Selecting only from the elements labelled on the periodic table above, write the letter (*A*, *B*, *D*, *E*, *L*, *G*, *J*, *M*, *Q*, *R*, *T* or *X*) corresponding to the element that is:

(a) a non-metal with four electrons in its outer shell

(b) in group 14, period 2

(c) a transition element

(d) a halogen with two fully occupied shells

(e) in period 2 and has a total of three electrons

(f) a halogen

(g) an alkaline earth metal. **7 marks**

CHAPTER

3 Ionic bonding

To construct new substances such as smart materials, self-repairing materials and nanotubes, scientists and engineers need to understand the way atoms are put together and the limitations of the bonding models. The way that the atoms in a material are bonded together directly affects its properties and performance. The following three chapters investigate some of the types of bonding between the atoms that shape our world.

YOU WILL EXAMINE:

- the formation of ions
- ionic bonding as the strong bonding found in solids such as sodium chloride and magnesium oxide
- the relationship between the structure and bonding of ionic materials and their properties
- the importance and limitations of the ionic bonding model
- chemical language, symbols and formulas
- data relating to the physical properties of ionic substances
- models that represent the ionic structure
- properties of ionic materials from a knowledge of their structure and bonding
- uses of ionic compounds.

> Simplicity is the ultimate sophistication.
>
> *Leonardo da Vinci*

Giant crystals of gypsum, some up to 11 metres long, found in the 'Cave of crystals' in Mexico. Note the size of the crystals compared with the human figure in the bottom right-hand corner.

Introducing chemical bonds

Metal and non-metal atoms are joined by ionic bonding.

Collecting crystals of minerals and gemstones is one of the most popular and fastest growing hobbies in the world. These substances are admired for their perfect symmetry, brilliant colour, lustre and size. In ancient times, the early Egyptians gathered attractive crystals to use as household decorations and jewellery. Many of these crystals consist of **ionic compounds** in which metal and non-metal atoms are joined by **ionic bonding**.

We are aware that all the varied substances around us are formed from atoms combining together. But what holds atoms together once they have combined? Why do some atoms join or bond with others, while some remain as individual atoms? Why do atoms combine in specific ratios?

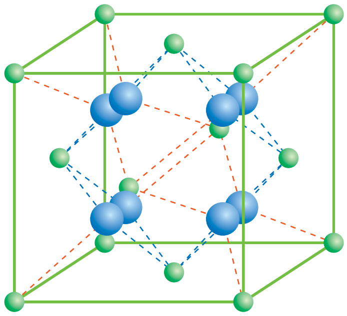
The forces between atoms in chemical bonds are electrostatic.

The answers lie in the nature of the **chemical bond**. All forces of attraction leading to chemical bonding between atoms are electrostatic in nature: that is, an attraction between positive and negative changes. Studying the electron structure of the atom helps us understand not only how atoms bond but also the characteristic properties of the substances that are formed when they do.

Most spontaneous changes that take place in nature occur in order to reach a more stable state. For example, things tend to roll downhill, but come to rest when they reach a stable position.

We seldom find free atoms in nature because atoms undergo changes in structure to become more stable. They do this by joining together, or bonding. Sometimes they bond with atoms of the same kind (for example, hydrogen gas, H_2) and sometimes with atoms of a different kind (for example, hydrochloric acid, HCl). Only the noble gases may exist as free atoms. This behaviour can be explained by an atom's electron configuration.

The mineral fluorite contains the ionic compound calcium fluoride, CaF_2, as its main component. Its beautiful crystals are often cubic in shape and appear in a variety of colours, pink being the rarest. What causes its crystal structure to be so regular?



Electron configuration and stability

We have already considered the electron configuration of atoms and have seen that the electrons may be found in certain shells around the nucleus. You may have noticed that some atoms have a complete outer shell of electrons. These atoms are helium, neon, argon, krypton and xenon, and are located in group 18 of the periodic table. They are known as the inert, or noble, gases because they rarely react or bond with any other substances. They are very stable elements and are able to exist as single atoms. Their stability is linked to the fact that each of them has a full outer shell of electrons.

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

Group 18 contains the noble gases.

Atoms with a full outer shell of electrons are very stable.

Atoms can gain or lose electrons to achieve a full outer shell.

Atoms of all other elements become more stable by gaining or losing enough electrons to achieve a complete outer shell configuration, like a noble gas. This does not mean that they have become noble gases, since they have not changed their nuclear structure (number of protons).

Atoms can become stable in one of three ways:

- by giving electrons to another atom
- by taking electrons from another atom
- by sharing electrons with another atom.

Types of chemical bond

Three types of bonding are ionic, metallic and covalent bonding.

When atoms combine to achieve more stable structures, three types of bonding are possible.

- An **ionic bond** results when metallic atoms combine with non-metallic atoms to form an ionic lattice.
- A **metallic bond** results when metallic atoms combine to form a metallic lattice.
- A **covalent bond** results when non-metallic atoms combine to form either molecules or covalent lattices.

From atoms to ions

Ions are charged atoms that have gained or lost electrons.

Cations have lost electrons and anions have gained electrons.

Atoms that gain or lose electrons to achieve more stable outer shell configurations are called **ions**. When an atom becomes an ion, it is no longer neutrally charged, since the number of electrons is not equal to the number of protons. Note that the number of protons remains the same. An atom that has lost electrons becomes *positively* charged and is called a **cation** (e.g. Na^+). An atom that has gained electrons becomes *negatively* charged and is called an **anion** (e.g. Cl^-).

TABLE 3.1 Common atoms and their ions

Atom/ion	Symbol	Number of protons	Number of electrons
sodium atom	Na	11	11
sodium cation	Na^+	11	10
chlorine atom	Cl	17	17
chloride anion	Cl^-	17	18

Electronegativity is the electron-attracting power of an atom.

Whether an atom gains or loses electrons to achieve stability depends on its electronegativity. The **electronegativity** of an atom is defined as its electron-attracting power, and is determined by a number of factors including the size of the atom, the charge on the nucleus and the number of electrons in the atom. Atoms with high electronegativities are able to attract electrons easily, whereas atoms with low electronegativities do not attract electrons as readily.

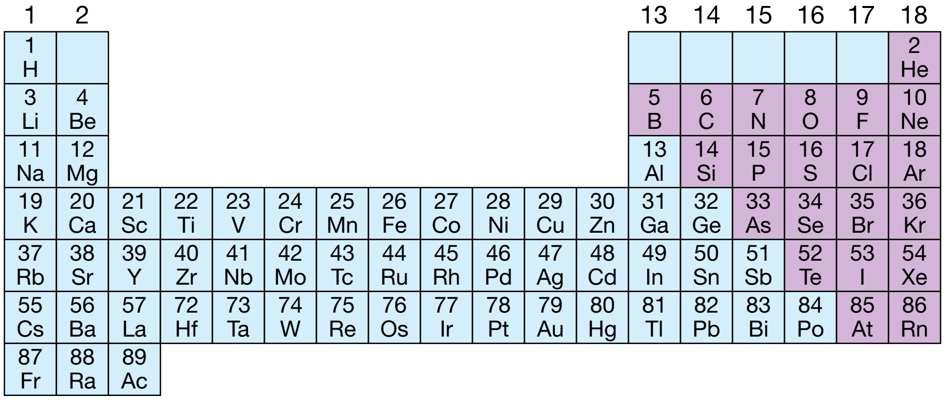
The valence shell is the outer shell of an atom.

The electrons in the outer shell of an atom are sometimes called the valence electrons. The number of outer shell electrons is related to the group number in the periodic table.

Electrovalency is the charge on an ion.

When an atom gains or loses electrons an ion is formed, in other words, the atom becomes charged.

The metallic elements are those on the left-hand side of the staircase in the figure below. These elements generally have low electronegativities. They can lose electrons to achieve a noble gas configuration in their outer shells.

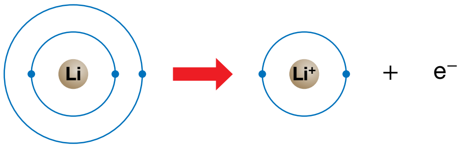


1	2											13	14	15	16	17	18
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43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

Periodic table (up to element 89) showing the division between metals and non-metals

The Li atom has:
+ + + 3 protons
– – – 3 electrons
= neutral charge.
The Li^+ ion has:
+ + + 3 protons
– – 2 electrons
= one proton unmatched
= 1+ charge.

For example, lithium is a very reactive group 1 metal with one outer shell electron and has the electron configuration 2, 1. In order to obtain the stable configuration of a full outer shell, the lone electron is lost. The electron configuration, 2, of the nearest noble gas, helium, results. Since the lithium cation has three protons but only two electrons, it has a net charge of 1+. Charges are written as superscripts above and to the right of the element symbol; thus the lithium atom is now written as Li^+. (*Note:* When an atom's net charge is 1, it is not necessary to include the numeral 1 in the superscript notation.) This process can be represented by electron shell diagrams or in the simple equation form below.



Formation of the lithium cation from the lithium atom

$$\text{lithum atom} \longrightarrow \text{lithium ion} + \text{one electron}$$
$$Li \longrightarrow Li^+ + e^-$$
$$2,1 \longrightarrow 2$$

Metal atoms lose electrons to form cations.

Note: The group 14 elements, carbon and silicon, do not form simple ions.

The group 2 and group 13 metals contain two and three valence electrons respectively. They lose their outer shell electrons to form ions with charges of 2+ and 3+ respectively.

The electron configurations for the simple ions formed by the metals in periods 2, 3 and 4 are identical to those of the closest noble gases. Each occupied energy shell contains the maximum number of electrons. Examples of their electron configurations are as follows:

Li^+	2
Na^+	2, 8
K^+	2, 8, 8.

Note that, when we name a metallic ion, we use the full name of the metal followed by the word 'ion' to distinguish it from the uncharged metal.

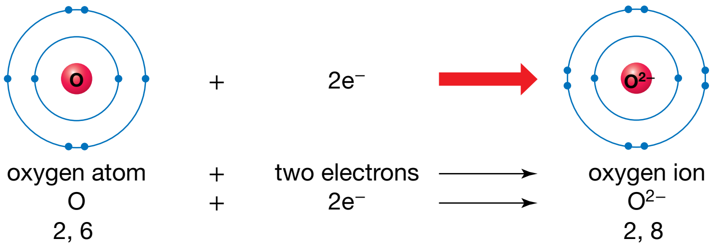
Non-metal atoms gain electrons to form anions.

Non-metallic elements are shown on the right-hand side of the purple section in the periodic table on the previous page. These elements generally have high electronegativities. They gain electrons to achieve a noble gas configuration of eight electrons in their outer shells (with the exception of hydrogen).

For example, oxygen in group 16 has six outer shell electrons and has the electron configuration 2, 6. It is too difficult to remove all six electrons to achieve a full outer shell, so the oxygen atom gains two electrons instead to become a stable anion, as shown in the following figure. An anion has more electrons than a neutral atom of the same element; here, the oxygen ion has eight protons and ten electrons, resulting in a net charge of 2–. The electron configuration of the oxygen anion is now similar to that of a neon atom: 2, 8. The oxygen atom has become an oxide ion. (It is a convention in chemistry to indicate the ions of non-metallic elements with the suffix *-ide*.)

The O atom has:
+ + + + + + + + 8 protons
– – – – – – – – 8 electrons
= neutral charge.
The O^{2-} ion has:
+ + + + + + + + 8 protons
– – – – – – – – – – 10 electrons
= 2 electrons unmatched
= 2– charge.

The oxygen atom has six valence electrons, and gains two electrons to form the oxide ion, which has a charge of 2–.



Sample problem 3.1

Write the symbol, charge and name of the ions you would expect atoms of the following elements to form:

(a) Mg (b) S.

Solution:

(a) Mg is a metal found in group 2. It has electron configuration 2, 8, 2. In order to become stable the Mg atom needs to lose two electrons. The charge of the resultant ion would therefore be 2+ and its symbol Mg^{2+}. It is called the magnesium ion.

(b) S is a non-metal found in group 16. In order to become stable the S atom needs two more electrons to complete its outer shell. The charge of the resultant ion would therefore be 2– and its symbol S^{2-}. It is called the sulfide ion.

Revision questions

1. **Explain why atoms such as Ca and Al form ions.**
2. **Use electron shell diagrams and simple equations to show how the following atoms form their corresponding ions:**
   - **(a) the metal atoms Ca and Al**
   - **(b) the non-metal atoms N and F.**
3. **For each of the following atoms, predict the charge of the ion formed, and write its name and symbol:**
   (a) Ba (b) K (c) P (d) Cl (e) S.
4. **Copy and complete the following table, predicting the general electrovalencies for period 2 and period 3 ions in each group.**

Period	Group 1	Group 2	Group 13	Group 15	Group 16	Group 17
2	1+					
3						

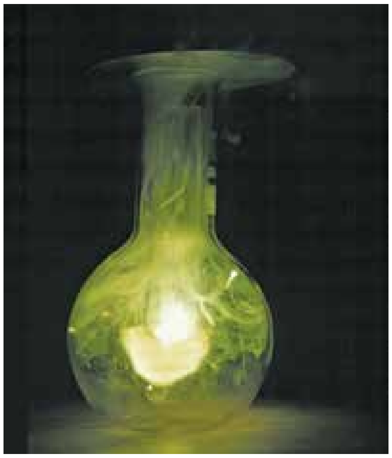
A transfer of electrons from a metal to a non-metal results in an ionic bond.

Ionic bonding

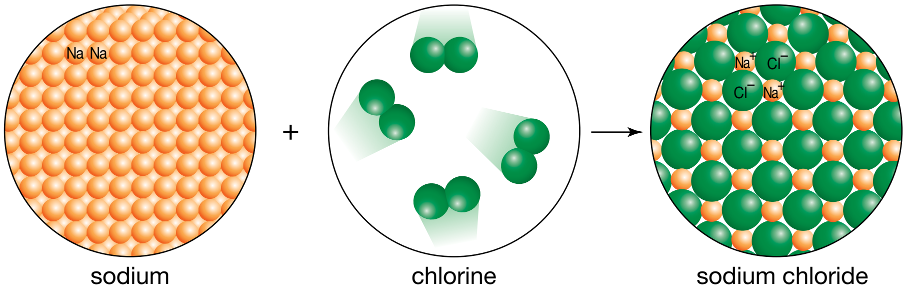
Metals and non-metals can react together because the electrons lost by the metals can be taken up by the more electronegative non-metals. This transfer of electrons results in an ionic bond between the metal cations and the non-metal anions.

For example, when a small piece of sodium metal is added to a flask containing chlorine gas, a chemical reaction occurs, as shown below. The sodium ignites, and the solid ionic compound sodium chloride forms. This involves a transfer of one electron from the sodium atom to the chlorine atom. The ways in which we can represent this process are shown in the solution to sample problem 3.2.

Sodium ions and chloride ions are oppositely charged and therefore they attract each other by **electrostatic attraction**. This strong force of attraction between positive and negative ions is what produces an ionic bond.



A small piece of sodium is melted in a metal spoon before being thrust into a flask containing chlorine. This produces an extremely violent reaction that releases a large amount of heat as the stable product, salt, is formed.



The formation of sodium chloride from solid sodium and gaseous chlorine

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Reaction of sodium with chlorine

Note that the bonded form is more stable than the unbonded form and has different properties from those of the original elements. Sodium is a very reactive metal; chlorine is a poisonous gas. Sodium chloride, however, is a non-toxic ionic compound familiar to us as table salt. The term 'salt', however, has a wider meaning in chemistry. The general name for an ionic compound, a substance formed from the reaction between a metal and a non-metal, is a **salt**.

A salt is the general name for a compound formed from a metal and a non-metal.

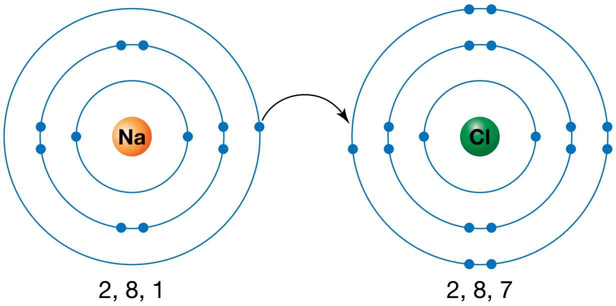
Sample problem 3.2

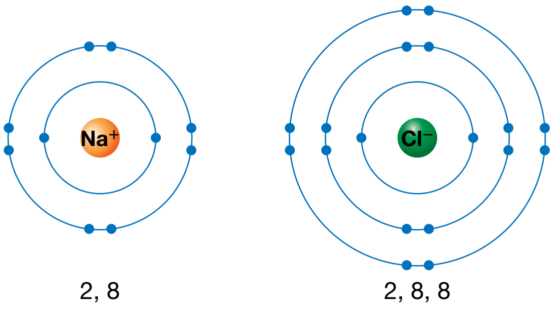
Consider the reaction between sodium and chlorine and explain it using:

(a) electron shell diagrams

(b) a simple equation.

Solution: (a) Sodium and chlorine can be represented as follows:





In bonding ionically with a chlorine atom, the sodium atom loses one electron (its entire outer shell) to become a positive ion, or cation. The chlorine atom gains one electron to complete its outer shell. It is now a negative ion, or anion. The sodium and chloride ions form the substance called sodium chloride, and can be represented as follows:

(b) $\mathrm{Na^+ + Cl^- \longrightarrow NaCl}$

$2, 8 \quad 2, 8, 8$

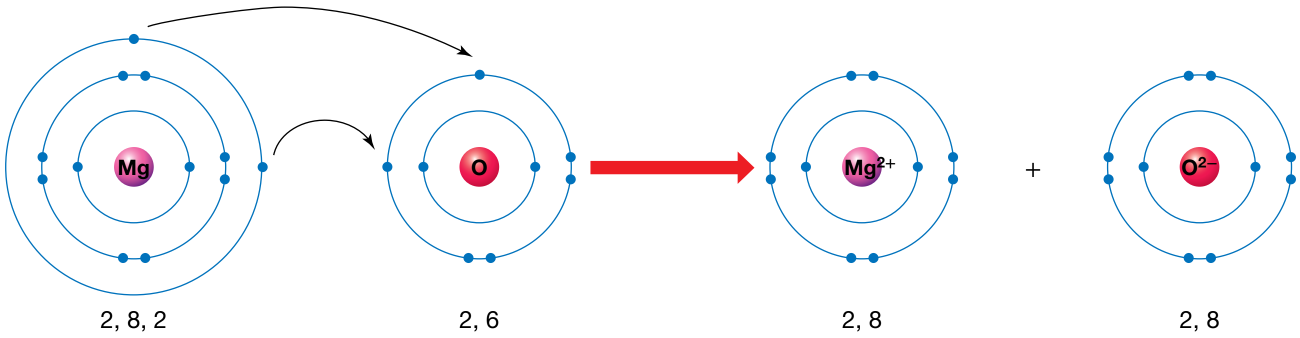
Sample problem 3.3

Consider the following ionic compounds and explain the bonds between their atoms using electron shell diagrams and simple equations:

(a) magnesium oxide

(b) magnesium fluoride.

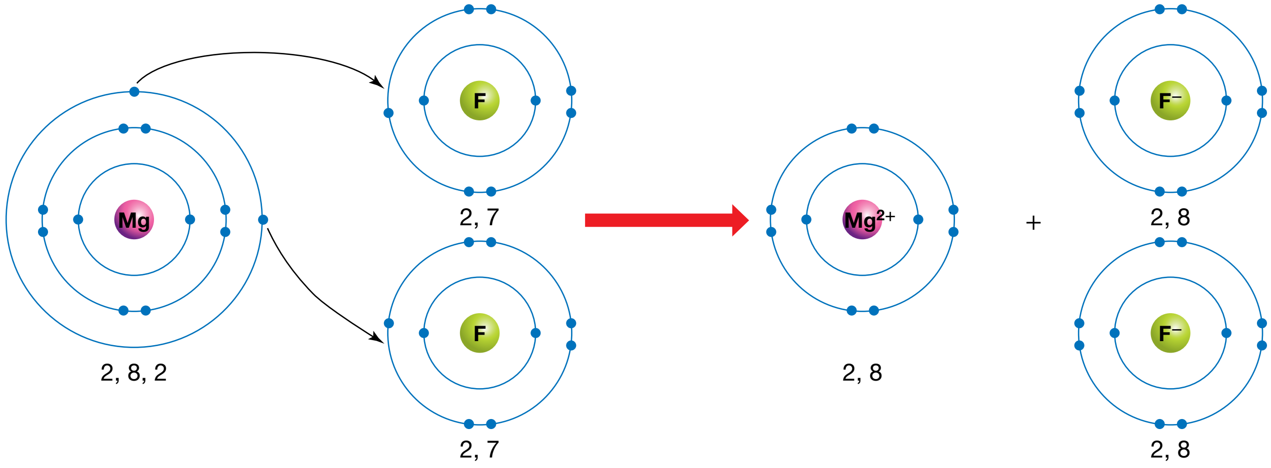
Solution: (a) A magnesium atom needs to lose two electrons to gain a stable noble gas configuration. In doing so, it becomes a doubly charged ion $\mathrm{Mg^{2+}}$. The two resulting electrons can be gained by a *single* atom, such as oxygen, which requires two electrons to complete its outer shell.



$\mathrm{Mg^{2+} + O^{2-} \longrightarrow MgO}$

$2, 8 \quad 2, 8$

(b) Each of the two electrons from the magnesium atom are accepted by *separate* fluorine atoms, which each require only one electron to fill their outer shells.



$\mathrm{Mg^{2+} + 2F^- \longrightarrow MgF_2}$

$2, 8 \quad 2, 8$

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Unit 1
AOS 1
Topic 4
Concept 2

Ionic bonding model
Summary screen and practice questions

Revision questions

5. **Draw electron shell structures and write simple equations to show the ionic bonding between the following:**
   - (a) **calcium and oxygen**
   - (b) **beryllium and chlorine**
   - (c) **lithium and fluorine**
   - (d) **aluminium and sulfur**
   - (e) **sodium and nitrogen**
   - (f) **magnesium and sulfur.**
6. **Write the electron configurations of:**
   - (a) **the cations potassium, calcium and aluminium**
   - (b) **the anions fluoride, oxide and nitride.**
7. **Name the noble gas that has the same electron configuration as the anions in question 6(b).**

Structure and properties of ionic substances

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Animation of NaCl structure

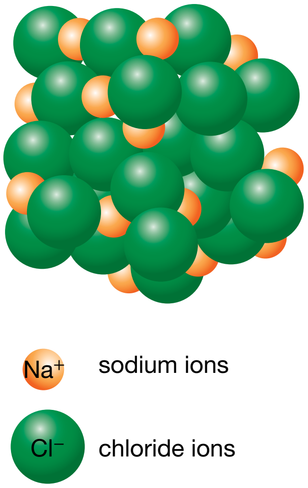
Like most ionic substances, sodium chloride is a hard, brittle crystalline solid at room temperature. It has a high melting point (800 °C). This indicates that its particles are arranged in repeating three-dimensional patterns and are strongly held together. When sodium chloride is dissolved in water or is melted, its crystal structure breaks down and in both these states it is able to conduct electricity, indicating that charged particles (ions) are free to move.

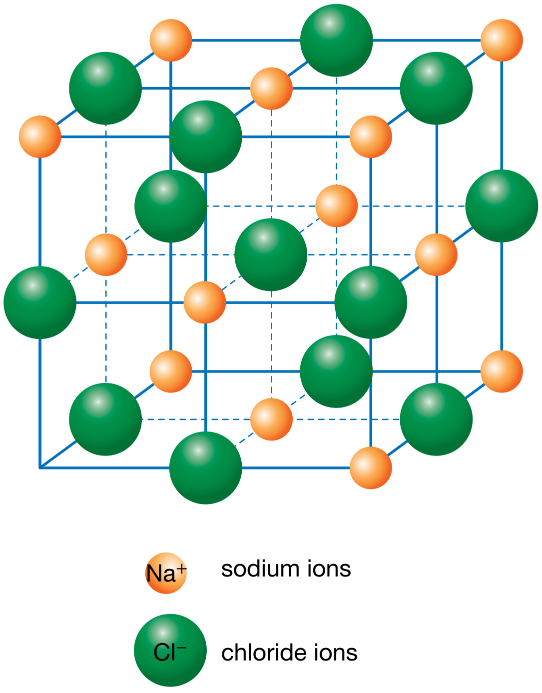
How can we explain the distinctive composition, melting point and conductivity characteristics of ionic substances? What is it about the structure of these substances that makes them behave so uniquely?

Ionic network lattices are three-dimensional arrays of cations and anions.

Ionic lattices

The major features of the current theory of the structure of ionic substances are as follows.





NaCl is represented here using both the space-filling and ball-and-stick molecular models.
- Which model is which?
- Which model do you consider more useful for the discussion of ionic network lattices?

- Ionic compounds form crystals that are composed of three-dimensional arrays of positive metal ions and negative non-metal ions.
- These arrays are called ionic network lattices and are held together by the strong electrostatic attraction — the ionic bond — between the oppositely charged ions.
- The ions are arranged in a regular repeating pattern throughout the crystal. They pack together in the way that will achieve the most stable arrangement, with oppositely charged ions as close together as possible, and similarly charged ions as far apart as possible. Each cation is thus surrounded by anions and each anion by cations.

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Interactivities
Ionic models
int-6351
Pass the salt
int-0675

- The relative numbers of cations and anions are fixed by the requirement that the solid is neutral.
- The relative sizes and numbers of the ions present determine the actual lattice structure.

Sodium chloride is a typical ionic compound. Its appearance and ionic structure are shown below. Its crystals are cubic, reflecting the regular arrangement of the ions in the lattice. Each sodium ion is surrounded by six chloride ions and each chloride ion is surrounded by six sodium ions. In this arrangement, each ion is strongly attracted to each of its neighbours. The large attractive forces result in a very stable structure.

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Unit 1	**Crystal structure**
AOS 1	Summary screen and practice questions
Topic 4	
Concept 3	

Sodium chloride crystals are cube-shaped owing to the regular cubic arrangement of the sodium and chloride ions. Each chloride ion is surrounded by six sodium ions, and each sodium ion is surrounded by six chloride ions.

The formula of the ionic substance sodium chloride is NaCl, which means that the ratio of positive to negative ions is 1:1. The formula NaCl does not represent a distinct molecule of sodium chloride but simply means that a sample of sodium chloride contains equal numbers of sodium ions and chloride ions. The formula $MgBr_2$ signifies that a sample of magnesium bromide contains twice as many bromide ions as magnesium ions.

The empirical formula is the simplest ratio of ions in the ionic lattice.

A chemical formula given for an ionic compound is called an empirical formula. This refers to the simplest ratio of ions in the lattice.

Revision question

8. Magnesium chloride, $MgCl_2$, and potassium chloride, KCl, are typical ionic compounds.
  (a) Describe how their ionic lattice structure is formed from their respective atoms.
  (b) Show why their empirical formulas have ratios of 1:2 and 1:1 respectively.
  (c) Explain what holds the ions close together in their lattice structures.

The properties of ionic compounds can be explained by their structures.

studyon

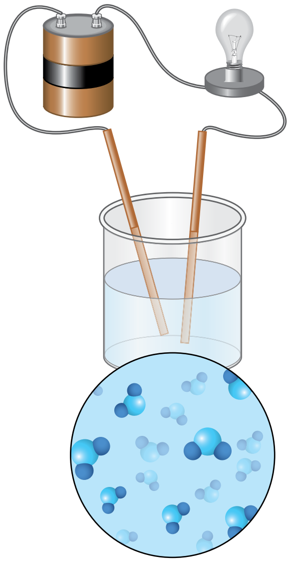
Unit 1
AOS 1
Topic 4
Concept 1

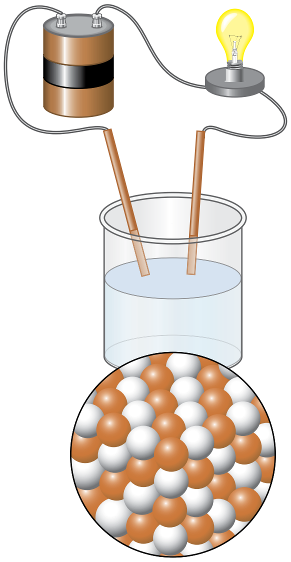
Properties of ionic compounds
Summary screen and practice questions

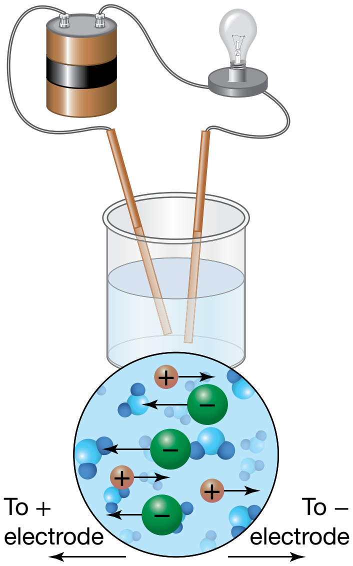
Connecting properties of ionic compounds to structure

Ionic compounds are composed of two or more different kinds of ions that have opposite charges. The electrostatic forces of attraction holding them together are strong. As a result of this, ionic compounds:

- are usually crystalline solids, owing to the arrangement of ions in repeating three-dimensional patterns
- have high melting and boiling points, as a large amount of energy is needed to separate the ions
- do not conduct electricity in the solid form, as the charged particles (ions) are not able to move (to conduct electricity, a substance must contain mobile charged particles)


A current cannot flow through distilled water.


A current cannot flow through a solid ionic compound.



A current can flow through a solution of positive and negative ions.

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Experiment 3.1 Investigating calcite crystals
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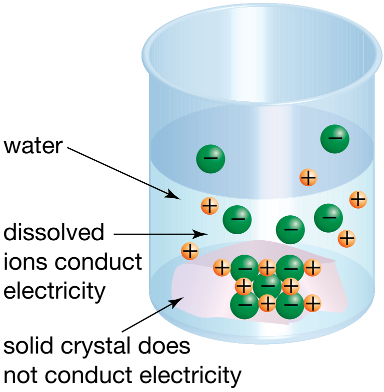
- are hard, since the surface of the crystal is not easily scratched due to the strong ionic bond holding the ions together
- often dissolve in water to form ions, because water molecules are able to move between ions and free them by disrupting the rigid crystal structure
- conduct electricity when molten or in aqueous solution (dissolved in water). In the molten form, ions are able to slide past one another and can therefore conduct electricity. When an ionic substance is dissolved in water, the ions dissociate from the lattice and can move freely to conduct an electric current. The solution is called an **electrolyte**.
- are brittle, since distortion of the crystal causes ions of like charge to come close together and the repulsion between these ions cleaves or shatters the crystal, as shown in the figure on the next page. If the distortion of the crystal is caused by a sharp blow along the plane of ions, cleavage (a clean split) along the plane occurs. If the blow is along different planes of ions, the crystal lattice shatters.

An electrolyte has ions that are free to move. It can be a molten ionic compound or an aqueous solution of an ionic compound.



Dissociation of sodium chloride to form a conducting solution called an electrolyte

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Experiment 3.2 Ionic models
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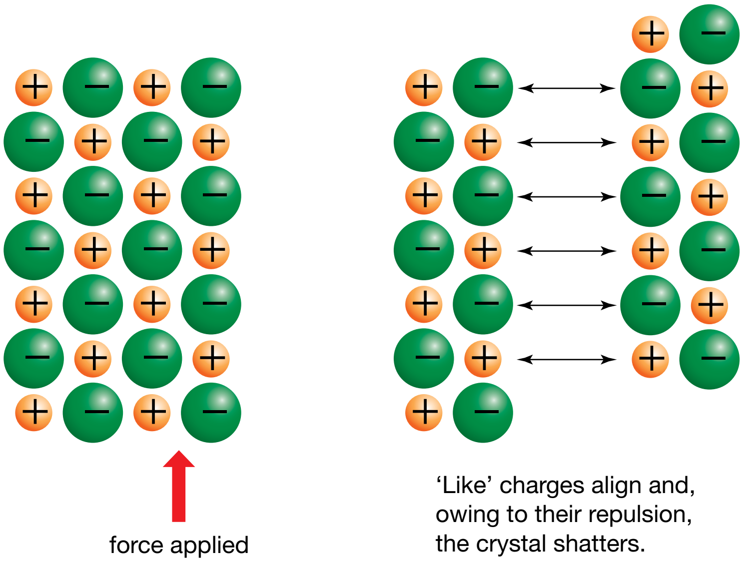


The brittle nature of an ionic compound, which cleaves or shatters when distorted

Revision questions

9. **Calcium chloride is a crystalline substance at room temperature. Predict whether it:**
  - **(a) has a low or high melting point**
  - **(b) shatters when pressure is applied**
  - **(c) conducts electricity in the solid or liquid state.**

  Justify your predictions.

10. **(a) Compare the ratio of positive to negative ions in the lattice of:**
  - **(i) sodium chloride**
  - **(ii) magnesium oxide.**

  (b) Which of these compounds would you expect to have the higher melting point? Give a reason for your choice.

11. **A solid substance has a high melting point, conducts electricity only in molten form and cleaves when struck with a sharp knife. However, it does not readily dissolve in water. Is this enough data to predict the chemical bonding in this substance? Explain in terms of the limitations of the bonding model.**

12. **Atom *X*, having one electron in its outer shell, combines with atom *Y*, which has six electrons in its outer shell.**
  - **(a) Write the formula of the compound formed.**
  - **(b) Name the type of bonding in the compound.**
  - **(c) Predict three general properties of the compound.**

Naming ionic compounds

Knowing how to write formulas is a skill that all chemists require. There are many known compounds and to memorise them all would be impossible. Chemists have developed a set of rules that allow us to predict the formulas of many compounds. Some compounds that have been known for many years also have common names, as listed in table 3.2.

TABLE 3.2 Common ionic compounds

Common name	Systematic name	Formula	Used to manufacture
salt	sodium chloride	$NaCl$	chlorine
soda ash	sodium carbonate	Na_2CO_3	glass
baking soda	sodium hydrogen carbonate	$NaHCO_3$	cake raising agent
lime	calcium oxide	CaO	mortar
limestone	calcium carbonate	$CaCO_3$	cement
potash	potassium nitrate	KNO_3	gunpowder
milk of magnesia	magnesium hydroxide	$Mg(OH)_2$	laxative or antacid
gypsum	calcium sulfate	$CaSO_4$	plaster
caustic soda	sodium hydroxide	$NaOH$	soap or drain cleaner

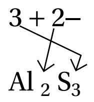
Formulas of binary ionic compounds

A binary compound contains only two elements.

In a **binary** ionic compound, only two elements are present. Examples include sodium chloride, $NaCl$, calcium fluoride, CaF_2, and potassium nitride, K_3N. Note that the word 'binary' does not refer to the relative number of ions or subscripts. When writing empirical formulas for binary ionic compounds, the following rules should be followed:

- Write the symbol for the cation first, followed by the symbol for the anion.
- Determine the lowest whole number ratio of ions that provides a net charge of zero.

The 'cross over' method is another way to determine formulas, using the charges of the ions. For example, to obtain the formula for aluminium sulfide, the symbol for the cation is written first, followed by the symbol for the anion: $Al^{3+}S^{2-}$. The charges are then crossed over so that the charge of the anion becomes the subscript of the cation and the charge of the cation becomes the subscript of the anion.



The formula is then written showing only the subscripts. Note that subscripts must always be reduced to the smallest possible whole numbers. (For example, the formula of magnesium oxide is MgO, not Mg_2O_2.)

Sample problem 3.4

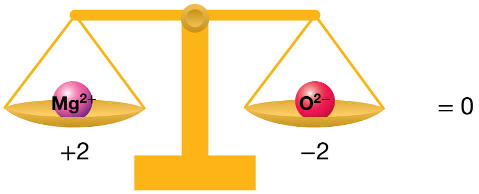
Write formulas for the following compounds:

(a) magnesium oxide
(b) potassium oxide
(c) aluminium sulfide.

Solution:

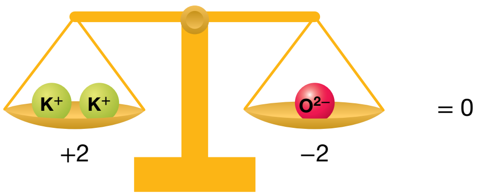
The number of positive charges must be balanced by an equal number of negative charges.

(a) The symbols for the cation and anion in the compound are Mg^{2+} and O^{2-} respectively. In order to have a net charge of zero, the number of positive charges must be balanced by an equal number of negative charges. Thus one Mg^{2+} ion is balanced by one O^{2-} ion and the ratio of ions is 1 : 1.



The formula is written as MgO.

(b) The symbols for the cation and anion in the compound are K^+ and O^{2-} respectively. In order to have a net charge of zero, two positively charged potassium ions are needed to balance the two negative charges of the oxide ion. The ratio of ions is 2 : 1.



The formula is written as K_2O. The subscript numeral '2' indicates that there are two potassium ions for every oxide ion.

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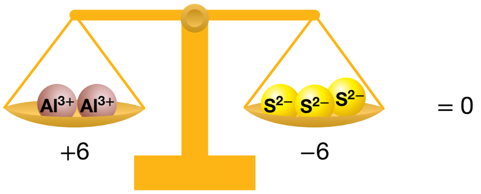
Interactive ionic compounds quiz

(c) The symbols for the cation and anion in the compound are Al^{3+} and S^{2-} respectively. In order to have a net charge of zero, two Al^{3+} ions, with a total charge of 6+, are needed to balance three S^{2-} ions, with a total charge of 6–. The ratio of ions is 2 : 3.



The formula is written as Al_2S_3.

Sample problem 3.5

Name the compound MgF_2.

Solution: First name the cation from the metallic element it came from: magnesium. Then add the name of the anion with the ending *-ide*: fluoride. The compound MgF_2 is known as magnesium fluoride.

Revision questions

13. **Write the formula for the ionic compound formed between each of the following sets of ions:**

- **(a) aluminium and chlorine**
- **(b) barium and oxygen**
- **(c) sodium and sulfur**
- **(d) magnesium and phosphorus.**

14. **Complete the table below by writing the formulas for the compounds formed when each cation is bonded to each anion.**

Ions	K^+	Ca^{2+}	Al^{3+}
F^-			
O^{2-}			
N^{3-}			

15. **Name the nine compounds formed in question 14.**

16. **Name the following compounds:**

- **(a) KCl**
- **(b) Ag_2S**
- **(c) Na_3N.**

Sample problem 3.6

Write the formula for the ionic compound formed between magnesium and chlorine (use table 3.3 on page 57).

Solution: Magnesium ions have a charge of 2+ and chloride ions have a charge of 1–, so two chloride ions are needed to balance the charge on magnesium. Therefore, the formula is $MgCl_2$.

Ions of variable charge

Atoms of transition metals can form ions with different charges: for example, Fe^{2+} and Fe^{3+}.

As the atomic number increases, atoms become larger and have more electrons. Some transition metals have quite complicated arrangements of electrons and may be able to form more than one type of ion. Such ions have different charges. This is very significant as ions of the same element with different charges have different physical and chemical properties. For example, a solution of Cr^{6+} is orange, whereas a solution of Cr^{3+} is yellow, as shown below.

Elements that have ions of variable charge can form two or more different binary ionic compounds containing the same element. Iron, for example, can form black FeO and brown Fe_2O_3, each with different physical and chemical properties. This means we must name them differently, since simply saying iron oxide is ambiguous. Therefore, we use roman numerals in brackets after the cation's name to denote its charge. For example, the black FeO contains Fe^{2+} ions, and so it is called iron(II) oxide. The brown Fe_2O_3 contains Fe^{3+} ions and is named iron(III) oxide.

Metal ions that form more than one cation include iron(II) and (III), copper(I) and (II), and mercury(I) and (II). Note that the roman numerals do not indicate the numbers of ions present in the compound, only the charge of the ion.

These solutions are different in colour. The orange flask contains Cr^{6+} ions, and the yellow flask contains Cr^{3+} ions. When chromium reacts with other chemicals, often Cr^{3+} changes to Cr^{6+}. The colour change can indicate that a reaction has occurred.

Sample problem 3.7

(a) Name the compound CuO.
(b) Give the formula for iron(II) chloride.

Solution:

(a) In order to determine the charge on the metal ion, we work backwards from the known charge on the anion. The charge on the oxide ion is 2–. Therefore, for a net charge of zero, the charge on the Cu ion must be 2+. The compound is therefore named copper(II) oxide.

(b) Iron(II) means that the Fe^{2+} ion is present in the compound. The chloride ion has a charge of 1–. To balance the 2+ of the Fe ion for a net charge of zero, two chloride ions are needed. The formula is $FeCl_2$.

Revision questions

17. **Name the following compounds:**

(a) FeS	(c) $CuCl_2$	(e) Cu_2O
(b) Fe_2S_3	(d) SnO	(f) $PbBr_2$.

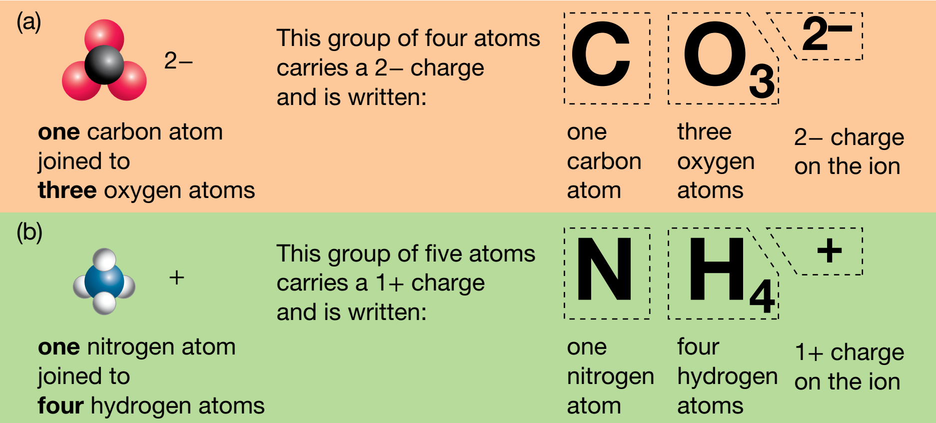
18. **Give the formula for each of the following compounds:**

(a) tin(IV) fluoride	(d) iron(III) nitride
(b) lead(II) sulfide	(e) copper(I) sulfide
(c) mercury(II) oxide	(f) tin(II) oxide.

Polyatomic ions

A polyatomic ion is a bound group of atoms with an overall charge.

A **polyatomic ion** is a group of tightly bound atoms that behaves as a single unit and carries an overall charge. It may be a positively or negatively charged ion. The carbonate ion, for example, is composed of one carbon atom and three oxygen atoms. The *whole group* of four atoms carries a negative two charge.



These diagrams show two polyatomic ions: (a) the anion carbonate and (b) the cation ammonium.

Ionic compounds consist of arrays of cations and anions. A table of charges may be used to help write the empirical formula of an ionic compound (see table 3.3).

Sample problem 3.8

Write the formula for the compound ammonium phosphate.

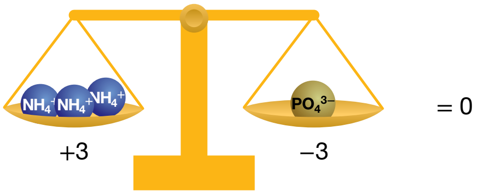
Solution: The method for writing the formula for this compound is the same as that for a binary ionic compound.

STEP 1

Write the symbol for the cation followed by the symbol for the anion. The cation is NH_4^+ and the anion is PO_4^{3-}.

STEP 2

Since the net charge must be zero, we need three NH_4^+ ions for every PO_4^{3-} ion.



STEP 3

Place brackets around the entire ammonium ion so that the subscript 3 applies to all of it. The formula is:

$(NH_4)_3PO_4$

Note that brackets need to be used whenever more than a single polyatomic ion is needed to balance a formula. This is the only time they are used.

TABLE 3.3 Anions and cations

Cations					
+1		+2		+3	
lithium	Li^{+}	magnesium	Mg^{2+}	aluminium	Al^{3+}
sodium	Na^{+}	calcium	Ca^{2+}	chromium(III)	Cr^{3+}
potassium	K^{+}	barium	Ba^{2+}	iron(III)	Fe^{3+}
caesium	Cs^{+}	iron(II)	Fe^{2+}		
silver	Ag^{+}	nickel	Ni^{2+}		
copper(I)	Cu^{+}	chromium(II)	Cr^{2+}		
ammonium	NH_4^{+}	copper(II)	Cu^{2+}		
		zinc	Zn^{2+}		
		tin(II)	Sn^{2+}		
		lead(II)	Pb^{2+}		
		manganese(II)	Mn^{2+}		
		mercury(II)	Hg^{2+}		
		strontium	Sr^{2+}		
Anions					
–1		–2		–3	
hydride	H^{-}	oxide	O^{2-}	nitride	N^{3-}
fluoride	F^{-}	sulfide	S^{2-}	phosphide	P^{3-}
chloride	Cl^{-}	sulfate	SO_4^{2-}	phosphate	PO_4^{3-}
bromide	Br^{-}	carbonate	CO_3^{2-}		
iodide	I^{-}	sulfite	SO_3^{2-}		
hydroxide	OH^{-}	dichromate	$Cr_2O_7^{2-}$		
nitrate	NO_3^{-}	chromate	CrO_4^{2-}		
hydrogen carbonate	HCO_3^{-}	thiosulfate	$S_2O_3^{2-}$		
hydrogen sulfate	HSO_4^{-}	hydrogen phosphate	HPO_4^{2-}		
chlorate	ClO_3^{-}				
hydrogen sulfite	HSO_3^{-}				
nitrite	NO_2^{-}				
permanganate	MnO_4^{-}				
hypochlorite	OCl^{-}				
dihydrogen phosphate	$H_2PO_4^{-}$				
cyanide	CN^{-}				

Revision questions

19. **Refer to table 3.3. Write the formula and name of each of the polyatomic ions that contains sulfur.**

20. **Using the charges provided in table 3.3, copy and complete the table below.**

Name of ionic compound	Valency of cation	Valency of anion	Empirical formula
silver chloride			
potassium sulfide			
magnesium oxide			
aluminium bromide			
iron(III) carbonate			
barium phosphate			
ammonium sulfate			

21. **Give the formulas for the following compounds:**
- (a) **sodium sulfite**
- (b) **calcium nitrite**
- (c) **copper(II) hydrogen carbonate.**

22. **Name the compound formed when each of the following pairs of ions is bonded, and write its formula.**
- (a) **Al^{3+} and CO_3^{2-}**
- (b) **sodium and nitrate**
- (c) **Hg^{2+} and PO_4^{3-}**
- (d) **lead(II) and sulfate.**

Uses of ionic compounds

Ionic compounds have a wide range of uses, as can be seen in table 3.2. Many metals are obtained from ionic compounds that have been extracted from different ores. Sodium hydrogen carbonate is used in cooking and also as an environmentally friendly cleaning agent. Ammonium nitrate is used to manufacture fertilisers as well as explosives. Compounds containing nitrate, nitrite and sulfite ions are used to preserve food. Our bodies can function only because of the presence of particular ions.

Revision question

23. **Write the formulas for the compounds and ions described in the paragraph above.**

As we sweat, we also lose valuable electrolytes (water-soluble ionic compounds); these may be replaced by sports drinks.

Why do athletes take sports drinks?

An electrolyte is a substance that dissociates to form ions. Athletes sometimes take electrolyte drinks after exercise to rehydrate, boost energy and replace important ions in the body that are lost during sweating. Adequate water consumption is essential but there is much discussion about the need for these drink supplements if individuals have an adequate diet. Table 3.4 provides a list of some ions that are essential for optimum functioning of our bodies.

TABLE 3.4 Common ions needed in the body

Name	Symbol	Functions in the body
sodium	Na^+	regulates fluid balance transmission of nerve impulses controls blood pressure by controlling blood volume acid–base balance of blood
potassium	K^+	controls the level of body fluids transmission of nerve impulses and muscle contraction including heartbeat important in reactions within cells
calcium	Ca^{2+}	needed for building teeth and bones involved in blood clotting transmission of nerve impulses and muscle contraction
magnesium	Mg^{2+}	needed for bone formation assists in energy production functioning of muscle and nerve tissue component of enzymes
chloride	Cl^-	acid–base balance of blood fluid balance in the body formation of hydrochloric acid in stomach

Revision question

24. **Research which foods are the main sources of each of the ions listed in table 3.4.**

Hydrated ionic compounds

Hydrated ionic compounds contain water molecules bonded within the crystal.

Note:

mono- = 1
di- = 2
tri- = 3
tetra- = 4
penta- = 5
hexa- = 6
hepta- = 7
octa- = 8
nona- = 9
deca- = 10

A number of ionic compounds, called **hydrates**, release water, which is part of their structure, when they decompose upon heating. When the formula of a hydrated compound is written, the number of water molecules is also included. For example, the formula for iron(II) sulfate octahydrate is written as $FeSO_4{\cdot}8H_2O$. This formula indicates that eight molecules of water are bonded within the ionic crystal for every one formula unit of $FeSO_4$.

The prefixes on the left are used to indicate the number of water molecules in a hydrated compound.

Sample problem 3.9

(a) Name the compound $ZnCl_2{\cdot}4H_2O$.
(b) Give the formula for the compound calcium sulfate dihydrate.

Solution:

(a) Name the binary ionic compound in the usual way, and then use the prefix *tetra-* for four followed by the word ‘hydrate’. The name of the compound is zinc chloride tetrahydrate.

(b) **STEP 1**

Write the symbols for the cation and anion, Ca^{2+} and SO_4^{2-}.

STEP 2

Balance the charges so the net charge is zero, keeping the polyatomic sulfate ion intact. Since the ions are equally charged, the formula for calcium sulfate is $CaSO_4$.

STEP 3

Determine how many water molecules there are from the prefix; *di-* means ‘two’, so the formula is $CaSO_4{\cdot}2H_2O$.

Revision question

25. **The table below shows the uses and formulas of some hydrated ionic compounds. Copy the table and fill in either the formula for the compound or its name. The first one is done for you.**

Name of hydrated ionic compound	Common name(s); use; description	Formula
copper(II) sulfate pentahydrate	blue vitriol; copper plating; blue solid	$CuSO_4{\cdot}5H_2O$
sodium carbonate decahydrate	washing soda, soda ash; water softener; white solid	(a)
(b)	Epsom salts; explosives, matches; white solid	$MgSO_4{\cdot}7H_2O$
magnesium chloride hexahydrate	disinfectants, parchment paper; white solid	(c)
(d)	dyeing fabrics, tanning leather; white solid	$BaCl_2{\cdot}2H_2O$

Growing ionic crystals

Ionic compounds form crystals of many interesting shapes depending on the arrangement of ions in the lattice, but the different sizes of crystals are caused by the diverse conditions in which they are formed. The largest natural crystals found have been in Chihuahua, Mexico, and these are over 12 metres long; the smallest fit into the nanoscale range.

Crystals of a more manageable size can be prepared in a school laboratory from saturated solutions, which are solutions in which no more solid (solute) will dissolve. To form a saturated solution, it is heated while the solute is added; this assists the dissolving process. As the solution cools, the water (or other solvent) evaporates and small crystals appear.

To produce a larger crystal, a regularly shaped crystal can be selected and suspended by a thread in a saturated solution and the crystal will grow. This crystal is called the seed crystal, and it provides a nucleation site, a point where crystallisation can begin. Dust particles or scratches on a glass container can also act as nucleation sites. The rate of evaporation affects the size of the crystal formed.

Recrystallisation processes are very useful to chemists for purifying substances, because impurities cannot form part of the growing crystal structure.

These vanadinite crystals are hexagonal in shape and are a source of the element vanadium.

Chapter review

studyon
Unit 1
AOS 1
Topic 4
Ionic bonding
Summary screens and practice questions

Summary

- Ions are formed when atoms lose or gain electrons to achieve the stable electron configuration of a noble gas.
  - Metals lose electrons in order to gain a full outer shell and become positively charged ions called cations.
  - Non-metals gain electrons to obtain a full outer shell and become negatively charged ions called anions.
  - Whether an atom gains or loses electrons depends on its electronegativity, or electron-attracting power.
- In the periodic table, the group numbers 1, 2, 13, 15, 16 and 17 have the charge 1+, 2+, 3+, 3–, 2– and 1– respectively. Group 14 and group 18 elements do not readily form ions.
- There are three types of chemical bond:
  - ionic bond
  - metallic bond
  - covalent bond.
- Ionic bonding is produced by the strong electrostatic attraction that results between ions when a metallic atom transfers electrons to a non-metallic atom. The ionic compound formed:
  - is known as a salt
  - is made up of an ionic network lattice of positive cations and negative anions
  - is neutrally charged; that is, the total positive charge of the cations must equal the total negative charge of the anions
  - can be represented by a formula unit, which shows the smallest whole number ratio of cation to anion, when charges are balanced. This is also called the empirical formula.
- The structure and bonding of ionic compounds has a direct influence on their properties. Ionic network lattices are rigid structures of cations bonded to anions. Ionic compounds:
  - do *not* conduct electricity in the solid state but become electrolytes when in aqueous or molten states, when their rigid structures have broken down sufficiently to allow free movement of their ions
  - are brittle as pressure causes like charges to align, resulting in the structure shattering
  - are often soluble in water, whose molecules disrupt the lattice and allow free mobile ions to form
  - are usually crystalline solids due to the close-packed, three-dimensional lattice structure
  - have high melting and boiling points due to the strong force of attraction between the ions.
- The key rules in naming ionic compounds and writing their formulas are:
  - When naming a binary ionic compound, always name the metal ion in full first and then add the non-metal ion with the ending *-ide* (e.g. lithium oxide).
  - To find the correct formula for an ionic compound, determine the lowest whole number ratio of ions that gives a net charge of zero.
  - When naming ionic compounds that contain metal ions with more than one charge, such as iron(II) and (III), copper(I) and (II), lead(II) and (IV), mercury(I) and (II), and tin(II) and (IV), use roman numerals to indicate which ion is present (CuS is copper(II) sulfide.) Note that the Roman numerals indicate the charge of the ion rather than the number of ions present in the compound.
  - When there is more than one polyatomic ion in the formula of an ionic compound, the usual rules of naming apply, but use brackets to separate the ions in the formula (e.g. ammonium carbonate is $(NH_4)_2CO_3$).
  - For hydrates, use the prefixes *mono-*, *di-*, *tri-*, *tetra-*, *penta-*, *hexa-*, *hepta-*, *octa-*, *nona-* and *deca-* to indicate the number of water molecules in the compound (e.g. $CuSO_4{\cdot}5H_2O$ is copper(II) sulfate pentahydrate).
- The size of a crystal depends on the conditions during its formation.

Multiple choice questions

1. When an atom loses an electron:
A a positive ion is formed
B a negative ion is formed
C the atomic number changes
D the atomic mass changes.

2. Sodium atoms and sodium ions:
A are identical in their chemical properties
B have the same number of electrons
C have the same number of protons
D have the same number of protons and electrons.

3. The elements sodium, caesium and lithium form ions with the charge in their compounds of:
A 1+ only
B 2+ only
C 3+ only
D 1+ or 2+ only.

4. Which of the following ions does not have the same electron configuration as an oxide ion, O^{2-}?
A N^{3-}
B F^-
C S^{2-}
D Al^{3+}

5. An unknown substance was found to combine with chlorine to form a compound of formula XCl_3. How many outer shell electrons does X have?

A 1 **C** 3
B 2 **D** 4

6. An atom of X, which has two electrons in its outermost shell, is in contact with an atom of Y, which has seven valence electrons. Which of the following is *incorrect*?

A X^{2+} and Y^- ions are present in the compound.
B The compound has the formula XY_2.
C The compound probably has high melting and boiling points.
D No compound can be formed between these two elements.

7. An element has an atomic number of 20. When the element reacts to form an ionic compound its electron configuration is:

A 2, 8, 8
B 2, 8, 8, 2
C 2, 8, 8, 4
D 2, 8, 10.

8. The electron configurations of four pairs of elements are given below. Which pair of elements is most likely to form an ionic bond?

A 2, 8, 1 and 2, 8, 2
B 2, 8, 1 and 2, 8, 7
C 2, 8, 2 and 2, 8, 4
D 2, 8, 7 and 2, 8, 6

9. The formation of an ionic compound from a reaction between the atoms of two elements involves:

A sharing of pairs of electrons between atoms
B donation of outer shell electrons to the entire crystal structure
C transfer of electrons between atoms
D ionisation of the atoms of some elements.

10. The structure of solid calcium chloride is best described as:

A a lattice consisting of diatomic chlorine molecules strongly bonded to calcium atoms
B a lattice of calcium and chloride ions, strongly bonded, in the ratio 1 : 2
C discrete molecules of calcium chloride with strong bonding within the molecule but weak bonding between molecules
D an infinite lattice in which calcium and chloride are linked by strong covalent bonds.

11. The formula of the compound ammonium phosphate is:

A $(NH_4)_3PO_4$ **C** NH_4PO_4
B NH_3P **D** $(NH_4)_2PO_3$.

12. The compound CuO is called:

A copper oxide **C** copper(II) oxide
B copper dioxide **D** copper(I) oxide.

13. The blue crystals of $CuSO_4{\cdot}5H_2O$ are called:

A copper sulfate pentahydrate
B copper sulfate hexahydrate
C copper(II) sulfate pentahydrate
D copper(I) sulfate hexahydrate.

Review questions

Electron configuration

1. Copy and complete the table below.

Name of atom	Symbol for atom	Electron configuration of atom	Name of ion	Symbol for ion	Electron configuration of ion
lithium					
beryllium					
nitrogen					
oxygen					
fluorine					
sodium					
magnesium					
aluminium					
phosphorus					
sulfur					
chlorine					
potassium					
calcium					

2. (a) What is the electron configuration of the calcium atom?
  (b) What is the electron configuration of the fluorine atom?
  (c) Describe what happens when the calcium and fluorine atoms react chemically.
  (d) Use an electron shell diagram to illustrate your answer to part (c).
  (e) What holds the ions together in a crystal?
  (f) What is the ratio of calcium ions to fluoride ions?

3. Cations form from metallic elements. Using an example, show how this occurs.

Ionic bonding

4. In a sample of potassium chloride (an important electrolyte in the human body), the number of K^+ ions and Cl^- ions are equal. In a sample of potassium oxide, however, the numbers of each ion are not equal. State the ratio of each ion in the potassium oxide and explain why they are not equal.

5. Write simple equations and electron shell diagrams to show what happens when the following pairs of atoms bond.
  (a) K and F
  (b) Al and O
  (c) Be and Cl

Properties and structure of ionic substances

6. Explain why ionic salts conduct electricity in the molten and aqueous states but not in the solid state.

7. Draw diagrams to illustrate the brittle nature of salt.

8. Explain why ionic substances have high melting points and are usually crystalline solids at room temperature.

9. A compound made up of *X* and *Y* atoms has the following properties.
- It has a high melting point and boiling point.
- It is very soluble in water but not in kerosene.
- It conducts electricity in the liquid or aqueous state but not in the solid state.
- The electron configuration of an atom of *X* is 2, 8, 2, while that of *Y* is 2, 6.

  (a) What is the formula for the compound?
  (b) What type of bonding must it have?
  (c) Write a simple equation showing how the bonding is attained.

Binary and polyatomic ionic compounds

10. Using table 3.3, write the formulas for the following compounds.
  (a) magnesium sulfate
  (b) zinc oxide
  (c) iron(II) hydroxide
  (d) silver chloride
  (e) aluminium nitride
  (f) calcium carbonate
  (g) calcium hydrogen carbonate
  (h) lead(II) iodide
  (i) potassium hydrogen sulfate
  (j) ammonium carbonate
  (k) silver sulfate
  (l) tin(II) chloride
  (m) potassium sulfate
  (n) sodium fluoride
  (o) barium nitrate
  (p) iron(III) hydroxide
  (q) sodium sulfide
  (r) aluminium oxide
  (s) calcium hydride
  (t) copper(II) sulfate
  (u) ammonium hydroxide
  (v) chromium(III) oxide
  (w) calcium nitrate
  (x) lithium chloride
  (y) potassium cyanide
  (z) sodium hydrogen phosphate

11. Give the names of the following compounds.
  (a) Fe_2O_3
  (b) $Al_2(SO_4)_3$
  (c) $CaCl_2$
  (d) $Mg(NO_3)_2$
  (e) $BaSO_4$
  (f) $ZnCl_2$
  (g) $Al_2(CO_3)_3$
  (h) Na_2SO_4
  (i) $AgNO_3$
  (j) NaOH
  (k) KNO_3
  (l) $PbCl_2$

12. Give the chemical name and formula for each of the following compounds.
  (a) Soda ash is the common name of a compound containing sodium ions and carbonate ions.
  (b) Baking soda is commonly used in baking cakes and is composed of sodium and hydrogen carbonate ions.
  (c) Chalk, marble and limestone are all composed of calcium ions and carbonate ions.

Hydrated ionic compounds

13. Write the formulas for the following compounds.
  (a) magnesium sulfate heptahydrate
  (b) sodium carbonate decahydrate
  (c) zinc chloride hexahydrate
  (d) barium chloride dihydrate

14. Name the following compounds.
  (a) $BaCl_2{\cdot}3H_2O$
  (b) $LiCl{\cdot}4H_2O$
  (c) $CoCl_2{\cdot}5H_2O$

15. Sodium chloride is used as a typical example of an ionic compound. Research the effects of too little or too much sodium chloride on the body.

16. Prepare a fact sheet outlining the advantages and disadvantages of sports drinks.

17. Find out about other ions and their importance to human health.

Exam practice questions

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Unit 1

AOS 1

Topic 4

Ionic bonding

Sit topic test

In a chemistry examination, you will be required to answer a number of short and extended response questions.

Multiple choice questions

1. Potassium chloride is an ionic compound. Which one of the following statements about potassium chloride is *incorrect*?

A It has a high melting point.

B It conducts electricity in all physical phases.

C It is soluble in water.

D Its ions are fixed in a crystal lattice that is brittle. **1 mark**

2. X and Y are elements. The ionic compound XY_2 is known to exist. If X^{2+} and Y^- both have the same electron configuration as the neon atom, then XY_2 is:

A magnesium fluoride

B magnesium chloride

C magnesium bromide

D calcium fluoride. **1 mark**

Extended response questions

3. Predict the charges on the ions formed when the following atoms react with other atoms.

(a) Ca

(b) P

(c) Al **3 marks**

4. Oxygen gas, which forms about 20% of the Earth's atmosphere and is essential to life, exists as O_2 molecules. Explain why ionic bonding does not exist between two atoms of oxygen. **2 marks**

CHAPTER

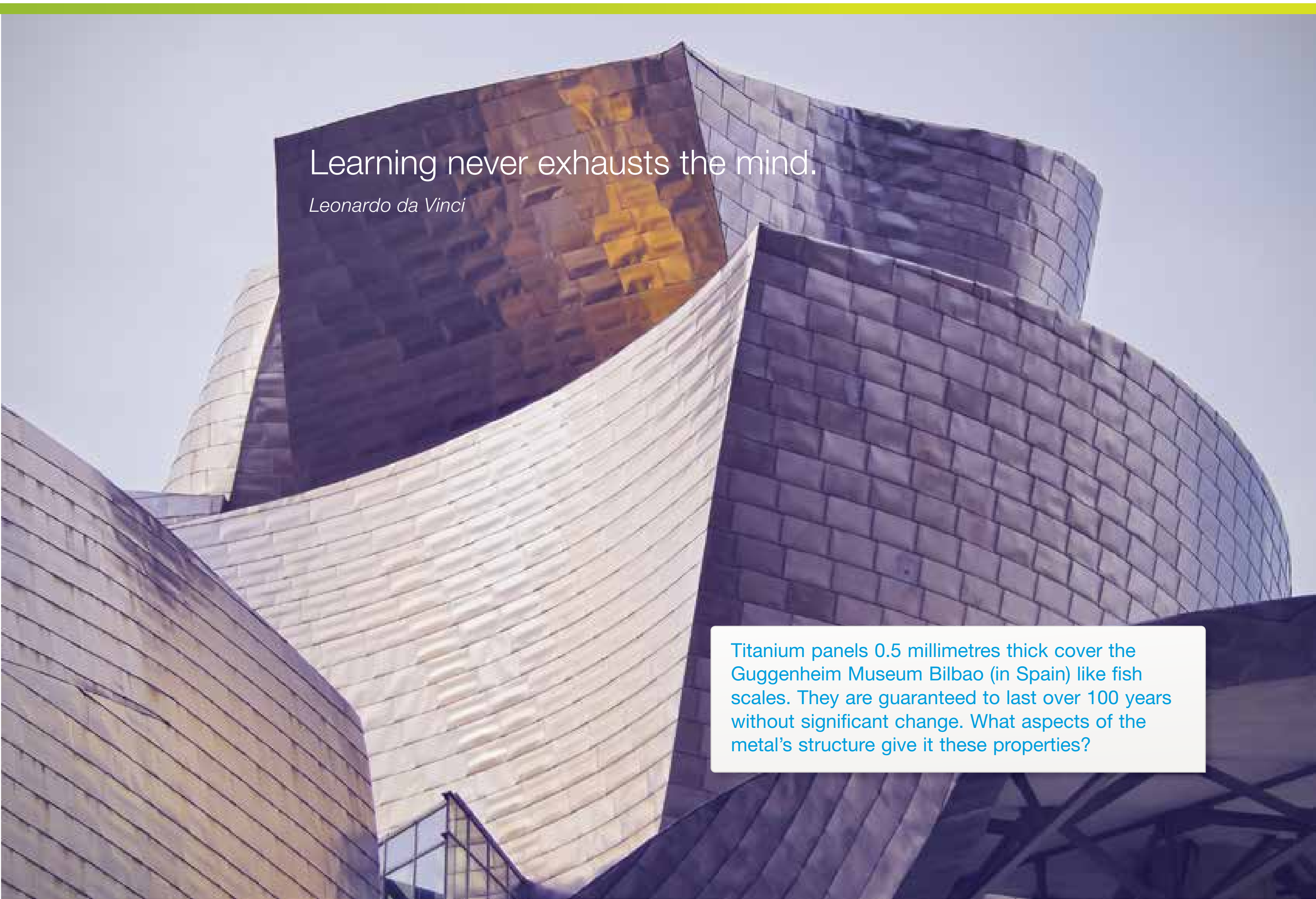
4 Metallic bonding

What are earrings made from? What is the best material for body piercings? Do you like gold, and, if so, do you prefer yellow, pink or white? Do you like silver, stainless steel or titanium? Some people's skin reacts chemically with different metals so they must choose jewellery carefully. Metals have different properties, so it is important to choose the right metal for the purpose. All metals, however, have a similar lattice and bonding structure that determines their properties and uses.

YOU WILL EXAMINE:

- metallic bonding as the strong bonding within metals
- the relationship between the structure and bonding of metals and their properties
- how the properties of transition metals compare with those of main-group metals
- the importance and limitations of the metallic bonding model
- the reactivity of metals
- the modification of metals by processes such as coating, alloying and heat treatment
- metallic nanomaterials and their properties.

> Learning never exhausts the mind.
>
> *Leonardo da Vinci*



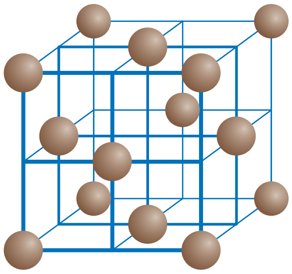
Titanium panels 0.5 millimetres thick cover the Guggenheim Museum Bilbao (in Spain) like fish scales. They are guaranteed to last over 100 years without significant change. What aspects of the metal's structure give it these properties?

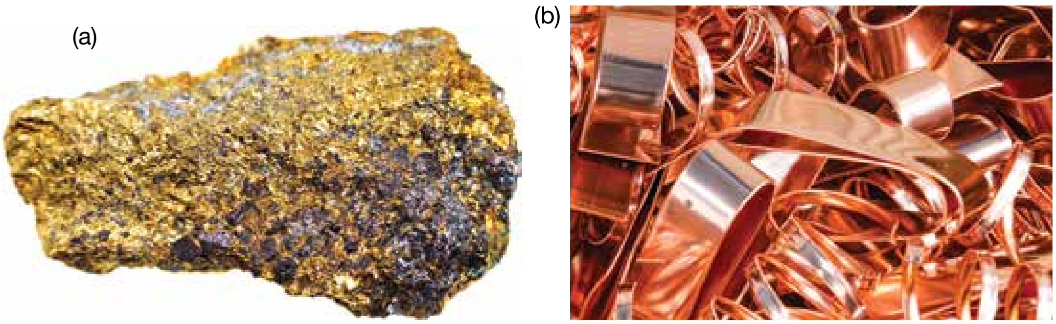
Metals

The largest group of elements is the metals.

It has been known for centuries that elements can be divided into two broad groups: metals and non-metals. The largest group is the **metals**. A metal is generally a hard, shiny solid that can be shaped, and is a relatively good conductor of heat and electricity. The discovery and use of metals moved civilisation beyond the Stone Age.

The first metal to gain widespread use was copper, which was first mined and used on the island of Cyprus around 5000 years ago. Copper's popularity was mainly due to its relatively low melting point (1084 °C) and the ease with which it could be extracted from its ores. In contrast, aluminium, which is the most common metal in the Earth's crust, was difficult and expensive to extract from its ore and has become widespread only in the last century due to technological advances.





Copper can be extracted from the ore chalcopyrite (a). In human history, copper ranks second only to iron in its importance. Today the main use of copper is in electrical cables, pipes and wires, shown here ready for recycling (b).

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Structure of metals

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Unit 1	**Properties of metals** Summary screen and practice questions
AOS 1	
Topic 5	
Concept 1	

Properties of metals

Since there are more than 80 metal elements, it is not surprising that some of their properties vary widely. Group 1 metals, such as lithium, sodium and potassium, must be stored in oil or kerosene because they react violently with water, whereas metals such as copper, silver and gold are chemically unreactive and are used in coins and jewellery. At room temperature, mercury is a liquid, but magnesium is a solid. Zinc is brittle, iron is hard, and lead sheet can be bent easily. Nevertheless, most metals have the following properties. Metals:

- are lustrous. They have a shiny surface when polished.
- are good conductors of heat
- are good conductors of electricity
- are malleable. They can be hammered, bent or rolled into sheets or other shapes.
- are ductile. They can be drawn out into wires.
- generally have a high density (mass per unit volume)
- have a range of melting points but most are quite high
- are often hard. They have high resistance to denting, scratching and bending.
- have high **tensile strength**, meaning that they offer high resistance to the stresses of being stretched or drawn out and, therefore, do not break easily.

Structure of metals

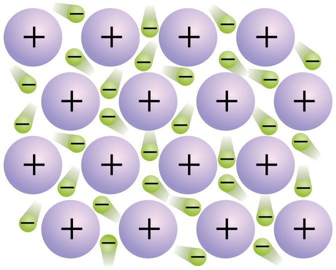
Metal atoms have low electronegativities and so lose electrons easily.

Metallic atoms have low electronegativities, which means they tend to lose their outer shell electrons easily. Once a metallic atom has lost its outer shell electron, it becomes a positively charged cation. This produces electron configurations similar to the group 18 elements (noble gases) and hence a more stable state.

The atoms in metals are arranged in a lattice structure surrounded by a sea of electrons.

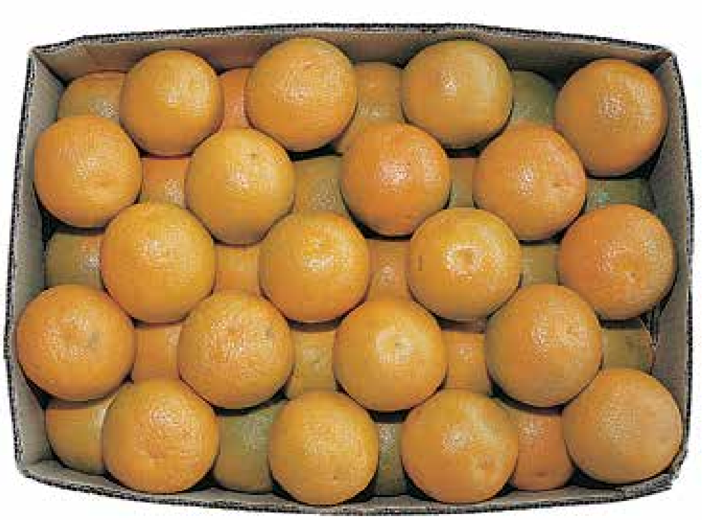
X-ray studies show that metals occur as crystal lattices. Chemists imagine these **metallic lattice structures** as being made up of an array of cations. The electrons from each metallic atom are found in a common pool and are free to move between all the cations. They are referred to as **delocalised** electrons. Electrostatic forces of attraction between the positively charged cations and the negatively charged electrons hold the lattice together.

A metal can therefore be imagined as a rigid framework of cations immersed in a 'sea' of electrons that serve as the cement that holds the three-dimensional cationic network together, as shown in the figure on the left. This type of bonding is known as **metallic bonding.**



Metallic bonding — fixed cations in a 'sea' of mobile, delocalised electrons

We think of a metal crystal as containing spherical atoms packed together, as modelled by the orange arrangement in the photo below. These atoms are bonded to each other equally in all directions. This means that metal atoms are hard to separate but relatively easy to move, provided the atoms stay in contact with each other. We can model this structure by packing uniform hard spheres in a way that uses the available space most efficiently, as the orange model demonstrates.



Metallic bonding may be compared with oranges packed in a box. The oranges represent the fixed, closely packed cations in the lattice. Electrons may be found moving freely in the spaces between the oranges.

Revision questions

1. **Lithium exists as a crystalline solid at temperatures less than 180 °C.**
   (a) **Write the electron configuration of a lithium atom.**
   (b) **Write the symbol and electron configuration of its cation.**
   (c) **Describe how lithium atoms are bonded together in the solid.**
2. **Which of the atomic numbers below belong to a metal?**
   (a) **2**
   (b) **11**
   (c) **17**
   (d) **19**
   (e) **25**
3. **Why is the structure of a metal stable? Illustrate your answer using the example of aluminium and its electron configuration.**

Metallic properties and lattice structure

The lattice structure of metals may be used to explain many of their properties. It is these properties that give metals their many applications in our society. These are discussed in the next section.

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Lattice structure

The metallic lattice structure explains many properties of metals.

Although the lattice structure of metals is strong, they can still react — as sodium does on contact with water.

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Experiment 4.1 Metallic trees
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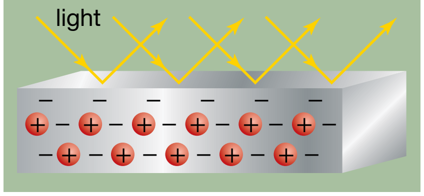
study on

Unit 1	**Metallic bonding model** Summary screen and practice questions
AOS 1	
Topic 5	
Concept 2	

The lattice structure of metals explains their key properties.

Lustre

The lustrous appearance of a metal is due to the mobile electrons within the lattice being able to reflect light, causing the metal to shine.



Density

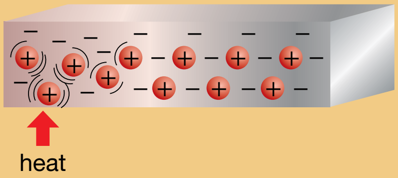
Most metals have relatively high densities because metallic lattices are close-packed.

Melting point and hardness

The generally high melting points and hardness of metals indicate that metallic bonding is quite strong. Melting points and hardness increase with an increase in the number of outer shell electrons, since there is a greater attractive force between the cations and the electrons.

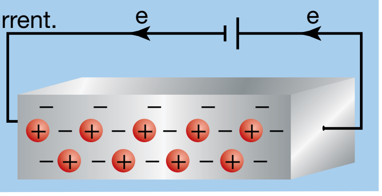
Conduction of heat

Electrons can gain kinetic energy in hotter areas of the metal and quickly transfer it to other parts of the metal lattice because of their freedom of movement. The heat causes the electrons to move faster, and the ‘bumping’ of these electrons with each other and the protons transfers the heat.



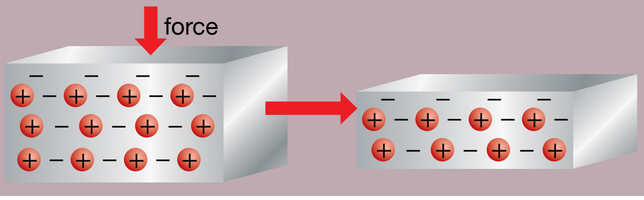
Conduction of electricity

When an electric field is applied to a metal, one end of the metal becomes positive and the other becomes negative. All the electrons experience a force towards the positive end. This movement of electrons is what we call an electric current.



Malleability and ductility

Metals are malleable and ductile, rather than brittle, as a result of the non-directional nature of metallic bonds. The attractive forces exerted by the positive metal ions for the mobile electrons occur in all directions. This means that layers of atoms can move past one another without disrupting the force between the positive ions and the negative sea of electrons. The nature of the metal does not change when the metal becomes thinner.



Revision questions

4. Explain why iron is:
   (a) hard
   (b) malleable and ductile
   (c) able to conduct electricity in solid and molten form.

5. **Magnesium is produced from the mineral dolomite, $CaCO_3 \cdot MgCO_3$.**
   - (a) **Describe the bonding in magnesium metal crystals.**
   - (b) **Magnesium crystals are silver-white and shiny. Explain this property of lustre in terms of structure and bonding.**

Alloys

The many properties of metals allow them to be used for a variety of purposes.

Not all the properties of a metal may be suited to the specific use we wish to make of it. By mixing a metal with other metals or some non-metals, we can change its properties. This process is called **alloying** and the resulting metal is called an **alloy**. For example, pure gold (24 carat) is a relatively soft element and is easily deformed. We therefore use the alloy 18 carat gold in jewellery, as it contains 75% (by mass) gold with the other 25% made up of silver and copper to make the product stronger and more durable. The term 'carat' is used to indicate the ratio of precious metal to base metal in an alloy. For example, 18 carat gold is $\frac{18}{24}$ or three-quarters gold to one-quarter base metal, and 24 carat gold is $\frac{24}{24}$ or pure gold.

Steel can be pressed and rolled into sheets, as at this factory. What metallic property is illustrated here?

Magnesium is unsuitable for industrial use because of its high reactivity. Yet, when alloyed with other metals, such as aluminium, zinc and manganese, it can be used for aircraft and guided missile parts where lightness and high tensile strength are essential. Other uses include aircraft and car wheels, artificial limbs and skis.

Alloying involves combining a metal with other metals or some non-metals.

Aluminium can be mixed with other metals to form very strong, light alloys that are ideal for making cars, aeroplanes, ships and bridges. In order to reduce fuel emissions from cars, fuel efficiency must be increased and this can be done by reducing the weight of the car. Aluminium alloys have thus been substituted for heavier metals such as steel.

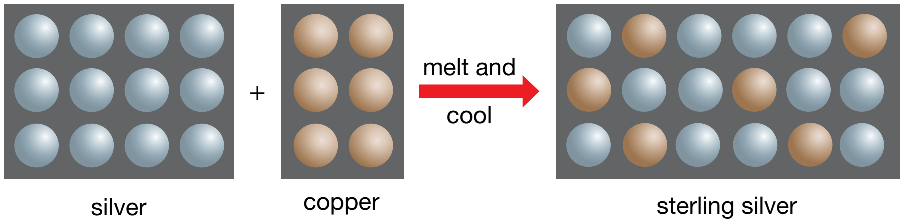
Magnesium alloys demonstrate considerable lightness and high tensile strength, hence their use in car wheels.

Today, intensive research is concentrated on producing higher quality magnesium alloys for lightweight magnesium car components as magnesium alloys are 35% lighter than aluminium and use less energy in the casting process.

Alloys are prepared by melting the metals or other materials together and cooling the mixture. Most of the metallic materials in everyday use are alloys rather than pure metals. Sometimes, non-metallic atoms such as carbon and silicon are used to form an alloy.

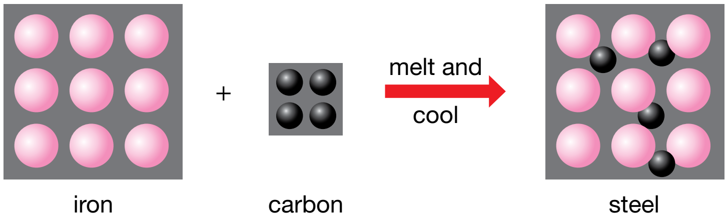
Two types of alloys are substitutional and interstitial alloys.

If the atoms of the metals being used to form an alloy are about the same size, they can replace each other in the metal crystals. This type of alloy is called a **substitutional alloy**. Sterling silver is a substitutional alloy made from silver and copper.



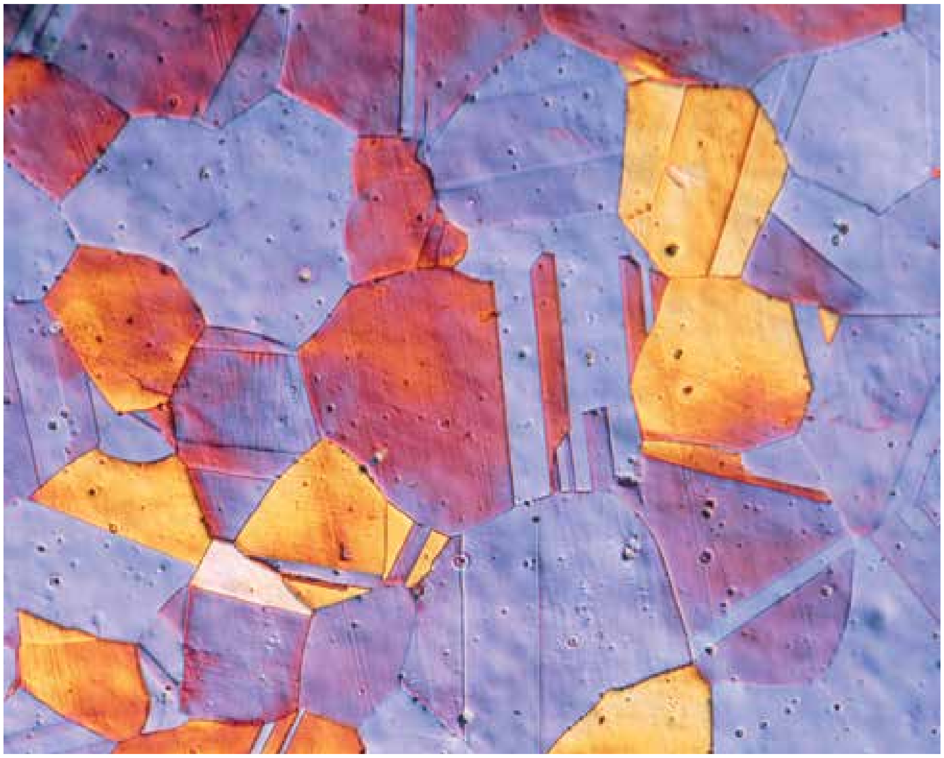
Sterling silver is a substitutional alloy.

Sometimes the atomic sizes of the metals in the alloy differ greatly. The smaller atoms may then fit into the spaces between the larger atoms, forming an **interstitial alloy**. Steel is an interstitial alloy made from iron and carbon.



Steel is an interstitial alloy.

Interstitial alloys may be modelled by using coins of different sizes to represent the different metals being alloyed. Note that the neat rows of one 'metal' have been spoilt by the new 'metal' introduced into the structure, creating more dislocations. This explains why alloys are usually harder than their component metals.



Polarised light micrograph of a thin section of common brass, a copper–zinc alloy, showing the distinct grain structure of this metal alloy. The grain boundaries — the three-dimensional surfaces that enclose grains — show up as fine lines around the more or less polygonal grains.

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Alloys

The properties of alloys differ from those of their component metals.

The properties of alloys differ from those of their component metals. For example, sterling silver is harder than pure silver, and stainless steel is stronger and more resistant to corrosion than iron. Solder is composed of 70% lead (melting point 327 °C) and 30% tin (melting point 232 °C). It is used to join metals together as it melts at 180 °C, lower than either of its components.

Some common metals and their alloys are shown in the figure on page 72.

Revision questions

6. **Solder (melting point 180 °C) is an alloy made from 30% tin (melting point 232 °C) and 70% lead (melting point 327 °C). It is commonly used to join metal surfaces or parts.**
  - **(a)** **Describe how solder could be made.**
  - **(b)** **The atoms of tin and lead are similar in size. Decide which coin model (interstitial or substitutional) better represents the structure of solder.**
  - **(c)** **What properties of solder differ from that of tin and lead? Relate this to its use.**

7. **Using the figure on page 72, predict which of the alloys shown might be suitable for making:**
  - **(a)** **a trumpet**
  - **(b)** **the fuel tanks of a space shuttle**
  - **(c)** **the hull of a ship**
  - **(d)** **rollerskates.**

8. **Automatic fire sprinkler plugs are used in most modern buildings. When they melt, they release water onto the fire. The plugs are made from an alloy of tin, lead, bismuth and cadmium, each of which melts at a temperature well over 200 °C. What properties would you expect this alloy to exhibit, and how do these properties differ from those of the metals from which it is made? Explain your answer.**

9. **Conduct an internet search to find three examples of metal alloys and their uses in society. Present your findings as a visual poster that clearly identifies the properties of each alloy and the metals used to make it.**

An alloy of mercury is called amalgam. Dental fillings are amalgams of mercury (50%), silver (35%) and tin (13%) with copper and zinc. It has been estimated that Australian dentists use over 4.5 tonnes of mercury in amalgam fillings per year. The amalgam filling starts as a pliable substance but undergoes a chemical reaction causing it to harden after it is used to fill a cavity.

Key

Name and composition of alloy
Properties
Uses

ALUMINIUM
Too soft and brittle to be useful

Alloys

duralumin 94.9% Al, 4% Cu, 0.6% Mn, 0.5% Mg	**mixture** of Al and Mn	**mixture** of Al and B
three times as hard as steel but one-third as heavy	resistant to tarnishing	increases electrical conductivity
aircraft; racing bikes	window frames; kitchen foil	electric cables

IRON
Hard but too brittle to be useful

Alloys

mild steel Fe with traces of carbon	**stainless steel** 73% Fe, 18% Cr, 9% Ni	**tungsten steel** 71% Fe, 5% Cr, 24% W (tungsten)
hard, strong and more resistant to corrosion	four times as strong as mild steel; very resistant to corrosion	very hard and very strong
cars; ships	surgical equipment; sinks; cutlery; razor blades	high-speed tools to work metals

COPPER
Soft and workable but corrodes easily

Alloys

bronze 90% Cu, 6% Sn, 4% Zn	**brass** 65–70% Cu, 35–25% Zn	**manganese bronze** 58% Cu, 39% Zn, 2% Fe, 1% Sn, traces of Mn
resistant to corrosion; harder and stronger	strong and workable; resistant to corrosion	highly resistant to wear
parts of ship's decks and propellers	wires; musical instruments; plumbing; door handles; beds	car clutch disks; valves; pumps

Jewellery made from gold alloys is harder than pure gold.

GOLD AND SILVER (precious metals)
Resistant to corrosion but soft

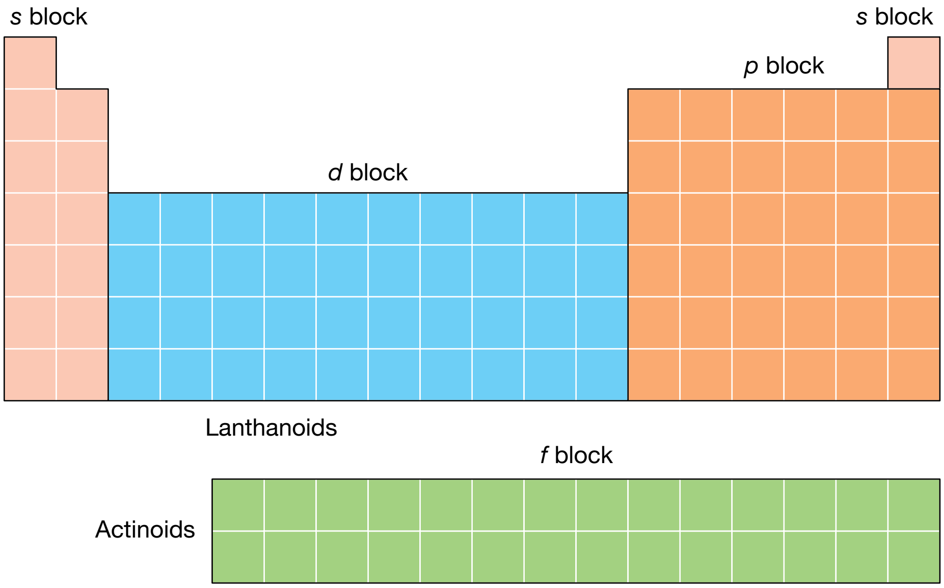
Gold alloys	Silver alloys
18 carat gold 75% Au, 12.5% Ag, 12.5% Cu	**sterling silver** 92.5% Ag, 7.5% Cu
harder	harder
jewellery	jewellery and silverware

Common metals and their alloys

Specific properties and uses of *s*-block metals

Metals are widely used in our society. Their different properties make them useful for particular purposes. The alkali metals are very reactive; they are not found in elemental form and so most uses involve their compounds. A knife can easily cut through these metals as they have such a low density. They have relatively low melting points due to having only one outershell electron to participate in bonding. Also, the atoms are larger than other atoms in the same period, and they are not as efficiently packed together. Alkaline earth metals are less reactive, harder and have higher melting points than alkali metals. Usually, these reactive elements are found as compounds. The elements and compounds have a variety of applications. *s*-block metals usually form compounds that are white.

s-block metals are reactive and generally have a low density.



Sodium is used as a coolant in nuclear reactors, in the manufacture of titanium and for street lights. Lithium is used in rechargeable batteries, and, because of its lightness, it is combined with other metals to make alloys. As already described, magnesium is alloyed because it produces superior qualities. The compound calcium carbonate is used to make calcium oxide for neutralising soils, glass making and as a component of cement.

TABLE 4.1 Properties of metals in the *s* block

Metal	Symbol	Density ($g\ mL^{-1}$)	Melting point (°C)	Thermal conductivity at 25 °C (watts $m^{-1}\ K^{-1}$)	Electrical conductivity at 25 °C ($ohm^{-1}\ m^{-1} \times 10^7$)	Hardness
lithium	Li	0.53	181	85	1.1	soft
sodium	Na	0.97	98	140	2.1	soft
potassium	K	0.66	64	100	1.4	soft
beryllium	Be	1.86	1278	190	3.1	hard
magnesium	Mg	1.75	649	160	2.3	medium
calcium	Ca	1.55	839	200	3.0	medium

Revision questions

10. (a) State the charge on ions formed from group 1 elements.
(b) State the charge on ions formed from group 2 elements.

11. Use the properties of sodium and potassium to predict the approximate values of the density, melting point, thermal conductivity, electrical conductivity and hardness of rubidium, the next element in group 1.

Specific properties and uses of *p*-block metals

The *p* block contains metals and non-metals. The properties of the metals are varied.

A can can be manufactured from aluminium or tin-coated steel. Aluminium and tin are metals found in the *p* block of the periodic table, which also contains metalloids and non-metals. Aluminium is the most abundant metal in the Earth's crust and is very versatile. It has a coating of aluminium oxide, Al_2O_3, and this prevents it from reacting with oxygen so it does not corrode further. As well as its uses in transport and building, it is a useful material for drink cans and saucepans. The good electrical conductivity and low density of aluminium make it ideal for use in overhead wires. Tin is used as a coating to prevent corrosion, such as in 'tin' cans, which are made from tin-coated steel. The alloys solder, bronze and pewter all contain tin.

Lead, another *p*-block metal, has been used since Roman times because it is easily worked and resistant to corrosion. Lead was used for pipes and to line aqueducts to bring water to homes and public baths. Roofing, pottery glazes and jewellery were also regular uses for lead. It is less commonly used now for pigments, as a petrol additive and in jewellery because it is hazardous to health. Current uses include car batteries, roofing, weights and bullets.

Beneath the outer layer of copper alloy is a core of lead. Lead is used because it is dense, readily available and easy to shape.

TABLE 4.2 Properties of metals in the *p* block

Metal	Symbol	Density (g mL^{-1})	Melting point (°C)	Thermal conductivity at 25 °C (watts m^{-1} K^{-1})	Electrical conductivity at 25 °C (ohm^{-1} m^{-1} × 10^7)	Hardness
aluminium	Al	2.7	660	236	3.8	medium
tin	Sn	7.3	232	66.8	0.9	medium
lead	Pb	11.3	327	35.3	0.5	soft

Revision question

12. Explain why aluminium is used for drink cans whereas 'tin' cans are used for other foods.

Specific properties and uses of *d*-block elements

The presence of transition metal compounds is responsible for the colourful stained glass windows in churches and other buildings. The formation of coloured compounds is an interesting characteristic of the *d*-block elements;

exceptions are scandium and zinc. The metals that display magnetism — iron, cobalt and nickel — are transition metals. As with other metals, most are used in the form of alloys. They can form ions with different charges; for example, iron forms two compounds with chlorine — $FeCl_2$, which is a green compound, and $FeCl_3$, which is an orange compound. Another valuable feature of transition metals is that they are very useful catalysts. A catalyst is a substance that increases the rate of a chemical reaction without itself being changed.

d-block elements form coloured compounds and are good catalysts.

studyon

Unit 1	**Main-group and transition metals**
AOS 1	Summary screen and practice questions
Topic 5	
Concept 3	

Transition elements form coloured solutions.

Transition metals are an ever-present part of our daily lives, from the iron in the oxygen transport molecule, haemoglobin, in our blood to the $2 coins in our pockets which are 92% copper. Iron is widely used in construction, often in alloy form as steel, which is less brittle and more resistant to corrosion than elemental iron. It is used in the manufacture of tools and vehicles and is a catalyst in the production of ammonia. Copper is used in electrical wiring and as bases on saucepans because of its good electrical and heat conduction. Interestingly, copper forms the basis of blood in crustaceans, which results in it being a blue colour. Silver and gold, as well as being used for jewellery, are good conductors of electricity and are used in electrical contacts. Apart from its presence in baked beans, the element nickel is used in making stainless steel, in batteries, in coins and as a catalyst in the manufacture of margarine.

TABLE 4.3 Properties of transition metals

Metal	Symbol	Density (g mL^{-1})	Melting point (°C)	Thermal conductivity at 25 °C (watts m^{-1} K^{-1})	Electrical conductivity at 25 °C (ohm^{-1} m^{-1} $\times 10^7$)	Hardness
iron	Fe	7.9	1535	75	1.0	hard
copper	Cu	8.9	1083	401	6.0	hard
silver	Ag	10.5	961	429	6.3	medium
gold	Au	19.3	1059	318	4.3	soft

Revision questions

13. **Give three reasons why silver and gold are used for jewellery.**
14. **How do (a) gold, (b) mercury and (c) zinc differ from the majority of transition metals?**
15. **Choose a metal that would be best suited for fabrication of:**
    (a) **car bodies**
    (b) **beer cans**
    (c) **household electrical wiring**
    (d) **fishing sinkers.**
    Justify your choices.
16. **Electricity transmission lines for industry are usually made from aluminium even though copper is a better electrical conductor. Suggest two reasons why aluminium is used.**
17. **Copy the table below and complete it by listing as many metals as you can and linking their uses to their properties.**

Metal name and symbol	Use	Property demonstrated

18. **Prepare a multimedia presentation that shows how metals are used in relation to their properties. A digital camera could be used to obtain examples from your environment.**

Other models that explain metallic properties

The lattice model does not explain all properties of metals.

Although many of the properties of metals can be explained using the lattice model, other properties are not explained.

- Why is iron magnetic and copper non-magnetic?
- Why is platinum twenty times as dense as sodium?
- Why is lead quite malleable whereas iron is tough?

Different models may be used to explain some of these properties. We will examine the **ball bearing model**.

Metal strength — the ball bearing model

Metal strength depends on the arrangement of atoms in their crystal grains.

Some metals bend easily whereas others are strong. When a metal bends, layers of atoms slide over each other. The more easily they slide, the more easily the metal bends.

A metal with perfect rows of atoms bends easily. If the rows are distorted, they cannot slide over each other as freely, so the metal does not bend easily.

Metals do not crystallise with perfect rows of atoms all through the lattice. They form areas of perfect close-packing called grains. The boundaries between the grains are called grain boundaries. Although the atoms in metal grains are packed in a regular order, the grains themselves are irregularly shaped crystals of the metal pushed tightly together.

The structure of a metal may be modelled by using ball bearings to show the arrangement of cations into grains. The model shows that:

- Metals with large grains have fewer dislocations and they bend easily.
- Metals with small grains have many dislocations and they do not bend easily.

The crystal structure, and therefore the properties, of metals can be changed in a number of ways. The next section discusses some of these.

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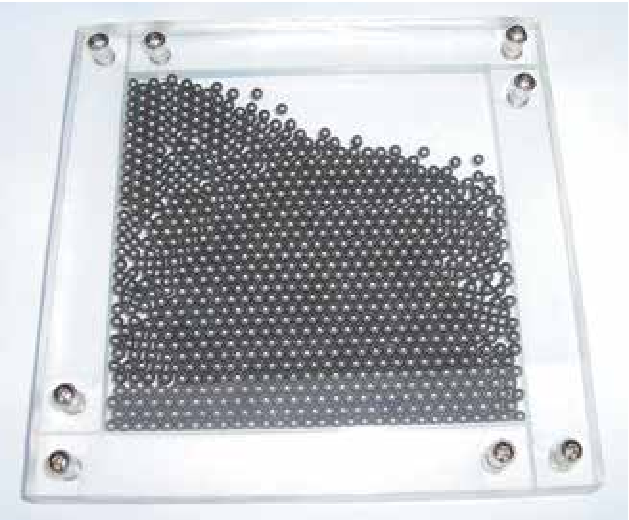
Weblink
Treatment of steel

(a)

(b)



(a) This ball bearing model shows small areas of perfect packing that simulate the arrangement of metal cations in a single crystal grain. This metal would not bend easily.

(b) Grains are large areas of regular arrangement; they represent a single metal crystal. Grain boundaries are narrow areas of disorder seen between one crystal grain and another. Vacant sites occur when an atom is missing from the crystal structure. Dislocations happen when a row of atoms is displaced and the regular packing stops.

Modifying metals

Modifying a metal by work hardening

The properties of metals can be modified by work hardening, heat treatment and alloying.

study on

Unit 1	**Modifying metals** Summary screen and practice questions
AOS 1	
Topic 5	
Concept 6	

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Digital document
Experiment 4.2 The effect of heat on metals
doc-15997

Bending or hammering cold metals causes their crystal grains to become smaller. Since bending is now made more difficult, the metal is toughened, or **work hardened**. This is easily seen by bending a metal coathanger and then trying to bend it back to its original shape. It does not bend back in the area that has been work hardened.

Work-hardened metals are usually more brittle owing to the increased number of dislocations in the crystal structure. Constantly bending a piece of metal backwards and forwards causes it to snap since more and more dislocations are being produced in the metal crystal.

Modifying a metal using heat

Annealing, quenching and tempering alter the properties of a metal.

The neatness of rows of atoms in a metal may be disrupted by heat. There are three main ways of doing this:

- **Annealed** metals are heated until they are red hot and then cooled slowly. Larger metal crystals are formed so that the metal produced is softer. Annealing is an intermediate step in metal-forming processes. It is used to restore ductility lost due to work hardening in items such as the coiled copper pipes in hot water systems.
- **Quenched** metals are heated until they are red hot and then allowed to cool quickly in cold water. Smaller crystals are formed, which makes the metal harder but more brittle. For example, blacksmiths formerly quenched horseshoes by plunging them into cold water.
- **Tempered** metals are produced when quenched metals are warmed again to a lower temperature and then allowed to cool slowly. This procedure reduces the brittleness of the metal while retaining the hardness. For example, metals used to make special wood-carving tools such as lathes and chisels, as well as axes and rock drills, are tempered.

Quenching, annealing and tempering alter the properties of a metal because these processes disrupt the metal lattice.

Revision questions

19. **What are some of the limitations of the 'delocalised sea' of electrons lattice model for metals?**

20. **How does the ball bearing model explain why some metals bend more easily than others?**

21. **Samurai swords were tempered to make them stronger and harder.**
  (a) **Describe how tempering alters the properties of a metal.**
  (b) **Name two other methods that use heat to modify the properties of a metal, and explain how they work.**

22. **Explain why work-hardened copper requires annealing before it may be further bent into a coiled shape.**

23. **When a metal paperclip is straightened out, it is difficult to bend it back to its original shape. Why?**

Coating metals

Metals may be coated for decorative, protective reasons or to make them more useful, such as non-stick frypans, chrome-plated pipes and jewellery. Generally, however, metals, particularly iron, are coated to prevent corrosion.

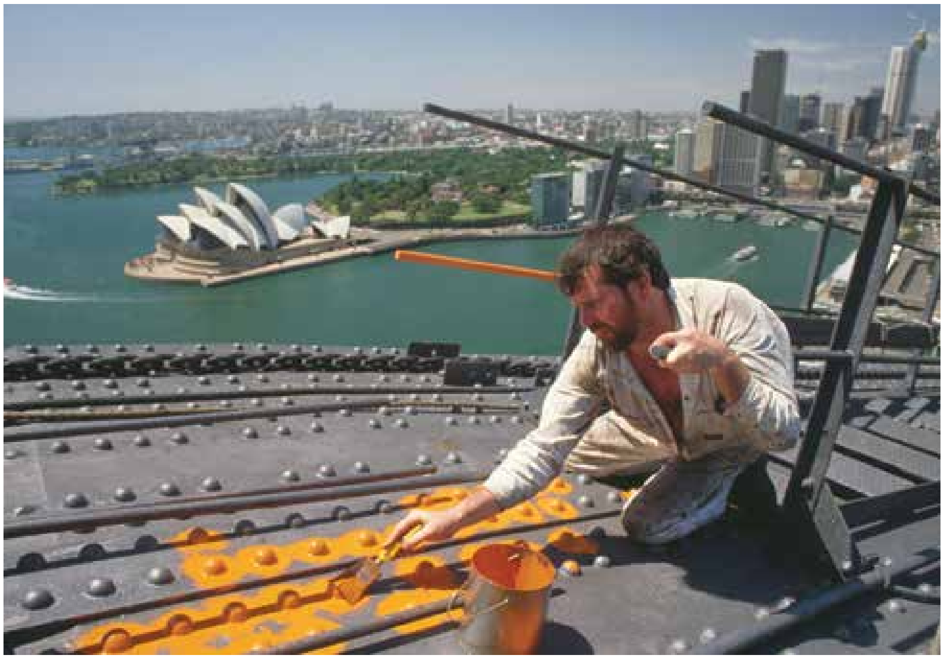
Surface protection

Surface protection is used to prevent the air and water from coming into contact with the metal. A number of methods may be used for surface protection of iron and steel:

- Plastic is the most common form of surface protection. It is used to cover items ranging from simple household coathangers to commercial cables and wiring.
- Paint is used for many large objects, such as ships and bridges, as shown below.
- Grease or oil may be used to coat the moving parts of machinery.
- Metal coatings are used on steel.

There are two types of such coatings.

Metals can be protected with plastic, paint, grease or another metal.



Painters must be permanently employed to ensure that the Sydney Harbour Bridge is protected against corrosion.

Sacrificial coating uses a coating of a more reactive metal.

A coating using a metal that is less reactive than steel is called noble coating, and it serves purely as a covering. Any scratch or break in the coating results in rapid corrosion of the exposed steel. Noble coating is used where scratching is unlikely and where zinc coating is not suitable, as zinc salts are poisonous. Steel food containers are plated with tin, and chromium is used as a noble coating on steel parts in cars and household items.

A coating using a metal that is more reactive than steel is called a sacrificial coating. A break in a sacrificial coating results in the formation of an electrochemical cell. The coating corrodes and the steel is protected. Zinc is the most common form of sacrificial coating. Objects coated in this way are said to be galvanised. When used on an iron roof, for example, the zinc coating reacts with carbon dioxide in air to form a layer of basic zinc carbonate, $Zn(OH)_2.ZnCO_3$, over its surface, thus protecting it and slowing its rate of corrosion. Prepainted galvanised steel sheets for roofs, walls and household appliances are now commercially available. Australians make extensive use of galvanised iron for roofing, guttering, pipes, rubbish bins, fencing wire and nails.

A worker pressing galvanised sheet steel at a steelworks. The zinc prevents the steel from rusting even if the surface is scratched or broken. This is because the more reactive zinc loses electrons, rather than the iron in the steel.

Reactivity of metals

The reactivity of metals can be determined experimentally by how readily they react with oxygen, water, steam and dilute acids. Using these reactions, the activity series of metals can be established. Potassium and gold are at opposite ends of the reactivity series of metals, shown in table 4.4 below.

TABLE 4.4 Reactivity series of metals

Element	Appearance of metal	Reaction with oxygen	Reaction with water or steam	Reaction with dilute acids
K	dull, stored under oil	oxidises in air at room temperature to give oxides	hydrogen formed from cold water	hydrogen formed violently
Na				
Ca	generally dull			hydrogen formed with cold acid
Mg			hydrogen formed with steam	
Al				
Zn				
Fe				
Sn		oxidises when heated in air or oxygen to give oxides	no reaction at Bunsen burner temperatures	
Pb				no visible reaction
Cu	generally shiny			no reaction
Hg			no reaction	
Ag		no reaction		
Au				

Note: Aluminium metal often appears to be less reactive than the reactivity series indicates because it has a coating of aluminium oxide and this protects it from reacting further with oxygen. This coating can be thickened by a process called anodising.

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Unit 1
AOS 1
Topic 5
Concept 4

Reactivity of metals
Summary screen and practice questions

Reactivity of metals and the periodic table

Reaction with oxygen

The most reactive metals are found on the very left of the periodic table, and reactivity increases down the group. Sodium reacts directly with air and so it is stored in oil to prevent it reacting in air or water.

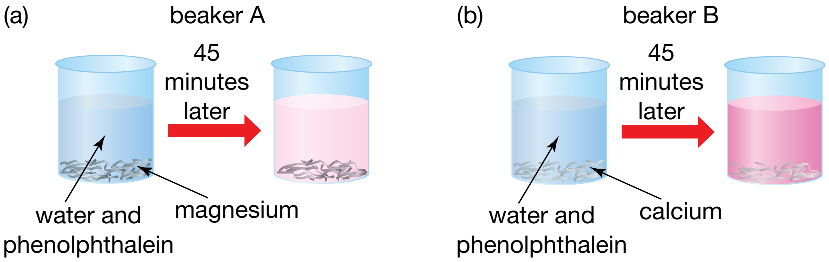
$$\text{metal} + \text{oxygen} \longrightarrow \text{metal oxide}$$
$$\text{sodium} + \text{oxygen} \longrightarrow \text{sodium oxide}$$
$$4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$$

Reaction with water

Group 1 metals are called alkali metals because they react readily with water to produce alkalis (soluble hydroxide compounds) and hydrogen gas.

$$\text{group 1 metal} + \text{water} \longrightarrow \text{group 1 metal hydroxide} + \text{hydrogen}$$
$$2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$$

The experiment illustrated below shows how the relative reactivity of magnesium and calcium can be observed. When magnesium and calcium react with water in beakers A and B as shown, hydrogen gas and a hydroxide compound are formed. The presence of the hydroxide compound is indicated by phenolphthalein. After 45 minutes, beaker B has a more intense pink colour, suggesting that more hydroxide compound has been produced by the reaction. We can deduce from this that calcium is more reactive with water than is magnesium.



Reaction with acid

Hydrogen is also produced when the more reactive metals react with acids.

$$\text{metal} + \text{acid} \longrightarrow \text{ionic salt} + \text{hydrogen}$$
$$\text{iron} + \text{hydrochloric acid} \longrightarrow \text{iron(II) chloride} + \text{hydrogen}$$
$$Fe(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2(g)$$

studyon

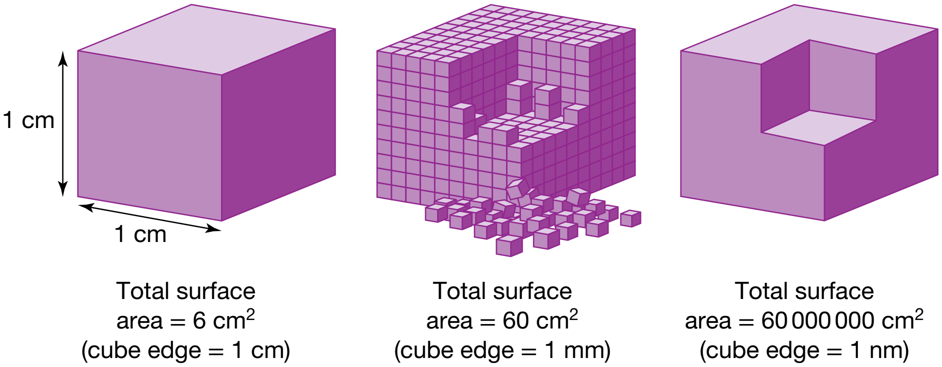
Unit 1
AOS 1
Topic 5
Concept 7

Metallic nanomaterials — silver and gold
Summary screen and practice questions

Metallic nanomaterials

The reactivity of a metal is influenced by the size of its particles. The very high surface area to volume ratio of nanoparticles increases reactivity and catalytic ability because of the greater contact with reactant particles.

As the size of the particles decreases, the surface area increases. This has significant effects on the properties of nanomaterials.



Because of their extremely small size and, therefore, large surface area, the properties of metallic nanoparticles are different from those of their corresponding bulk materials.

Socks that are permeated with nanosilver and can be worn for extended periods of time may or may not be appealing to many people. Due to the presence of silver nanoparticles that have antibacterial properties, much of the odour is eliminated. Some clothing, food packaging, surfaces, appliances, bedding and even baby bottles have also undergone the same treatment. The antimicrobial properties of nanosilver make it useful for bandages, dressings and surfaces in hospitals. Its proven effectiveness is due to its extremely large surface area, a characteristic of nanoparticles in general. However, there is evidence of increasing bacterial resistance to nanosilver and concern about the effects of its use on the environment.

Nanoparticles have a diameter between 1 and 100 nm. In just 1 millimetre, there are 1 million nanometres. Metallic nanomaterials are not new. Nanogold has been used in Roman pottery and in medieval stained glass windows, as have nanocopper and nanosilver. Scientists are interested in their huge potential in biomedical science and engineering. The properties discussed so far in this chapter have referred to what is described as bulk materials or at the macro level. At the macro level, solid samples can contain billions of atoms whereas a nanoparticle has between 10 and 70 000 atoms. Metallic nanomaterials have the same metallic structure but the number of atoms in a sample is very much smaller; this results in different properties. At the nanoscale, many of the classical laws of physics do not operate; it is here that quantum effects apply.

Nanosilver is a useful antibacterial agent.

Nanomaterials look different from and react faster than bulk samples. They may be stronger and better conductors of electricity; they are good catalysts. In bulk metals, the high density of electrons on the surface causes light to be fully reflected. When the particles are much smaller, some parts of visible light are absorbed, resulting in the appearance of a colour that depends on the size of the nanoparticles. Nanosilver is yellow; nanogold is red for particles of about 20 nm diameter but can be orange or even blue if the particle size is different. Bulk gold boils at 1064 °C but nanogold boils at a few hundred degrees. Nanogold is used in electronic chips and in diagnosis and treatment of disease. It is also an effective catalyst.

The colour of nanogold depends on the size of the clusters of particles.

In the future, metallic nanotechnology has significant potential to contribute to smaller and faster electronics, biotechnology, the environment and industry. Understanding the implications of nanotechnology is essential to minimising risks to health and the environment.

Connecting chemistry to society

Titanium bicycles — a material advantage?

Materials research has become an important component of competition preparation in many sports. Track and field athletes look for running shoes that provide stability, support and track grip. Tennis players seek the ultimate racquet that will give them a larger 'sweet spot' (the optimum position on the mesh from which to hit the ball), greater power and more control. Bicycle frames have recently been built from steel, aluminium and carbon fibre composites. A new material has now emerged as the preferred frame material — titanium. Each frame material has its advantages and disadvantages, as shown in table 4.5.

Titanium is preferred by cyclists for use in bicycle frames due to its combination of toughness, elasticity and resilience. Bicycles that are built stiffly, so that they resist pedalling forces, often provide a harsh, uncomfortable ride. Although titanium frames are very stiff against high pedalling forces, they seem to transmit much less road shock than bicycles made of other materials. Steel, for example, has a significantly higher density than titanium, but shock waves travel faster in steel than in titanium. Despite the fact that titanium's high vibration-dampening properties are not yet fully understood, we do know that titanium provides three things crucial to cyclists: low weight, stiffness and a smooth ride.

Titanium is a lustrous, silvery metal with a high melting point (1667 °C). Found as an ore, titanium is quite abundant in the Earth's crust, ranking ninth of all the elements. Metals are usually extracted from their ores by heating them with carbon. The extraction of titanium, however, is quite difficult since carbon cannot be used due to the formation of titanium carbides. Titanium's unusual ability to stretch makes it hard to shape by machine. It tends to push away from even a very sharp cutting blade, giving a rather unpredictable final edge.

However, the superior bicycle that results is worth all these difficulties.

Revision questions

24. **Choose a sport that uses metals in its equipment. Draw a timeline to show how the materials used to make the equipment have been modified over time.**

TABLE 4.5 Materials used in bicycle frames

Material	Advantages	Disadvantages
aluminium	• light • stiff	• low resistance to repeated stresses
steel	• strong • economical • easily shaped to required specifications	• rusts
carbon fibre composites	• high strength-to-mass ratio • good shock-dampening power • good vibration-dampening power	• very expensive
titanium	• less than half the density of steel • excellent resistance to metal fatigue • high corrosion resistance	• expensive • difficult to shape

25. **Use the internet to find out how titanium is modified for use in different applications. Tabulate your findings then use your table to produce a pie chart or graph. The following key words could be used in your search:**
- **titanium usage research**
- **sporting materials.**

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Unit 1	**Extracting a metal from its ore**
AOS 1	
Topic 5	Summary screen and practice questions
Concept 5	

Extraction of metals

The method used to extract metals from their ores depends on their reactivity.

The reactivity of a metal also influences how easily it is extracted. Silver and gold, for example, are unreactive and so are found naturally in their elemental state, which is why they have been known and used since ancient times. More reactive metals must be extracted from their ores. Ores are rocks that are mixtures of different compounds and from which it is economically viable to separate out the metal.

TABLE 4.6 Reactivity metals and type of extraction

Reactivity	Metal	Extraction
very reactive	K, Na, Ca, Mg, Al	electrolysis of molten compound, using electricity to cause a chemical reaction
reactive	Zn, Fe, Sn, Pb	heating the metal ore with carbon
less reactive	Cu, Hg, Ag	heating the metal ore in air
unreactive	Ag, Pt, Au	occur as free elements

Extraction of iron

Our society is highly dependent on the use of iron. Iron is needed to manufacture steel for transport and construction and to a lesser degree for machinery, containers and appliances. About 5% of the Earth's crust is iron; it is the fourth most abundant element after oxygen, silicon and aluminium. Because iron is a metal of medium reactivity, it is extracted from its ores by reaction with carbon in a blast furnace. Iron can be extracted from the following compounds, which are obtained from the ores in parentheses: Fe_3O_4 (magnetite), Fe_2O_3 (haematite), FeS (iron pyrite), $Fe_2O_3.3H_2O$ (cimonite), $FeCO_3$ (siderite).

The blast furnace

The first stage in the manufacture of iron or steel is preparing the ore (the rock that contains the iron) for the blast furnace. It has to be crushed to fist-sized lumps of rock. These rocks are heated to drive off any water or other impurities that will evaporate.

Molten iron from a blast furnace

The second stage involves the reduction of iron oxides to iron in the blast furnace. To do this, carbon in the form of coke and limestone (calcium carbonate) are added.

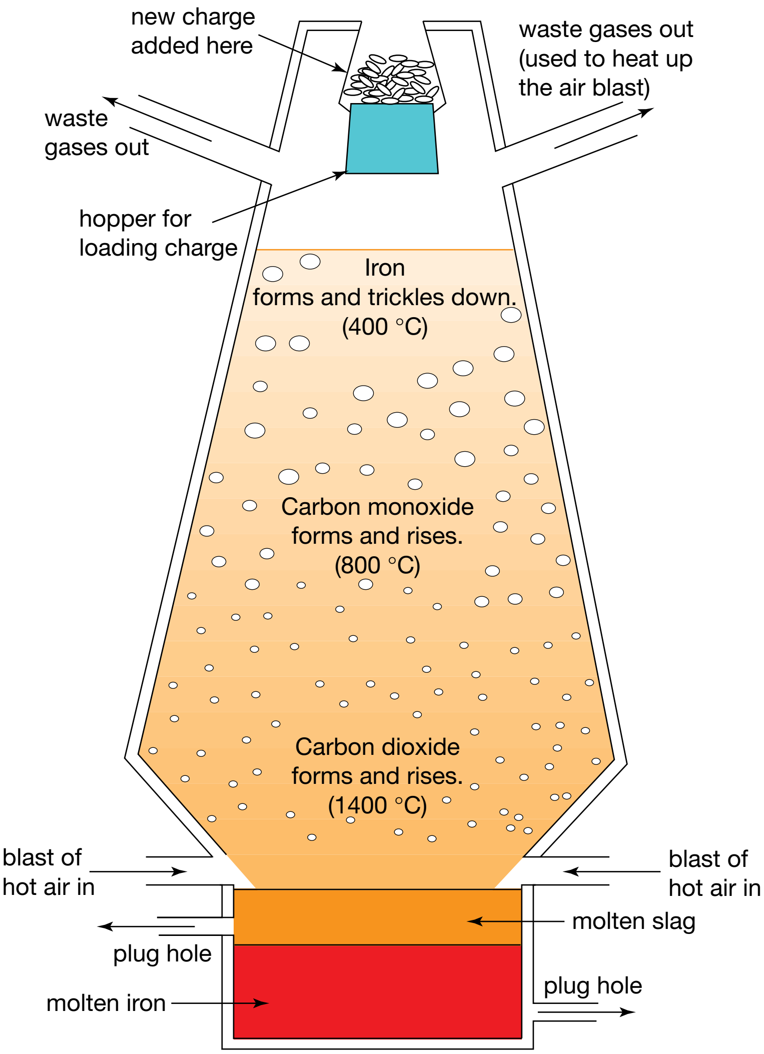
The furnace itself is a tapered cylindrical tower about 40 metres high. It is made of steel and is lined with bricks.

A mixture of iron ore, coke and limestone is added at the top of the furnace in such a way that no gas escapes while it is being added. At the same time, blasts of hot air at about 720 °C are blown into the furnace through small holes near the base. The purpose of the air is to burn the coke to form carbon monoxide. The carbon monoxide reduces the iron oxide to iron in a reaction that also produces carbon dioxide and great quantities of heat.

The iron falls to the bottom of the furnace and eventually melts at about 1500 °C. It is drained off every few hours.

The role of the limestone is to remove impurities inside the furnace. Impurities such as silica, SiO_2, and alumina, Al_2O_3, are removed. The substances formed are known as 'slag'. The slag floats on top of the iron and is drained off at a different level.

The iron obtained from a blast furnace is not completely pure and is called 'pig-iron'. It is hard and brittle and melts at about 1300 °C.



(a) The coke reacts with oxygen in the air, giving carbon dioxide:

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

(b) The limestone decomposes to calcium oxide and carbon dioxide:

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

(c) The carbon dioxide reacts with more coke, giving carbon monoxide:

$$C(s) + CO_2(g) \longrightarrow 2CO(g)$$

(d) This reacts with iron oxide in the ore, giving liquid iron:

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(l) + 3CO_2(g)$$

The iron trickles to the bottom of the furnace.

(e) Calcium oxide from step (b) reacts with sand in the ore, to form calcium silicate or slag:

$$CaO(s) + SiO_2 \longrightarrow CaSiO_3(l)$$

The slag runs down the furnace and floats on the iron. The slag and iron are drained from the bottom of the furnace.

Iron production in a blast furnace

Revision question

26. **State the chemical name for the iron compounds obtained from the different iron ores.**

Carbon in the form of coke (made from coal) and limestone (calcium carbonate) reacts with iron oxides in the blast furnace to produce iron.

Impacts of iron production on society, the economy and the environment

The benefits to society of iron production include not only providing essential materials but also offering employment in mining, production, transport and sales. Exporting iron ore contributes to the economy and therefore the wealth of the nation. These benefits must be balanced with the costs of obtaining and transport of raw materials and the energy requirements for the process. The cost to the environment may also be considerable. The destruction of the landscape for open-cut mining, the significant waste materials that must be disposed of, and the particulate matter and waste gases all contribute to environmental damage. The waste carbon dioxide contributes to the greenhouse effect, and sulfur dioxide gas contributes to acid rain. Recycling iron can reduce some of these problems. Less energy is used to recycle metals than the original process, and resources are conserved, resulting in lower costs and less pollution.

Revision questions

27. **Find out what methods are used by industry to minimise the effects on the environment of iron production.**
28. **What differences would you notice in your daily life if iron were not available?**

Australia is one of the world's largest producers of iron ore; most of it is mined in Western Australia.

Chapter review

studyon

Unit 1	**Metallic bonding**
AOS 1	Summary screens and practice questions
Topic 5	

Summary

- When metallic atoms combine, a lattice is formed consisting of cations in a sea of delocalised electrons. The electrostatic attraction between the cations and electrons is called metallic bonding.
- The metallic lattice structure influences the properties of metals, which, in turn, determine their use.
- Metals are generally:
  - lustrous due to the mobile electrons being able to reflect light
  - good conductors of heat due to the mobile electrons
  - able to transfer the kinetic energy in hotter areas of the metal to cooler areas
  - good conductors of electricity due to the electrons being able to move when an electric field is applied
  - malleable and ductile because the attractive forces that bond the cations and mobile electrons occur in all directions, so that one layer of metal atoms can easily slip over another, without breaking the bonds
  - high-density substances because their lattices are close packed
  - hard and tough and have high tensile strength and high melting points due to the strong attraction between the electrons and cations.
- Alloys are made by melting metals or other materials together and cooling the mixture.
  - The alloy has different properties from that of the original materials.
  - There are two types of alloy: substitutional alloys, where the metal atoms are the same size, and interstitial alloys, where the metal atoms differ in size.
- The properties of metals in different blocks of the periodic table show variations from the general properties.
  - *s*-block metals are reactive and some have lower density.
  - *p*-block metals have varied properties.
  - *d*-block metals usually form coloured compounds and are useful as catalysts.
- The ball bearing model was developed to explain different metal strengths.
  - Metals with large areas of perfect close packing (grains) have fewer dislocations (flaws in the packing) and they bend easily.
  - Metals with small grains have many dislocations and do not bend easily.
- Metals can be modified by:
  - work hardening. Work-hardened or toughened metals are more brittle because the crystal grains are smaller and have more dislocations.
  - different heat treatment
  - alloying
  - coating. Coating is used to improve usefulness and prevent corrosion.
- The three types of heat-modified metal are:
  - annealed metals, which are heated and then cooled slowly to produce larger crystals and hence a softer metal
  - quenched metals, which are heated until red hot and then cooled quickly in cold water to produce smaller crystals and hence a harder but more brittle metal
  - tempered metals, which are made by warming a quenched metal and allowing it to cool slowly to reduce brittleness but retain hardness.
- Metals vary in their reactivity. Reactivity decreases across the periods and increases down a group of the periodic table.
- Because of their extremely small size and resulting large surface area, the properties of metallic nanoparticles are different from those of the corresponding bulk materials, including colour and reactivity.
- Nanosilver is a useful antibacterial agent.
- Nanogold has beneficial applications in disease detection and treatment.
- The method of metal extraction from their ores depends on their reactivity. Iron is a reactive metal and is extracted by reaction with carbon in a blast furnace.
- Iron is an essential material for construction and manufacturing, but its production has impacts on the environment, including land degradation and pollution. These effects may be reduced by collecting waste gases and revegetation.

Multiple choice questions

1. A characteristic of metals is that their atoms:

A are smaller than those of non-metals
B readily form ionic bonds with atoms of other metals
C have few electrons in their outer shells, and these are readily lost in chemical reactions
D have high electronegativity.

2. The structure and bonding in solid sodium metal at room temperature can best be described as:

A a network lattice of closely packed sodium atoms, held together by strong ionic bonds
B a lattice of Na_2 molecules
C a network lattice of sodium ions, held together by a 'sea' of electrons
D a lattice of Na_2 molecules held together by weak bonds.

3. Which of the following properties is *not* characteristic of metals? They are:
- **A** good conductors of heat
- **B** capable of being polished
- **C** relatively dense solids at room temperature
- **D** hard and brittle.

4. The property of metals that enables them to conduct an electric current is:
- **A** the outer electrons of metal atoms are not firmly bound to the atom
- **B** metal atoms are better suited as charge carriers than non-metal atoms
- **C** ions in the metal can move freely through the metallic lattice
- **D** metal atoms are not as firmly bonded to each other as are non-metallic atoms.

5. The physical properties of solid metals can best be explained by proposing that:
- **A** each metal atom is held in the crystal lattice by covalent bonds
- **B** positive metal ions are arranged in an orderly way, with valence electrons able to move freely throughout the crystal lattice
- **C** positive and negative metal ions are arranged in an orderly way, with valence electrons able to move freely through the crystal lattice
- **D** each metal atom is surrounded by a variable number of valence electrons, which complete a 'noble gas' electronic structure in the crystal lattice.

6. The term 'malleable' is applied to a metal that:
- **A** can be hammered into thin sheets
- **B** resists corrosion
- **C** can be drawn out to form thin wire
- **D** can be used to form alloys.

7. Both copper and sodium chloride have high melting points. Copper, in its solid and liquid states, is a good conductor of electricity while sodium chloride conducts electricity only when molten or dissolved in water. This difference in electrical conductivity can be explained by the fact that:
- **A** there are much stronger bonding forces between atoms in sodium chloride than between those in copper
- **B** electrons in solid copper have much greater mobility than those in solid NaCl
- **C** chlorine is a poor electrical conductor
- **D** copper has only one kind of atom.

8. The structure of a solid formed by heating a mixture of two metallic elements at high temperature is likely to be:
- **A** a metallic solid called an alloy
- **B** an ionic solid
- **C** a mixed ionic lattice and metallic lattice
- **D** separate molecules of each element.

9. Which of the following is *not* a possible mechanism for the formation of an alloy of two metals?
- **A** Atoms of one metal fit into the interstitial spaces between atoms of the other.
- **B** A new element is formed.
- **C** Atoms of one metal replace atoms of the other in the crystal lattice.
- **D** Separate crystals of one metal are dispersed throughout the other.

10. Which element can be found in nature in elemental (uncombined) form?

A	K	**C**	Au
B	Ca	**D**	Al

11. Aluminium, although a fairly reactive metal, does not react with oxygen gas in air because:
- **A** it is protected by small amounts of its own impurities
- **B** its oxide forms a hard protective layer on the metal
- **C** nitrogen and carbon dioxide gas in the air prevent oxidation
- **D** it forms a hard protective layer by a reaction with sulfur in the air.

12. Iron rubbish bins coated with a complete layer of zinc do not rust because:
- **A** the zinc acts as a sacrificial metal, reacting to prevent the iron rusting
- **B** iron is a more reactive metal than zinc
- **C** the zinc combines with the iron to form a new compound that does not rust
- **D** particles of rust are unable to stick to the zinc surface.

13. Consider the following:
- Metal G slowly reacts with cold water.
- Metal H is generally shiny.
- Metal E readily reacts with oxygen.
- Metal F produces hydrogen when it reacts with acid.

What is their reactivity order, from the most to the least reactive?
- **A** H > G > F > E
- **B** H > E > F > G
- **C** F > H > G > E
- **D** E > G > F > H

14. Which is formed by the reaction of zinc with hydrochloric acid?
- **A** zinc hydroxide + hydrogen
- **B** zinc sulfate + chlorine
- **C** zinc chloride + hydrogen
- **D** zinc chloride + water

15. When a piece of sodium metal is carefully added to water, it reacts vigorously. A piece of caesium metal would:
- **A** react with about the same vigour
- **B** react more vigorously
- **C** react less vigorously
- **D** not react at all with the water.

16. Stainless steel is made by alloying iron, nickel and:
- **A** aluminium
- **B** copper
- **C** magnesium
- **D** chromium.

17. Which of the following is a metal that reacts with dilute acids but not cold water?
- **A** Potassium
- **B** Gold
- **C** Iron
- **D** Carbon

Review questions

Properties and structure of metals

1. The electron configurations of elements *A*, *B*, and *C* are as follows.
- *A*: 2, 6
- *B*: 2, 8, 1
- *C*: 2, 8, 5

Which of these elements is a metal?

2. Which of the following substances contain metallic bonding?
- (a) lithium
- (b) sulfur
- (c) sodium bromide
- (d) mercury(II) fluoride
- (e) calcium
- (f) argon

3. Explain the following properties of metals in terms of the metallic structure.
- (a) Metals are conductors of heat and electricity.
- (b) Metals are malleable.
- (c) Metals are ductile.
- (d) Metals are shiny.
- (e) Metals have high density.

Properties and uses of metals

4. Woks used in cooking can be made from stainless steel or normal steel. It is believed that normal steel enhances the flavour of the food. The wok is washed without an abrasive, dried, and then stored with a thin coating of oil on its inner surfaces.

Regular saucepans, on the other hand, are usually made of the much more expensive stainless steel and often have a copper base.
- (a) Discuss the properties of these metals with reference to their use in the kitchen.
- (b) Why do people coat the woks with oil?
- (c) What other metals are used in the kitchen?
- (d) Draw a table that relates each of the metals you named in part (c) to its best use in the kitchen.

5. (a) Why was gold one of the first metals to be used?
- (b) Why was it not used for weapons?

6. Over the years, dentists have used metals for tooth fillings: first gold and then an amalgam of mercury. This amalgam is made by shaking a powdered mixture of silver, tin, copper and zinc with mercury. Within seconds, the mercury dissolves to form an amalgam that the dentist packs into the cavity. This alloy hardens in two hours.

Copy the following table and use arrows to match the property of each metal with those of the alloy used in a dental amalgam.

Metal	Property
silver	• can be poured into the tooth easily
tin	• does not tarnish or react with food or drink
mercury	• bonds to the mercury and helps the amalgam set
copper and zinc	• helps the amalgam stay free from tarnish
gold	• gives the amalgam strength

Modifying metals

7. (a) What is the difference between a metal and an alloy?
- (b) Give examples of three alloys and indicate what advantages they have over the metal.

8. State, with reasons, whether each of the following substances is an element, a compound or an alloy.
- (a) copper
- (b) brass
- (c) gold
- (d) 18 carat gold
- (e) silver bracelet
- (f) iron
- (g) bronze
- (h) copper oxide
- (i) rust
- (j) oxygen
- (k) solder
- (l) stainless steel
- (m) tap water

9. Although mercury and mercuric compounds are toxic, mercury can safely be used in a tooth amalgam. What does this suggest about mercury and the amalgam?

10. Steel is an alloy of iron and carbon. Using steel as an example, explain why alloys are desirable.

11. Identify from which block (*s*, *p* or *d*) each of the following metals is found.
- (a) It corrodes easily, forming a green compound.
- (b) It has a partially filled *p* subshell and low density.

  (c) It has a low melting point and can be cut easily with a knife.
  (d) It is shiny and grey and, in nano form, has good antibacterial properties.
  (e) It forms a compound that is white and has an subshell electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$.
  (f) It is a reddish orange metal that forms blue compounds and is a good conductor of electricity.

12. Explain why it is difficult to straighten a bent metal coathanger.

13 Explain the difference between tempering and annealing a metal.

14. Explain, in terms of their bonding, why quenched metals are harder than unquenched metals.

15. Describe a situation where each of the following coatings may be used.
  (a) Plastic
  (b) Oil
  (c) Noble
  (d) Sacrificial

16. Describe an experiment that could be used to find the position of nickel in the activity series of metals.

17. Write fully balanced equations, including states, for the following reactions.
  (a) potassium + oxygen
  (b) calcium + water
  (c) tin + nitric acid, HNO_3

18. Complete the following table by placing ticks in the appropriate column.

Property	Nanogold	Bulk gold	Both	Neither
(a) reactive				
(b) magnetic				
(c) spreads through water				
(d) range of colours				
(e) good conductor				
(f) very large surface area				
(g) insoluble				
(h) unreactive				
(i) does not conduct				
(j) transition metal				

19. Iron is an important metal for industry and consumers. It is used in greater quantity and variety than any other metal.
  (a) List four important uses of iron.
  (b) List four properties of iron that are relevant to the uses in part (a).
  (c) Why is iron used so extensively?
  (d) Why is iron usually alloyed with other metals?
  (e) Outline the reactions in the blast furnace involving carbon.

Exam practice questions

In a chemistry examination, you will be required to answer a number of short and extended response questions.

studyon

Unit 1
AOS 1
Topic 5

Metallic bonding

Sit topic test

Extended response questions

1. The solders A, B and C have the following properties:

A is a liquid above 180 °C

B melts gradually from 180 °C to 230 °C

C has a very high melting point.

Which solder would you use for each of the following jobs?

(a) Mending car radiators

(b) Joining wires in an electric circuit

(c) Filling gaps in car bodies **3 marks**

2. Iron is an important metal for industry and consumers. It is used in greater quantity and variety than any other metal. There are several reasons for this:

- Iron is malleable and ductile when hot, so that means that different articles may be made from it.
- Iron forms a variety of alloys that may have vastly different properties.
- Although pure iron is relatively soft, when it is alloyed with chromium, its hardness and strength increase.
- Iron is the cheapest metal.

(a) Suggest why iron is the cheapest of all metals. **2 marks**

(b) What is the name of the alloy described in the third bullet point, and what is another property of this alloy? **2 marks**

(c) List four important uses of iron. **2 marks**

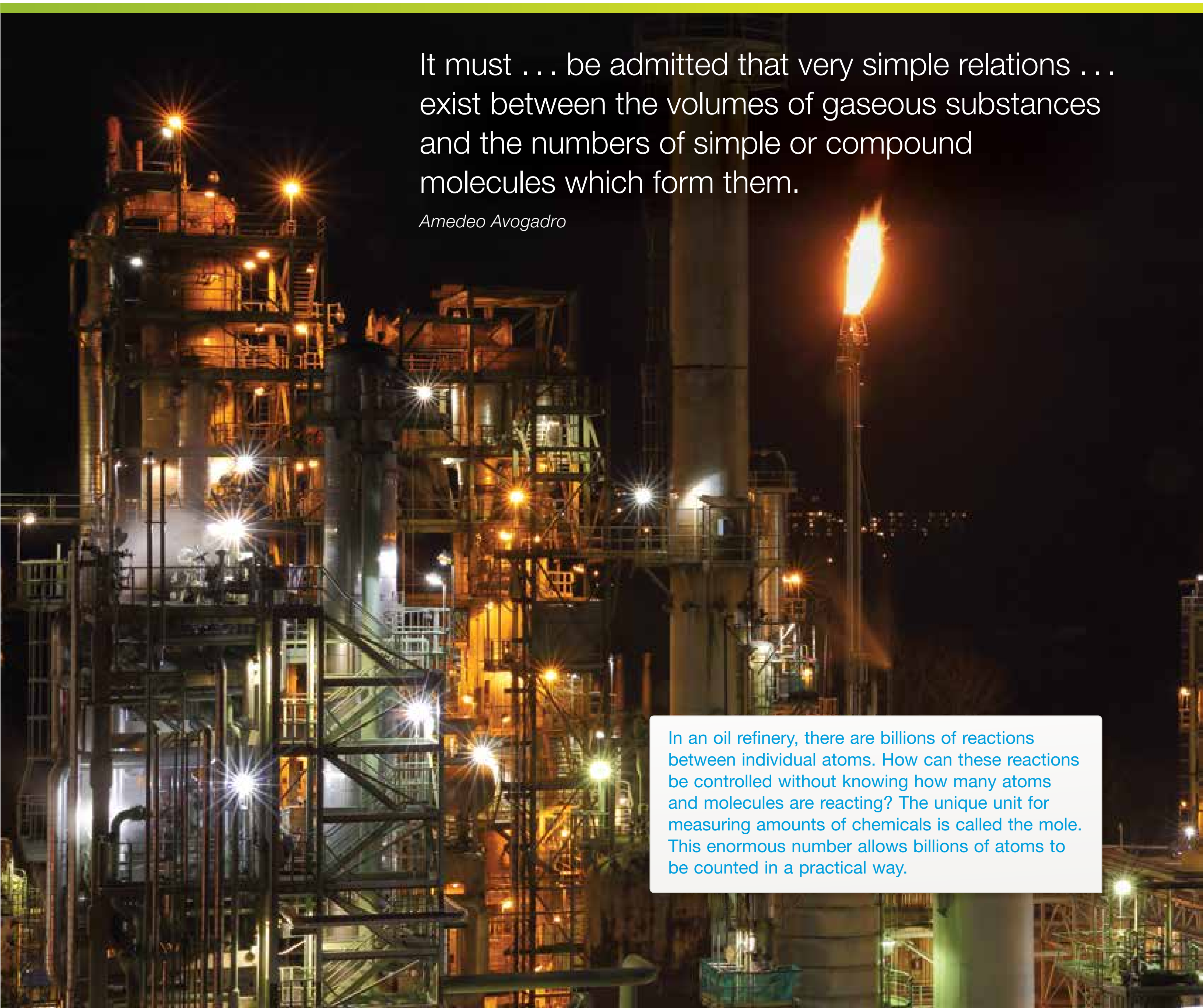
CHAPTER

5 Quantifying chemistry

YOU WILL EXAMINE:

- the use of mass spectrometry in identifying isotopes of elements
- how to calculate relative atomic masses
- the mole concept
- Avogadro's constant and its relationship to the mole concept
- how to calculate molar masses
- how to determine percentage composition and empirical formulas.

> It must . . . be admitted that very simple relations . . . exist between the volumes of gaseous substances and the numbers of simple or compound molecules which form them.
>
> *Amedeo Avogadro*



In an oil refinery, there are billions of reactions between individual atoms. How can these reactions be controlled without knowing how many atoms and molecules are reacting? The unique unit for measuring amounts of chemicals is called the mole. This enormous number allows billions of atoms to be counted in a practical way.

John Dalton devised his own set of symbols for the known elements. Note the relative weights of the elements that he calculated based on hydrogen having a mass of 1.

Measuring atoms

Atoms are extremely tiny. Even though the air is full of oxygen and nitrogen molecules, you cannot see them. We know a lot about atoms and molecules, and this knowledge is invaluable when explaining the properties of substances. But, how do we measure atoms? The scale of atomic size means that chemists rarely deal with atoms one at a time. What is needed is a convenient way to compare and measure masses and amounts of particles in elements and compounds. The mole concept, which is central to most chemical calculations, helps us to work with the vast numbers of atoms that are present in different types of substances.

Atoms are extremely small but they still have mass. The problem is to find a way of measuring that mass. In 1803, English chemist John Dalton described matter as made up of particles that were solid, indivisible and having weight. He prepared a table of atomic weights by experimentally determining relative weights of elements in compounds; he based his figures on hydrogen having a mass of 1. We now use the term 'mass' instead of 'weight'.

This reference standard was later changed from hydrogen to oxygen, but this led to a disagreement between chemists, who used natural oxygen as the standard, and physicists, who used the oxygen-16 isotope only. Having two slightly different lists of atomic masses caused many problems. An agreement was reached in 1961 to change the standard to carbon-12, and this settled the dispute; this change had the added advantage that carbon-12 can be measured very accurately because it is stable and abundant. The development of an instrument called the mass spectrometer allowed scientists to accurately compare the masses of all atoms.

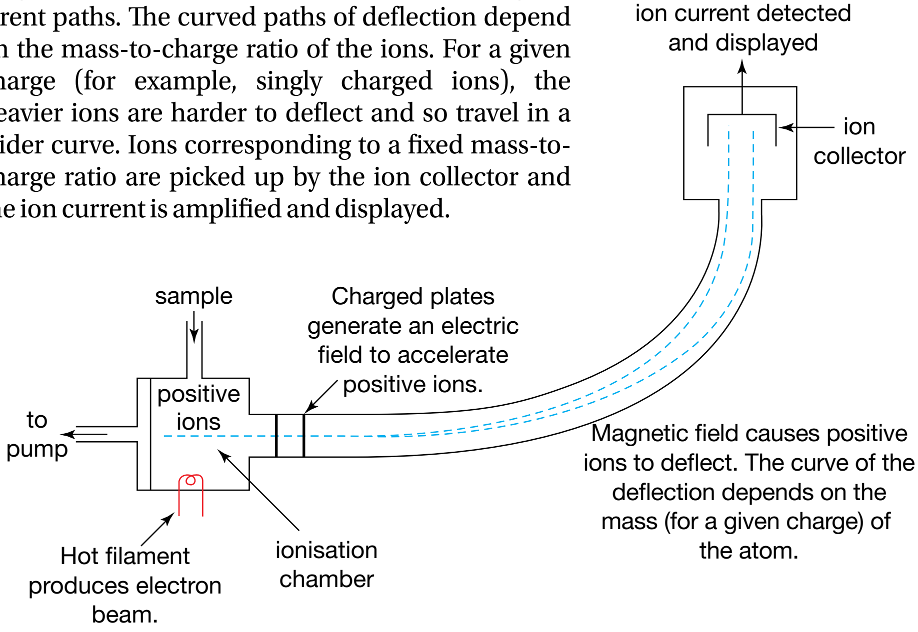
Carbon-12 is used as the standard for comparing relative isotopic masses.

Measuring masses of atoms

In 1913, Joseph John Thompson, an English physicist, discovered that some elements can have isotopes. Francis Aston, another English scientist, developed Thompson's equipment into a **mass spectrometer**. Francis Aston was later awarded the Nobel Prize for identifying the 212 naturally occurring isotopes by comparing the relative masses of atoms.

A mass spectrometer deflects the path of the lightest element the most. Similarly, it deflects the most highly charged particle the furthest.

The figure below shows the main features of a mass spectrometer. A sample of the element to be analysed is injected as a gas into the ionisation chamber, where the atoms are ionised by bombardment with electrons produced by the hot filament. The positive ions formed are accelerated through an electric field and deflected in a magnetic field that forces the ions to travel along different paths. The curved paths of deflection depend on the mass-to-charge ratio of the ions. For a given charge (for example, singly charged ions), the heavier ions are harder to deflect and so travel in a wider curve. Ions corresponding to a fixed mass-to-charge ratio are picked up by the ion collector and the ion current is amplified and displayed.



A mass spectrometer. Different types of atoms of the same element can be separated on the basis of their mass:charge ratio.

The mass spectrometer provides us with information about:

- the number of isotopes in a given sample of an element
- the relative isotopic mass of each isotope
- the percentage abundance of the isotopes.

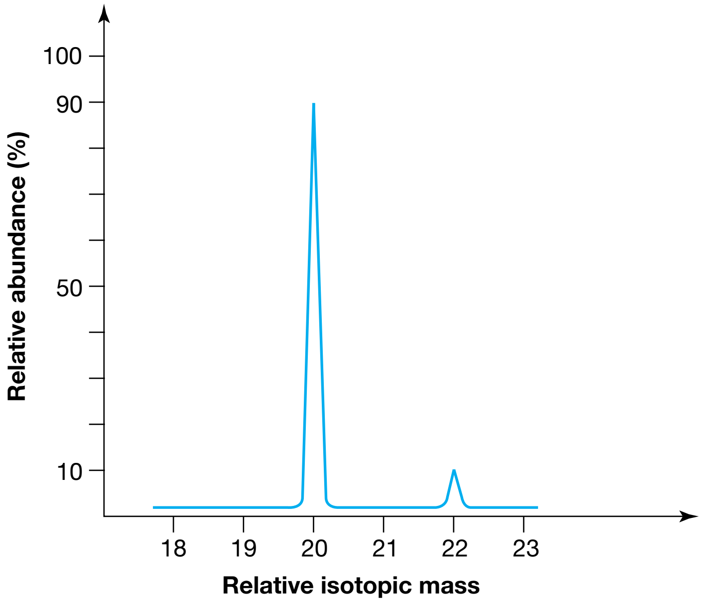
The **relative isotopic mass** is the mass of a single isotope and is determined by comparing the mass of ions of the isotope to the value of a standard, $^{12}_{6}\mathrm{C}$ (which has been assigned a mass of 12 exactly).

A mass spectrum shows information about the relative isotopic mass and percentage abundance of each isotope.

The figure below shows the mass spectrum for neon. The two peaks in the trace represent the two isotopes of neon. Both isotopes of neon have an atomic number of 10 (that is, the number of protons is 10), but one isotope has a relative isotopic mass of 20 and the other has a relative isotopic mass of 22. The atoms of neon-20 have 10 protons and 10 neutrons, and atoms of neon-22 have 10 protons and 12 neutrons. The number of protons added to the number of neutrons in an atom is called its **mass number**.

study on

Unit 1	**Relative isotopic mass and relative atomic mass**
AOS 1	Summary screen and practice questions
Topic 3	
Concept 1	



A mass spectrum for the element neon. The element neon is made up of two isotopes of different proportions. 90% of the element neon is made up of the isotope with relative mass of 20, while the remaining 10% is the isotope with relative mass of 22.

study on

Unit 1	**Using mass spectrometry to find relative atomic mass**
AOS 1	Summary screen and practice questions
Topic 3	
Concept 2	

Relative atomic mass

The structure of an atom is identified using the convention $^{A}_{E}E$, where E is the symbol of the element, A is the mass number (number of protons + neutrons), and Z is the atomic number (number of protons). For example, the two isotopes of neon are represented as $^{20}_{10}\mathrm{Ne}$ (atomic number, $Z = 10$; mass number, $A = 20$) and $^{22}_{10}\mathrm{Ne}$ ($Z = 10$; $A = 22$).

The relative abundance of an isotope measures how much of that isotope is present compared with the other isotopes in an element.

$^{A}_{Z}E, A =$ mass number

$Z =$ atomic number

Most elements consist of a mixture of isotopes. The **relative atomic mass** (A_r) of an element represents the average mass of one atom, taking into consideration the number of isotopes of the element, their relative isotopic mass (RIM) and their relative abundance. Using data from the mass spectrum, one could calculate the relative atomic mass for an element by using the following method of calculation:

$$A_r = \frac{(\text{RIM of first isotope} \times \text{abundance}) + (\text{RIM of second isotope} \times \text{abundance}) + \ldots}{100}$$

$$A_r = \frac{\sum \text{RIM} \times \text{abundance}}{100}$$

Because A_r is a relative mass, a unit is not required; however, the unit 'unified atomic mass' (symbol u) can be used. This replaces the old unit 'amu' (atomic mass unit).

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Interactivity
Relative isotopic mass
int-6353

For example, the two isotopes of neon, $^{20}_{10}Ne$ and $^{22}_{10}Ne$, have relative isotopic masses of 20 and 22 respectively. Their relative abundances are 90.0% and 10.0% respectively (see the graph on the previous page).

$$\begin{aligned} A_r(\text{Ne}) &= \frac{(20 \times 90) + (22 \times 10)}{100} \\ &= 20.2 \end{aligned}$$

The relative molecular mass (M_r) of a molecule is the sum of the relative atomic masses, as shown in the periodic table, of elements in the formula. For example:

$$\begin{aligned} M_r(\text{NH}_3) &= A_r(\text{N}) + (3 \times A_r(\text{H})) \\ &= 14.0 + (3 \times 1.0) \\ &= 17.0 \end{aligned}$$

Revision questions

1. **Why is carbon used as a standard to determine relative atomic masses?**

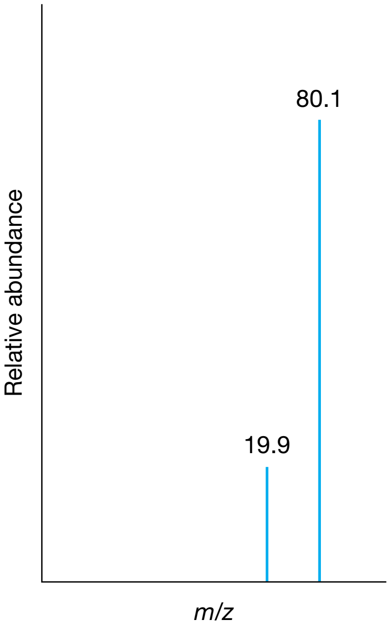
2. **Lithium consists of two isotopes. One isotope, $^{6}_{3}Li$, has a relative isotopic mass of 6.01 and an abundance of 7.42%. The other isotope, $^{7}_{3}Li$, has a relative isotopic mass of 7.01 and an abundance of 92.58%.**
  - **(a)** **Is the relative atomic mass closer to 6 or 7? Explain your answer.**
  - **(b)** **Calculate the relative atomic mass of lithium.**

3. **Copper is widely used for electrical wiring. It has two isotopes: $^{63}_{29}Cu$ and $^{65}_{29}Cu$. The lighter isotope has an abundance of 69.2%. Calculate the relative atomic mass of copper. (Note that, if accurate values of the relative isotopic masses are not provided, use mass numbers.)**

4. **Three isotopes of magnesium and their relative abundances are $^{24}_{12}Mg$ (78.8%), $^{25}_{12}Mg$ (10.2%) and $^{26}_{12}Mg$ (11.0%).**
  - **(a)** **Sketch on a graph the mass spectrum for magnesium.**
  - **(b)** **Calculate the relative atomic mass of magnesium.**

5. **Gallium has two isotopes. One isotope is $^{69}_{31}Ga$ and has a relative abundance of 60.50%. The relative atomic mass of gallium is 69.70. Find the relative isotopic mass of the other isotope.**

6. **The green colour of fireworks can be produced using the element boron. Boron's two isotopes are $^{10}_{5}B$ and $^{11}_{5}B$. Use the mass spectrum on the left to calculate the relative atomic mass of boron.**



Relative abundance of boron isotopes

7. **Calculate M_r of each of the following compounds.**
  - **(a)** **CO_2**
  - **(b)** **NaCl**
  - **(c)** **H_2O_2**
  - **(d)** **H_2SO_4**
  - **(e)** **$C_6H_{12}O_6$**

Counting atoms

In one gram of sugar there are about 1 760 000 000 000 000 000 000 or 1.76×10^{21} sugar molecules. If we count in dozens, that corresponds to 140 000 000 000 000 000 000 dozen molecules. Clearly, in chemistry, we need a measurement that can manage these large numbers of atoms and molecules. The mole concept helps us count atoms and is fundamental to chemical calculations.

Chemists use a unit called the mole to count atoms, ions and molecules. The mole contains a very large number of particles. This number is 6.02×10^{23}.

The mole concept

The term **mole** (symbol mol) represents a number. It is a unit of measurement, just as the term 'dozen' represents 12 and 'kilo' represents 1000. Since there are billions of atoms in any substance, the mole must represent a very large number. Thus, a mole of hydrogen atoms means a certain large number of hydrogen atoms.

A mole is defined as the amount of substance that contains as many particles (atoms, ions or molecules) as there are atoms in exactly 12 g of the $^{12}_{6}\text{C}$ isotope.

The number of carbon atoms in 12 g of $^{12}_{6}\text{C}$ isotope has been experimentally estimated to be 602 000 000 000 000 000 000 000 atoms, or 6.02×10^{23} atoms. This is often called **Avogadro's number (N_A)**. We can therefore say:

The number of atoms in one mole of an element is 6.02×10^{23}.

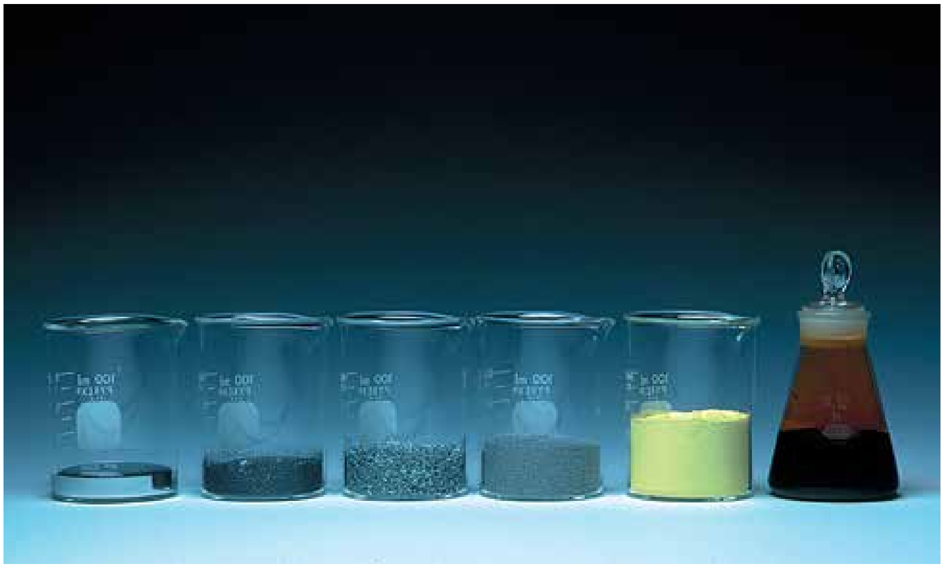
To find the mass of one mole of an element, simply add 'g' to the relative atomic mass of that element.

You may have noticed that the definition of the mole, like the definition of relative isotopic mass (see chapter 2), is based on the $^{12}_{6}\text{C}$ isotope. There is a good reason for this. The relative isotopic scale allows us to deal with atoms 'one at a time'. The mole concept allows us to deal with much larger groups 'one mole at a time'. The common reference to $^{12}_{6}\text{C}$ means that, to find out the mass of one mole (molar mass) of any element, all you have to do is add 'g' to the relative atomic mass.

study on

Unit 1
AOS 1
Topic 3
Concept 3

The mole
Summary screen and practice questions



One mole each of mercury, zinc, silicon, aluminium, sulfur and bromine. This means that the quantity of each of the substances shown contains 6.02×10^{23} particles.

Molar mass (*M*)

The molar mass (M) of an element is defined as the mass of 1 mol of the element: that is, 6.02×10^{23} atoms of the element. The unit is grams per mole (g mol^{-1}). For example:

Molar mass of carbon atoms = mass of 1 mol of C atoms = 12.0 g mol^{-1}

$\therefore$ 12.0 g of carbon contains 6.02×10^{23} atoms of carbon.

Molar mass of oxygen atoms = mass of 1 mol of O atoms = 16.0 g mol^{-1}

$\therefore$ 16.0 g of oxygen contains 6.02×10^{23} atoms of oxygen.

Molar mass of oxygen *molecules* = 2 × mass of 1 mol of O atoms
= 2 × 16.0 g mol^{-1} = 32.0 g mol^{-1}

$\therefore$ 32.0 g of oxygen contains 6.02×10^{23} molecules of oxygen.

32.0 g of oxygen contains 1.20×10^{24} atoms of oxygen.

study on

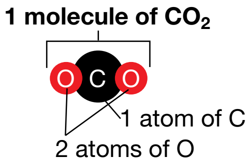
Unit 1
AOS 1
Topic 3
Concept 4

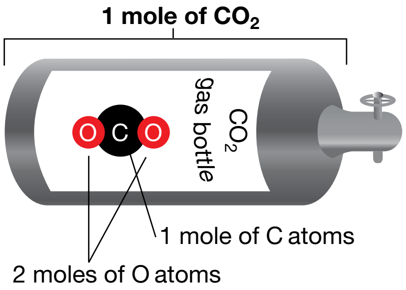
Molar mass
Summary screen and practice questions

Compounds and molar mass

The molar mass of a compound is defined as the mass of 1 mol of the compound expressed in grams per mole (g mol^{-1}).

The molar mass of a compound is the mass of one mole of that compound. It is found by adding together all the relative atomic masses for the atoms in its formula and adding ‘g mol^{-1}’ as the unit.





Molecular compounds

For compounds consisting of molecules, the molar mass (M) is numerically equal to the relative molecular mass (M_r) expressed in grams per mole (g mol^{-1}). For example:

- Molar mass of water molecules = mass of 1.00 mol of H_2O molecules
  = 18.0 g mol^{-1}
  ∴ 18.0 g of water contains 6.02×10^{23} molecules of water.
- Molar mass of chlorine gas = mass of 1.00 mol of Cl_2 molecules
  = 71.0 g mol^{-1}
  ∴ 71.0 g of chlorine gas contains 6.02×10^{23} molecules of chlorine.

Ionic compounds

The M_r of an ionic compound is found by adding the A_r of each atom in the formula of the compound. For example:

$$\begin{aligned} M_r(CuSO_4) &= A_r(Cu) + A_r(Cu) + (4 \times A_r(O)) \\ &= 63.5 + 32.1 + (4 \times 16.0) \\ &= 159.6 \end{aligned}$$

The molar mass (M) of $CuSO_4$ is 159.6 g mol^{-1}.

$$M(CuSO_4) = 159.6 \text{ g mol}^{-1}$$

Whether a substance is made up of atoms, molecules or ions, the same principle applies: one mole of any substance always contains 6.02×10^{23} particles. This number is always 6.02×10^{23}.

Revision question

8. Calculate the molar masses of the following.

(a) ozone, O_3	**(c) calcium oxide, CaO**
(b) iodine, I_2	**(d) hydrochloric acid, HCl**

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Interactivity
The mole concept
int-6354

Weblink
What is mole day?

Changing moles to numbers of particles

We have described a relationship between the number of particles and the mass of the substance. This means that, when we measure the mass of a substance equal to its molar mass (e.g. 18.0 g of H_2O), we are also indirectly counting out 6.02×10^{23} particles of that substance (e.g. water molecules). If the molecule or formula unit contains more than one kind of atom, we can also calculate the number of different atoms present in one mole of the substance. For example:

1.00 mol of carbon dioxide molecules contains:

- 1.00 mol of CO_2 molecules
- 6.02×10^{23} CO_2 molecules
- 1.00 mol of C atoms
- 6.02×10^{23} C atoms
- 2.00 mol of O atoms
- 1.20×10^{24} O atoms.

The molar mass of CO_2 is 44.0 g mol^{-1}.

These observations suggest the following:

$$\text{mass} \underset{\text{multiply by molar mass}}{\overset{\text{divide by molar mass}}{\rightleftarrows}} \text{moles}$$

or

$$n = \frac{m}{M}$$

where n = number of moles, m = mass, M = molar mass

The formula $n = \frac{m}{M}$, where n is the number of moles, m is the mass of the sample and M is the molar mass, can be used to calculate the number of moles of an element or a compound.

The formula

$$n = \frac{\text{number of particles}}{6.02 \times 10^{23}}$$

can be used to calculate the number of moles.

and

$$\text{number of particles } (N) \underset{\text{multiply by } 6.02\times10^{23}}{\overset{\text{divide by } 6.02\times10^{23}}{\rightleftarrows}} \text{moles}$$

or

$$n = \frac{\text{number of particles}}{6.02 \times 10^{23}} = N/N_A$$

Sample problem 5.1

A gas balloon contains 5.5 mol of helium atoms. How many helium atoms are present?

Solution:

$$\begin{aligned}\text{number of particles} &= n(\text{He}) \times 6.02 \times 10^{23}\\ &= 5.5 \times 6.02 \times 10^{23}\\ &= 3.3 \times 10^{24} \text{ atoms of helium}\end{aligned}$$

Revision questions

9. **Write a brief explanation of the mole concept, and explain the importance of Avogadro's number.**

10. **Calculate the number of chlorine atoms or chloride ions in:**
  - **(a)** **2.3 mol of chlorine atoms**
  - **(b)** **15.8 mol of chlorine molecules, Cl_2**
  - **(c)** **3.5 mol of sodium chloride, NaCl**
  - **(d)** **0.50 mol of magnesium chloride, $MgCl_2$.**

11. **Determine how many:**
  - **(a)** **mol of ethanoic acid molecules**
  - **(b)** **molecules of the acid**
  - **(c)** **mol of oxygen atoms**
  - **(d)** **atoms of oxygen**

  are present in 16.2 g of ethanoic acid, CH_3COOH.

Sample problem 5.2

Pure liquid ammonia is called anhydrous ammonia and is used extensively as a fertiliser as it has a high nitrogen content. If we have 4.6×10^{28} molecules of ammonia, NH_3, in a fertiliser plant store, how many moles does this represent?

Solution:

$$\begin{aligned}\text{number of moles } (NH_3) &= \frac{\text{number of } NH_3 \text{ molecules}}{6.02 \times 10^{23}}\\ &= \frac{4.6 \times 10^{28}}{6.02 \times 10^{23}}\\ &= 7.6 \times 10^{4}\end{aligned}$$

Therefore, 7.6×10^4 mol of ammonia molecules is stored.

Revision question

12. **Calculate the number of moles of each of the following particles present in 5.2×10^{24} molecules of methane, CH_4.**
  - **(a)** **methane molecules**
  - **(b)** **carbon atoms**
  - **(c)** **hydrogen atoms**

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Weblink
How big is a mole?

Changing mass to moles and moles to mass

Since one mole of a substance refers to both a *mass* of the substance and a number of particles of that substance (moles), we can calculate one from the other. There are many different ways to refer to the composition of a substance.

Sample problem 5.3

Ethanol, C_2H_5OH, is found in alcoholic beverages. If one such beverage contains 4.6 g of ethanol, how many moles does this represent?

Solution:

$$\begin{aligned} n(\text{ethanol}) &= \frac{\text{mass(ethanol)}}{M(\text{ethanol})} \\ &= \frac{4.6}{46.0} \\ &= 0.10 \text{ mol} \end{aligned}$$

Revision question

13. **Calculate the amount, in moles, of:**
    - (a) **46 g of water, H_2O**
    - (b) **2.4 g of carbon dioxide, CO_2**
    - (c) **67 g of chlorine gas, Cl_2**
    - (d) **2.0 g of sodium chloride, NaCl**
    - (e) **128 g of copper(II) sulfate pentahydrate, $CuSO_4{\cdot}5H_2O$**
    - (f) **38 kg of iron(III) oxide, Fe_2O_3.**

Sample problem 5.4

Pure ethanoic acid, CH_3COOH, can be used to make vinegar when dissolved in water. If 3.5 moles of ethanoic acid was used, what mass was weighed out?

Solution:

Since $n = \frac{m}{M}$, $m = n \times M$.

$$\begin{aligned} \text{mass(ethanoic acid)} &= n(\text{ethanoic acid}) \times M(\text{ethanoic acid}) \\ &= 3.5 \times 60.0 \\ &= 210 \text{ g} \end{aligned}$$

Therefore, 2.1×10^2 g of ethanoic acid was weighed out.

Revision questions

14. **Calculate the mass of:**
    - (a) **0.41 mol of carbon monoxide, CO**
    - (b) **12.0 mol of sulfur dioxide, SO_2**
    - (c) **3.84 mol of sucrose, $C_{12}H_{22}O_{11}$**
    - (d) **58.2 mol of iron, Fe**
    - (e) **0.0051 mol of silver chloride, AgCl**
    - (f) **2.53 mol of magnesium phosphate, $Mg_3(PO_4)_2$.**
15. **Which of the following substances has the greatest mass?**
    - (a) **2.5 mol of hydrogen gas, H_2**
    - (b) **0.2 mol of zinc, Zn**
    - (c) **11.56 g of calcium chloride, $CaCl_2$**

Sample problem 5.5

What is the mass of 3.01×10^{23} molecules of hydrochloric acid, HCl?

Solution:

$$n(\text{HCl}) = \frac{3.01 \times 10^{23}}{6.02 \times 10^{23}}$$

$$= 0.500 \text{ mol}$$

$$\text{Since } n = \frac{m}{M},\ m = n \times M$$

$$\therefore \text{mass(HCl)} = n \times M$$

$$= 0.500 \times 36.5$$

$$= 18.3 \text{ g}$$

Revision questions

16. **What is the mass of each of the following?**

- **(a)** **5.25×10^{24} molecules of glucose, $C_6H_{12}O_6$**
- **(b)** **1.83×10^{21} molecules of nitrogen dioxide, NO_2**
- **(c)** **3.56×10^{14} molecules of carbon dioxide, CO_2**
- **(d)** **4.13×10^{28} molecules of carbon disulfide, CS_2**
- **(e)** **3.62×10^{24} molecules of dinitrogen tetraoxide, N_2O_4**

17. **Which of the following substances has the greatest mass?**

- **(a)** **200 g of magnesium**
- **(b)** **5.00 mol of sulfur**
- **(c)** **1.2×10^{24} atoms of helium**
- **(d)** **3.5×10^{22} molecules of alanine, $C_3H_7O_2N$**

18. **Use a spreadsheet program to produce a spreadsheet that converts amounts of substances to moles. Set up your spreadsheet according to the following template, and then complete the table. Use a formula and then the 'fill' function.**

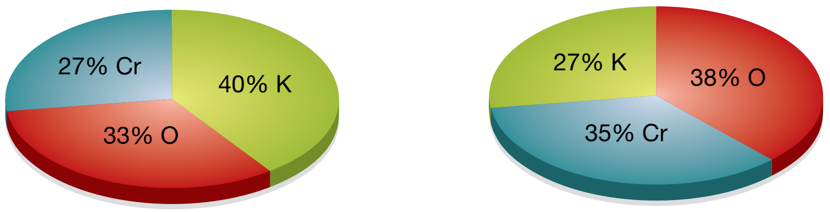
Conversion of amounts of substances to moles

	A	B	C	D	E	F	G
	Substance	**Amount (g)**	**Molar mass (*M*)**	**Number of atoms in the molecule**	**Number of moles (*n*) of substance**	**Number of molecules**	**Total number of atoms**
1	water, H_2O	3.2					
2	methane, CH_4	2.7					
3	ammonia, NH_3	0.056					
4	ethanoic acid, CH_3COOH				27.3		
5	benzene, C_6H_6				0.56		
6	octane, C_8H_{18}				2.34		
7	ethanol, CH_3CH_2OH					6.0×10^{24}	
8	ozone, O_3					1.27×10^{21}	
9	sulfuric acid, H_2SO_4						3.0×10^{26}
10	carbon dioxide, CO_2						7.5×10^{25}

Percentage composition

The percentage that each element contributes to the total mass of a compound is called the percentage composition.

The chemical formula of a compound tells us about the relative numbers of atoms in its constituent elements. However, the composition of a compound is often expressed in terms of the percentage that each element contributes to its mass. This is called **percentage composition**. For instance, although potassium chromate, K_2CrO_4, and potassium dichromate, $K_2Cr_2O_7$, have the same number of potassium atoms, the percentage that potassium contributes to the mass of a unit of the chromate is higher than its contribution to a unit of the dichromate.



(a) potassium chromate, K_2CrO_4 (b) potassium dichromate, $K_2Cr_2O_7$

The percentages of K, Cr and O in (a) potassium chromate, K_2CrO_4 and (b) potassium dichromate, $K_2Cr_2O_7$

The following formula can be used to calculate percentage composition:

$$\begin{array}{c}\%\text{ of } x \text{ in compound}\\ \text{containing } x\end{array} = \frac{\text{mass of } x \text{ in 1 mol of compound}}{\text{molar mass of compound}} \times \frac{100}{1}$$

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Interactivity
Percentage composition
int-6355

Sample problem 5.6

Stannous (tin) fluoride, SnF_2, is an active ingredient in some toothpastes. Find the percentage composition by mass of each element in the compound.

Solution: One formula unit of SnF_2 contains one Sn atom and two F atoms.

$$\begin{aligned} M_r(SnF_2) &= 118.7 + (2 \times 19.0) \\ &= 156.7 \end{aligned}$$

Since Sn contributes 118.7 units toward this mass:

$$\begin{aligned} \%\text{ mass Sn} &= \frac{118.7}{156.7} \times \frac{100}{1} \\ &= 75.7\% \end{aligned}$$

F_2 contributes 2 × 19.0 units, so we get:

$$\begin{aligned} \%\text{ mass F} &= \frac{2 \times 19.0}{156.7} \times \frac{100}{1} \\ &= 24.3\% \end{aligned}$$

study on

Unit 1	**Percentage composition by mass** Summary screen and practice questions
AOS 1	
Topic 3	
Concept 5	

Revision question

19. **Calculate the percentage composition of each element in the following compounds.**
    - **(a) propane, C_3H_8**
    - **(b) sodium hydrogen sulfate, $NaHSO_4$**
    - **(c) calcium acetate, $Ca(CH_3COO)_2$**
    - **(d) hydrogen cyanide, HCN**

Calculating percentage composition of hydrated compounds

Ionic compounds often incorporate water into their lattice structure to form hydrated compounds. For many hydrated ionic compounds, this number of water molecules is very specific, and is called the 'water of crystallisation'. The actual number of water molecules in the formula is called the 'degree of hydration'.

Some ionic compounds crystallise from an aqueous solution to form a **hydrated** ionic compound. In these compounds, water molecules are included in the crystal lattice structure. This water is called **water of crystallisation**. Hydrated copper(II) sulfate, for example, appears as a blue crystalline solid and has the formula $CuSO_4{\cdot}5H_2O$. This means that, for each Cu^{2+} ion and SO_4^{2-} ion in the crystal lattice structure, five water molecules are also included. Heating the copper(II) sulfate crystals removes the water molecules to leave a white powder known as **anhydrous** copper(II) sulfate. The formula of anhydrous copper(II) sulfate is $CuSO_4$. The mass of the water of crystallisation, and its percentage contribution to the mass of the ionic compound, can be calculated when the masses of both the hydrated and anhydrous compound are known or if the degree of hydration is shown in the formula.

Hydrated copper(II) sulfate, $CuSO_4{\cdot}5H_2O$

Many structures have water molecules included in their composition. When the structures are dehydrated, the water is released. Dehydrated copper sulfate is white in colour. When water is added, it turns the characteristic blue colour of hydrated copper sulfate.

Sample problem 5.7

Calculate the percentage of water by mass in $CuSO_4{\cdot}5H_2O$.

Solution: **STEP 1**

Find the molar mass, making sure to include the water present in the formula.

$$\begin{aligned} M(CuSO_4{\cdot}\,5H_2O) &= 63.5 + 32.1 + (4\times16.0) + 5([2\times1.0] + 16.0) \\ &= 249.6\text{ g mol}^{-1} \end{aligned}$$

STEP 2

Calculate the mass of the water.

$$\begin{aligned} M &= 5([2\times1.0]+16.0) \\ &= 90.0\text{ g mol}^{-1} \end{aligned}$$

STEP 3

Calculate the percentage mass of water.

$$\begin{aligned} \%\text{ mass water} &= \frac{90.0}{249.6}\times\frac{100}{1} \\ &= 36.06\% = 36.1\% \end{aligned}$$

Revision question

20. **Calculate the percentage of water in each of the following compounds.**
  - **(a)** **nickel(II) sulfate hexahydrate, $NiSO_4{\cdot}6H_2O$**
  - **(b)** **sodium carbonate decahydrate, $Na_2CO_3{\cdot}10H_2O$**
  - **(c)** **magnesium chloride hexahydrate, $MgCl_2{\cdot}6H_2O$**

Empirical formulas

The empirical formula of a compound is the simplest whole number ratio of atoms or ions that are present in that compound. Empirical formulas can be calculated from percentage compositions or molecular formulas.

The **empirical formula** of a compound gives the simplest whole number ratio of the atoms or ions present in the compound and can be found only by experiment.

To determine the empirical formula of a compound, an experimentally determined *ratio of elements by mass* must be converted to a *ratio of elements by numbers*. This is done by calculating the number of moles of each element.

The steps involved in finding an empirical formula are:

1. Write down the *symbols* of the elements present.
2. Assume that the *mass* of the sample is 100 g and all percentages become grams.
3. Convert masses to *moles*.
4. Find the simplest whole number *ratio* of the atoms by dividing all numbers of moles by the smallest number of moles.
5. If necessary, multiply by a *factor* to convert all numbers to whole numbers.

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Unit 1
AOS 1
Topic 3
Concept 6

Empirical formula of an ionic compound
Summary screen and practice questions

Sample problem 5.8

A compound of sulfur contains 2.4% hydrogen, 39.0% sulfur and 58.6% oxygen. Find the empirical formula of the compound.

Solution: Follow the steps outlined above.

STEP 1	symbols	H	S	O
STEP 2	masses	2.4 g	39.0 g	58.6 g

STEP 3 moles

$$\frac{2.4}{1.0} = 2.4 \qquad \frac{39.0}{32.1} = 1.2 \qquad \frac{58.6}{16.0} = 3.7$$

STEP 4 ratio/divide by smallest

$$\frac{2.4}{1.2} = 2.0 \qquad \frac{1.2}{1.2} = 1.0 \qquad \frac{3.7}{1.2} = 3.1$$

The empirical formula of the compound is H_2SO_3.

Empirical formulas can also be calculated when experiments reveal the actual mass of each element that is present in a sample of a compound. In this situation, the masses involved are simply written in at step 2 in the sequence of steps on the previous page.

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Digital document

Experiment 5.1 Empirical formula of magnesium oxide

doc-15998

Revision questions

21. **Aspirin is a drug used extensively for pain relief. Chemical analysis of an aspirin tablet determined that it was composed of 57.7% carbon, 37.5% oxygen and 4.8% hydrogen. Calculate the empirical formula of aspirin.**

22. **Hydroquinone is a liquid that is used as a photographic developer. Determine the empirical formula of hydroquinone given that it contains 65.4% carbon and 29.1% oxygen with the remainder being hydrogen.**

Sample problem 5.9

Washing soda crystals may be used to bleach linen. When crystallised from water, washing soda (sodium carbonate, Na_2CO_3) forms crystals of a hydrated ionic compound. When 5.00 g of washing soda crystals were dried in a desiccator, 1.85 g of sodium carbonate remained. Calculate the empirical formula of the hydrated compound.

Solution: **STEP 1**

Calculate the mass of water present in the hydrated compound. Since 1.85 g of sodium carbonate was obtained on dehydration of the crystals, the amount of water is 3.15 g, or the difference in mass between the hydrated and dehydrated compounds (5.00 – 1.85) = 3.15 g.

STEP 2

The mole ratio of sodium carbonate to water is:

$$\frac{1.85}{106.0} : \frac{3.15}{18.0}$$
$$= 0.0175 : 0.175$$
$$= 1 : 10$$

So the empirical formula of the washing soda crystals is $Na_2CO_3{\cdot}10H_2O$.

Sodium carbonate can be used to bleach linen.

Revision questions

23. **A 1.124 g sample of $CaSO_4$ crystals was heated to drive off the water of crystallisation. When completely dry, a residue of 0.889 g was obtained. Determine the empirical formula of the hydrated compound.**

24. **A 0.942 g sample of $MgSO_4$ crystals was heated to drive off the water of crystallisation. When completely dry, a residue of 0.461 g was obtained. Determine the empirical formula of the hydrated compound.**

Molecular formulas

The **molecular formula** of a compound represents the actual composition of a compound that is made up of molecules. A molecular formula is either the same as its experimentally determined empirical formula, or is a whole-number multiple of it. Table 5.1 shows the empirical and molecular formulas of some common hydrocarbons. Although their empirical formulas may be the same, compounds with different molecular formulas may have very different properties. For example, acetylene is a gas that is used in welder's torches, whereas benzene is a highly flammable liquid that is a recognised carcinogen. Although both compounds have the same empirical formula, CH, the molecular formula of acetylene is C_2H_2, and the molecular formula of benzene is C_6H_6.

The molecular formula of a compound may be determined from its empirical formula only if its molar mass is also known.

$$n \times (\text{empirical formula}) = \text{molecular formula}$$

where *n* represents a whole number.

The molecular formula of a compound is the actual number of atoms that are present in a molecule of that substance. It can be equal to the empirical formula, or it can be a whole number multiple of the empirical formula.

Note that only molecular compounds can have a molecular formula.

Four structural representations of a benzene molecule. A benzene molecule is the basis of a group of chemicals called aromatics.

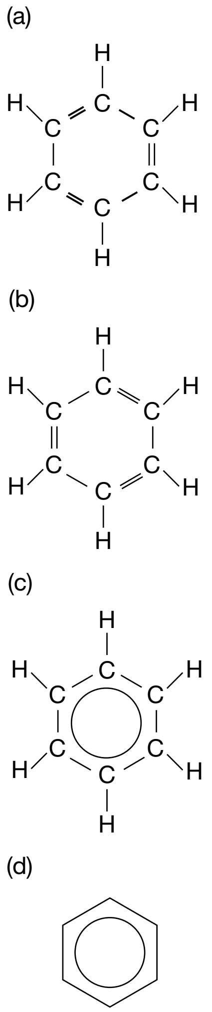

TABLE 5.1 Comparison of empirical and molecular formulas

Name of molecule	Empirical formula (simplest ratio of atoms in molecule)	Molecular formula (actual number of atoms in molecule)
ethyne (acetylene)	CH	C_2H_2
benzene	CH	C_6H_6
formaldehyde	CH_2O	CH_2O
ethanoic acid (acetic acid)	CH_2O	$C_2H_4O_2$
glucose	CH_2O	$C_6H_{12}O_6$

Sample problem 5.10

Benzene has the empirical formula CH and its molar mass is 78 g mol^{-1}. Find the molecular formula.

Solution:

STEP 1

If the molar mass is 78.0 g mol^{-1}, then its relative molecular mass (the sum of the relative atomic masses according to the molecular formula) is also 78.0. If the empirical formula is written as C_xH_y, then the molecular formula is $(C_xH_y)n$, where *x*, *y* and *n* are whole numbers. The value of *n* may be determined by comparing the relative molecular mass and the empirical formula mass.

STEP 2 The empirical formula is CH.

$$\begin{aligned}\text{empirical formula mass} &= A_r(\text{C}) + A_r(\text{H})\\ &= 12.0 + 1.0\\ &= 13.0\end{aligned}$$

STEP 3

$$\frac{\text{molecular mass}}{\text{empirical formula mass}} = \frac{78.0}{13.0}$$

$$\text{ratio} = 6$$

STEP 4

$$\begin{aligned}\therefore \text{molecular formula} &= 6 \times \text{empirical formula}\\ &= 6 \times \text{CH}\\ &= C_6H_6\end{aligned}$$

Revision questions

25. **During the decay of animal tissues, a noxious compound called putrescine may be produced. Putrescine has the empirical formula C_2H_6N and a relative molecular mass of 88.0. What is the molecular formula of putrescine?**

26. **Calculate the molecular mass of each compound listed in table 5.1. What is the mathematical relationship between the empirical and molecular formulas of a substance?**

27. **Nicotine, the main active chemical in tobacco, has the empirical formula C_5H_7N and a relative molecular mass of 162.0. Determine the molecular formula of nicotine.**

Sample problem 5.11

The common insect repellent sold commercially as ‘mothballs’ is the organic compound naphthalene. It is a hydrocarbon containing 93.7% carbon and 6.3% hydrogen and has a molar mass of 128.0 g mol^{-1}. Find the empirical and molecular formulas of naphthalene.

A molecular formula is calculated in the same way as an empirical formula. However, it needs one additional piece of information — the molar mass. This allows a comparison to be made with the empirical formula to determine the whole number multiple.

Solution: Calculate the empirical formula using the five steps listed on page 102; then calculate the empirical formula mass and use it to find the molecular formula.

STEP 1	H	C
STEP 2	6.3 g	93.7 g
STEP 3	$\frac{6.3}{1.0}$ $= 6.3$	$\frac{93.7}{12.0}$ $= 7.81$
STEP 4	$\frac{6.3}{6.3}$ $= 1.0$	$\frac{7.81}{6.3}$ $= 1.24$
STEP 5	1.0×4 $= 4.0$	1.24×4 $= 5.0$

The empirical formula of naphthalene is C_5H_4.

STEP 6

$$\begin{aligned}\text{empirical formula mass} &= A_r(\text{C}) + A_r(\text{H})\\ &= (5 \times 12.0) + (4 \times 1.0)\\ &= 64.0\end{aligned}$$

STEP 7

$$\frac{\text{molecular mass}}{\text{empirical formula mass}} = \frac{128.0}{64.0}$$

$$\text{ratio} = 2$$

STEP 8

$$\begin{aligned}\text{molecular formula} &= 2 \times \text{empirical formula}\\ &= 2 \times C_4H_5\\ &= C_8H_{10}\end{aligned}$$

Revision questions

28. **Caffeine is a stimulant that is found naturally in coffee, tea and chocolate. Analysis of caffeine shows that it contains 49.5% carbon, 28.9% nitrogen, 16.5% oxygen and 5.1% hydrogen by mass. Determine the molecular formula of caffeine given that its molar mass is 194.2 g mol^{-1}.**

29. **The compound methyl butanoate smells like apples. Its percentage composition is 58.8% C, 9.8% H and 31.4% O. If its molar mass is 102.0 g mol^{-1}, what is its molecular formula?**

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Unit 1
AOS 1
Topic 3
Quantifying chemistry
Summary screens and practice questions

Chapter review

Summary

- The relative atomic mass (A_r) of elements shown in the periodic table is calculated using the weighted mean of the relative isotopic masses:

$$A_r = \frac{(\text{RIM of first isotope} \times \text{abundance}) + (\text{RIM of second isotope} \times \text{abundance}) + \ldots}{100}$$

  where A_r is the relative atomic mass and RIM is the relative isotopic mass.
- Relative isotopic masses are determined using a mass spectrometer and are masses compared with the carbon-12 isotope.
- The carbon-12 isotope is defined as having a mass of 12 g exactly. It is used as a standard as it is a stable and abundant isotope, which means that its mass can be determined accurately.
- The mass spectrometer is an instrument that can separate isotopes of an element based on their mass-to-charge ratio.
- A mass spectrum of an element shows the number of isotopes present in an element, their relative isotopic masses and their proportions.
- The mole concept is an important chemical idea, enabling us to 'count' particles present in a quantity of a substance.
  - The unit that measures the amount of a substance is the mole (mol).
  - A mole of any substance is composed of Avogadro's number (6.02×10^{23}) of representative particles.
  - The molar mass (M) of an element or compound is the mass of 1 mole of the substance (numerically equal to the relative molecular mass (M_r) and is expressed in g mol^{-1}.
  - The number of particles in a sample of substance may be determined by the formula:

$$\text{number of particles} = \text{number of moles of substance} \times \text{Avogadro's number}$$

- The percentage composition of a compound is the percentage by mass of each element in the compound:

$$\%\text{ of } x \text{ in compound containing } x = \frac{\text{mass of } x \text{ in compound}}{\text{molar mass of compound}} \times \frac{100}{1}$$

- Hydrated ionic compounds contain water molecules as part of their lattice structure. This water is called water of crystallisation. When the mass of both the hydrated and anhydrous (dried) compounds is known, the percentage composition by mass of water in the hydrated compound can be determined.
- Formulas for compounds take two forms:
  - An empirical formula is the simplest whole-number ratio of atoms of the elements in the compound and is determined experimentally.
  - A molecular formula is the same as, or some simple multiple of, an empirical formula and specifies the exact number of each type of atom in a molecule of the compound.

Multiple choice questions

1. The particles in a mass spectrometer are deflected according to their:
A abundance
B mass only
C charge only
D mass and charge.

2. Rubidium has two stable isotopes. Rubidium-85 has an abundance of 72% and rubidium-87 has an abundance of 28%. The mass spectrum of rubidium has:
A 2 peaks: a smaller peak at RIM 85 and a larger peak at RIM 87
B 2 peaks: a smaller peak at RIM 87 and a larger peak at RIM 85
C 1 peak at about RIM 85.5
D 1 peak at about RIM 86.

3. When the element fluorine, F_2, is passed through the mass spectrometer the particles detected are:
A F_2 ions
B F ions
C F^- and F^+ ions
D F^+ ions.

4. A certain element has two isotopes of relative isotopic mass 203 and 205. Given that its A_r value is 204.4, which of the following is the most likely percentage abundance of the heavier isotope?
A 7%
B 30%
C 50%
D 70%

5. The difference between the relative molecular mass (M_r) of carbon dioxide and the molar mass of carbon dioxide is that the relative molecular mass and the molar mass of carbon dioxide are, respectively:
A 44.0 g and 1 mol
B 44.0 and 44.0 g mol^{-1}
C 44.0 g and 44.0 g mol^{-1}
D 44.0 and 44.0 g.

6. The number of moles of oxygen atoms in 143 g of sodium carbonate decahydrate, $Na_2CO_3{\cdot}10H_2O$, is:
A 0.50
B 1.5
C 2.0
D 6.5.

7. The *smallest* number of molecules would be contained in:

- **A** 1.0 g of N_2
- **B** 1.0 g of O_2
- **C** 1.0 g of NO
- **D** 1.0 g of NO_2.

8. The *largest* number of oxygen atoms would be found in:

- **A** 300 g of water, H_2O
- **B** 3.2 mol of hydrated copper(II) sulfate, $CuSO_4{\cdot}5H_2O$
- **C** 3 kg of a fat having the molecular formula $C_{57}H_{110}O_6$
- **D** 7.35×10^{24} molecules of nitrogen dioxide, NO_2.

9. Polymers may be made from the three monomers acrylic acid, $C_3H_4O_2$, methyl acrylite, $C_4H_6O_2$, and acrylonitrile, C_3H_3N. The order of increasing percentage carbon content of these monomers is:

- **A** acrylic acid, methyl acrylite, acrylonitrile
- **B** acrylic acid, acrylonitrile, methyl acrylite
- **C** acrylonitrile, acrylic acid, methyl acrylite
- **D** acrylonitrile, methyl acrylite, acrylic acid.

10. Which of the following lists of substances shows only empirical formulas?

- **A** H_2O, CH_3COOH, HCl, $Na_2Cr_2O_7$
- **B** H_2SO_4, H_2O, $Al(NO_3)_2$, $Al_2(SO_4)_3$
- **C** HCl, $C_6H_{12}O_6$, NH_3, $Al(NO_3)_2$
- **D** $Al_2(SO_4)_3$, O_2, HCl, H_2SO_4

11. A substance, on analysis, was found to contain 27.3% carbon and 72.7% oxygen. Which of the following is the empirical formula of the substance?

- **A** CO_2
- **B** CO_3
- **C** C_2O_4
- **D** C_3O_7

12. A hydrocarbon that forms the major component of liquefied petroleum gas was determined experimentally to contain 16.3% hydrogen by mass. If the hydrocarbon has a relative molecular mass of 86, its molecular formula may be represented as:

- **A** C_3H_7
- **B** C_6H_{14}
- **C** C_7H_2
- **D** C_7H_{16}.

Review questions

Relative atomic mass

1. Silicon-containing ores have three isotopes: 92% silicon-28, 5% silicon-29 and 3% silicon-30.

(a) Draw the mass spectrum for silicon.
(b) What is the atomic number of silicon?
(c) What are the relative isotopic masses of the three silicon isotopes?
(d) What is the relative atomic mass of silicon?

2. Silver is an unreactive metal that is used to produce jewellery. Its relative atomic mass is 107.9, and it consists of two naturally occurring isotopes: silver-107 and silver-109. Calculate the abundance of the lighter isotope.

Mole calculations

3. In 2.0 moles of $(NH_4)_3PO_4$ (fertiliser) there are:

(a) _____ moles of nitrogen atoms
(b) _____ moles of hydrogen atoms
(c) _____ moles of oxygen atoms
(d) _____ moles of phosphorus atoms
(e) _____ (total number) nitrogen atoms
(f) _____ grams of phosphate ions
(g) _____ grams of nitrogen atoms.

4. Copy and complete the table below.

Name	Formula	Molar mass (M)	Mole (n)	Mass (m)
sodium hydroxide	NaOH			3.41 g
	CCl_4		1.40 mol	
sodium carbonate		106 g mol^{-1}	1.00 mol	
	KCl		0.25 mol	
ammonium phosphate				8.46 g

5. Find the number of moles *and* the number of atoms present in the following samples.

(a) 14.6 g acetylene (ethyne), C_2H_2
(b) 0.48 g propane, C_3H_8
(c) 485 g ethanol, C_2H_5OH
(d) 8.6 g carbon dioxide
(e) 67 g iodine, I_2

6. Calculate the mass of each of the following.

(a) 160 mol Fe
(b) 0.075 mol silicon dioxide
(c) 4.23 mol NO_2

7. Sodium fluoride is thought to reduce tooth decay, especially in children. It is therefore added to some brands of toothpaste. If a tube of toothpaste contains 0.013 g of sodium fluoride,

(a) how many moles of sodium fluoride does this represent
(b) how many fluoride ions does this represent?

8. To prevent a gum disease called scurvy, the minimum daily requirement of vitamin C, $C_6H_8O_6$, is 60 mg.

(a) How many moles of vitamin C is this?
(b) How many molecules is this?

(c) If 10 g of spinach is found to contain 1.2×10^{-5} g of vitamin C, how much spinach must be eaten to attain the minimum daily requirement?

9. A typical vitamin C tablet contains 500.0 mg of ascorbic acid. (The molecular formula of vitamin C is $C_6H_8O_6$.)

(a) How many moles of vitamin C does a typical vitamin C tablet contain?

(b) How many molecules of vitamin C does a typical vitamin C tablet contain?

10. Too much cholesterol, $C_{27}H_{46}O$, is associated with heart disease, although the body produces its own cholesterol in the liver. A 250 g sample of cholesterol was removed from the arteries of a patient.

(a) How many moles of cholesterol are present in this sample?

(b) How many moles of carbon atoms does this represent?

(c) How many grams of carbon atoms are present in the sample?

(d) How many grams of oxygen atoms are present in the sample?

11. Diamond is a naturally occurring form of pure carbon. The mass of a diamond is measured in a unit called a carat, where 1.00 carat = 0.200 g. How many atoms of carbon are in:

(a) a 1.00-carat diamond

(b) a 3.15-carat diamond?

Percentage composition

12. Calculate the percentage composition of all the elements in:

(a) ethanoic acid, CH_3COOH

(b) hydrated magnesium chloride, $MgCl_2{\cdot}6H_2O$

(c) $Fe_2(SO_4)_3$.

13. Photocells use a semiconducting material that produces an electric current on exposure to light. Compounds of cadmium such as CdS, CdSe and CdTe are used in many common photocells. Calculate the percentage, by mass, of cadmium in:

(a) CdS (b) CdSe (c) CdTe.

14. A number of different compounds may be formed when oxygen and nitrogen react together. Calculate the percentage, by mass, of nitrogen and oxygen in each of the following oxides of nitrogen.

(a) NO, a colourless gas that is formed in internal combustion engines

(b) NO_2, a brown gas that is mainly responsible for the brown colour of the photochemical smog that hangs over many industrialised cities

(c) N_2O_4, a colourless liquid that is used as a fuel in space shuttles

(d) N_2O, a colourless gas (commonly called laughing gas) that is used as a dental anaesthetic

15. The synthetic narcotic methadone is used for treatment of heroin addiction and has the molecular formula $C_{21}H_{27}NO$. Calculate:

(a) the molar mass of methadone

(b) its percentage composition.

16. Nitrogen is essential for plant growth. Ammonium nitrate, NH_4NO_3, and urea, CON_2H_4, can be used as fertilisers since each contains a significant proportion of nitrogen.

(a) Calculate the percentage, by mass, of nitrogen in each of the fertilisers.

(b) Would you expect all compounds with a high nitrogen content to be suitable for use as a fertiliser? Justify your response.

Empirical formula

17. Coffee contains the stimulant caffeine. Analysis shows it consists of 49.48% carbon, 5.19% hydrogen, 28.85% nitrogen and 16.48% oxygen by mass. Calculate the empirical formula of caffeine.

18. The amino acid cysteine contains the elements carbon, hydrogen, nitrogen, oxygen and sulfur. Analysis of a 1.210 g sample of cysteine shows it to contain 0.0704 g of hydrogen, 1.80×10^{22} carbon atoms, 0.0100 mol of nitrogen and equal masses of sulfur and oxygen. Determine the empirical formula of cysteine.

19. Methanol is an alternative fuel to petrol that has been used in experimental cars. Determine the empirical formula of methanol, given that it is composed of 49.9% oxygen, 37.5% carbon and 12.6% hydrogen.

20. Borax is a naturally occurring compound that is used in the manufacture of optical glasses. Analysis of borax shows that it is made up of three elements: sodium (22.8%), boron (21.5%) and oxygen. Determine the empirical formula of borax.

Molecular formula

21. A compound of Na, S and O contains 17.04% Na and 47.41% S. The M_r of the compound is 270.0. Calculate the empirical formula and the molecular formula.

22. The taste of sour milk is due to lactic acid. The percentage composition of lactic acid by mass is 40.00% carbon, 6.71% hydrogen and 53.29% oxygen, and the molar mass is 90.0 g mol^{-1}. Find the empirical formula and molecular formula of lactic acid.

23. A compound contains 12.8% carbon and 2.13% hydrogen, the rest being bromine. The relative molecular mass of the compound is 188.0. Calculate the empirical formula and the molecular formula of the compound.

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Unit 1 Quantifying chemistry

AOS 1

Topic 3 Sit topic test

Exam practice questions

In a chemistry examination you will be required to answer a number of short and extended response questions.

Multiple choice question

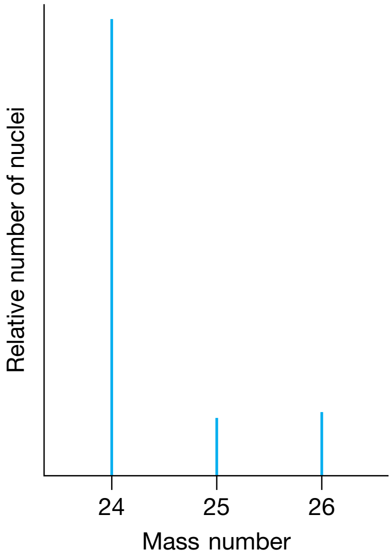
1. The mass spectrum of magnesium is shown on the right. Which of the following statements is true?

A Three isotopes are shown with masses in the ratio 24 : 25 : 26 and relative abundances of about 8%, 1% and 1%.

B The three isotopes have the same mass number but different atomic numbers.

C Isotopes 25 and 26 have relative abundances of about 10%.

D Isotope 24 has a relative abundance of about 10%. **1 mark**



Extended response questions

2. Octane, C_8H_{18}, is a major component of petrol.

(a) What is the M_r of octane?

(b) What is the mass of 3.20 mol of octane?

(c) How many molecules are there in 2.5 g of octane?

(d) How many individual atoms are there in 5.0 g of octane?

(e) What is the percentage of hydrogen in octane?

(f) What mass of carbon would be present in 150 g of octane? **6 marks**

3. The odour of rancid butter is caused by butyric acid, which contains 54.5% carbon and 9.1% hydrogen, the rest being oxygen. Given that the M_r of butyric acid is 88.0, calculate the empirical and molecular formulas of the acid. **4 marks**

4. An oxide of copper is heated in a stream of hydrogen until only the copper remains, according to the equation:

$$Cu_xO(s) + H_2(g) \longrightarrow xCu(s) + H_2O(l)$$

The data for the experiment is given in the table below. Calculate the empirical formula of the oxide of copper.

Item	Mass (g)
crucible	27.002
crucible plus contents before heating	27.128
crucible plus contents after heating	27.114

3 marks

5. 4.6 g of anhydrous zinc sulfate with an M of 161.5 g mol^{-1} was obtained by driving the water from 8.2 g of the crystalline hydrated salt with empirical formula $ZnSO_4{\cdot}xH_2O$. Calculate the value of x. **3 marks**

6. Insects of a particular species can identify their mate by using special chemicals called pheromones that transmit chemical messages. The pheromone that serves as a sex attractant for gypsy moths is called disparlure and contains the elements C, H and O. Analysis of disparlure shows that 0.282 g contains 16.00×10^{-3} g of O atoms and 0.228 g of C atoms. The M_r of disparlure is 282 g mol^{-1}. Determine the molecular formula of disparlure. **3 marks**

7. Use examples to explain the difference between the terms 'mass number', 'relative atomic mass' and 'molar mass'. **6 marks**

CHAPTER

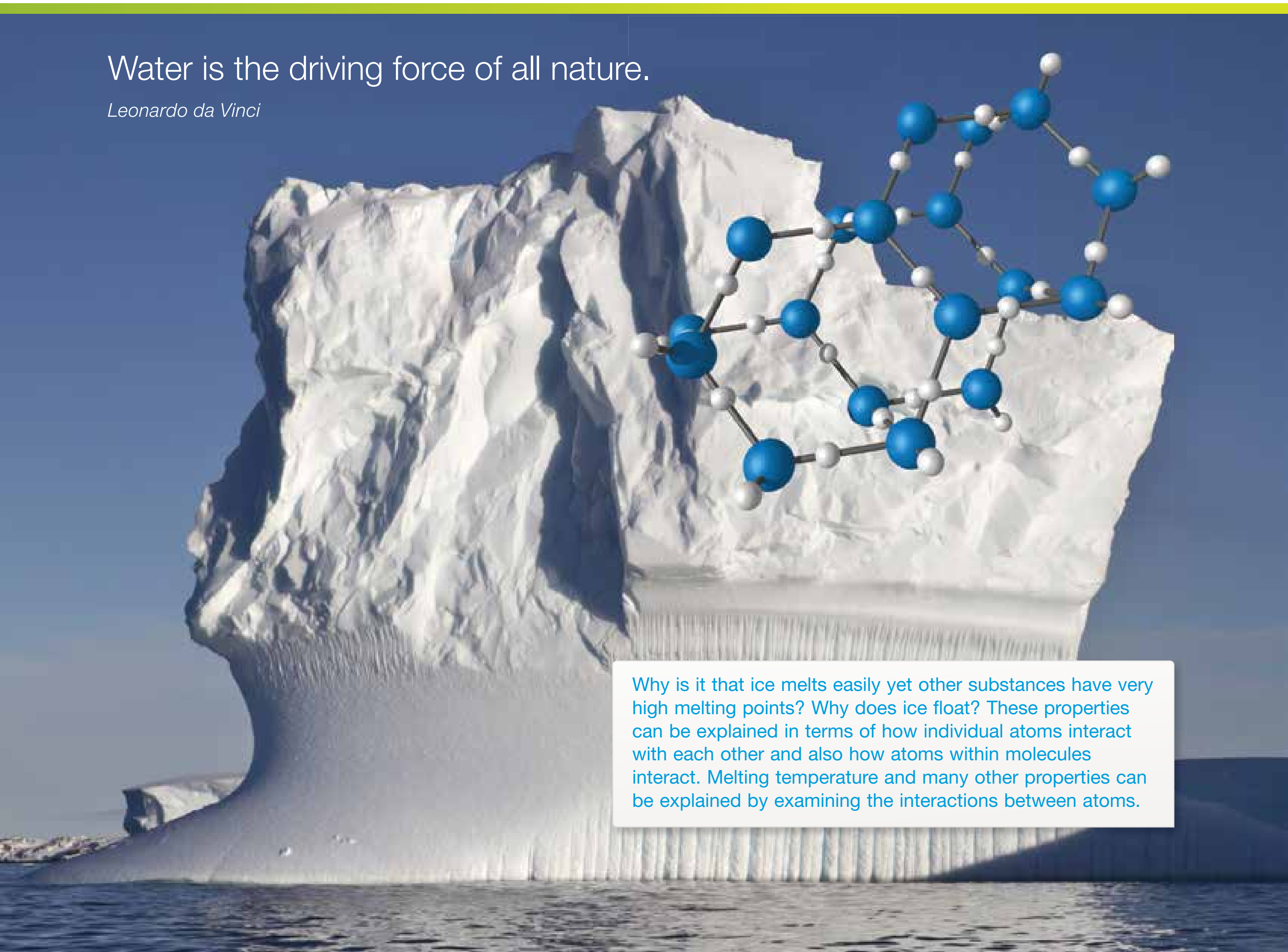
6 Covalent bonding

This chapter deals with the formation and naming of molecules, outlines the attractive forces between molecules in the formation of solids, and explores the relationship between the bonding within and between molecules to explain their properties. The end of the chapter provides a review of the three main types of bonding — ionic bonding, metallic bonding and covalent bonding.

YOU WILL EXAMINE:

- covalent bonding as the strong bonding that exists within discrete molecules such as hydrogen, oxygen and water
- the relationship between the structure and bonding of covalent materials and their properties
- the shapes of molecules with up to four electron pairs around a central atom
- weak forces such as dispersion forces and permanent dipole–dipole interactions, including hydrogen bonding
- bonding in and between water molecules.

> Water is the driving force of all nature.
>
> *Leonardo da Vinci*



Why is it that ice melts easily yet other substances have very high melting points? Why does ice float? These properties can be explained in terms of how individual atoms interact with each other and also how atoms within molecules interact. Melting temperature and many other properties can be explained by examining the interactions between atoms.

Introducing covalent bonds

Covalent bonds are formed when non-metal atoms share electrons.

Our Earth is a complex and fascinating world of chemicals. The air we breathe consists mainly of the gases nitrogen and oxygen, which are examples of covalent molecular elements. Drinking water, regardless of whether it comes from a local river, lake or distilled sea water is a covalent molecular compound. Rocks, sand and minerals found in the solid part of the Earth contain substances such as quartz and diamond, which are classified as covalent network lattices. All these substances are composed of non-metals only and behave very differently from the metallic and ionic substances we have already studied. This different behaviour is due to the unique structure that results when non-metal atoms are held together by a **covalent bond**. These bonds are formed when atoms, rather than gaining or losing electrons, share them. Covalent bonds differ from ionic bonds in that there is no transfer of electrons involved.

Covalent molecular substances

Covalent molecular substances have low melting and boiling temperatures. They are poor conductors of electricity.

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Unit 1	**Molecules and covalent bonding** Summary screen and practice questions
AOS 2	
Topic 1	
Concept 1	

Covalent molecular substances are held together by covalent bonds and are all around us. How can substances that look so different and behave in such different ways have the same type of bonding between atoms? The answer lies in some of their common properties.

Fuels are an important group of covalent molecular substances. The gas that is piped to homes is mainly methane, CH_4, while bottled (LPG) gas used for barbecues and camping is a mixture of propane, C_3H_8, and butane, C_4H_{10}. Octane, C_8H_{18}, is a major component of petrol. Methanol, CH_3OH, and ethanol, C_2H_5OH, are fuels used in high-performance racing cars. These substances all share the ability to ignite easily and burn in oxygen to release large amounts of heat. These compounds are all gases or liquids at room temperature and pressure. This means that they have low boiling and melting temperatures. They are also poor conductors of electricity.

Most other non-metallic elements and compounds of non-metals also have low melting and boiling temperatures. From these properties, we can make the following deductions about their structures:

- The forces of attraction between the particles are very weak.
- No charged particles are free to move through the substance.

A theory for the structure of covalent molecular substances

The major features of the current theory of the structure of covalent molecular substances are as follows:

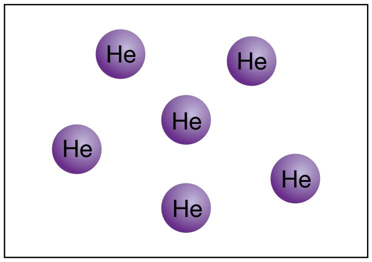
- The basic units of covalent molecular substances are groups of atoms called **molecules**. All molecules within a substance are identical.
- Adjacent atoms within a molecule share electrons in order to achieve a full outer shell.
- Electrical attraction between the nuclei of adjacent atoms and the shared electrons causes the atoms in a molecule to be held together. This force of attraction is called covalent bonding.
- The overall charge on each molecule is zero and so adjacent molecules are held together by weak intermolecular forces.

The attraction of adjacent nuclei for the shared electrons is called covalent bonding.

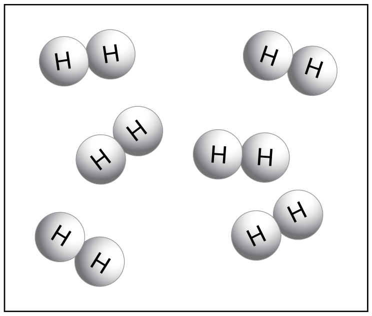
The molecule

A molecule is a group of non-metallic atoms held together by covalent bonds. The atoms are combined in a fixed ratio and are electrically neutral.

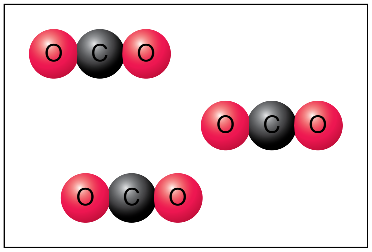
A **covalent molecular element** is made up of identical atoms held together by covalent bonds. Such elements may be found in various arrangements, including **diatomic** molecules. These molecules are described as **discrete**, as each molecule is separate and distinct from the others.



Atoms of helium gas



The molecules of hydrogen gas are diatomic.



Carbon dioxide is a covalent molecular compound formed from atoms of carbon and oxygen.

The noble gases, helium, neon, argon and krypton, all exist as discrete atoms. Since they have full outer shells, they are unreactive and few of their compounds are known.

Diatomic molecules are composed of two atoms of the same element. For example, the covalent molecular element hydrogen exists as a gas made up of discrete diatomic molecules. As shown at left, each molecule is composed of two hydrogen atoms. This is represented by the symbol H_2.

Hydrogen is one of several non-metallic elements that are made up of diatomic molecules. Other molecular elements include oxygen, O_2, nitrogen, N_2, phosphorus, P_4, fluorine, F_2, and sulfur, S_8.

In addition to these elements, a large number of compounds also exist as molecules. In a molecular compound, atoms of *different* elements share electrons with each other. These compounds are often called **covalent molecular compounds** because they contain covalent bonds. Water, carbon dioxide and methane are examples of these compounds.

What is covalent bonding?

Atoms with full outer shells are stable.

Neighbouring atoms share electrons to achieve a full outer shell.

We have observed that atoms of elements with a full outer shell, such as the noble gases helium, neon and argon, are particularly stable. Other atoms tend to achieve noble gas configuration — that is, a full outer shell — by losing or gaining electrons to form ionic bonds, or *sharing* electrons to form covalent bonds. The hydrogen atom needs two electrons in its outer shell for stability. Most of the elements in the second and third periods of the periodic table need eight outershell electrons for stability.

Non-metallic atoms generally have high **electronegativities**. This means they can attract electrons easily. Except for hydrogen, non-metals have four or more outershell electrons. When two non-metallic atoms react together, both of them need to gain electrons in order to complete their outer shells. Since both cannot become negative ions, they share pairs of electrons to achieve eight electrons in the outer shell. This sharing of electron *pairs* between two non-metallic atoms produces the covalent bond. The atoms bond in order to reach a more stable state.

Atoms try to achieve a stable noble gas configuration of eight outer shell electrons.

To better understand this process, we can use electron shell diagrams. Consider the two chlorine atoms shown in the figure at the top of the following page. Chlorine, as well as all members of group 17, exists in nature as diatomic molecules (two atoms covalently bonded together). Each chlorine atom has seven valence electrons. In order to attain a noble gas configuration (a full outer shell of eight electrons), each chlorine atom needs one electron.

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Ionic and covalent bonding

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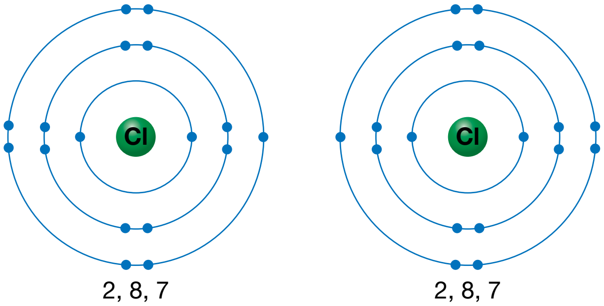
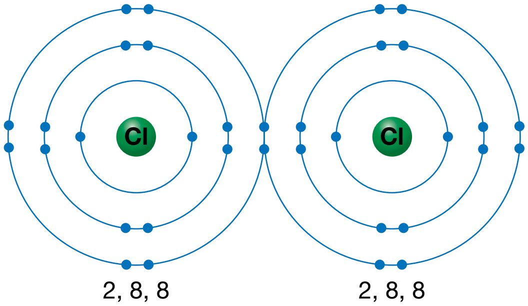


By sharing the single, unpaired electrons, both atoms can complete their outer shells, as shown in the figure below. A covalent bond results from the simultaneous attraction of the two positive chlorine nuclei to the same shared pair of electrons. This bond holds the two atoms together. A molecule of chlorine forms and is represented by the symbol Cl_2.



A molecule of Cl_2, showing how the two atoms share electrons to complete their outer shells

Hydrogen — a special case

Hydrogen atoms need only two electrons to complete their outer shells.

The hydrogen atom is an exception. This is because each of two hydrogen atoms contains one electron in its first (and outer) shell. They both need another electron to attain the stable electron configuration, 2, of the nearest noble gas, helium. This is achieved by each atom gaining a share in the electron of the other and forming a covalent bond, as shown in the figure below. In this way, the diatomic hydrogen molecule, H_2, is formed.

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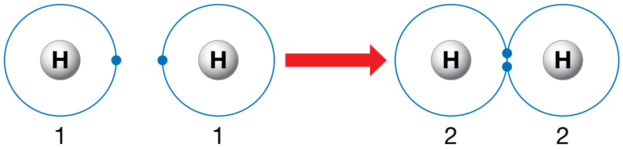
Covalent bonding 1



The formation of a hydrogen molecule

Electron dot diagrams

Electron dot diagrams may be used to show how electrons are shared.

In order to simplify drawings of atoms, **electron dot diagrams** (sometimes called Lewis diagrams) may be used to represent the outershell electrons. These can be drawn for atoms and molecules and are a useful way to understand the sharing of electrons in covalent bonds.

Electron dot diagrams for atoms

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Topic 1	
Concept 2	

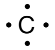
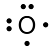
In an electron dot diagram, the atom's nucleus and all innershell electrons are replaced by its element symbol. The outershell electrons are represented by dots or small crosses around the symbol in a square arrangement. These dots are arranged in pairs if more than four outer electrons are present. (*Note*: The innershell electrons do not participate in the bonding.)

For example, nitrogen has five outershell electrons and can be represented by the electron dot diagram at right.

Non-bonding electrons are known as lone pairs.

The two dots arranged as a pair represent **non-bonding electrons** or **lone pairs**. The three that appear as single dots (unpaired) are *available for sharing* and are called **bonding electrons**.

TABLE 6.1 Electron dot diagrams for non-metals in the second period of the periodic table

Group	Element	Valence electrons	Electron dot diagram	Lone pairs	Bonding electrons (8-group number)
14	C	4		0	4
15	N	5		1	3
16	O	6		2	2
17	F	7		3	1
18	Ne	8		4	0

Elements of the same group in the periodic table have electron dot diagrams with the same number of dots or crosses. The halogens (group 17), for example, all have seven dots around their element symbols.

Revision question

1. **Copy and complete the following table.**

Group	Element	Valence electrons	Electron dot diagram	Lone pairs	Bonding electrons
	S				
	Si				
	P				
	Cl				
	Br				
	Se				
	Ar				

The rules for drawing electron dot diagrams for molecules may be simplified as follows:

1. Draw the dot diagrams for each of the atoms in the molecule.
2. Unpaired electrons (bonding electrons) are available for sharing to form a covalent bond.
3. Paired electrons (lone pairs) do not participate in the bonding.
4. The atom with the most bonding electrons (called the **central atom**) is placed in the centre with the other atoms bonded to it.
5. In the resulting diagram, all electrons must be paired and each atom except hydrogen must be surrounded by eight electrons.

In a finished electron dot diagram, every atom except hydrogen must be surrounded by eight electrons (either in lone pairs or shared with other atoms).

Structural formulas

Electron dot diagrams themselves may be simplified by substituting a dash for each shared electron pair. Often, lone pairs are omitted, particularly for non-central atoms. The resulting representation is called a **structural formula**. The structural formula does not generally represent the shape of the molecule; it merely represents the bonding that occurs. For example, the structural formula for the hydrogen molecule is H—H, signifying that the two hydrogen atoms are bonded by one pair of shared electrons.

Electron dot diagrams for molecular elements

A structural formula substitutes a dash for each electron pair.

The structural formula for the chlorine molecule is represented by the symbol Cl—Cl. Even though dots (·) and crosses (×) are used to represent electrons, that doesn't mean there are different types of electrons. All electrons are identical.

Each chlorine molecule shares one pair of electrons to form a **single bond**. Each chlorine atom in a chlorine molecule contains three pairs of unbonded electrons.

There are several useful ways of representing molecules. The electron dot diagrams above left show two chlorine atoms bonding to form a chlorine molecule. This molecule can be modelled three-dimensionally as shown in the centre, or represented by the structural formula Cl—Cl above right.

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Covalent bonding 2

Multiple covalent bonds

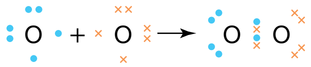
Many molecular substances are held together by **multiple bonds**: that is, bonds formed when two atoms share two or more pairs of electrons. If two atoms share two pairs of electrons, the covalent bond is called a **double bond**. Double bonds are found in molecules such as carbon dioxide, CO_2, and oxygen O_2. When two atoms share three pairs of electrons a **triple bond** is formed, as in the case of nitrogen, N_2.

Multiple bonds are formed when two or three pairs of electrons are shared.

Sample problem 6.1

Show the bonding between two oxygen atoms.

Solution: Oxygen molecules have the symbol O_2. This means that an oxygen molecule contains two oxygen atoms. Each oxygen atom has only six outershell electrons; therefore, it needs to share two more in order to attain a complete outer shell, as shown below.



$O{=}O$

Since the oxygen atoms share two pairs of electrons, the bond between them is a double covalent bond, or just a double bond. This is shown by the structural formula.

Revision questions

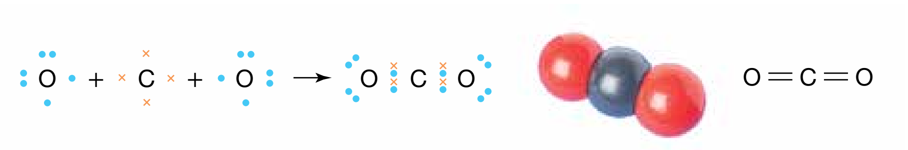
2. **Nitrogen molecules are diatomic molecules. Draw an electron dot diagram of the nitrogen molecule and use it to answer the following questions.**
   - (a) **How would you describe the covalent bond between the two nitrogen atoms?**
   - (b) **How many lone pairs of electrons does the nitrogen molecule contain?**
   - (c) **Draw the structural formula of the nitrogen molecule.**
3. **Draw the electron dot diagram and structural formula of the fluorine molecule.**
4. **Which of the following substances contain discrete diatomic molecules?**
   - (a) **magnesium oxide**
   - (b) **iodine**
   - (c) **aluminium**
   - (d) **neon**
   - (e) **phosphorus**
   - (f) **carbon dioxide**

Electron dot diagrams for molecular compounds

Carbon dioxide

The formula for carbon dioxide is CO_2. In each carbon dioxide molecule, carbon forms a double bond with each of the oxygen atoms.

The central carbon atom in a carbon dioxide molecule now has eight valence electrons, and so does each oxygen atom.



Electron dot diagrams showing carbon and oxygen atoms bonding covalently to form a molecule. The molecular model and structural formula offer other ways of understanding the molecule.

Water

The formula for water is H_2O. In each water molecule, an oxygen atom shares electrons with two hydrogen atoms so that they all achieve complete outer shells.

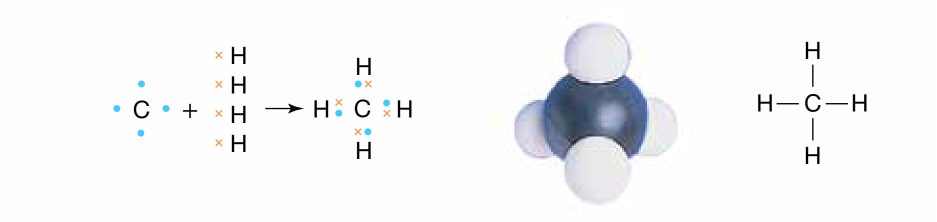
Notice that the central oxygen atom in a water molecule has two pairs of unbonded electrons. These are called lone pairs. Although they do not participate in the covalent bond, they affect the shape of the molecule, and are shown in the structural formula by the two 'unbonded' dashes.

Lone pairs must be considered when determining the shape of a molecule.

Electron dot diagrams, molecular model and structural formula for water

Methane

The formula for methane is CH_4. In each methane molecule, carbon forms a single bond with each of the four hydrogen atoms. All atoms have achieved complete outer shells.



Electron dot diagrams, molecular model and structural formula for methane

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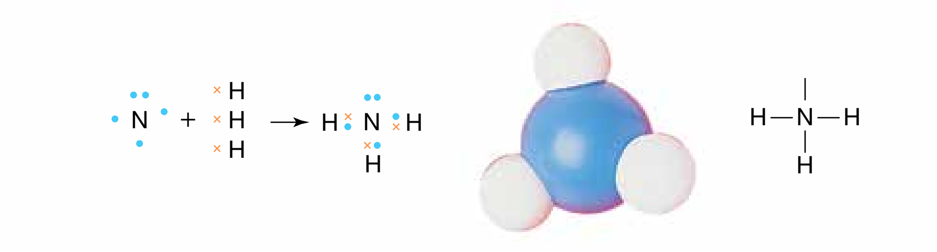
Ammonia

The formula for ammonia is NH_3. In each ammonia molecule, nitrogen forms a single bond with each of the three hydrogen atoms in addition to having one lone pair. This arrangement achieves a stable full outer shell for each atom.



Electron dot diagrams, molecular model and structural formula for ammonia

Revision questions

5. **Use electron dot diagrams and structural formulas to illustrate the bonds between each of the following pairs of atoms.**
  - **(a) hydrogen and bromine**
  - **(b) oxygen and fluorine**

6. **Draw the electron dot diagram and the structural formula for nitrogen trifluoride, NF_3, in which all three fluorine atoms are bonded to the nitrogen atom.**

7. **Draw electron dot diagrams and structural formulas for the following molecules.**
  - **(a) CS_2**
  - **(b) HCl**
  - **(c) C_2H_6**
  - **(d) C_2F_4**
  - **(e) I_2**
  - **(f) C_3H_8**
  - **(g) HCN**

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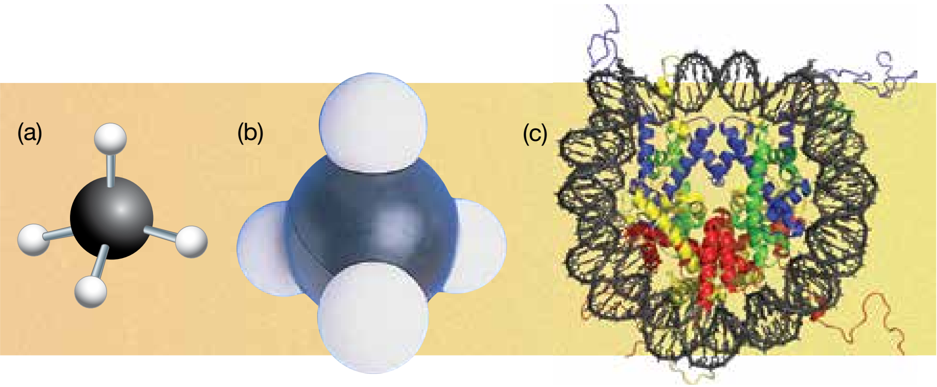
Unit 1	**Representing molecules with models**
AOS 2	Summary screen and practice questions
Topic 1	
Concept 10	

Molecular models

Physical models of molecules help us visualise their shapes so that we can better understand their behaviour. However, it must be remembered that models have limitations.

There are three main types of three-dimensional model used by chemists to represent molecules: **ball-and-stick** models, **space-filling** models and **computer-generated** models. Examples of these are shown below.

Molecules may be represented by ball-and-stick, space-filling or computer-generated models.



Molecular models allow us to visualise molecule shapes. They range from the (a) ball-and-stick and (b) space-filling models of methane to (c) complex computer-generated molecules, such as this genetic material.

Electron dot diagrams and structural formulas do not show the three-dimensional shapes of molecules.

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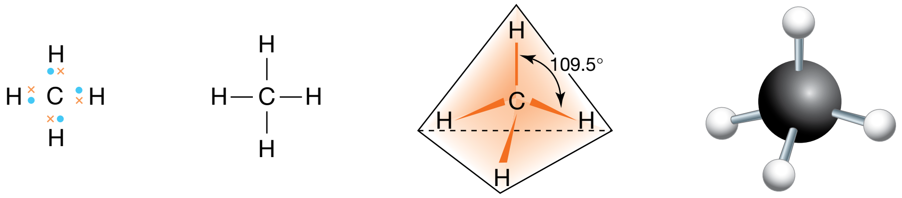
Unit 1	**VSEPR hypothesis and molecular shape** Summary screen and practice questions
AOS 2	
Topic 1	
Concept 5	

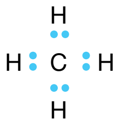
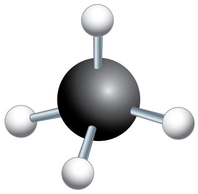
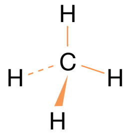
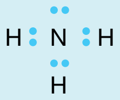
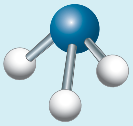
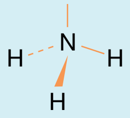
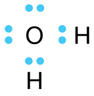
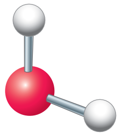
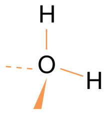
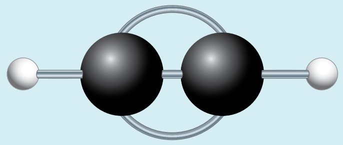
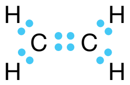
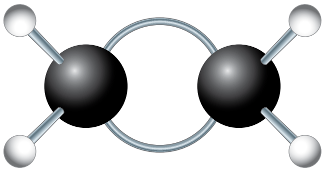
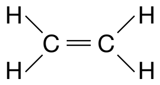
Predicting molecular shape

All discrete molecules have a definite three-dimensional shape. Electron dot diagrams and structural formulas fail to represent the three-dimensional shapes of molecules. For example, the electron dot diagram and structural formula of methane, CH_4, show the molecule in only two dimensions. In reality, methane molecules exist in three dimensions. The hydrogen atoms are arranged in a tetrahedral shape around the central carbon atom as shown by the ball-and-stick diagram in the figure below.



The electron dot diagram and structural formula of methane do not show its three-dimensional structure, which is tetrahedral. This structure is better represented using a ball-and-stick model.

TABLE 6.2 Shapes of some common molecules

Compound	Lewis diagram	Number of lone pairs around central atom	Number of bonding electron groups (pairs) around central atom	Shape	Ball-and-stick model	Shape diagram
methane $CH_4(g)$		0	4	tetrahedral (around C atom)		
ammonia $NH_3(g)$		1	3	pyramidal (around N atom)		
water $H_2O(l)$		2	2	V-shaped (around O atom)		
ethyne $C_2H_2(g)$		0 for each C atom	2 for each C atom	linear (around each C atom)		$H-C\equiv C-H$
ethene $C_2H_4(g)$		0 for each C atom	3 for each C atom	planar (around each C atom)		

The shapes of molecules can be predicted using VSEPR theory.

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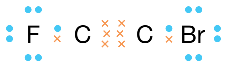
The **valence shell electron pair repulsion theory (VSEPR theory)** provides a relatively simple basis for understanding and predicting molecular geometry. The theory requires only that the number of outershell electrons for each atom in the molecule be known. The electron pairs in the molecule repel each other and take up positions as far from one another as possible. When determining the shapes of molecules, the electron pairs of a multiple bond count as only one pair of electrons for prediction purposes. Examples of the five most common molecular shapes are given in table 6.2. Note that, in the shape diagrams, a dotted line indicates that the bond is directed into the plane of the paper and the wedge indicates that the bond is directed out of the plane of the paper. The solid lines represent bonds that are in the plane of the paper.

Sample problem 6.2

Use table 6.2 to predict and draw the shape diagram of C_2FBr.

Solution:

STEP 1

The electron dot diagram of C_2FBr is 

STEP 2

There are two central carbon atoms in this molecule, so we must count the number of bonding pairs and lone pairs around each carbon atom. (Note that, for purposes of determining shape, the electrons of multiple bonds are counted as a single bonding pair). So around each carbon atom there are two bonding pairs and no lone pairs.

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Lewis structures

STEP 3

From table 6.2, we see that the best orientation to minimise repulsion is a linear arrangement of the other atoms around each carbon atom, as shown below.

$$F-C\equiv C-Br$$

Revision question

8. **Copy and complete the following table.**

Compound	Electron dot diagram of molecule	Number of lone pairs around central atom	Number of bonding pairs around central atom	Shape diagram and name of shape
NI_3				
CF_4				
OF_2				
HOF				
C_2F_2				
C_2Cl_4				

Naming covalent compounds

Molecular elements

Names of elements do not indicate how many atoms are bonded in a molecule.

Not all non-metallic elements are molecular. The noble gases, for example, exist as single atoms, and carbon and silicon exist as continuous arrays of atoms. However, all molecular elements are non-metallic elements. These non-metallic molecules are known simply by their elemental name. N_2, for example, is nitrogen. When we talk about hydrogen gas, we mean H_2 molecules, *not* H atoms. It is not necessary to indicate the number of atoms that are bonded together when naming elements.

Molecular compounds

Molecular compounds contain discrete molecular units. They consist of more than one type of non-metallic atom. **Binary** molecular compounds are those that contain two elements. Naming these is similar to naming binary ionic compounds. However, it is quite common for one pair of elements to form several different compounds. To avoid confusion in naming the compounds, Greek prefixes are used to indicate the number of atoms of each element present. Carbon and oxygen, for example, form carbon monoxide, CO, and carbon dioxide, CO_2.

There are four rules for naming binary covalent compounds:

1. The first element in the formula is named in full.
2. The second element is named as if it was an anion. The name is shortened and given an -*ide* suffix.
3. The number of each type of atom in the molecule is indicated by a prefix. These prefixes are listed in table 6.3. Note that the prefix *mono*- is never used for the first element.
4. If the name of the second element begins with a vowel and the prefix for that name ends with *a* or *o*, these letters are dropped to make pronunciation easier.

Some molecular compounds have trivial names, or common names, by which they are more conveniently known. For example, H_2O is known as water, rather than dihydrogen oxide. Table 6.4 lists some other examples of molecular substances with common names.

TABLE 6.3 Greek prefixes

Prefix	Meaning
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

TABLE 6.4 Common names and uses of molecular substances

Formula	Common name	Use
NH_3	ammonia	refrigerant
H_2O_2	hydrogen peroxide	bleach
O_3	ozone	blocks harmful UV light

Sample problem 6.3

Name the following compounds.
(a) SO_3
(b) N_2O_5

Solution: (a) Since one atom of sulfur and three oxygen atoms are present, the compound is called sulfur trioxide.
(b) Two nitrogen atoms and five oxygen atoms are present, so the compound is called dinitrogen pentoxide. Note that the *a* is omitted in *penta*-.

Sample problem 6.4

Write the formulas for the following molecular compounds.
(a) silicon tetrafluoride
(b) nitrogen tribromide

Solution: (a) Since one silicon atom and four fluorine atoms are present, the formula is SiF_4.

(b) One nitrogen atom and three bromine atoms are present, so the formula is NBr_3.

(*Note:* The order of elements is the same in both the molecule's formula and its name.)

Revision questions

9. **Use prefixes to name the compounds whose common names are shown in table 6.4.**

10. **Name the following molecular compounds.**
(a) NCl_3 (b) SF_6 (c) Cl_2O (d) SiH_4 (e) PH_3

11. **Write chemical formulas for the following molecular compounds.**
(a) **carbon monoxide**
(b) **silicon tetrachloride**
(c) **diphosphorus pentoxide**
(d) **nitrogen monoxide**
(e) **sulfur trioxide**

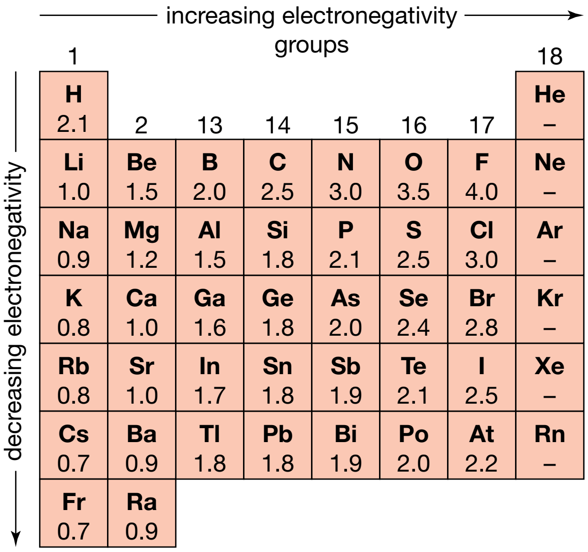
12. **Copy and complete the following table.**

Chemical formula	Name of compound
	dinitrogen monoxide
HCl	
P_4O_{10}	
	hydrogen fluoride

Electronegativity is the electron-attracting power of an atom.

Electronegativity

Covalently bonded atoms usually exhibit unequal attractions for shared electrons. It has been found that different atoms have different electron-attracting abilities. The relative attraction that an atom has for shared electrons in a covalent bond is known as its electronegativity. A scale of electronegativities was developed by Linus Pauling in which the most electronegative atom, fluorine, is assigned a value of 4.0. Fluorine attracts electrons almost twice as well as hydrogen, which is given an electronegativity value of 2.1. No values are assigned for the noble gases.



1	2	13	14	15	16	17	18
H 2.1							He –
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne –
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar –
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr –
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe –
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn –
Fr 0.7	Ra 0.9						

The Pauling scale of electronegativities shows electronegativity trends in the periodic table.

The Pauling scale of electronegativities

The following electronegativity trends may be seen in the periodic table.

1. Electronegativities increase from left to right within a period.
2. Electronegativities decrease from top to bottom within a group.
3. Metals generally have lower electronegativities than non-metals.

Non-polar and polar covalent bonds

Covalent bonds in which the bonding electron pair is shared equally and is uniformly distributed between the nuclei of two bonded atoms are called **non-polar covalent bonds**. Such bonds can result only when two atoms of the same element or two atoms of equal electronegativity simultaneously attract a shared pair of electrons.

TABLE 6.5 Non-polar covalent bonding in molecules

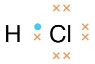
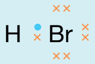
Molecule	Electron dot diagram	Structural formula
hydrogen		H — H
chlorine		Cl — Cl
oxygen		O = O
carbon disulfide		S = C = S

A non-polar covalent bond is formed between atoms with the same electronegativity.

A polar covalent bond is formed between elements with different electronegativities.

Covalent bonds in which the bonding electrons are unequally shared, and therefore unsymmetrically distributed between the nuclei of two bonded atoms, are called **polar covalent bonds**. Such bonds occur between atoms of different electronegativities. The shared pair of electrons move closer to the more electronegative atom. This means that the atom that has greater control of the electron pair becomes slightly negatively charged (δ–), while the atom that lost some control of the electron pair becomes slightly positively charged (δ+). Polar covalent bonds can be said to have a charge separation, or a **bond dipole**.

TABLE 6.6 Polar covalent bonding in molecules

Molecule	Electron dot diagram	Structural formula
hydrogen chloride		δ+ H — Cl δ–
hydrogen bromide		δ+ H — Br δ–

The type of bond formed depends on the difference in electronegativity between the atoms.

A polar covalent bond is not purely covalent, since there is not an *equal* sharing of electrons. It has some characteristics of ionic bonding, although the transfer of electrons from one atom to the other is not complete, as it is in a purely ionic bond. In fact, if the difference in electronegativity between two cations is 2 or greater on the Pauling scale, an ionic bond does form. The relationship between electronegativity and bond type is shown in table 6.7.

TABLE 6.7 Relationship between electronegativity and bond type

Difference in electronegativity between bonding atoms	Type of bond formed
zero (e.g. H (2.1) and H (2.1))	non-polar covalent
medium (e.g. H (2.1) and Cl (3.0))	polar covalent
large (e.g. Na (0.9) and F (4.0))	ionic

Revision questions

13. **Use the periodic table on page 121 to predict whether the following bonds are polar or non-polar. Show the bond dipoles where they are present using δ+ and δ–.**
(a) H—F (b) O—H (c) C—H (d) N—H (e) C—C

14. **Classify the following bonds as ionic, polar covalent or non-polar covalent.**
(a) HI (b) KCl (c) F_2

15. **Arrange the following bonds in order from non-polar covalent to ionic.**
Li—F, N—N, F—N, Cl—N

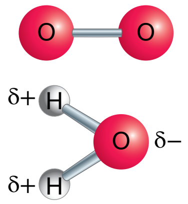
Non-polar and polar molecules

Molecules containing only non-polar bonds are non-polar molecules.

In a polar molecule, one end is slightly negative and the other end is slightly positive.

Molecules containing only non-polar bonds are **non-polar molecules**, such as O_2.

In a **polar molecule**, one end of the molecule is slightly negative and one end is slightly positive. Such a molecule is sometimes called a dipolar molecule, or molecular dipole. All polar molecules contain polar bonds. However, some molecules with polar bonds are non-polar. This occurs when the individual bond polarities are arranged in such a way that they cancel each other out. So having polar bonds in a molecule does not necessarily mean that the molecule is polar overall. The polarity or non-polarity of the molecule depends, rather, on the *direction* of the bond dipoles in the molecule, which is determined by the molecule's *shape*. The figure below shows examples of both a non-polar and polar molecule.



In the non-polar oxygen molecule, the oxygen atoms have equal electronegativity and share electrons equally. Water molecules, however, are polar. Since the oxygen atom has greater electronegativity than the hydrogen atoms, it therefore becomes the more negatively charged end of the molecular dipole.

To determine the polarity of a molecule the following steps should be followed:

1. Draw an electron dot diagram of the molecule.
2. Apply the VSEPR rules to draw a shape diagram of the molecule.
3. Use electronegativities to determine bond dipoles.
4. Use the shape diagram along with bond dipoles to determine whether the molecule is polar or non-polar. If bond dipoles cancel each other out, the molecule is non-polar. If bond dipoles do not cancel each other out, then one side of the molecule attracts electrons more than another side and the molecule is polar.

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Molecular models

The polarity can be determined by examining the shape of the molecule.

Sample problem 6.5

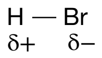
Predict the polarity of the following molecules.

(a) iodine
(b) hydrogen bromide
(c) carbon dioxide
(d) water

Solution: First draw the structural formula and, if necessary, shape diagram for each molecule and then decide if the dipoles cancel each other out.

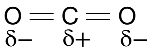
I — I

(a) Iodine is a linear molecule as there is only one bonding pair of electrons between the two iodine atoms. There is no bond dipole present as there is no electronegativity difference between the two identical iodine atoms. Hence, the molecule is non-polar.



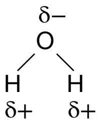
$\overset{\delta+}{H}-\overset{\delta-}{Br}$

(b) Hydrogen bromide is a linear molecule as there is only one shared pair of electrons between the hydrogen and bromine atoms. The bromine atom has a higher electronegativity than the hydrogen atom, resulting in a dipole. The molecule is polar.



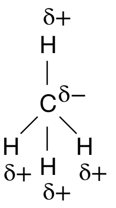
$\overset{\delta-}{O}=\overset{\delta+}{C}=\overset{\delta-}{O}$

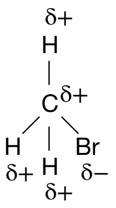
(c) Carbon dioxide, CO_2, is a linear molecule that has two polar bonds. However, the bond polarities cancel each other out because they are in opposite directions. Electrons are not attracted preferentially to one side or the other of the molecule. Carbon dioxide is therefore a non-polar molecule.



(d) Water, H_2O, is a V-shaped molecule due to the effect of the two lone pairs. The two O—H bond dipoles are equal but since they are not at an angle of 180° to one another (that is, aligned in opposite directions) they do not cancel each other out. A water molecule is therefore polar.

The polarity of a molecule can be predicted from a knowledge of its shape. The following general rules can be applied.





- Linear, planar and tetrahedral molecules with equal polar bonds that cancel each other out are non-polar. If the bond dipoles are not equal or do not cancel each other out, the molecule is polar. For example, CH_4 is non-polar whereas CH_3Br is polar because the C—Br bond dipole is not cancelled by the C—H bond dipoles.
- V-shaped or pyramidal molecules are polar as their polar bonds do not cancel.

CH_4 is non-polar as the bond dipoles cancel each other out. CH_3Br, however, is polar.

study on

Unit 1
AOS 2
Topic 1
Concept 6

Polar bonds and polar molecules
Summary screen and practice questions

Revision question

16. **Sketch the shapes of the following molecules and use your knowledge of electronegativity to draw bond dipoles for each and predict whether it is polar or non-polar.**

(a) HI
(b) SiH_4
(c) CS_2
(d) H_2S
(e) CH_3Cl
(f) SF_2
(g) NF_3
(h) CH_3F
(i) $SiBr_4$
(j) HCl
(k) N_2

Intermolecular attractions

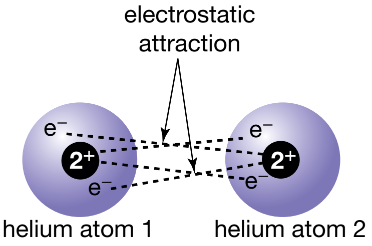
The existence of molecular crystals indicates that molecules can be held together in an orderly array. How do these molecules 'stick' together?

In addition to covalent bonds *within* molecules, called **intramolecular** bonding, there are attractions *between* molecules, called **intermolecular** forces of attraction, which hold molecules to each other. These attractive forces are weaker than either covalent or ionic bonds, but they determine whether a molecular compound exists in the solid, liquid or gaseous states. The temperature at which a molecular substance melts or boils, therefore, depends on the strength of the intermolecular forces. The stronger the intermolecular forces the higher the melting or boiling point.

Intramolecular bonding is bonding within molecules.

The force of attraction between molecules is called an intermolecular force.

Intermolecular forces include dispersion forces and dipole–dipole interactions.



There are three types of intermolecular forces. In some molecular substances there may be more than one force operating between the molecules. The weakest attractions between molecules are **dispersion forces** and **dipole–dipole interactions**. The third and strongest intermolecular force is called **hydrogen bonding**.

When sufficiently cooled, helium atoms are very weakly attracted to each other by the electrostatic attraction of one nucleus to the electrons of other atoms.

Dispersion forces

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AOS 2
Topic 1
Concept 3

Intramolecular and intermolecular bonding
Summary screen and practice questions

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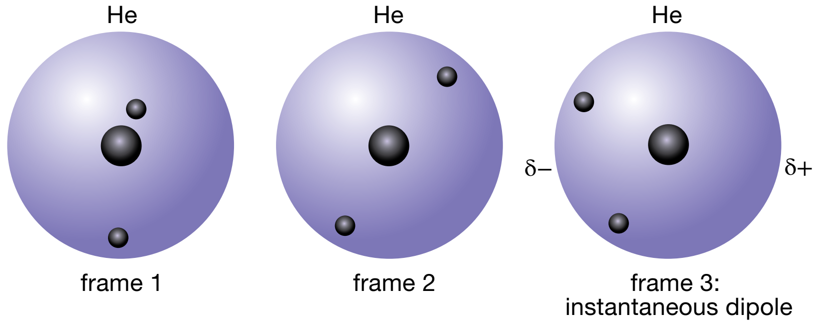
Unit 1
AOS 2
Topic 1
Concept 7

Intermolecular bonding: dispersion forces
Summary screen and practice questions

In a sample of molecules, the nuclei of atoms in one molecule are able to attract the electrons of atoms in neighbouring molecules (in addition to attracting their own electrons in forming a covalent bond). All electrons are attracted by all neighbouring nuclei.

Dispersion forces are the weakest of the intermolecular forces. To be condensed to a liquid, the noble gas helium must be cooled to –269 °C so that the atoms are travelling slowly enough for the dispersion forces to have a better chance of pulling these two atoms together. The helium nucleus in one atom attracts the electrons in a neighbouring atom. If further cooled to –272 °C, the liquid helium solidifies.

Dispersion forces can also occur between instantaneous dipoles. Electrons are moving constantly. At any point in time there is a chance that the electrons may be found on one side of an atom or molecule, and it can develop an 'instantaneous dipole.' This temporary dipole can cause a shift in the distribution of electrons in neighbouring atoms or molecules, resulting in 'induced dipoles.' Very small forces of attraction result between the particles.

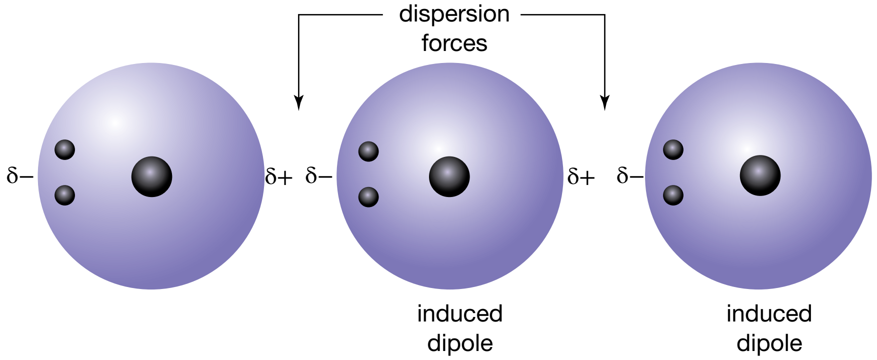


The electrons in this helium atom are moving all the time; sometimes they create an instantaneous dipole.

Dispersion forces occur when nuclei attract electrons from neighbouring atoms.

Two factors influence the strength of dispersion forces:

1. The number of electrons in the molecules. In general, the more electrons the molecules of a substance have, the stronger the dispersion forces between them.
2. The shapes of the molecules. Shape affects how closely the molecules may approach each other in the solid and liquid states. The closer the molecules can get, the stronger the attraction is. All covalent molecular substances have dispersion forces between their molecules.



The atom on the left has caused a temporary induced dipole in the middle atom, which in turn has caused another temporary induced dipole in the atom on the right.

Consider the noble gases He and Ne, or small molecules such as F_2, H_2 and CH_4. These atoms or molecules have relatively few electrons and so have very weak dispersion forces. Such substances have low boiling points and exist as gases at room temperature. Larger compounds, such as octane, C_8H_{18} (a component of petrol), exist as liquids. However, candle wax, $C_{25}H_{52}$, is a solid with a low melting point.

This trend can be illustrated using the halogens, as shown in table 6.8. As we go down the group, the molecules become larger and there is a progression from a gaseous to liquid and ultimately solid state. Boiling and melting points of halogens show a corresponding increase.

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Unit 1 **AOS 2** **Topic 1** **Concept 8**

Intermolecular bonding: dipole–dipole interactions
Summary screen and practice questions

TABLE 6.8 Properties of group 17 elements

Name	Formula	State	Melting point (°C)	Boiling point (°C)
fluorine	F_2	gas	–220	–188
chlorine	Cl_2	gas	–101	–34
bromine	Br_2	liquid	–7.3	59
iodine	I_2	solid	113	184

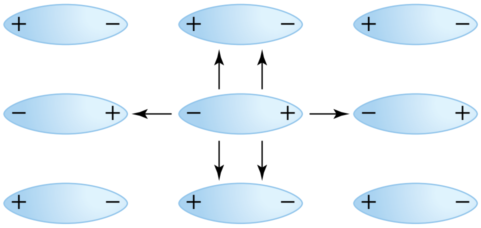
The strength of dispersion forces depends on the number of electrons and the shape of the molecule.

Dipole–dipole interactions

If the molecules in a sample are polar, the presence of molecular dipoles causes simultaneous intermolecular attraction. The positive side of one molecule attracts the negative side of another molecule, which attracts the next, and so on to the limits of the sample.

Polar molecules have two forces of attraction operating between their molecules. If two molecules are of similar size — that is, have a similar number of electrons — then the dispersion forces acting on them are similar. However, if one of the molecules is polar, it will also be affected by dipole–dipole attractions. This results in a stronger overall intermolecular force and a higher boiling point. Consider argon (boiling point –186 °C) and HCl (boiling point –83.7 °C). Both substances have 18 electrons, yet HCl has a much higher boiling point. This is because HCl is a polar molecule and so has an extra intermolecular force of dipole–dipole interactions. Such a compound is generally more likely to exist as a liquid or solid at room temperature than a non-polar compound of similar size.

Dipole–dipole interactions occur when the positive side of one molecule attracts the negative side of another molecule.



In dipole–dipole interactions, the central polar molecule is attracted to other polar molecules around it. They, in turn, are attracted by their neighbours.

Hydrogen bonding

Hydrogen bonding is a special case of dipole–dipole attraction. Hydrogen is an atom containing one proton in the nucleus and one electron revolving around it.

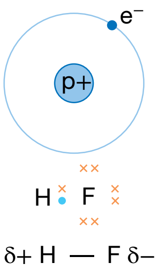
When hydrogen bonds to a more electronegative atom such as nitrogen, oxygen or fluorine, its electrons move slightly toward that atom.

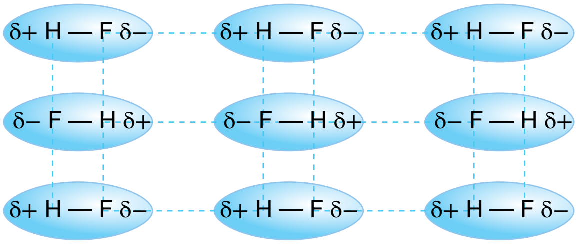
This causes the hydrogen nucleus to be exposed or unshielded. The molecule that forms is a dipole.

The negative end of one of these dipoles is attracted to the positive end of another.

Because the hydrogen (positive end) is unshielded, the other dipole can approach far more closely. The closer the dipoles get, the stronger the bond that forms. The bond between the dipoles is called a hydrogen bond. These bonds are represented in the figure below by dotted lines, indicating that they are a weaker bond type than a covalent bond. Hydrogen bonds form only *between* molecules and only when hydrogen has been bonded to fluorine, oxygen or nitrogen.

Hydrogen bonding occurs between two molecules in which hydrogen is bonded to nitrogen, oxygen or fluorine.





The negative ends of the dipoles are attracted to the positive ends, causing a hydrogen bond to form.

Hydrogen bonds are stronger than other dipole–dipole bonds and result in materials with higher melting and boiling points than would otherwise be expected.

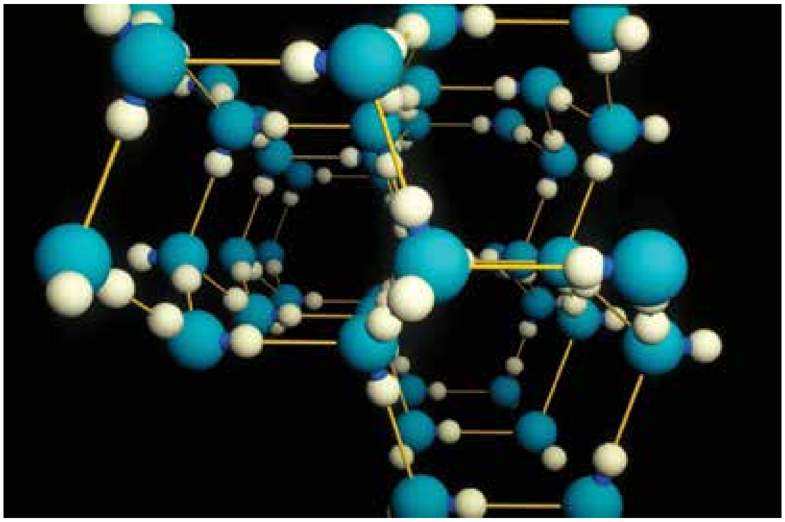
Hydrogen bonding is the strongest intermolecular force.

Effects of hydrogen bonding on the properties of water

Unusual properties of water include its relatively high melting and boiling temperature and its expansion upon freezing. Its ability to form hydrogen bonds also explains why it is such a good solvent for other polar substances as well as many ionic substances.

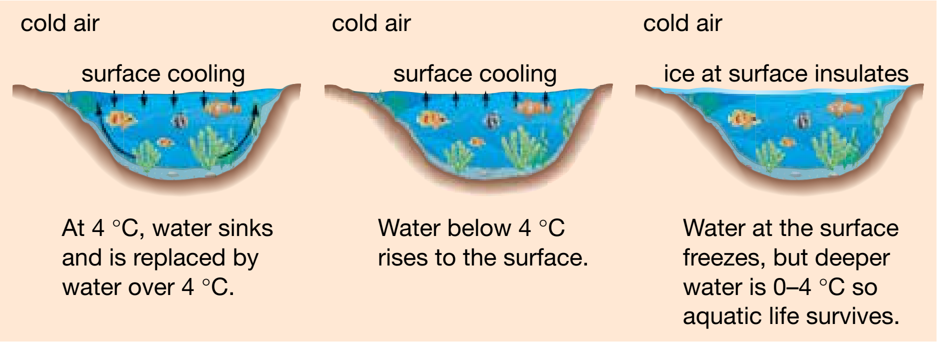
Hydrogen bonding significantly affects the physical properties of water. The existence of hydrogen bonding between water molecules results in:

- the relatively high melting and boiling points of water compared with other substances. More heat is required to enable the molecules to gain sufficient kinetic energy to break free of the hydrogen bonds, which are stronger than dispersion forces alone. For example, H_2O is a liquid at room temperature with a boiling point of 100 °C, whereas H_2S, which has more electrons, is a gas at room temperature and boils at –61 °C. The behaviour of H_2O as it changes state between liquid water and solid ice is also governed by the hydrogen bond.
- the expansion of water upon freezing. Generally, as substances are heated they expand. However, water behaves a little differently. The density of water is greatest at 4 °C. (Density refers to how much mass is in a given volume; for example, oil is less dense than water so it floats on top of water.) Water does expand when heated from 4 °C but, unusually, it also expands as the temperature decreases from 4 °C. As the temperature continues to decrease, water at 4 °C sinks and eventually the temperature at the surface becomes 0 °C and freezes. The open, hexagonal crystalline lattice of ice places the water molecules further apart than occurs in the liquid state. Because water expands on freezing, it is less dense as a solid. Ice, therefore, floats on water.



Snow and ice are both the same molecular solid with a unique structure. Each water molecule has four hydrogen bonds that connect it to four other water molecules. The molecules in liquid water are hydrogen bonded but free to move about, while those in snow are held further apart and in a more orderly arrangement than water owing to the extensive hydrogen bonding. A three-dimensional model shows the hexagonal structure of ice.

If ice was denser than water, some bodies of water would freeze solid during winter and the aquatic life would die. Icebergs form when sections of the frozen icecap break off and float in the sea. When ice forms on the surface of water, it acts as an insulator, preventing the water below from freezing. This means that aquatic life can survive, even in sub-zero conditions.



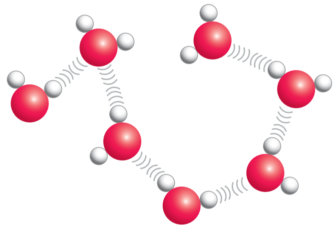
At 4 °C, water sinks and is replaced by water over 4 °C.

Water below 4 °C rises to the surface.

Water at the surface freezes, but deeper water is 0–4 °C so aquatic life survives.

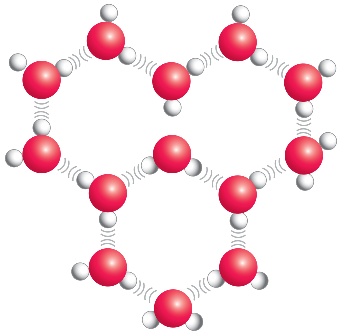
The density of water is greatest at 4 °C.

(a)



Water molecules move too quickly when in liquid form to form regular structures; however, they still remain attached.

(b)



When in solid form (ice), water molecules join together to form hexagonal structures.

Hydrogen bonding in water

Sample problem 6.6

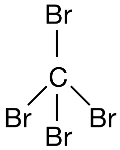
Name the type(s) of intermolecular forces that exist between molecules in each of the following substances.

(a) HBr (b) CBr_4 (c) H_2 (d) CH_3OH

Solution:

(a) HBr is a polar molecule, and therefore has both dispersion and dipole–dipole forces between the molecules.

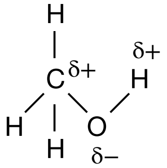
(b) CBr_4 is a non-polar tetrahedral molecule. There are only dispersion forces between these molecules.



(c) H_2 is a non-polar linear molecule. It has only dispersion forces between its molecules.

H — H

(d) CH_3OH has a tetrahedral shape around the carbon atom and a V-shape around the second central atom, oxygen, resulting in a polar molecule. It also has a highly electronegative atom (oxygen) bonded to a hydrogen atom. So, dispersion forces and hydrogen bonding exist between these molecules.



study on

Unit 1
AOS 2
Topic 1
Concept 9

Hydrogen bonding in water
Summary screen and practice questions

Revision questions

17. **Name the type(s) of intermolecular forces that exist between molecules in each of the following species.**

- **(a)** CH_4
- **(b)** Cl_2
- **(c)** HCl
- **(d)** CH_3F
- **(e)** NH_3

18. **The boiling points of HCl, HBr and HI are, respectively, −85 °C, −67 °C and −35 °C. Explain the difference in these boiling points with reference to the forces that exist between the molecules of each substance.**

Properties of molecular substances in relation to structure

Although the covalent bonding holding the atoms together in a molecule is strong, the forces between molecules are usually weak. The physical properties of a compound depend on the type of bonding it displays. Ionic compounds

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Intermolecular forces

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Unit 1 **AOS 2** **Topic 1** **Concept 4**

Molecular solids
Summary screen and practice questions

Molecular substances do not conduct electricity in the solid or molten states/and have low melting and boiling temperatures.

Polar substances dissolve in polar solutions.

are crystalline solids, but molecular substances may exist as solids, liquids or gases. A great variety of physical properties occurs among molecular substances due largely to the differences in strength of the intermolecular attractions. Even so, a few generalisations can be made.

- Molecular substances do not conduct electricity in the solid or molten form because the molecules are electrically neutral.
- If dissolved in water without reacting with it, molecular substances do not conduct electricity, again because the molecules are electrically neutral.
- Some molecules dissolve in water to produce ions. These molecules are said to ionise when they are dissolved in water, and can conduct electricity owing to the movement of the ions produced.
- Molecular substances vary in their solubility in water and other solvents. Generally, polar compounds are soluble in polar solvents such as water but insoluble in non-polar solvents such as tetrachloromethane. Non-polar solvents tend to be more soluble in non-polar solvents than polar solvents. This is sometimes described as the 'like dissolves like' rule.
- Molecular compounds have low melting and boiling points since the forces between the molecules are weak, and relatively little energy is required to break them. Many molecular substances are gases or liquids at room temperature.
- Most molecular substances are soft and easily scratched.

Sample problem 6.7

The compounds Br_2 and ICl have the same number of electrons, yet Br_2 melts at –7.2 °C, and ICl melts at 27.2 °C. Explain this in terms of intermolecular forces.

Solution: Br_2 is a non-polar molecule and has only dispersion forces between its molecules. ICl is a polar molecule and has dipole–dipole forces as well as dispersion forces between its molecules. This extra force causes the higher melting point as more heat energy is required to separate the molecules.

Revision questions

19. **N_2 has 14 electrons and Cl_2 has 34 electrons. Predict which of these substances would have the higher boiling point. Justify your answer in terms of intermolecular forces.**

20. **Two noble gases, helium and argon, have boiling points of –269 °C and –186 °C respectively. Explain the large difference in their boiling points.**

21. **Both Kr (boiling point –152 °C) and HBr (boiling point –67 °C) have the same number of electrons. Explain what factors could affect intermolecular forces to cause the difference in boiling points between Kr and HBr.**

22. **HCl has more electrons than HF so we would expect it to have the higher boiling point. However, this is not the case, as the boiling points of HCl and HF are –83.7 °C and 19.4 °C respectively. Explain what factors could account for this reversal in trend.**

23. **Explain why glucose, $C_6H_{12}O_6$, dissolves in water but does not conduct electricity in solid or aqueous form.**

24. **Explain why F_2, O_2 and N_2 are all gases at room temperature.**

25. **Predict whether I_2 dissolves more readily in non-polar tetrachloromethane or in polar water. Explain your answer.**

26. **Would you expect candle wax, $C_{25}H_{52}$, to be:**
(a) soft or hard (b) soluble in water (c) an electrical conductor?
Explain your predictions using your knowledge of structure and bonding.

Chapter review

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Unit 1 **AOS 2** **Topic 1**

Covalent bonding
Summary screens and practice questions

Summary

- Covalent molecular substances are substances held together by covalent bonds and include elements and compounds. These substances are made up of molecules.
- A molecule is a discrete (separate) group of non-metallic atoms held together by covalent bonds. The atoms are combined in a fixed ratio and are electrically neutral.
- A covalent molecular element is made up of molecules of the *same* element. The molecules may contain varying numbers of atoms. These include diatomic molecules, which are composed of two atoms that share electrons (e.g. Cl_2).
- A covalent molecular compound is made up of atoms of *different* elements (e.g. CH_4).
- Atoms tend to lose, gain or share electrons in order to achieve a stable noble gas configuration of eight electrons in the valence shell. The exception to this rule is the hydrogen atom, which needs only two electrons to complete its outer shell.
- Covalent bonds are produced when non-metallic atoms react together, *sharing* pairs of electrons so that each may get eight outershell electrons. These bonds are the result of the force of attraction between shared electrons and the nuclei of the non-metal atoms in the bond.
- Covalent bonds may be single bonds (formed by one shared pair of electrons) or multiple bonds.
  - In a double bond, two electron pairs are shared by two atoms.
  - In a triple bond, three electron pairs are shared by two atoms.
- There are several ways of representing molecules in order to understand their structures. We use ball-and stick, space-filling and computer-generated models to represent the three-dimensional structure of molecules. Two-dimensional models include:
  - electron dot diagrams
  - structural formulas, which show each pair of shared electrons as a dash (—).
- Though electrons are identical, when representing a covalent bond using a Lewis diagram they may be thought of as either bonding or non-bonding electrons.
  - Bonding electrons are the single electrons that are available for sharing, and they determine the number of bonds an atom can form with other non-metals.
  - Non-bonding electrons or lone pairs are paired electrons that are not available for sharing.

  The atom with the most bonding electrons is known as the central atom.
- Molecule shapes can be determined using electron dot diagrams and VSEPR theory, which allows us to determine the best shape for minimum repulsion between the electron pairs around the central atom. In general, the molecule shape is:
  - linear if the central atom has *one* bonding pair or one pair on either side (e.g. H_2, CO_2)
  - V-shaped if the central atom has *two* bonding pairs and *two* lone pairs (e.g. H_2O)
  - planar if the central atom has *three* bonding pairs and no lone pairs (e.g. C_2H_4).
  - pyramidal if the central atom has *three* bonding pairs and *one* lone pair (e.g. NH_3)
  - tetrahedral if the central atom has *four* bonding pairs and *no* lone pairs (e.g. CH_4).
- When naming binary molecular compounds:
  - name the first element in full
  - use the Greek prefix before the second element to indicate how many atoms of that element are in the molecule
  - name the second element, ending with *-ide*.
- Non-metallic atoms have high electronegativities, which means that they have a strong attraction for the shared electrons in a covalent bond. Differing electronegativities cause electrons to be unequally shared, and can affect the polarity of the bond, as can the shape of the molecule.
  - Non-polar covalent bonds occur in molecular elements, which share electrons equally as their atoms have the same electronegativity. Linear, planar and tetrahedral molecules are non-polar molecules when equal polar bonds cancel each other out.
  - Polar covalent bonds result when the two atoms in the bond have different electronegativities and the pair(s) of electrons are shared unequally. These bonds are called a bond dipole. Pyramidal or V-shaped molecules are polar molecules.
- In addition to the intramolecular forces that work within molecules to hold them together, weak bonds exist between molecules and are called intermolecular forces.
  - Dispersion forces are found in all atoms or discrete molecules and increase with the size of the atom or molecule and the corresponding number of electrons.
  - Dipole–dipole interactions are found only in polar discrete molecules.
- Hydrogen bonding is a special case of dipole–dipole interaction. It is found only in molecules where a

hydrogen atom is directly bonded to a more electronegative fluorine, oxygen or nitrogen atom. These bonds are stronger than other dipole–dipole bonds and result in higher melting and boiling points.

- Hydrogen bonding significantly affects the physical properties of water. Water has relatively high melting and boiling points and has the unusual property that the density of solid water is less than the density of liquid water.
- Discrete covalent molecular substances have similar properties. Molecular substances:
  - have low melting and boiling points due to weak intermolecular forces
  - are usually liquid or gaseous at room temperature due to weak intermolecular forces
  - do not conduct electricity as the molecules are electrically neutral
  - are soluble in water if polar, and soluble in nonpolar solvents if non-polar (the 'like dissolves like' rule).

Multiple choice questions

Covalent bonding

1. The sharing of electrons in bond formation *always* involves:

A the formation of positive and negative ions
B the formation of polar molecules
C shared electrons being attracted more by one atom than another
D the bonded atoms having greater stability than the unbonded atoms.

2. Which one of the following electron dot diagrams best represents the bonding in the nitrogen molecule?

A


C


B


D


3. Which of the following is a characteristic property of a covalent molecular compound?

A Relatively low melting point
B Malleable and ductile
C High melting point
D Conducts electricity when molten but not when solid

4. The structure of C_2H_2 is:

A linear
B V-shaped
C planar
D pyramidal.

5. Which of the following best describes the molecular shape, polarity of bond and molecular polarity of the H_2S molecule?

A Linear, polar covalent, non polar
B Linear, polar covalent, polar
C V-shaped, ionic, polar
D V-shaped, polar covalent, polar

6. Which of the following substances might one expect to exhibit the weakest intermolecular forces?

A HCl **C** NH_3
B He **D** H_2O

7. At room temperature, methane is a gas while tetrachloromethane is a liquid. Choose the statement that best explains this.

A There is appreciable hydrogen bonding in tetrachloromethane but not in methane.
B The tetrachloromethane molecule is polar, while the methane molecule is non-polar.
C Tetrachloromethane has an appreciably higher molecular mass than methane.
D The bonds in tetrachloromethane are polar covalent.

8. The compound tetraphosphorus decoxide has the formula:

A P_3O_2 **C** P_4O_{10}
B P_6O_{10} **D** P_3O_6.

9. The correct order of increasing attractive strength for weak intermolecular forces is:

A hydrogen bonding, dipole–dipole interaction, dispersion forces
B hydrogen bonding, dispersion forces, dipole–dipole interaction
C dipole–dipole interaction, hydrogen bonding, dispersion forces
D dispersion forces, dipole–dipole interaction, hydrogen bonding.

10. Which one of the following substances forms strong hydrogen bonds?

A HCOOH **C** SiH_4
B CH_3CN **D** CH_3OCl

11. Which of the following *best* explains the relatively low melting point of covalent molecular substances?

A Covalent molecular materials depend on weak electrostatic forces holding the ions together.
B The intermolecular forces between the molecules are weaker than ionic or covalent bonds.
C The metals involved create uneven bonding with the non-metals.
D The similar electronegativity of the atoms causes repulsions between the molecules.

12. The correct name for SO is:

A sulfur oxide
B sulfur monoxide
C sulfur dioxide
D sulfur oxygen.

13. What is the maximum number of double bonds that a hydrogen atom can form?

A 0 **C** 2
B 1 **D** 3

14. Using the VSEPR model, the molecular shape of the central atom in CF_4 is:

A trigonal planar
B tetrahedral
C V shaped
D trigonal pyramidal.

15. Which of the following is *not* a covalent compound?

A CaO
B SO_3
C NH_3
D CH_3Cl

Review questions

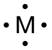
Covalent bonding

1. (a) Why do the noble gases *not* form molecules such as Ne_2 and He_2?
(b) In your own words, explain what a filled outer shell is.

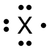
2. Describe the bonding in a molecule of:
(a) fluorine
(b) hydrogen.
How are the outer shells different?

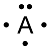
Lewis diagrams and structural formulas

3. Give the group number of the element that each of the following electron dot diagrams represents.

(a) 

(b) 

(c) 

(d) 

4. Use electron dot diagrams to draw the structural formulas of the following.
(a) H_2S
(b) NCl_3
(c) CH_4
(d) Br_2
(e) HOCl
(f) H_2CO

5. Give five examples of molecular elements, and draw their structures.

6. Give five examples of molecular compounds, and draw their structures.

Predicting shapes of compounds

7. Write the structural formulas for the following and name the shape of each molecule.
(a) O_2
(b) $CHCl_3$
(c) Br_2
(d) NH_3
(e) CH_3OH
(f) HCN
(g) OF_2
(h) CCl_4
(i) C_2H_2
(j) PCl_3
(k) N_2

8. Draw and name the shape diagrams of the following compounds.
(a) CF_4
(b) PF_3
(c) HCl
(d) H_2S
(e) NF_3
(f) C_2Br_4

Polarity

9. Predict whether the following bonds are non-polar covalent, polar covalent or ionic.
(a) H—H
(b) H—S
(c) C—S
(d) Cl—F
(e) Na—N

10. (a) Draw a polar molecule and explain what this means.
(b) Why are molecular elements always non-polar?
(c) Why are molecular compounds sometimes polar and sometimes non-polar?

11. State whether each molecule in questions 7 and 8 is polar or non-polar.

12. Predict whether the following molecules are polar or non-polar.
(a) HI
(b) SiH_4
(c) CS_2
(d) SF_2
(e) CH_3Cl
(f) C_2H_4

13. Explain why the ammonia molecule is polar but carbon dioxide molecules are not.

Intermolecular forces

14. (a) The boiling point of iodine is 184 °C, whereas fluorine has a boiling point of –188 °C. Explain this difference.
(b) What states do fluorine and iodine assume at room temperature? Explain.

15. CH_3F has a boiling point of –78 °C, whereas CH_3OH has a boiling point of 65 °C. Explain this difference.

16. Which of the following sets of compounds has the stronger intermolecular forces? Explain your answer.
(a) CO_2 or OCS
(b) HF or HBr

17. Arrange the following compounds in order of increasing boiling point: HCl, Ar, F_2.

18. Which of the following exhibits hydrogen bonding?
(a) H_2O
(b) C_2H_6
(c) CF_3H
(d) CH_3OCH_3
(e) CH_3OH
(f) CH_3CH_2OH
(g) NH_3
(h) HF

19. Methanol and ethanol have the formulas CH_3OH and CH_3CH_2OH. Predict which has the higher boiling point and give clear reasons why.

Properties in relation to structure

20. Pure liquid HCl is virtually a non-conductor of electricity, whereas a solution of this compound in water is an excellent conductor. Explain these observations.

21. Copy and complete the following to indicate:
    (a) whether the melting point of each substance is high or low
    (b) whether or not it conducts electricity.

Substance	Chemical formula	Melting point (high or low)	Electrical conductivity		
			Solid	Molten	Dissolved in water
magnesium nitrate	$Mg(NO_3)_2$				
copper	Cu				insoluble
neon	Ne				
methanol	CH_3OH				
chlorine	Cl_2				

Naming molecular compounds

22. Copy and complete the following table. The first line is done for you.

Molecular formula	Description/use	Name
NO	pollutant (car exhaust)	nitrogen monoxide
NO_2	pollutant (smog)	
	pollutant (smog)	sulfur dioxide
	colourless, odourless poison	carbon monoxide
	good in upper atmosphere	ozone
P_4	used in matches	
	cleaning agent when dissolved in water	ammonia
CH_4	85–95% of natural gas	
	rotten egg gas	dihydrogen sulfide

23. Name the following compounds.
    (a) NBr_3
    (b) N_2O_4
    (c) N_2O
    (d) N_2O_5
    (e) N_2O_3
    (f) PCl_3
    (g) PCl_5
    (h) P_2O_5
    (i) SF_6
    (j) SCl_2

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Unit 1
AOS 2
Topic 1
Covalent bonding
Sit topic test

Exam practice questions

In a chemistry examination you will be required to answer a number of short and extended response questions.

Multiple choice questions

1. A water molecule is polar because:

A hydrogen and oxygen atoms have the same electronegativity

B the molecule is linear with a net dipole

C oxygen atoms have a higher electronegativity than hydrogen atoms and it is non-linear

D the hydrogen atoms acquire a slight negative charge while the oxygen atom acquires a slight positive charge. **1 mark**

2. An element *X* whose atoms contain 13 electrons is chemically bonded to element *Y* whose atoms contain 9 electrons. The formula of the compound is:

A a covalent compound, XY

B an ionic compound, XY_3

C a covalent compound, XY_3

D an ionic compound, XY. **1 mark**

3. Consider the following four compounds: HF, HCl, HBr and HI. On the basis of the intermolecular forces operating between the molecules of each compound in the liquid state, which compound is most likely to have the highest boiling point?

A HF **B** HCl **C** HBr **D** HI **1 mark**

Extended response questions

4. Draw simple diagrams to show how electrons are arranged in the following covalent molecules.

(a) H_2S (b) Cl_2 (c) HF **3 marks**

5. The amount of CO_2 in the atmosphere is increasing.

(a) Use dot diagrams to determine the structure and shape of a CO_2 molecule.

(b) What is the shape of the CO_2 molecule?

(c) How many lone pairs are there in a molecule of carbon dioxide?

(d) Are the bonds in the molecule polar or non-polar?

(e) Is the molecule polar or non-polar?

(f) What type of interactions would you expect between the molecules in solid carbon dioxide? **6 marks**

6. Ammonia, NH_3, is a gas at room temperature. It has a characteristic pungent odour.

(a) Show how the shape of the molecule is determined.

(b) Discuss the polarity of the molecule.

(c) Explain in terms of the intermolecular bonding why ammonia is a gas at room temperature. **5 marks**

7. Water has some unique properties and we cannot live without it.

(a) Use a labelled diagram to explain why the water molecule is V shaped.

(b) Explain the intramolecular and intermolecular forces in water.

(c) Why is the density of ice less than the density of water?

(d) Describe an experiment to show the difference in density between water and ice. **9 marks**

CHAPTER 7

Networks and nanomaterials

The lead in the ‘lead’ pencil that you use to record your notes is not actually made of lead; it is made of a form of carbon called graphite that is mixed with clay. When you write, the pencil makes contact with the paper and many layers of carbon atoms are deposited. Although lead was used by the Romans to make marks on papyrus, it was not until the 1500s that graphite was used for writing. At that time, sticks of graphite were wrapped in string so that they were easy to hold. It was mistakenly thought that the shiny graphite was a form of lead. Graphite is also a source of a useful nanomaterial graphene, which has some remarkable properties. It is another fascinating fact of chemistry that diamonds are also made of carbon atoms, but the atoms in diamond are arranged quite differently.

YOU WILL EXAMINE:

- **bonding and properties in network lattices such as diamond**
- **bonding and properties in graphite**
- **the structure and properties of carbon nanomaterials**
- **applications of carbon nanomaterials**
- **the strength of the bonds in different types of lattice structures.**

> The noblest pleasure is the joy of understanding.
>
> *Leonardo da Vinci*

Diamonds are beautiful gemstones, but you may be surprised to know that, in chemistry, diamonds are regarded as unstable and are eventually converted into another allotrope of carbon, graphite. Fortunately, the time span for this conversion is incredibly large so there is no need to discard any diamonds that you may have.

studyon

Unit 1
AOS 2
Topic 2
Concept 1

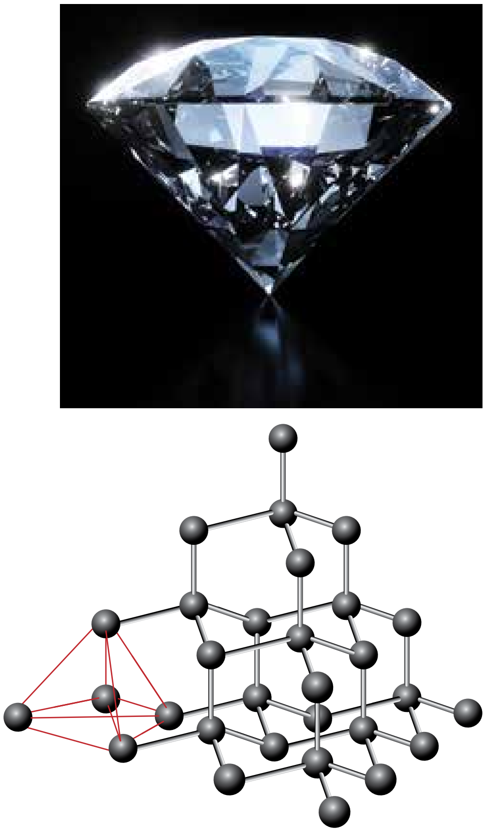
Diamond
Summary screen and practice questions

Carbon

What do you think is the most important element? Carbon should probably be at the top of the list. It has been known from ancient times and yet there are still many exciting discoveries of new materials based on carbon. Carbon with its four bonding electrons is present in all forms of life and is known to form millions of compounds. We saw in chapter 6 that this element forms simple, small molecular compounds such as methane, CH_4, and carbon dioxide, CO_2. We breathe out carbon dioxide, which is an essential reactant in photosynthesis, the process by which plants make food. Carbon dioxide produced as a result of human activity is increasing and so contributing to global warming. In chapter 4, you met ionic metal carbonate compounds that are components of many rocks.

Chapter 8 focuses on organic chemistry, which is a branch of chemistry that is devoted entirely to compounds of carbon from the small to very large. In chapter 9, you will see that most of the materials that are generally called plastics, but scientists call polymers, are also based on large molecules with carbon backbones. Organic compounds are obtained from fossil fuels that have been formed from the decayed remains of plants and formed over millions of years. Fossil fuels also provide us with energy for households, transportation and manufacturing. Carbon in its different forms is a major part of our lives.

How can both one of the hardest and one of the softest materials known consist of the same element? How can one material be an insulator and another a conductor? It is all about the arrangement of the atoms. Diamond and graphite are naturally occurring allotropes of carbon; 'allotrope' is the term used to describe different forms of the same element. In this chapter, carbon's ability to form both network lattices and layer lattices is discussed as well as new developments in nanotechnology.



Diamonds are prized for their brilliance, durability and hardness. These properties are a consequence of the ordered and rigid internal structure in which each carbon atom is covalently bonded to four other carbon atoms. Millions of carbon atoms can combine in a three-dimensional covalent network lattice to form a diamond crystal.

Revision questions

1. **(a) In what group and period would you find the element carbon?**
   (b) Draw and label a carbon atom $^{12}_{6}C$ showing protons, neutrons and electrons.
   (c) What is carbon's electron configuration?
2. **Why is carbon able to form such a variety of compounds?**

Same but different

Covalent network lattices

Some non-metals form giant structures in which no individual molecules exist. They consist of countless numbers of atoms covalently bonded to each other, forming a three-dimensional network lattice. Common examples of covalent network lattices include diamond, silicon carbide, silicon dioxide (quartz) and tungsten carbide.

Diamond — the hardest known substance

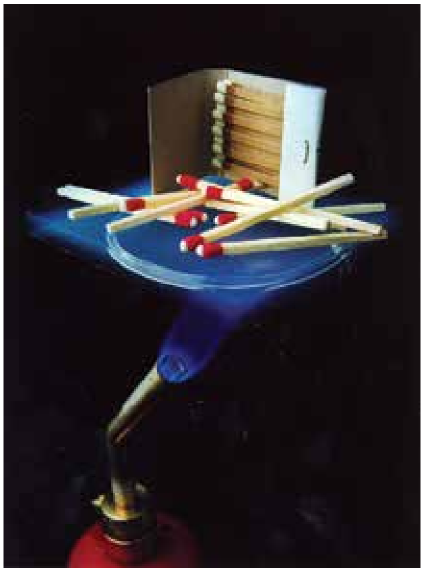
Diamond is the hardest substance known because of the bonding pattern of its carbon atoms. Hardness is measured on the Mohs scale (see table 7.1). The Mohs scale gives the order of hardness, but this is not linear; the difference between corundum and diamond is only 1 on this scale but diamond is many times harder than corundum. Diamond is used as an abrasive for sawing, cutting and grinding hard substances such as glass, stone and porcelain.

TABLE 7.1 Mohs hardness scale

Hardness	Mineral	Scratch test	Other materials
1	talc	very easily scratched by fingernail	sulfur, sodium, chalk
2	gypsum	scratched by fingernail	lead, **graphite**
3	calcite	very easily scratched with knife	gold, silver, copper
4	fluorite	easily scratched with knife	platinum, nickel
5	apatite	scratched with knife	tooth enamel, glass (5.5)
6	orthoclase	scratches glass	manganese, steel
7	quartz	scratches glass easily	vanadium
8	topaz	scratches glass very easily	cubic zirconia
9	corundum	cuts glass	tungsten carbide, ruby, sapphire
10	**diamond**	scratches all other materials	

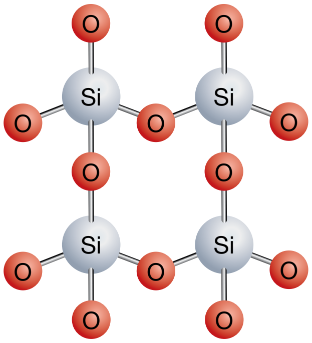
Carbon atoms in diamond are strongly bonded in a three-dimensional tetrahedral lattice structure.

A diamond is really one 'giant' molecule. It is made up of carbon atoms bonded together by strong covalent bonds. Each carbon atom has four covalent bonds around it and is surrounded by four other carbon atoms. The covalent bonding is three-dimensional since the bonds are arranged tetrahedrally around each atom. To scratch diamond it is necessary to separate the carbon atoms. This is very difficult to do as they are held together by four covalent bonds.



Matches on aerogel over a Bunsen burner flame. This shows both the low density and the effective insulating properties of this material. Aerogel, commonly referred to as 'frozen smoke', is a silicon-based solid comprising up to 99.9% air. It is 1000 times less dense than glass and 39 times more insulating than fibreglass. This material has uses as building insulation in walls and windows, improving energy efficiency.

Another covalent network lattice is silicon dioxide. Silicon dioxide is found as quartz (below left) in granite rock and in sandstone. Quartz, SiO_2, is a three-dimensional covalent network lattice (below right). Each silicon atom is bonded to four oxygen atoms, and each oxygen atom is bonded to two silicon atoms. The resulting substance is a hard, crystalline rock. Beach sand also contains silicon dioxide.



Covalent network lattices are generally hard, non-conductors of electricity, inert and insoluble in water.

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Allotropes of carbon
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Properties of covalent network lattices

Covalent network lattices have the following general properties:

- Since strong covalent bonding extends through the crystal structures of covalent network lattices, the particles are held quite rigidly. These substances are therefore very hard, difficult to scratch and have high melting points and boiling points.
- Since there are no free ions or electrons in the structure, covalent network lattices are usually non-conductors of electricity in the solid and liquid states.
- Covalent network lattices are brittle and must be cut in a specific way or they can shatter. When the covalent bonds break, the lattice is distorted.
- Covalent network lattices are chemically inert and are insoluble in water and most other solvents.

Covalent network lattices are generally hard, non-conductors of electricity, inert and insoluble in water.

Covalent layer lattices

Covalent layer lattices consist of countless atoms held strongly together in planes by covalent bonds. Covalent bonding is therefore present in two dimensions only. These layers are held together by weaker dispersion forces. Graphite is a common example of a covalent layer lattice.

Carbon atoms in graphite are arranged in a covalent layer lattice.

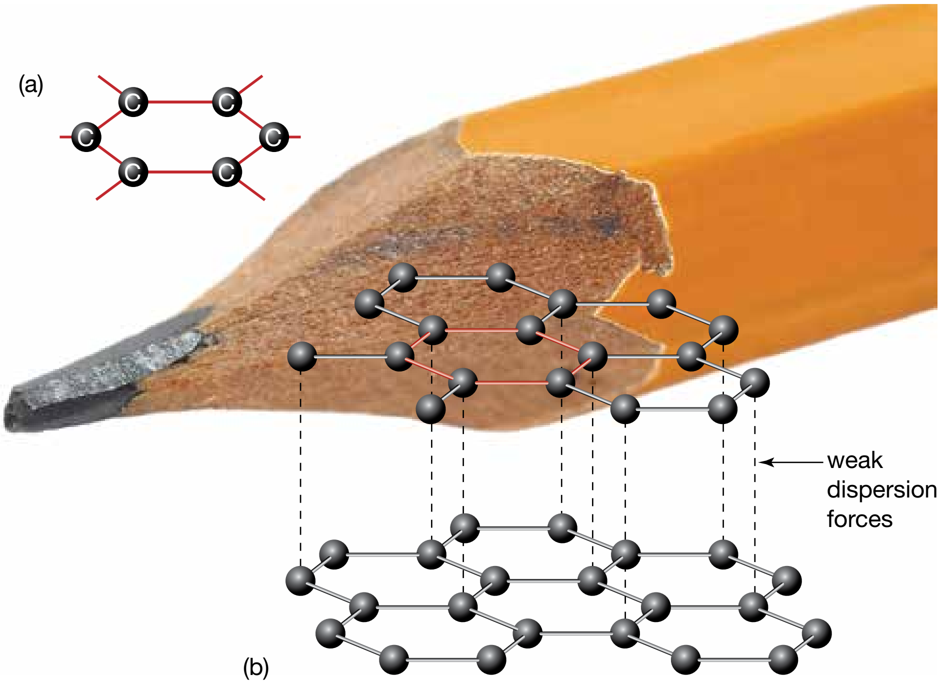
study on

Unit 1	**Graphite** Summary screen and practice questions
AOS 2	
Topic 2	
Concept 2	

Graphite

Graphite is an oily, black, opaque solid with a metallic sheen. Like diamond, graphite is made solely of carbon atoms.

Graphite, however, looks and behaves quite differently from diamond, and this can be explained by its structure. Each carbon atom in graphite forms three covalent bonds with three other carbon atoms within the same plane. This forms layers of hexagonal rings with strong covalent bonds in two dimensions. Since only three of the four outershell electrons are used in the covalent bonds, one outershell electron remains. This electron from each atom becomes delocalised and can move across the graphite layers. Layers are stacked on top of each other in the crystal lattice and are held together by weak dispersion forces.



Graphite is made of flat sheets of carbon atoms. Each carbon atom covalently bonds to three other carbon atoms. This produces rings of six atoms that join to form flat sheets, as shown in (a). These sheets of carbon atoms lie on top of each other, held together by dispersion forces, as shown in (b). The substance produced by these covalent bonds has many uses, among them the common pencil. Industrial-grade 'synthetic' diamonds are made by subjecting graphite to high temperatures and pressures.

Revision questions

3. **The density of diamond is 3.5 g cm^{-3} and that of graphite is 1.9–2.3 g cm^{-3}. Explain why they are different.**
4. **Explain why the following substances are used in the way described:**
   - (a) **graphite in pencils**
   - (b) **diamond in oil drilling.**
5. **Graphite is used to make electrodes in batteries. Explain why graphite is a good conductor of electricity and diamond is not.**
6. **Using a molecular model kit (or plasticine and toothpicks), build models of:**
   - (a) **graphite**
   - (b) **diamond.**

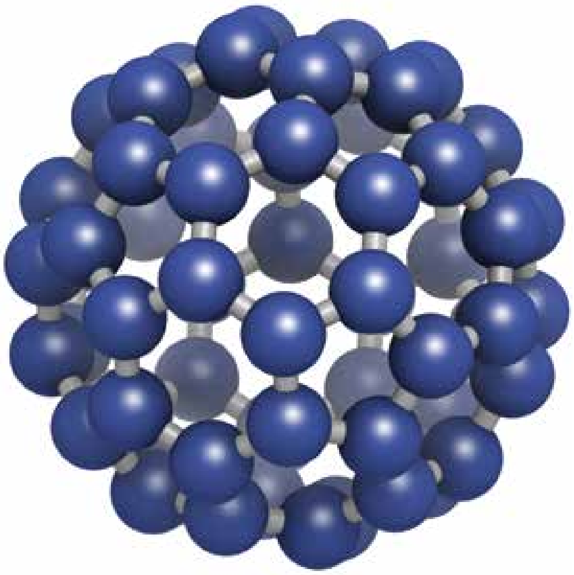
Fullerenes

Buckyballs

Buckyballs are molecules of carbon arranged in a hollow sphere.

Buckminsterfullerenes (or 'buckyballs' for short) are a recently discovered form of carbon. Small quantities of these unusual molecules exist in nature; however, in 1985 they were synthetically produced by vaporising graphite with a laser. They are large cage-like molecules of carbon, but they are not considered giant lattice structures. The most common is C_{60}, which consists of a curved surface in the shape of a soccer ball made up of alternating hexagonal and pentagonal rings of covalently bonded carbon atoms. Other forms include C_{70}, C_{76} and C_{84} molecules.

Buckyballs are stable solids at room temperature, soft and resistant to collision, and they have low melting points. They are insoluble in water but soluble in methyl benzene. In different forms, C_{60} can act as an insulator, a conductor, a semiconductor or, when doped with another element at low temperatures, a superconductor.



Buckyballs are cage-like molecules of carbon.

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Weblinks
Buckyballs
Allotropes of carbon

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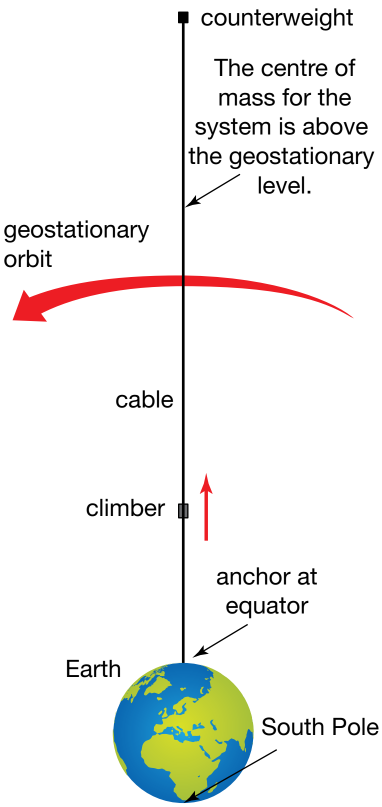
Unit 1	**Carbon nanomaterials**
AOS 2	Summary screen and practice questions
Topic 2	
Concept 3	

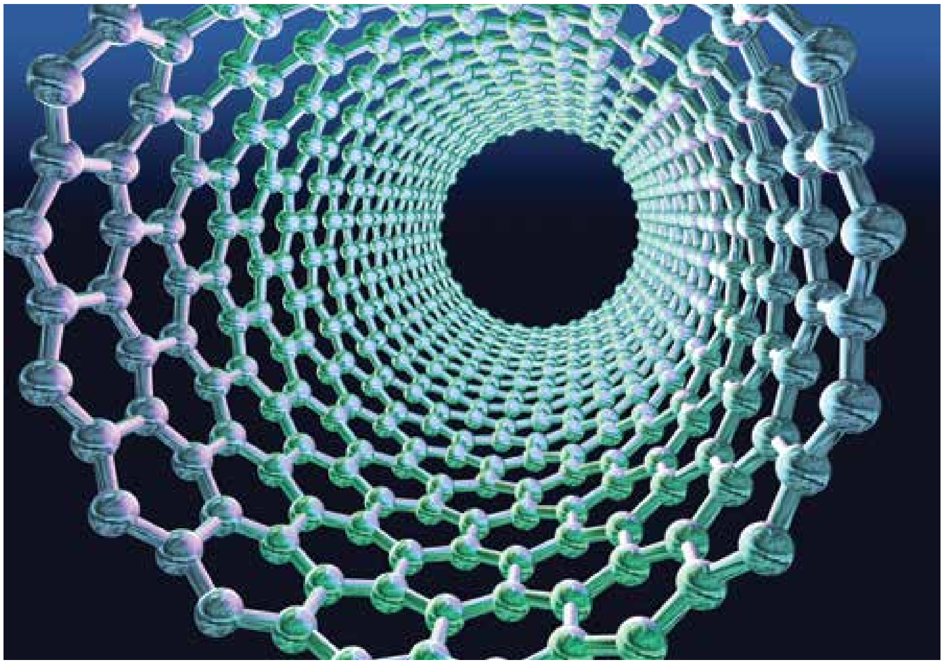
Nanotubes

Carbon nanotubes are tubes made up of sheets of hexagonally arranged carbon atoms formed into cylinders.

An elevator into space is not as unrealistic as it might appear if suitable materials are available. Carbon nanotubes would be ideal for this purpose. They are continuous tubes made up of flat sheets of hexagonal rings of carbon atoms. These sheets are similar to those in graphite but rolled into a cylinder; they may form into a single layer or multiple layers. A human hair is about 10 000 times thicker than one of these very strong fibres. Nanotubes are about 100 times stronger than steel with one-fifth the density. Once established, the elevator could transport materials into space at a fraction of the cost of the massive amount of fuel used in rockets.



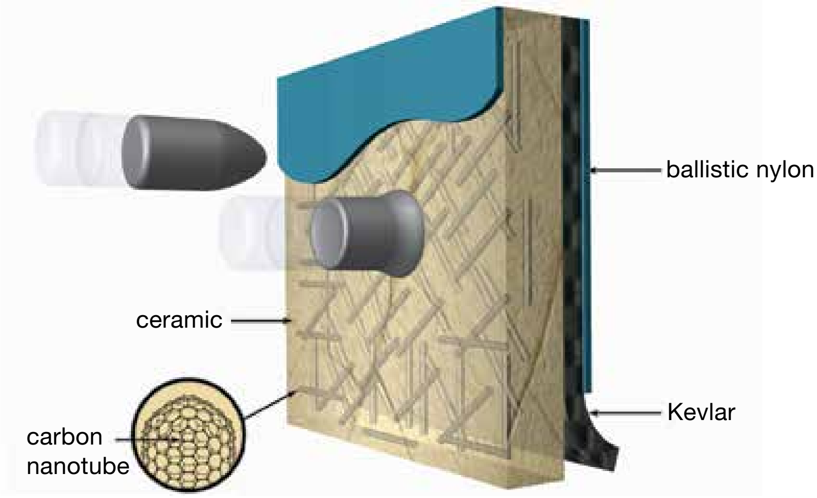
A space elevator made of **carbon nanotubes** anchored to an offshore sea platform would stretch to a small counterweight approximately 100 000 km away in space.



A single-walled nanotube

Nanotubes have extensive applications in many fields. They can be used in batteries to extend their energy storage capacity, and their huge surface area and low electrical resistance increase efficiency of electrodes. Because of their strength they would be beneficial for bulletproof vests and composite materials for transport; their small size and even walls enable easier water purification by filtration. Health applications include possible conduction of nerve impulses, artificial muscle tissue, sensors and targeted drug delivery.

Carbon nanotubes have a tubular graphite-like structure that is incredibly strong, yet light. Because they are on a nanoscale, many millions can be combined together into a small surface area. When incorporated into nylon fabric, they are able to resist the impact of a bullet. The fabric will not tear but will pass the force on to the Kevlar^{TM} layer where it will be dissipated.



Graphene

Imagine a mobile phone that is flexible enough to be worn around your wrist. This could indeed be possible with a material that was discovered in 2004. Graphene is similar to nanotubes but is a flat layer of carbon that is only one atom thick. This wonder material has some remarkable properties that make it superior to nanotubes for use in composite materials. Graphene's large surface area allows it to have closer contact with other materials and its rippled surface enables interlocking with surrounding polymer atoms to readily form very useful composites.

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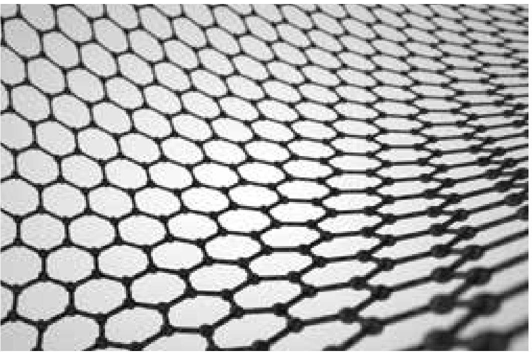
Weblink
Nanotechnology



Graphene is a flat layer of carbon that is only one atom thick.

Graphene is obtained from readily available graphite; it is much stronger than steel but light, transparent and flexible. It conducts electricity and heat more effectively than any metal. Potential applications based on these properties include lower cost mobile phone screens. Phones could be charged in minutes because of its ability to store and release energy quickly. Although impermeable, tiny pores could be introduced in graphene sheets and water could be economically desalinated or purified. Composite materials can use the lightness and strength of graphene for protective clothing or to make transport lighter and more fuel efficient.

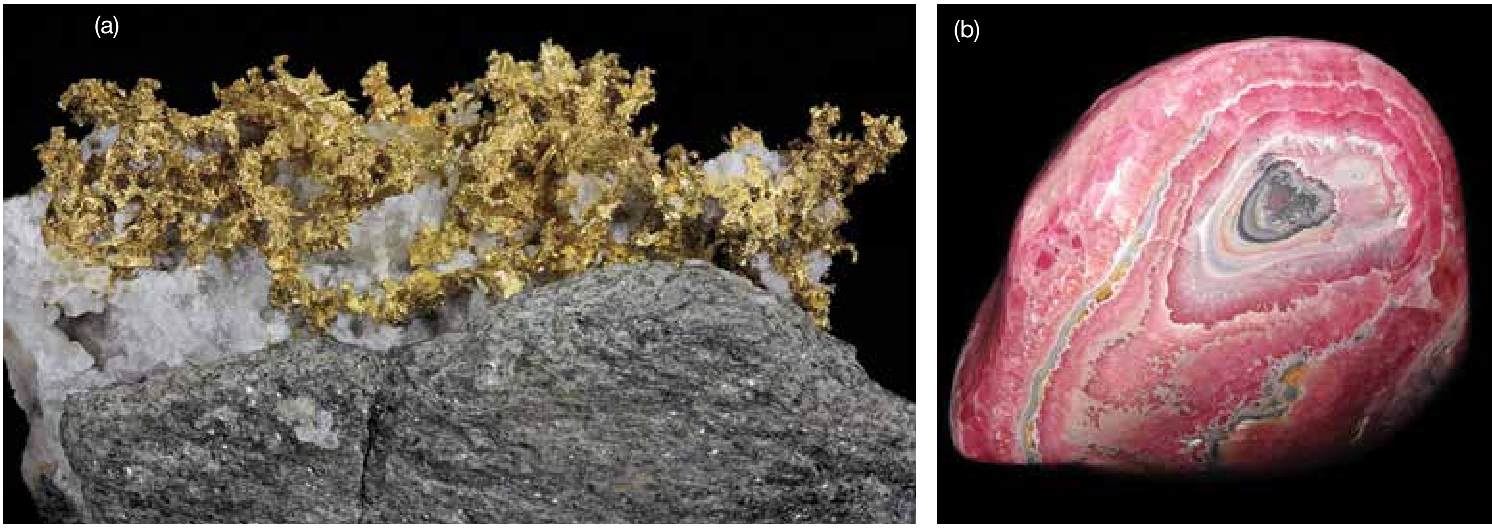
Revision questions

7. **Both graphite and carbon nanotubes are made up of layers containing hexagonal rings of carbon atoms. Explain why some of their properties are so different.**
8. **Explain the difference in structure between buckyballs, nanotubes and graphene.**
9. **Use the internet to research the latest developments in the use of graphene.**

Comparing structures

Bonding and modelling

This section reviews the different types of bonding between atoms, the different models for the structure of materials, and methods of classification using properties and composition.



(a) Rare gold crystals and (b) red crystals of rhodochrosite (manganese carbonate)

Bonding

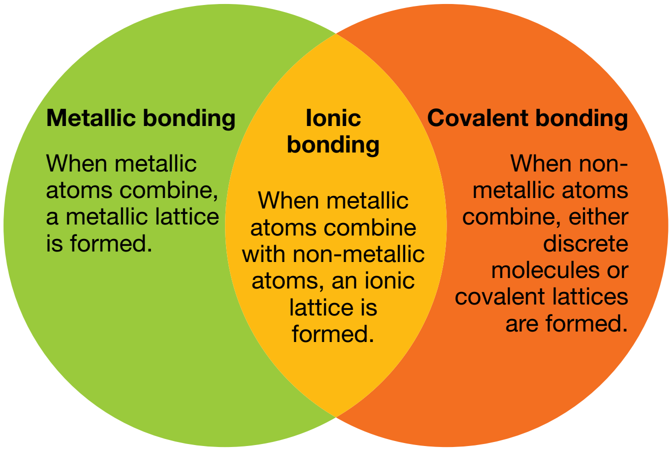
We have seen from the previous chapters that only the noble gases exist as separate atoms since they have complete outer shells of electrons. The atoms of all other elements form chemical bonds with each other in order to attain a more stable outershell configuration.

Atoms can become stable in one of three ways:

- by giving electrons to another atom
- by taking electrons from another atom
- by sharing electrons with another atom.

When atoms combine to achieve more stable structures, three types of bonding are possible. This is illustrated at the top of the next page.

Atoms become stable by giving, taking or sharing electrons.

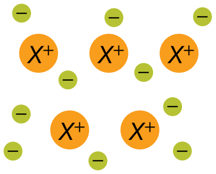
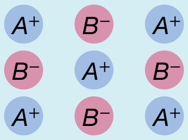
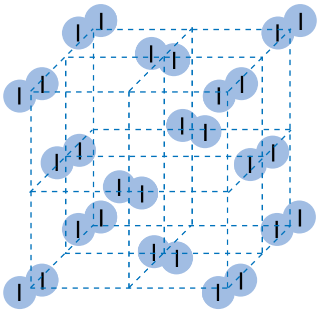
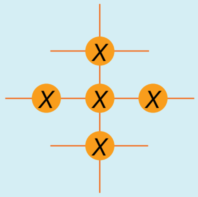
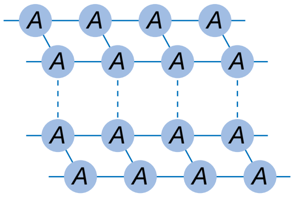


The three main types of bonding

Lattice structures

Atoms may bond together to form crystalline solids. The bonding between atoms gives rise to one of five lattice types as seen in table 7.2.

TABLE 7.2 The structures of different crystalline lattices

Type of lattice structure	Type of substance	Particles in the structure	Pictorial representation of the structure	Examples
metallic	element or alloy that contains metals only	cations and electrons		Fe, Ca
ionic	compound that contains both a non-metal and a metal	cations and anions		NaCl, KF
covalent molecular	elements or compounds that contain non-metals only	molecules		I_2
covalent network	elements or compounds that contain non-metals only	atoms		SiO_2, C (diamond)
covalent layer	elements or compounds that contain non-metals only	atoms		C (graphite)

Identifying the bonding in different substances

A material's properties depend on its structure, and its structure is determined by the nature of the constituent particles and the bonds holding them together. We can therefore identify the type of bonding in different substances using:

1. the properties of a substance
2. the composition of a substance.

Properties

The properties of different crystalline solids are a consequence of the type of bonding present in the lattice (see table 7.3).

TABLE 7.3 The properties of different crystalline lattices

Type of lattice structure	Melting point	Hardness	Electrical conductivity		Solubility in	
			Solid	Liquid (molten or aqueous)	Water (polar)	Petrol (non-polar)
metallic	high	varies	good	good	insoluble	insoluble
ionic	high	hard	poor	good	most are soluble	most are insoluble
covalent molecular	low	soft	poor	poor	soluble if polar	soluble if non-polar
covalent network	very high	hard	poor	N/A	insoluble	insoluble
covalent layer	very high	soft	good	N/A	insoluble	insoluble

Sample problem 7.1

Identify the bonding in substance *A* and substance *B*, given the following information: Substance *A* is hard, melts at 890 °C and conducts electricity in the molten form but not in the solid form. Substance *B* is soft, melts at –183 °C, and is a poor conductor of electricity in solid or liquid form.

Solution: Using the information in table 7.3, substance *A* has an ionic lattice structure with ionic bonding, and substance *B* has a molecular lattice structure with covalent bonding within the molecules.

Revision questions

10. **Classify each of the following as covalent molecular, ionic or metallic.**

- **(a)** **A yellow solid that melts at 105 °C to form a clear yellow liquid. Both solid and liquid are poor conductors of electricity.**
- **(b)** **A solid that melts at 99 °C to form a silvery liquid. Both solid and liquid are good conductors of electricity.**
- **(c)** **A dark shiny solid that sublimes to form a vapour. It is a poor conductor of electricity and heat.**
- **(d)** **A white solid that melts at 872 °C to form a colourless liquid. The solid does not conduct electricity, whereas the liquid does.**

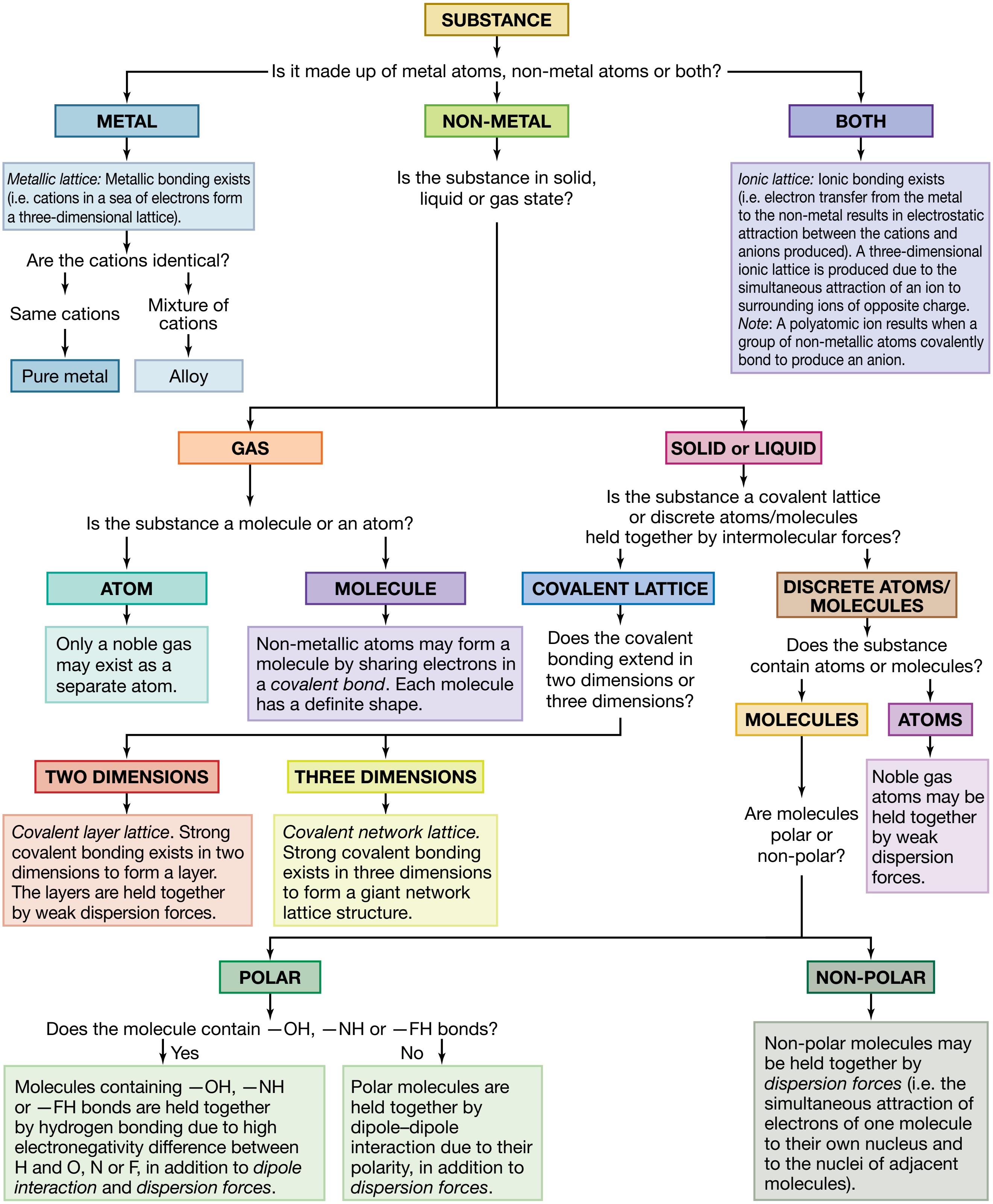
11. **Consider the substances in the table on the left.**

- **(a)** **Which substance is metallic?**
- **(b)** **Which substance is ionic?**
- **(c)** **Which substance has a covalent molecular structure?**
- **(d)** **Which substance has a covalent network structure?**

Substance	Melting point (°C)	Electrical conductivity		
		Solid	Molten	Dissolved in water
L	–115	—	—	—
M	1260	—	conducts	conducts
N	3600	—	—	insoluble
O	2468	conducts	conducts	insoluble

Composition

The flow chart shown below is designed to help you classify a substance into its bond type, using composition.



Metallic lattice: Metallic bonding exists (i.e. cations in a sea of electrons form a three-dimensional lattice).

Ionic lattice: Ionic bonding exists (i.e. electron transfer from the metal to the non-metal results in electrostatic attraction between the cations and anions produced). A three-dimensional ionic lattice is produced due to the simultaneous attraction of an ion to surrounding ions of opposite charge. *Note*: A polyatomic ion results when a group of non-metallic atoms covalently bond to produce an anion.

Only a noble gas may exist as a separate atom.

Non-metallic atoms may form a molecule by sharing electrons in a *covalent bond*. Each molecule has a definite shape.

Noble gas atoms may be held together by weak dispersion forces.

Covalent layer lattice. Strong covalent bonding exists in two dimensions to form a layer. The layers are held together by weak dispersion forces.

Covalent network lattice. Strong covalent bonding exists in three dimensions to form a giant network lattice structure.

Molecules containing —OH, —NH or —FH bonds are held together by hydrogen bonding due to high electronegativity difference between H and O, N or F, in addition to *dipole interaction* and *dispersion forces*.

Polar molecules are held together by dipole–dipole interaction due to their polarity, in addition to *dispersion forces*.

Non-polar molecules may be held together by *dispersion forces* (i.e. the simultaneous attraction of electrons of one molecule to their own nucleus and to the nuclei of adjacent molecules).

Classification of substances according to composition

Sample problem 7.2

Identify the bonding in (a) magnesium (a solid metal), (b) magnesium chloride (a solid salt) and (c) chlorine gas.

Solution:

(a) Using the classification of substances flow chart, magnesium is a pure metal, so the bonding is metallic.

(b) Magnesium chloride is made up of both metal and non-metal elements; according to the flow chart, the bonding is ionic.

(c) Chlorine gas is composed of diatomic molecules of non-metallic atoms. According to the flow chart, the bonding is therefore covalent within the chlorine molecule. As it is non-polar, there are also dispersion forces between the chlorine molecules.

Revision questions

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Interactive molecules

12. **Use the flow chart to identify the bonding in the following substances.**
    - (a) copper
    - (b) copper(II) chloride
    - (c) argon
    - (d) carbon tetrachloride liquid
    - (e) sulfur dioxide
    - (f) steel
    - (g) diamond
    - (h) graphite
    - (i) ammonia gas
    - (j) iron(II) sulfate
13. **Classify the type of lattice structure found in the following solids.**
    - (a) sulfur
    - (b) silicon dioxide
    - (c) aluminium fluoride
    - (d) bronze
    - (e) tungsten
    - (f) dry ice
14. **Use the flow chart to classify the bonding within and between the particles that make up the following substances.**
    - (a) CO_2
    - (b) NH_3
    - (c) CH_4
    - (d) CH_3OH
15. **Draw the molecular structures of the compounds given in question 14.**
16. **The bond with the strongest dipole is:**
    - (a) C—H
    - (b) O—O
    - (c) N—Cl
    - (d) H—F.
17. (a) **Discuss the bonding in sodium chloride (table salt).**
    - (b) **Why are salt crystals brittle?**
    - (c) **What property of sodium chloride makes it useful as a flavouring? Relate this property to its structure and bonding.**
18. **Explain why, in terms of structure and bonding and their relationship to properties, the following materials have the uses given.**
    - (a) **Gold is used to make jewellery.**
    - (b) **Graphite is used as a lubricant in locks.**

Chapter review

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Unit 1	**Networks and nanomaterials**
AOS 2	Summary screens and practice questions
Topic 2	

Summary

- Carbon is a unique and ubiquitous element. It is essential to life and found in rocks, organisms, the atmosphere and fuels. It forms millions of compounds because it bonds readily with itself and other elements. It can be found in small and large molecules as well as lattices.
- A covalent network lattice is a three-dimensional network lattice where many atoms are covalently bonded to each other. Examples are diamond and silicon dioxide. Such substances have similar properties.
- Covalent network lattices:
  - exist as solids and have very high melting points due to the strong covalent bonds between atoms
  - do not conduct electricity as they have no free electrons
  - are hard and brittle
  - are chemically inert
  - are insoluble in water.
- Covalent layer lattices, such as graphite, are made up of many atoms held strongly in two-dimensional layers by covalent bonds. These layers are held together by weak dispersion forces.
- Graphite:
  - is solid with a high melting point due to strong covalent bonds
  - conducts electricity due to the presence of delocalised electrons
  - is slippery to the touch as weak dispersion forces between layers allow the layers to slide over each other
  - has a metallic sheen due to the interaction between light and the delocalised electrons.
- Diamond, graphite and fullerenes are allotropes of carbon and have very different properties and uses.
- Fullerenes, such as buckyballs, are spherical molecules of carbon that are shaped like a soccer ball. Each carbon is covalently bonded to three other carbon atoms. The most common type of buckyball is C_{60}.
- Carbon nanotubes consist of one or more layers of carbon atom rings rolled into a sheet. They are good conductors and strong.
- Graphene consists of a single layer of carbon atoms arranged similarly to graphite but it is only one atom thick. It is two dimensional and almost transparent and it is also an excellent conductor and stronger than steel.
- Comparison of networks
  - Metallic: cations and electrons, high melting point, hard, malleable, ductile, good conductor in solid or liquid state
  - Ionic: cations and anions, high melting point, brittle, non-conductor as solid, conductor in liquid state
  - Covalent molecular: molecules, low melting point, soft, non-conductor in any state
  - Covalent network: atoms, very high melting point, hard, non-conductor
  - Covalent layer: atoms, very high melting point, soft, conductor

Multiple choice questions

1. Diamond is an electrical insulator because:

A it is a crystalline substance
B all the outershell electrons are delocalised
C it consists of non-conducting carbon atoms only
D all the outershell electrons are involved in single covalent bonds.

2. Crystalline diamond, C, and crystalline quartz, SiO_2, are alike in some respects. Which one of the following statements is *incorrect*?

A Both substances are above average in hardness.
B At room temperature, neither substance dissolves in water.
C In pure form, both substances may be colourless and transparent.
D Diamond has covalent bonds between carbon atoms, and quartz has covalent bonds between silicon atoms only.

3. The addition of powdered graphite to a lock that has been sticking often allows the key to turn more easily. Graphite has this effect because:

A each carbon atom has already formed four bonds, and hence can form an unreactive film over which metal can slide
B its infinite lattice structure is very hard and strong, enabling moving parts to slide over a thin layer of graphite
C the tetrahedral arrangement of bonds around each carbon atom causes neighbouring carbon atoms to slide past one another
D the carbon atoms are bonded into two-dimensional sheets that can slide freely over one another.

4. What makes carbon unique?

A Carbon has two allotropes.
B Carbon forms four bonds.
C Carbon forms covalent compounds.
D Carbon can bond to itself to form straight chains, branched chains and rings more than any other element.

5. Diamond, graphite and graphene are all:

A isotopes of carbon
B isomers of carbon
C allotropes of carbon
D compounds of carbon.

6. How does the structure of nanotubes make them suitable as catalysts?
A They have a large surface area.
B They are made from reactive carbon atoms.
C They have strong covalent bonds.
D They are tubular in shape.

7. Graphite can be described as:
A made of carbon and hydrogen atoms only
B a metal
C hard and crystalline
D soft and slippery.

8. Why might fullerenes be used in new drug delivery systems?
A They are made from carbon atoms.
B They are hollow.
C They are very strong.
D They are allotropes of carbon.

Bonding — an overview

9. Which of the following statements would not be true of the compound magnesium fluoride?
A It is brittle and crystalline in the solid state.
B It is soluble in non-polar solvents.
C It conducts electricity in the molten state.
D It is a poor conductor of electricity in the solid state.

10. Which of the following sets of substances have only covalent bonds?
A copper(II) oxide, ammonia, methane, silicon dioxide
B copper(II) oxide, sodium chloride, calcium fluoride, hydrogen chloride
C ammonia, methane, silicon dioxide, hydrogen chloride
D ammonia, calcium fluoride, methane, silicon dioxide

11. Metals conduct electricity in the solid state because they contain:
A many mobile electrons
B equal numbers of cations and electrons
C many mobile anions
D equal numbers of cations and anions.

12. Graphite is:
A a molecular compound
B an ionic lattice compound
C a covalent layer solid
D an electrical conductor in the molten state only.

13. If a material is described as 'malleable', it means that the substance:
A is very flexible
B has a high density
C can be hammered into thin sheets
D can be drawn out into wires.

14. Which of the following substances is the best electrical conductor when molten?
A diamond
B carbon tetrachloride
C potassium chloride
D silicon dioxide

15. The electron configurations of elements *P*, *Q*, *R* and *S* are given below.
P: 2, 8, 2
Q: 2, 8, 6
R: 2, 7
S: 2, 8, 8
Which one of the following pairs of elements is most likely to react to form:
(i) an ionic compound
(ii) a covalent compound?
A *P* and *S*
B *P* and *Q*
C *Q* and *R*
D *Q* and *S*

16. Chemical bonds form for all of the following reasons *except*:
A a tendency to achieve a more stable electron configuration
B attractions between nuclei and electrons
C a tendency for atoms to fill their valence shell with electrons
D a tendency of bonded atoms to react more readily than unbonded atoms.

17. The situation that does not involve intermolecular bonding is:
A methane molecules attracting methane molecules
B the sulfur atom in an H_2S molecule attracting the hydrogen atoms
C the hydrogen atom of an HCl molecule attracting the chlorine atom of another HCl molecule
D forces of attraction holding molecules of H_2O together in ice.

18. Sodium chloride melts at 801 °C and cyclohexane melts at 6.6 °C. Such evidence supports the hypothesis that:
A the melting of sodium chloride involves the conversion of ionic bonds to covalent bonds
B sodium chloride is decomposed by melting
C the melting of cyclohexane involves the conversion of covalent bonds to hydrogen bonds
D ionic bonds are much stronger than dispersion forces.

19. A difference between ionic and molecular compounds is that ionic compounds:
A dissolve in water, but molecular compounds are insoluble
B usually form crystals, whereas molecular compounds usually do not
C conduct electricity in the solid state, whereas molecular compounds usually do not
D usually melt at higher temperatures than molecular compounds.

20. In which of the following substances would you expect the bonding to be most ionic?

A Ice

B Solid ammonia

C Solid lead bromide

D Solid silicon dioxide

Review questions

Covalent network and layer lattices

1. (a) Discuss the structure of diamond and graphite.
  (b) What are the similarities and differences in these two structures?
  (c) Why are they called ‘giant’ molecules?
  (d) Give three uses of graphite and three uses of diamond.

2. A zipper is rubbed with a pencil to make it move freely.
  (a) What property of graphite is this illustrating?
  (b) Why does graphite have this property?

3. Graphite powder is mixed with clay to make ‘lead’ pencils. Pencils vary in colour and hardness. What is the purpose of the clay and why is the ratio of clay to graphite important?

4. What property of diamond enables it to be used to cut glass?

5. Diamonds are brittle. Discuss this.

6. When wax from a candle vaporises, some molecules burn in the tip of the flame, producing soot that contains buckyballs. Describe the structure of the C_{60} buckyball.

Bonding — an overview

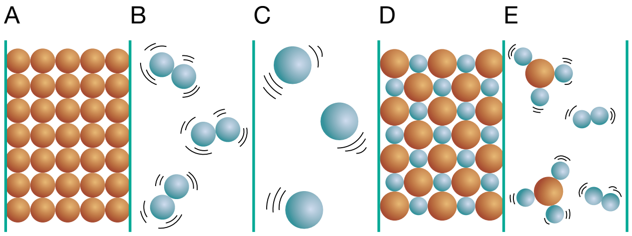
7. The atoms of a noble gas are stable. Explain.

8. Explain in terms of electron configuration why the metals Na ($Z = 11$) and K ($Z = 19$) have similar chemical properties.

9. Name and give the symbols of three ions that have the same stable configuration as that of neon.

10. An element *X* whose atoms contain 12 protons is chemically bonded to another element, *Y*, whose atoms contain 9 protons.
  (a) What type of bonding would be present between *X* and *Y*?
  (b) What is the formula of the compound formed from *X* and *Y*?

11. The figure below shows various arrangements of atoms, molecules or ions. Which of these diagrams represents (a) a noble gas, (b) a mixture of gases, (c) a metallic solid, (d) an ionic solid and (e) hydrogen gas?



12. In which of the following substances do molecules exist?
  (a) calcium oxide, CaO
  (b) nitric oxide, NO
  (c) calcium, Ca
  (d) ammonia, NH_3
  (e) ammonium nitrate, NH_4NO_3

13. Briefly discuss the bonding in the following substances. Draw structures where appropriate (s = solid, l = liquid, g = gas).
  (a) $CaCl_2(s)$
  (b) $CCl_4(g)$
  (c) $MgO(s)$
  (d) $H_2O(s)$
  (e) $CO_2(g)$
  (f) $CO_2(s)$
  (g) C (diamond)
  (h) Cu(s)
  (i) He(s)
  (j) $C_2H_2(g)$
  (k) $HNO_3(l)$
  (l) KCl(s)
  (m) $H_2S(g)$
  (n) $CaCO_3(s)$
  (o) KOH(s)
  (p) Sterling silver
  (q) 18 carat gold

14. Classify the following as metallic, ionic or covalent molecular lattice structures.
  (a) iodine, I_2
  (b) silver, Ag
  (c) silver nitrate, $AgNO_3$
  (d) sulfur, S
  (e) sulfur dioxide, SO_2
  (f) copper(II) oxide, CuO
  (g) ammonia, NH_3
  (h) nitrogen trichloride, NCl_3
  (i) aluminium chloride, $AlCl_3$
  (j) oxygen, O_2
  (k) titanium, Ti
  (l) hydrogen iodide, HI
  (m) phosphine, PH_3

15. Discrete molecules (simple molecular substances) often have a smell. Metallic, ionic and covalent network molecules do not have a smell. Discuss.

16. An experiment is conducted in order to classify some compounds. Each compound is melted and then its electrical conductivity is tested.

Compound	Electrical conductivity
benzoic acid	poor
magnesium sulfate	good
sodium fluoride	good
paraffin wax	poor
aluminium	good
sugar	poor

(a) (i) Use the electrical conductivity data to classify each substance as either ionic, metallic or covalent molecular.
   (ii) Would any of the substances conduct electricity in the solid state? Explain.
   (iii) What types of bond or force are breaking when these substances melt?
   (iv) Suggest whether you expect the melting point of each substance to be high or low. Give reasons for your answers.
(b) Explain why some substances tested did not conduct electricity in the molten state.
(c) Testing the electrical conductivity of a substance when it is added to water is *not* a reliable method of classification. Give reasons to explain why.

17. Account for the fact that at room temperature metals are good conductors of electricity, but ionic network solids are not.

18. (a) Name the type of lattice structure formed for each of the following substances.
   (i) NaF
   (ii) CH_3OH
   (iii) Fe
   (iv) SiO_2
(b) Arrange the substances in order of increasing melting point.

19. Which of the two substances, ammonia, NH_3, and potassium iodide, KI, has a higher melting point? Explain your reasoning.

20. Select from the following list the substances that would be (a) good electrical conductors when molten and (b) poor electrical conductors when molten.
(a) P_4O_{10}
(b) CCl_4
(c) KF
(d) SiO_2
(e) $MgBr_2$
(f) CaO
(g) C_2H_4

21. Explain why copper is malleable but calcium chloride is brittle. Make reference in your answer to the bonding between the particles.

22. Diamond and tungsten both have extremely high melting points. Contrast the bonding and structure of each and explain why they both have a high melting point.

23. A compound *AB* has a melting point of −114 °C and does not dissolve in water. It does not conduct electricity in solid or liquid form. Another compound *CD* has a melting point of 801 °C and dissolves in water. It conducts electricity in the liquid form but not the solid form.
(a) What type of bonding and structure would be expected in *AB*?
(b) What type of bonding and structure would be expected in *CD*?

24. Some of the properties of the pure substances *E*, *F*, *G* and H are given in the table below.

Substance	Melting point (°C)	Electrical conductivity of solid	Electrical conductivity of liquid	Electrical conductivity of solution
E	−110	nil	nil	nil
F	21	nil	nil	(insoluble)
G	810	nil	high	high
H	1640	high	high	(insoluble)

(a) Which substance(s) could have ionic bonding?
(b) Which substance(s) could have metallic bonding?
(c) Which substance(s) could have covalent bonding?

25. Classify the following substances as metals, ionic substances, covalent network substances or covalent molecular substances.
- Substance *A* has a low melting point and a negligible electrical conductivity in both the solid and molten states.
- Substance *B* has a high melting point and a high electrical conductivity in both the solid and molten states.
- Substance *C* has a high melting point. It does not conduct electricity in the solid state, but the molten state has high electrical conductivity.
- Substance *D* has a very high melting point and a negligible electrical conductivity in both the solid and molten states.

26. With reference to structure and properties, discuss why the following materials are used for the applications given.
(a) Aluminium in aircraft parts
(b) Diamond as an industrial cutting tool
(c) A solution of copper(II) sulfate as an electrolyte (a solution that conducts electricity)
(d) Tungsten (melting point 3400 °C) in filaments inside light globes

27. Produce concept maps using the following clusters of words.
(a) anion; cation; atom; electron; covalent; ionic; metallic; molecule; non-metallic
(b) magnesium chloride; ionic lattice; ions; brittle; molten; high melting points; conductivity
(c) chlorine; ammonia; covalent bonding; electrons; low melting point; polar bond; non-polar bond
(d) graphite; diamond; lattice structure; melting point; brittle; covalent
(e) copper; lattice structure; conductor; shiny; malleable

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Unit 1
AOS 2
Topic 2

Networks and nanomaterials

Sit topic test

Exam practice questions

In a chemistry examination you will be required to answer a number of short and extended response questions.

Multiple choice questions

1. Which of the following statements best describes the structure of solid carbon dioxide?

A It is an array of positive carbon ions and negative oxide ions.

B Each carbon atom is joined by strong covalent bonds to four oxygen atoms to give an array of many thousands of atoms.

C Molecules, each containing one carbon atom covalently bonded to two oxygen atoms, are held together by weak intermolecular forces.

D Carbon atoms and covalently bonded oxygen molecules are held together by weak intermolecular forces. **1 mark**

2. The melting points of ionic compounds are usually:

A higher than those of molecular compounds

B lower than those of molecular compounds

C equal to those of molecular compounds

D higher or lower than molecular compounds **1 mark**

3. Covalent bonds are *least* likely to be formed between:

A elements in period 1 and period 3

B an element in group I and an element in group VII

C atoms of the same element

D atoms of different elements on the right of the periodic table. **1 mark**

4. Which one of the following statements is *incorrect* about the bonding within metals?

A The delocalised electrons can move freely through the lattice of cations.

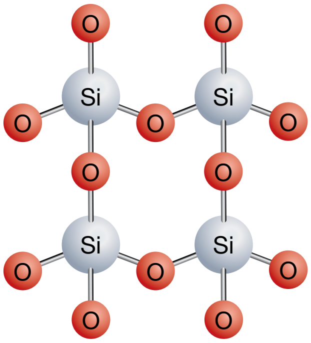
B The forces between the cations and delocalised electrons are weak.

C Strong electrostatic forces exist between the cations and delocalised electrons

D When a force is applied, the metal ions slide over each other without altering the basic structure.

Extended response questions

5. Silica, SiO_2, is a solid with a high melting point, while CO_2 at room temperature is a gas. Both silicon and carbon are found in group 14 of the periodic table, and hence you would expect similar properties in their compounds. Account for the differences in physical states of SiO_2 and CO_2 in terms of the structures of the two substances. **3 marks**



Exam practice questions *(continued)*

6. Consider the following processes:

(i) sublimation (solid ⟶ gas) of iodine when heated
(ii) $I_2(g) \longrightarrow 2I(g)$
(iii) shattering a crystal of salt
(iv) shattering a diamond crystal
(v) bending a piece of copper wire until it breaks
(vi) melting pure gold
(vii) melting ice.

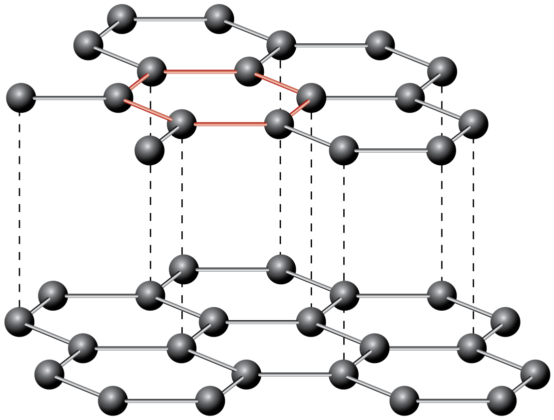
For each of the processes above, state:

(a) the types of bond or force being broken
(b) the types of particle (molecules, atoms or ions) being separated. **7 marks**

7. Consider each of the following solids: Na, Si, H_2S, He, HF, KF. Which would be an example of:

(a) a solid in which hydrogen bonding exists between molecules
(b) a solid which is a good electrical conductor
(c) a solid which is a poor electrical conductor, but conducts on melting
(d) a solid consisting of atoms held together by weak dispersion forces
(e) a solid in which the atoms are covalently bonded together into a network lattice
(f) a solid containing a V-shaped, polar molecule? **6 marks**

8. Graphite (shown below) and graphene consist of continuous rings of carbon atoms. Why are their properties so different? **4 marks**



CHAPTER

8 Organic chemistry

YOU WILL EXAMINE:

- the structures, names, uses and formulas of the homologous series of alkanes, alkenes, alkynes, alcohols, carboxylic acids and esters up to C_{10}
- physical and chemical properties of hydrocarbons
- properties of alcohols, carboxylic acids and esters
- saturated and unsaturated hydrocarbons
- isomers and their naming
- alcohol, carboxylic acid and ester functional groups
- formation of esters
- the origin of hydrocarbons.

> Where the senses fail us, reason must step in.
>
> *Galileo Galilei*

Firefighters battle to put out an oilwell fire. The organic chemicals, including natural gas, oil and petrochemicals, are highly flammable. Oil products are vital to our chemical, transport and fuel industries.

Organic chemistry

Organic chemistry is the study of the compounds of carbon.

Organic chemistry is the systematic study of the compounds of carbon, excluding oxides of carbon, carbonates, carbides and cyanides. These exceptions plus the compounds of the remaining elements are considered to be inorganic.

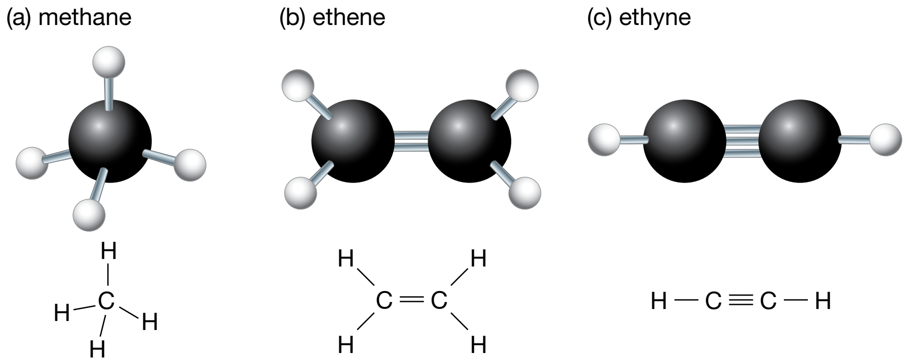
Buckminsterfullerene, a soccerball-shaped organic molecule used for research in the pharmaceutical industry. 'Buckyballs' such as these are produced at the tip of a candle flame, where C—C bonds form in ring structures.

What are organic compounds?

The large number of compounds that carbon forms can be explained by its ability to make four bonds and its ability to bond to itself using single, double and triple bonds. Carbon also has the ability to form cyclic or ring structures.

Organic compounds are a class of chemical compounds that occur naturally (although they are also synthesised in laboratories). They comprise over 90% of all known chemicals. They include not only those compounds that were part of or were derived from plants and animals, but also all carbon compounds except for those mentioned above. Of all the elements in the periodic table, carbon is the only one that has properties that make it possible for living systems to develop. The main reason for carbon's unique ability to form a wide range of chemicals is that carbon–carbon bonds are strong. Carbon normally forms four covalent bonds. These strong bonds form bonds with other non-metals. Carbon can chemically bond with itself using single, double or triple bonds to form long chain-like structures and even rings.

Hydrocarbons are compounds that contain only carbon and hydrogen atoms. The carbon–carbon bonds can be single, double or triple. If these bonds are all single, the hydrocarbon is described as saturated. If double or triple bonds are present, it is termed unsaturated.

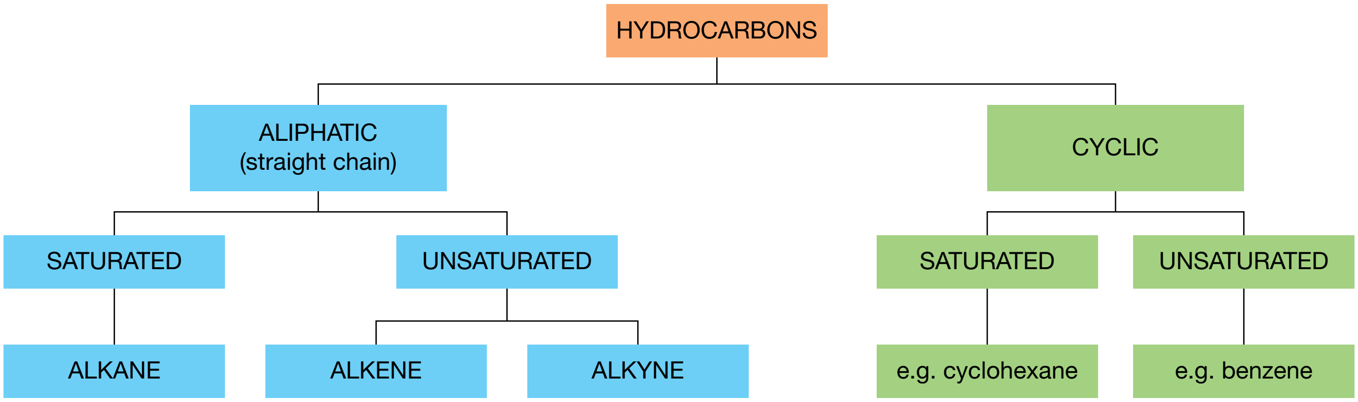

(a) methane
(b) ethene
(c) ethyne

Carbon can form double and triple bonds with itself. The simplest hydrocarbons with carbon–carbon double and triple bonds are ethene and ethyne respectively.

Hydrocarbons

A good place to begin a study of organic chemistry is with **hydrocarbons**. These are the simplest organic compounds and yet are among the most useful. Hydrocarbons are found in crude oil, and processing the oil yields many useful products. We will look at this process in more detail and the products that are formed later in this chapter. Hydrocarbons are molecules made up of only hydrogen and carbon.

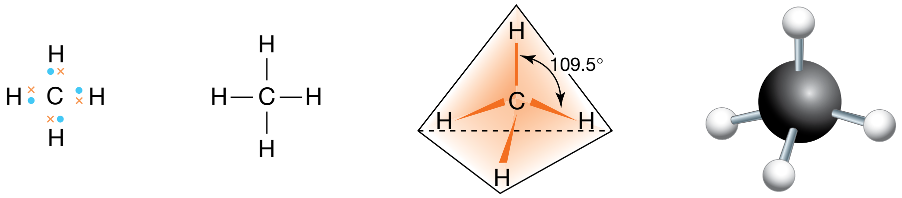
There are different classes of hydrocarbons, based on the carbon bonding involved. These are indicated below.



The classes of hydrocarbons

Saturated hydrocarbons contain only single carbon–carbon bonds. **Unsaturated hydrocarbons** contain double or triple carbon–carbon bonds. In **cyclic compounds**, the carbon atoms are arranged in a ring.

The simplest hydrocarbon is a carbon atom bonded to four hydrogen atoms. Carbon can form four equivalent bonds that are 109.5° apart, forming a tetrahedral shape. This molecule is called methane. The structure of a methane molecule was described in chapter 6.



Lewis structure, structural formula, shape, and ball-and-stick model of methane

Properties of hydrocarbons

Compounds containing C—H bonds are generally insoluble in water and do not react with it. They are more likely to be soluble in non-polar solvents. This is because the electronegativities of carbon and hydrogen atoms are similar. This means that the C—H bond is nearly non-polar, because of the symmetry of the molecule, and hydrocarbon molecules are non-polar.

This non-polar nature of hydrocarbons also explains why the boiling point of hydrocarbons increases with the length of the chain of carbon atoms. Only dispersion forces hold the molecules together, and these forces increase as the size of the molecules increases. The increase in boiling point with size reflects the increased dispersion forces between the molecules. If a molecule is branched instead of a straight chain, this lowers the boiling point. Branches prevent the molecules from coming closer together and so the dispersion forces have less effect. Table 8.1 shows this trend in the alkanes.

Hydrocarbons are insoluble in water.

The boiling point of hydrocarbons increases with the length of the molecule.

The alkanes

The **alkanes** are a family of hydrocarbons containing only **single bonds** between the carbon atoms. Notice how the name of each alkane has the same ending: -*ane*. This is how alkanes are identified. The different prefixes, *meth*-, *eth*-, *prop*- etc., show how many carbon atoms are present in each carbon chain (see table 8.1).

Alkanes are hydrocarbons that contain only single carbon–carbon bonds. All alkanes have the formula C_nH_{2n+2}. As each successive member differs from the one before it by CH_2, they are an example of a homologous series.

TABLE 8.1 Common straight chain alkanes

Formula	Name	Phase	Typical use
CH_4	methane	gas	natural gas
C_2H_6	ethane	gas	refrigerant
C_3H_8	propane	gas	liquid petroleum gas
C_4H_{10}	butane	gas	manufacture of synthetic rubber
C_5H_{12}	pentane	liquid	solvent
C_6H_{14}	hexane	liquid	filling for thermometers
C_7H_{16}	heptane	liquid	testing engine knocking
C_8H_{18}	octane	liquid	automobile petroleum
C_9H_{20}	nonane	liquid	gasoline
$C_{10}H_{22}$	decane	liquid	gasoline
$C_{16}H_{34}$	hexadecane	liquid	lubricating oil
$C_{20}H_{42}$	icosane	solid	wax candles
$C_{28}H_{58}$	octacosane	solid	tar

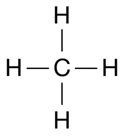
study on

Unit 1
AOS 2
Topic 3
Concept 10

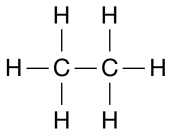
Empirical and molecular formulas for organic molecules
Summary screen and practice questions

Any series of organic compounds in which each successive member differs by CH_2 from the previous one is called a **homologous series**. They can be represented by the general formula C_nH_{2n+2}, where n is the number of carbon atoms in the molecule.

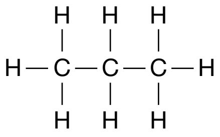
Structural and molecular formulas of the first three alkanes



methane, CH_4



ethane, C_2H_6



propane, C_3H_8

Sample problem 8.1

Find the molecular formula of an alkane with six carbon atoms.

Solution: $n = 6$

Number of hydrogen atoms $= 2n + 2$

$= (2 \times 6) + 2 = 14$

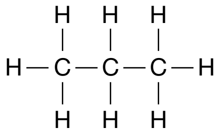
The formula is therefore C_6H_{14}.

Revision questions

1. **Draw the Lewis structures for ethane and propane.**
2. **Give the formula for an alkane with 16 carbon atoms.**
3. **Give the formula for an alkane with 52 hydrogen atoms.**
4. **Draw the structure of pentane.**

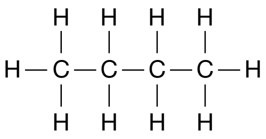
5. **Which of the following compounds would you expect to have the higher boiling point? Give reasons for your answer.**

A



propane, C_3H_8

B



butane, C_4H_{10}

Reactions of alkanes

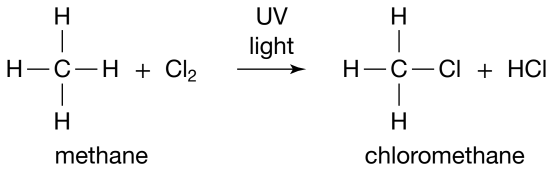
The main reactions of alkanes are combustion reactions and substitution reactions. In complete combustion reactions, the products are always carbon dioxide and water.

Alkanes burn in oxygen to form carbon dioxide and water. This is known as an oxidation or combustion reaction. For example:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

Alkanes react with chlorine or fluorine in a reaction called a **substitution** reaction. In this type of reaction, one or more of the chlorine or fluorine atoms take the place of one or more of the hydrogen atoms. The products of these reactions are known as **haloalkanes**. For example:

This reaction may continue with a further series of hydrogen substitutions if there is enough chlorine present. As each remaining hydrogen atom is removed from CH_3Cl and replaced by a chlorine atom, a mixture of CH_2Cl_2, $CHCl_3$ and CCl_4 is formed.



methane

chloromethane

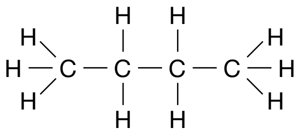
Revision questions

6. **Complete and balance the following equations.**
   (a) $C_2H_6(g) + O_2(g) \rightarrow$ **(b)** $C_5H_{12}(g) + O_2(g) \rightarrow$
7. **Write a structural chemical equation for the substitution reaction between methane and fluorine gas, $F_2(g)$.**
8. **Write a balanced chemical equation for the combustion of an alkane that has six carbon atoms.**

Structural isomers

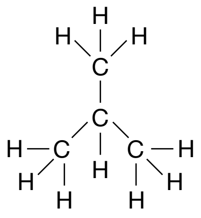
Isomers are compounds that have the same molecular formula but different structural formulas. Butane, C_4H_{10}, is the simplest alkane to show this property.

Isomers are organic molecules with the same molecular formula but different structural formulas. For example, C_4H_{10} can be drawn in two ways as shown below.



butane

The two isomers of C_4H_{10}



methylpropane

Structural isomers have difference names and chemical properties because the molecules have different shapes.

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Unit 1 **AOS 2** **Topic 3** **Concept 9**

Structural isomers
Summary screen and practice questions

Revision questions

9. **Use a molecular modelling kit to form a model of butane. Rearrange your model to form methylpropane.**
10. **Form a model of pentane. How many isomers of C_5H_{12} can you form?**
11. **Draw each of the isomers of C_5H_{12} that you formed in question 10.**
12. **Suggest a reason why the boiling point of methylpropane (–11.7 °C) is lower than that of butane (–0.5 °C).**

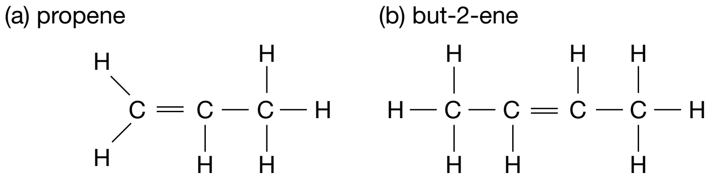
The alkenes

The **alkenes** make up a family of hydrocarbons that each contains one **double bond** between two carbon atoms.

Notice that each of these names ends with *-ene*. This is the characteristic ending used for the alkenes.

The alkenes have a general formula C_nH_{2n} and are another example of a homologous series.

Structural formulas of two alkenes

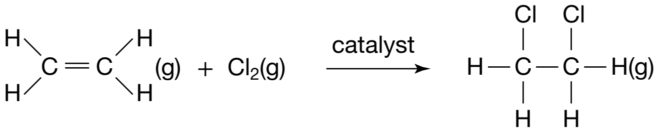


Reactions

Alkenes undergo combustion reactions with oxygen and, like alkanes, form carbon dioxide and water.

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(g)$$

Alkenes react with hydrogen or the halogens in addition reactions. The hydrogen or halogen attacks the double bond and breaks it. When chlorine is added to ethene in the presence of a catalyst the following reaction takes place:



Alkenes are hydrocarbons that contain a carbon–carbon double bond. They have the general formula C_nH_{2n}. Alkenes are another example of a homologous series, as each alkene varies from the one before it by a $—CH_2—$ group. The main reactions of alkenes are combustion reactions and addition reactions.

Alkenes can undergo self-addition in which the alkene molecules join together to form long chains called **polymers**. This will be discussed in chapter 9.

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Homologous series of alkenes
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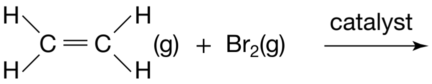
Revision questions

Complete the following chemical reactions:

13. (a)



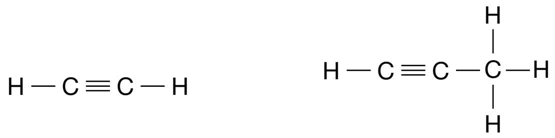
(b) $H_2C = CH_2(g) + HCl(g) \xrightarrow{\text{catalyst}}$

14. **Explain the terms 'saturated' and 'unsaturated' in relation to hydrocarbons. Give two examples of each.**

The alkynes

Alkynes are hydrocarbons containing one **triple bond** between two carbon atoms.

Structural formulas of two alkynes



$H-C\equiv C-H$

$H-C\equiv C-\overset{\overset{\large H}{|}}{\underset{\underset{\large H}{|}}{C}}-H$

(a) ethyne (also known by its old name, acetylene)

(b) propyne

Alkynes contain a carbon–carbon triple bond. They have the general formula C_nH_{2n-2}.

Note that the ending in each case is -*yne*. This ending shows that there is a triple bond in the molecule.

The general formula for the alkynes is C_nH_{2n-2}. The alkynes are also another example of a homologous series.

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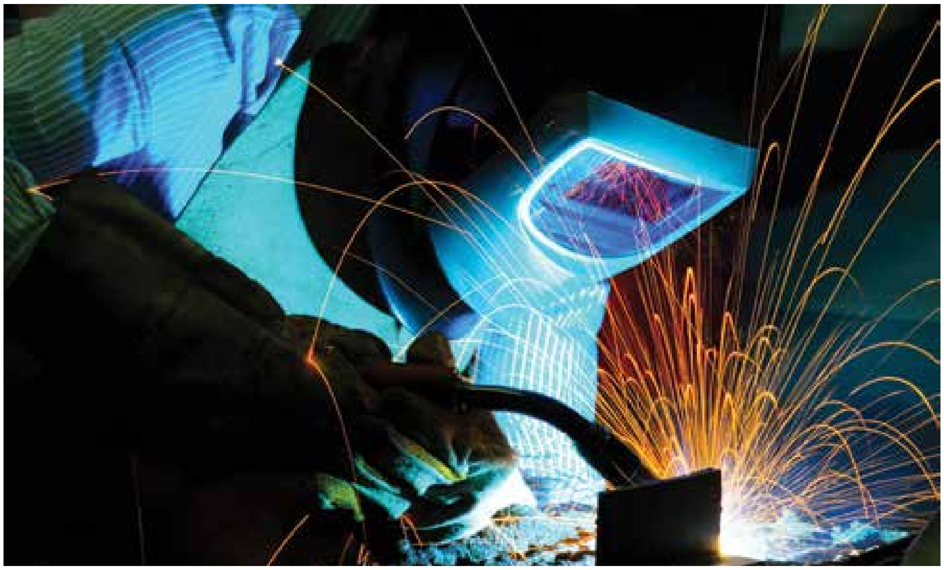
Unit 1	**Hydrocarbons: alkanes, alkenes and alkynes**
AOS 2	Summary screen and practice questions
Topic 3	
Concept 4	

Reactions

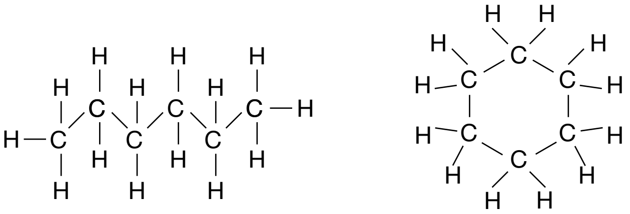
Alkynes can undergo combustion (oxidation) and **addition reactions**.



The combustion of ethyne (acetylene) is used in oxyacetylene welding. The extremely high temperature generated when ethyne is burned with pure oxygen can melt metal and allow surfaces to be welded together. Like all hydrocarbons, the products of this reaction are carbon dioxide and water. The equation for the reaction is: $2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(g)$.

Cyclic hydrocarbons

The carbon backbone of some hydrocarbons can form a ring. Such compounds are known as cyclic hydrocarbons and have the prefix *cyclo*- before their name.



(a) hexane

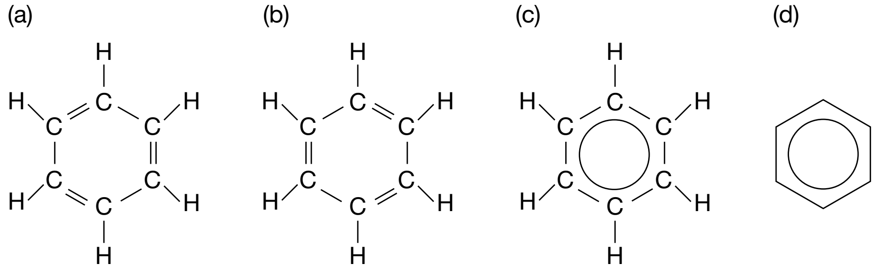
(b) cyclohexane

Structural formulas for straight chain and cyclic hexane

An important category of the hydrocarbons is the unsaturated cyclic compounds making up the aromatic series. The simplest member of this series is benzene, C_6H_6. This molecule has been shown to have two equivalent forms as shown in diagrams (a) and (b) below.

Benzene is found naturally in crude oil but is usually synthesised from other hydrocarbons when used in making pharmaceuticals.

Four structural representations of benzene



Many hydrocarbons are formed by carbon atoms joining to form cyclic structures such as benzene, C_6H_6.

It has been shown experimentally that all carbon–carbon double bonds in benzene are identical and somewhere between single and double bonds in length and bond strength. The equivalence of the bonds is shown by the alternative structure shown in diagram (c). The circle represents the even distribution of the six electrons around the benzene ring. The six electrons are said to be delocalised. Diagram (d) shows a shorthand representation of benzene.

Revision questions

15. **Draw the Lewis structures and structural formulas of ethane and ethyne, and describe their shapes.**
16 **Give the structural formula for each of the following.**
  (a) **propane** **(c)** **propene**
  (b) **octane** **(d)** **propyne**
17. **What are the products of the complete combustion of propene?**
18. **Write a balanced chemical equation for the reaction between propene and oxygen.**
19. **What is the molecular formula of an alkene having 21 carbon atoms?**
20. **Give the molecular formula of an alkene having 12 carbon atoms.**
21. **Alkenes are more reactive than alkanes. In terms of chemical bonding, explain why this is the case.**
22. **Name two reactions that alkenes can undergo.**
23. **Complete and balance the following combustion reactions.**
  (a) $C_4H_8(g) + O_2(g) \longrightarrow$ **(b)** $C_6H_{12}(g) + O_2(g) \longrightarrow$
24. **Give a structural equation for the reaction of ethene with fluorine gas.**
25. **Draw two possible isomers of C_4H_8.**
26. **As the molecular size of the alkanes increases, there is a change from the gaseous phase to the solid phase. Explain why this occurs.**

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Unit 1
AOS 2
Topic 3
Concept 5

Systematic naming of alkanes, alkenes and alkynes
Summary screen and practice questions

Naming organic substances

Over the years, as organic chemistry grew and more and more compounds were either discovered or synthesised, naming them became more and more of a problem. To solve this problem, the International Union of Pure and Applied Chemistry (IUPAC) developed a set of rules by which any organic substance could be named depending on the structure of its molecules. This system has widespread use today, although many common substances still retain their 'old' or 'trivial' names in everyday use. For example, ethyne, mentioned earlier in this chapter, is still called acetylene in everyday use.

Organic compounds are named according to their structures. A series of rules can be followed to accurately name any organic compound. For naming hydrocarbons, chain length, branching and type of bonding are important considerations.

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Unit 1
AOS 2
Topic 3
Concept 6

Systematic naming of branched hydrocarbons
Summary screen and practice questions

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Rules for naming hydrocarbons

There are a number of rules for naming hydrocarbons.

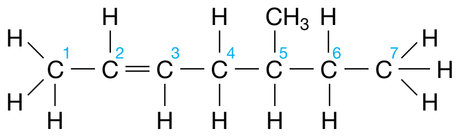
- Rule 1: Determine the longest chain of carbon atoms.
- Rule 2: Determine which end is nearest to a branch, a double bond or a triple bond. (A double or triple bond takes precedence over a branch if they are equidistant from either end of the chain.)
- Rule 3: Number the carbon atoms from the end chosen.
- Rule 4: Name any branches first with the ending *-yl* (for example, methane becomes *methyl* and ethane becomes *ethyl*), then the longest chain, then any single or double bond.
- Rule 5: When two or more branches occur on the same carbon atom, the number of the carbon atom is indicated for each branch, with the names given in alphabetical order (ignoring the prefixes described in rule 6).
- Rule 6: When two or more identical branches occur on different carbon atoms, the prefixes *di-*, *tri-* and *tetra-* are used.

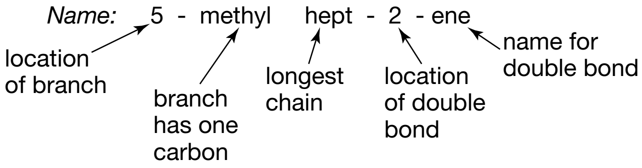
Avoid the following common errors when naming hydrocarbons:

- not identifying the longest chain possible
- not listing the side branches in alphabetical order
- omitting the prefixes *di-*, *tri-* and *tetra-* when they are required.

Sample problem 8.2

Name the following compound using rules 1 to 4 above.





Solution: **STEP 1**

The longest chain has seven carbon atoms, so the *hept* name is given.

STEP 2

The double bond is nearest the left-hand end.

STEP 3

Number from the left-hand end.

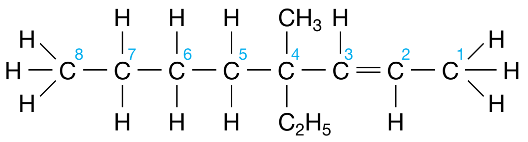
STEP 4

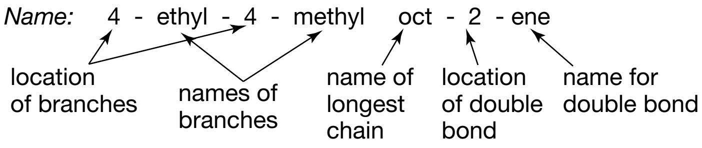
The branch on carbon atom number 5 contains one carbon atom, so it is given the name *5-methyl*. There is a double bond on carbon atom number 2 and this is represented by *2-ene*.

So, the hydrocarbon is called 5-methylhept-2-ene.

Sample problem 8.3

Name the following compound using rules 1 to 5 above.





Solution: **STEP 1**

The longest chain has eight carbon atoms, so the *oct* name is given.

STEP 2

The double bond is nearest the right-hand end.

STEP 3

Number from the right-hand end.

STEP 4

When writing names, separate numbers by commas and separate numbers and words by hyphens.

The upper branch on carbon atom number 4 contains one carbon atom, so it is given the name *4-methyl*. The lower branch contains two carbon atoms, so it is named *4-ethyl*. The double bond on carbon atom number 2 is represented as *2-ene*.

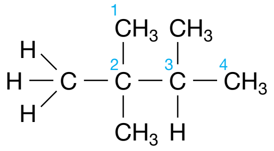
STEP 5

The two branches on carbon atom number 4 both have the number indicated and are written alphabetically as *4-ethyl-4-methyl*.

So the hydrocarbon is called 4-ethyl-4-methyloct-2-ene.

Sample problem 8.4

Name the following compound.



Name: 2,2,3-trimethylbutane

Solution:

1. The longest chain has four carbon atoms so the name *but-* is given.
2. There are methyl groups on carbon atoms number 2 and 3. Two of them occur on carbon atom number 2 and one on carbon atom number 3.
3. Number from the left-hand end.
4. Give the number of the carbon atom each time the methyl group occurs on that carbon atom.
5. Use the prefix *tri-* to indicate the total number of methyl groups.

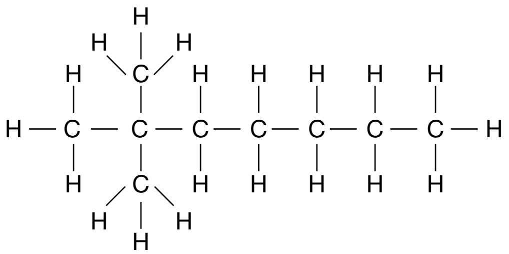
So, the hydrocarbon is called 2,2,3-trimethylbutane.

Revision questions

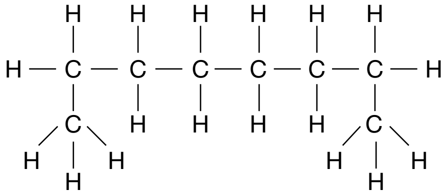
27. **Write structural formulas for:**
 (a) **hex-2-ene**
 (b) **methylpropene**
 (c) **pent-2-yne.**

28. **Name each of the following.**

(a)



(b)



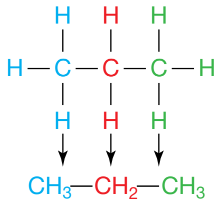
29. **Draw structural formulas for:**
 (a) **4-methylpent-2-ene**
 (b) **2,4-dimethylhexane (one of the components of unleaded fuel).**

Condensed formulas

The condensed formula is a way of representing the structural formula of an organic compound in a single line.

Hydrocarbons can be represented by condensed formulas. Another name for condensed formulas is semi-structural formulas. Ethane, C_2H_6, can be represented by the condensed formula CH_3CH_3. The condensed formula of propane, C_3H_8, is $CH_3CH_2CH_3$. When abbreviating a structural formula to its condensed formula equivalent, go 'carbon by carbon'.

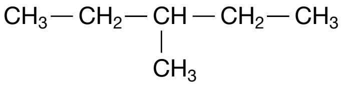
full structural formula



condensed formula

Brackets are used:

(a) to indicate side chains, which are written after the carbon to which they are attached — for example

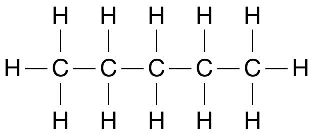


$$\begin{array}{c} CH_3—CH_2—\underset{\displaystyle CH_3}{\underset{|}{CH}}—CH_2—CH_3 \end{array}$$

3-methylpentane

$CH_3CH_2CH(CH_3)CH_2CH_3$

(b) if there are repeating CH_2 groups — for example:



pentane

$CH_3CH_2CH_2CH_2CH_3$ or $CH_3(CH_2)_3CH_3$

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Unit 1
AOS 2
Topic 3
Concept 2

Molecular, structural and condensed formulas
Summary screen and practice questions

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Molecular representations of butan-1-ol
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Revision questions

30. **Write the condensed formula for each of the compound(s) in the previous revision questions: Q27, Q28 and Q29.**

31. **The condensed formula for methylpropane is $CH_3CH(CH_3)_2$ and that of 2-methylpentane is $CH_3CH(CH_3)CH_2CH_2CH_3$. Write a condensed formula for 2,3-dimethylhexane.**

Functional groups

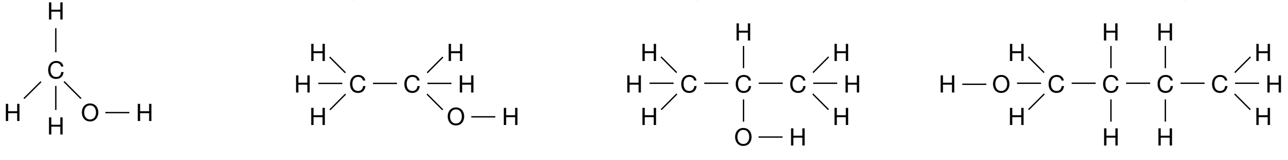
A functional group is an atom or group of atoms, or a bond, that gives an organic molecule its specific properties.

The bond, atom, or group of atoms that give a molecule its specific properties is called its functional group. For example, the functional group of the alkenes is C═C; the functional group of the alcohols is —OH. The oxygen–hydrogen bond in alcohols and carboxylic acids gives them significantly different properties from those of hydrocarbons.

Alcohols contain a hydroxy group, OH, as their functional group. They are named on the basis of the alkane that they resemble; for example, ethanol is the alcohol derived from ethane.

Alcohols

Alcohols are carbon chains containing one or more —OH groups. The —OH group is called the 'hydroxy' group. Names of alcohols have the ending *-ol*.



(a) methanol — solvent for varnishes, windscreen washer fluid

(b) ethanol — component of alcoholic beverages

(c) propan-2-ol — rubbing alcohol, solvent

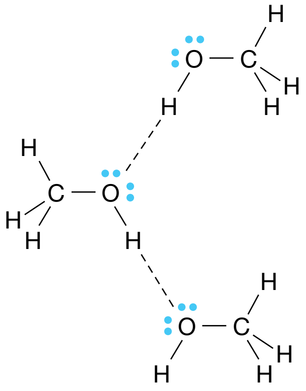
(d) butanol — solvent, hydraulic fluid

The first four alcohols

Properties of alcohols

Hydrogen bonding in alcohols affects boiling point and solubility in water.

In the smaller alcohols, the hydrogen bonding between the alcohol molecules is stronger than the dispersion forces between hydrocarbon molecules in alkanes, alkenes and alkynes. This results in higher boiling points than the corresponding alkane or alkene. More energy is need to separate methanol molecules than methane molecules.

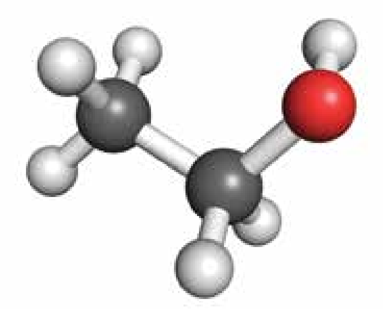


Hydrogen bonds between molecules of ethanol are represented by dotted lines.

In addition, the boiling points of alcohols increase with size because the non-polar portion of the molecule increases; this diminishes the effect of hydrogen bonding but increases the influence of the dispersion forces.

Smaller alcohols are soluble in water; hydrogen bonds form between the water molecules and the alcohol molecules. Solubility decreases with increasing size due to the increasing non-polar section of the alcohol molecules.

Reactions



Ball-and-stick model of ethanol

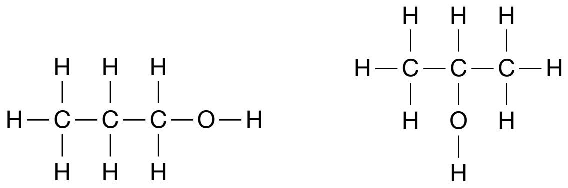
Alcohols also undergo combustion reactions in air. They can be used as fuels, so research is currently being done on so-called 'alcohol blends'. These involve adding alcohols (usually ethanol) to petrol to reduce our dependence on fossil fuels. Research is also being done on the viability of pure alcohols as transport fuels. A typical example of this combustion process is the burning of ethanol to produce carbon dioxide and water.

$$CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$

Naming of alcohols

The functional group is —OH. The number of the carbon atom to which the —OH has bonded should be given.

There are two isomers of C_3H_8O. Both have the same molecular formula but their structural formulas are different. Their condensed formulas are $CH_3CH_2CH_2OH$ (propan-1-ol) and $CH_3CH(OH)CH_3$ (propan-2-ol).



(a) propan-1-ol

(b) propan-2-ol

Structural formulas for two different isomers of this alcohol

Revision questions

32. **Write molecular, structural and condensed formulas for butan-1-ol.**
33. **Write the molecular, structural and condensed formulas for butan-2-ol.**

Carboxylic acids

Carboxylic acids are a further example of a homologous series. They contain the carboxy functional group —COOH.

Carboxylic acids are another example of a homologous series. These compounds contain the functional group —COOH at the end of a chain. The —COOH group is called the carboxy group. When naming a carboxylic acid,

Painful bee and ant stings contain methanoic acid, HCOOH.

add the ending *-oic acid* to the alkyl prefix. Remember to count the carbon atom in the —COOH group. For example, methane becomes methanoic acid.

Properties of carboxylic acids

These acids have typical acid properties. They are classed as weak acids, but concentrated ethanoic acid has a very strong pungent odour. When acting as an acid, the proton that is donated is the one that is attached to the oxygen atom in the —COOH group. The hydrogen bonding due to the —OH in the carboxy group is the reason for the relatively high boiling points and solubility of the smaller carboxylic acids; this is consistent with the alcohols.

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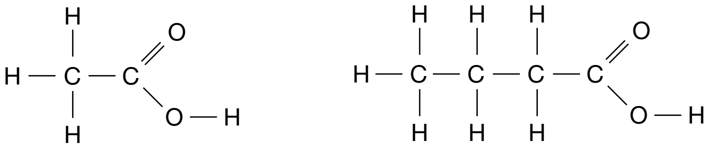
Unit 1
AOS 2
Topic 3
Concept 7

Alcohols and carboxylic acids
Summary screen and practice questions



(a) ethanoic acid (acetic acid) (b) butanoic acid

Structural formulas for two carboxylic acids

TABLE 8.2 Sources and uses of some common organic acids

Name	Sources and uses
methanoic acid (formic acid)	responsible for sensation caused by nettle, bee and ant stings; used in medicine, food preservation and textile industry
ethanoic acid (acetic acid)	main component of vinegar; used as a solvent; salts of the acid used in insecticides and fungicides
propanoic acid (propionic acid)	used as an antifungal agent in the baking industry and in ointments (either in salt or acid form)
butanoic acid (butyric acid)	odour-causing component of rancid butter, rotten socks and body odour
2-hydroxybenzoic acid (salicylic acid)	used as a food preservative
2-acetoxybenzoic acid (acetylsalicylic acid, aspirin)	used for relief of fever, pain and rheumatic conditions

Note: Common names are given in brackets.

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Unit 1 **AOS 2** **Topic 3** **Concept 8**

Esters
Summary screen and practice questions

Revision questions

34. **Draw the Lewis structure and structural and condensed formulas for methanoic acid.**

35. **Write molecular, structural and condensed formulas for:**
  (a) **ethanoic acid**
  (b) **propanoic acid**
  (c) **butanoic acid.**

36. **Explain why the carboxylic acid series can be described as a homologous series. Provide names of members of the series and structural formulas to support your explanation.**

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Experiment 8.1 Formation of esters
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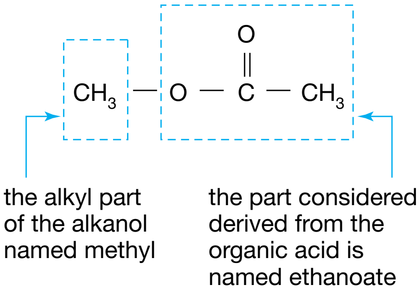
Reactions

The carboxylic acids react with alcohols to form **esters**. Esters are a group of compounds that give the pleasant ‘fruity’ smell to various fruits. This is called a condensation or esterification reaction. Concentrated sulfuric acid is used to link the alcohol and carboxylic acid together by removing a water molecule.

$$\text{carboxylic acid} + \text{alcohol} \xrightarrow{H_2SO_4} \text{ester} + \text{water}$$

$$CH_3COOH(l) + CH_3OH(l) \xrightarrow{H_2SO_4} CH_3COOCH_3(l) + H_2O(l)$$

$$\text{ethanoic acid} + \text{methanol} \xrightarrow{H_2SO_4} \text{methyl ethanoate} + \text{water}$$



Full name of ester: methyl ethanoate

Naming methyl ethanoate

Esters have relatively low boiling points because the intermolecular attraction is dipole-dipole (and dispersion forces) instead of the hydrogen bonding that occurs in carboxylic acids.

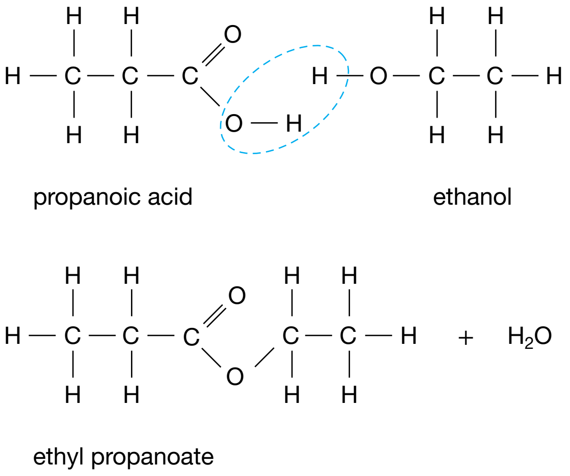
Naming an ester

When naming an ester, the alcohol (alkyl) name is given first, followed by the acid part. (For example, the alcohol methanol becomes *methyl*). The -*oic* ending of the acid is replaced with -*oate*. (The carboxylic acid called ethanoic acid becomes *ethanoate*.) So, methanol reacting with ethanoic acid forms the ester methyl ethanoate.

Sample problem 8.5

Name and draw the expanded structural formula of the ester formed when ethanol reacts with propanoic acid.

Solution:



An important reaction of carboxylic acids is their reaction with alcohols to produce esters. Esters are named according to the particular alcohol and carboxylic acid used to form them.

Revision questions

37. **Name the ester formed from ethanol and ethanoic acid.**
38. **Give the expanded structural formula of ethyl pentanoate.**
39. **What is the role of concentrated sulfuric acid in the formation of an ester?**
40. **What is the functional group in the carboxylic acids?**
41. **Give an expanded structural formula for pentanoic acid.**
42. **Which of the carboxylic acids would you be most likely to eat? Explain.**

Where do hydrocarbons come from?

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Unit 1
AOS 2
Topic 3
Concept 1

Organic compounds and their origins
Summary screen and practice questions

Hydrocarbons are found in crude oil, and processing the oil yields many useful products. The hydrocarbons described so far in this chapter are largely synthesised or produced from fossil fuels. Fossil fuels also form the major part of our fuel resources. As the name suggests, they are derived from the fossil remains of living organisms, which have been altered by heat and pressure.

Petroleum

Petroleum is also known as crude oil. The word ‘petroleum’ is derived from two Latin words, *petro* meaning ‘rock’ and *oleum* meaning ‘oil’. So it is literally ‘rockoil’. It was given this name because it was originally found seeping through rocks to the Earth’s surface. Petroleum takes many thousands of years to form. In ancient times, warm seas covered much of the area that is now land. These areas contained plant and animal life that died, fell to the sea floor and became covered with sediment. Increased pressure and temperature over a long period of time caused these substances to change into petroleum. Later, movements of the Earth’s crust pushed these deposits closer to the surface and trapped pockets of natural gas, crude oil and salt water.

Refining crude oil

Many mixtures of liquids, including crude oil, may be separated into its components, called fractions, using the process known as **fractional distillation**. This process is used to separate one liquid from a mixture of liquids with different boiling temperatures. In industry, this process is used on a large scale to separate crude oil into a number of fractions based on boiling temperatures. This is the first stage of refining. It represents a ‘rough sorting out’ of the thousands of different compounds that are in the crude oil.

In this process, hot crude oil (which is essentially a very complicated mixture of alkanes) enters the base of the fractionating tower and almost all of it is immediately vaporised. The small portion that remains liquid (due to the high boiling temperature caused by its large molecules) is removed from the base to make products such as bitumen for roads. As the vapours rise up the tower, they cool and re-liquefy at different heights depending on their boiling (or condensation) temperatures. The lighter fractions have the smallest molecules and, therefore, the lowest condensation temperatures; they are collected near the top of the tower. Moving down the tower, the fractions collected become ‘heavier’; that is, they contain larger and larger molecules. Special arrangements inside the tower called bubble caps let vapours rise while allowing any condensed liquids to be collected on special trays.

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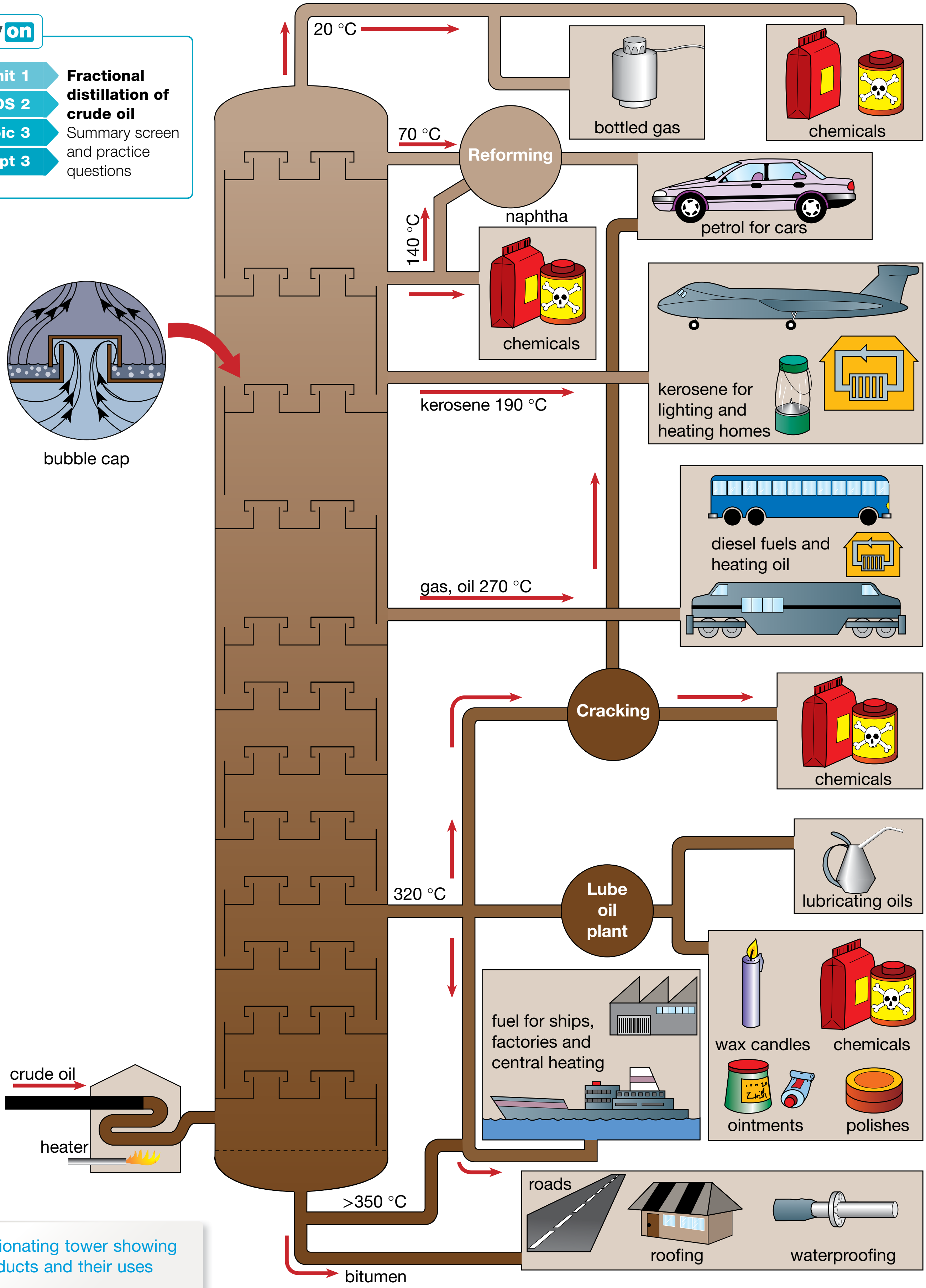
Unit 1	**Fractional distillation of crude oil**
AOS 2	Summary screen and practice questions
Topic 3	
Concept 3	



A fractionating tower showing oil products and their uses

Distillation is used to separate the complicated crude oil mixture into fractions. These fractions contain molecules of roughly equal size that therefore have similar boiling temperatures.

TABLE 8.3 Products from the fractional distillation of crude oil

Product	Number of carbon atoms	Boiling point (°C)	Use
gas	1–4	<40	bottled gas, plastics, chemicals
petrol	4–12	40–75	fuel for vehicles, chemicals
kerosene	9–16	150–240	jet fuel, chemicals
diesel oil	15–25	220–250	fuel
lubrication oil	20–70	250–350	lubricants, waxes
bitumen residue	>70	>350	road surfaces

Revision questions

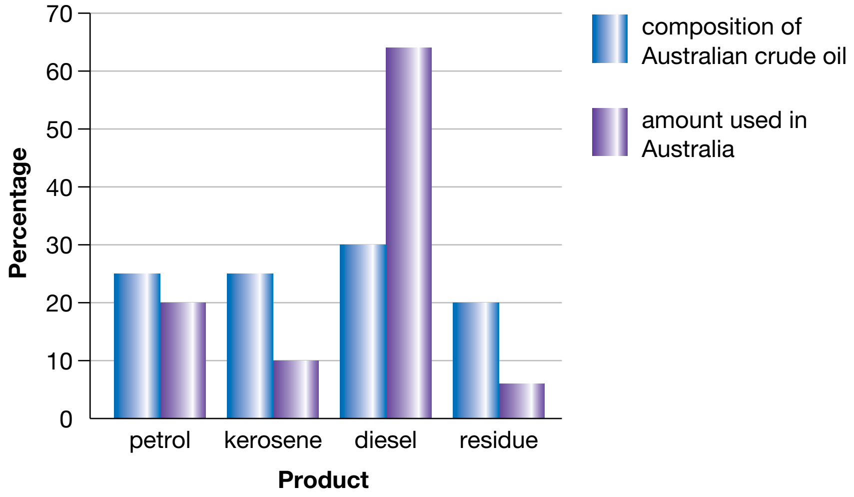
43. **What property of the lighter fractions of crude oil makes them burn more easily than the heavier ones?**

44. **Viscosity refers to how well a liquid flows; honey is described as a viscous liquid. Why do you think the fractions that are recovered later in the fractional distillation process have increasing viscosity?**

45. **Explain why the fractions with the fewest carbon atoms in their molecules rise to the greatest height in the fractionating tower. Use table 8.3.**

Cracking

We need far more diesel fuel than the amount obtained from the fractional distillation of Australian crude oil. This process also produces far more kerosene than we need.



A number of different fractions can be obtained from crude oil. Unfortunately, demand does not equal the output of the various fractions.

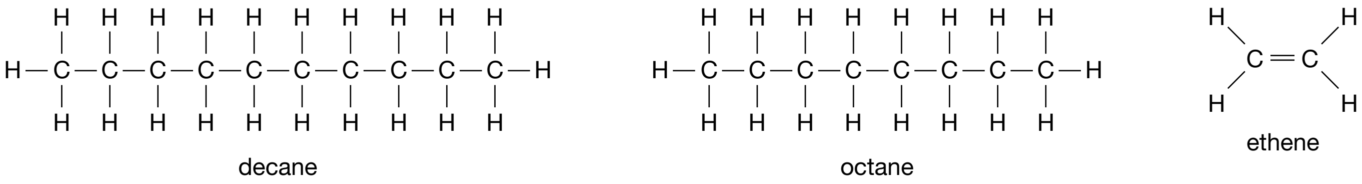
Cracking converts larger and less useful alkane molecules into smaller, more useful ones. Small alkene molecules are usually formed at the same time. Cracking may be done either at high temperatures (thermal cracking) or at lower temperatures in the presence of a catalyst (catalytic cracking).

In order to obtain more diesel and other smaller molecules, a process has been developed whereby some of the excess larger hydrocarbons are broken into smaller ones — the ones that, generally speaking, are more useful. This process is called 'cracking'. There are two main ways that this can be done. The first uses high temperatures and is termed **thermal cracking**. The second uses a catalyst at much lower temperatures to produce the same effect. This is called **catalytic cracking**. Both methods have their own advantages and disadvantages, and both are used at a number of locations around Australia.

An important by-product of cracking is the industrially important class of hydrocarbons called alkenes — the raw materials of the plastics industry.

An example of cracking is the breakdown of decane (a large hydrocarbon) into octane (used in motor fuel) and ethene (used for the manufacture of plastic).

$$C_{10}H_{22}(g) \xrightarrow{\text{catalyst}} C_8H_{18}(g) + C_2H_4(g)$$



This is just one possible way in which decane could crack. It could also crack at other points to produce different molecules.

Revision questions

46. **Write a chemical equation to show one way in which dodecane, $C_{12}H_{26}$, could crack when subjected to catalytic cracking. Compare your answers with others in the class. Who is correct?**

47. **One kind of catalyst used in catalytic cracking is made of silicon. What is the role of the catalyst in catalytic cracking?**

48. **Give two advantages of using a catalyst when cracking large hydrocarbons.**

Chapter review

studyon

Unit 1
AOS 2
Topic 3

Organic chemistry
Summary screens and practice questions

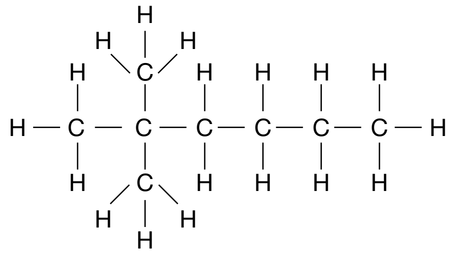
Summary

- Organic chemistry is the study of the compounds of carbon. The oxides of carbon, carbonates, carbides and cyanides, as well as compounds of the remaining elements, are considered to be inorganic substances.
- Carbon in organic compounds always forms four covalent bonds.
- Hydrocarbons are organic compounds made up of only carbon and hydrogen atoms. Dispersion forces are involved in intermolecular bonding in hydrocarbons. This means that boiling temperatures are low. Hydrocarbon molecules are non-polar so they are insoluble in water.
- Saturated hydrocarbons contain only single bonds; unsaturated hydrocarbons contain double or triple bonds.
- Alkanes are hydrocarbons containing only single bonds between carbon atoms. Alkanes can undergo oxidation (combustion) and substitution reactions. They burn in oxygen to give off carbon dioxide and water.
- Haloalkanes contain one or more halogen atoms.
- Alkenes are hydrocarbons containing one double bond between two carbon atoms. Alkenes can undergo combustion, addition and polymerisation reactions. These reactions involve, respectively, reacting with oxygen to form carbon dioxide and water; reacting with hydrogen (or the halogens) so that the double bond between the two carbon atoms is broken; and forming long chains called polymers as a result of alkene molecules joining together.
- Alkynes are hydrocarbons containing one triple bond between two carbon atoms. Alkynes can undergo combustion and addition reactions.
- In a homologous series, successive members differ by a —CH_2— group.
- Isomers of a compound have the same molecular formula but different structural formulas.
- A functional group is the bond, atom or group of atoms that gives a group of molecules its specific properties.
- Alcohols contain one or more hydroxy (—OH) groups. Due to the presence of hydrogen bonding, alcohols have a higher boiling temperature than the corresponding alkane. Smaller alcohols are soluble in water.
- Carboxylic acids contain a carboxy (—COOH) group at the end of a chain. Due to the presence of hydrogen bonding, carboxylic acids have a higher boiling temperature than the corresponding alkane. Smaller carboxylic acids are soluble in water.
- Esters form when alcohols react with carboxylic acids. Esters are named by changing the name of the alcohol to the alkyl form and replacing the ending of the name of the carboxylic acid with *–oate*. For example, the ester formed from propanol and ethanoic acid is propyl ethanoate.
- The three main groups of fossil fuels are natural gas, petroleum and coal.
- Hydrocarbons are obtained from petroleum (crude oil).
- Petroleum is separated using fractional distillation. Fractional distillation separates components according to boiling points. In a fractionating tower, components with lower boiling points are drawn up closer to the top.
- Catalytic and thermal cracking are used to increase the number of industrially important smaller molecules produced from the fractional distillation of crude oil.

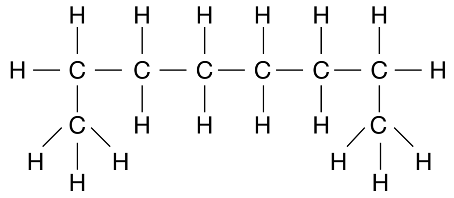
Multiple choice questions

1. The structure that is *not* an isomer of C_8H_{18} is:

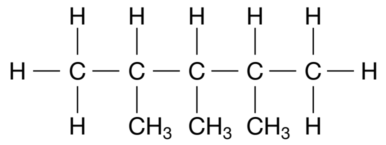
A



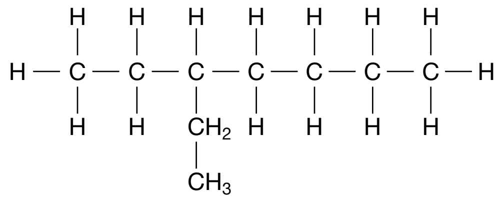
B



C



D



2. Which of the following is *not* the formula of an alkane?

A C_4H_{10}
B C_7H_{16}
C $C_{11}H_{22}$
D $C_{14}H_{30}$

3. The process of changing large hydrocarbon molecules into smaller hydrocarbon molecules is called:

A distillation
B polymerisation
C cracking
D esterification.

4. The bonds between carbon atoms in benzene are:

A single bonds
B double bonds
C alternate double and single bonds
D intermediate between double and single bonds.

5. Which of the following statements does *not* represent a basis for the existence of a large number of carbon compounds?

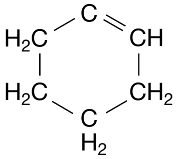
A A single carbon atom can form up to four covalent bonds.
B Carbon can link to a number of other elements by strong bonds.
C Carbon forms a great variety of chains and rings.
D Carbon exists in a number of structural forms in the solid state.

6. Which of the following statements best describes a saturated compound?

A A saturated compound, when used as a solvent, does not dissolve any more solute at a particular temperature.
B A saturated compound contains carbon and hydrogen atoms only.
C A saturated compound has only single carbon–carbon bonds.
D A saturated compound does not react with chlorine.

7. Which one or more of the following compounds is saturated?

A $H_2C=CH_2$
B $HC\equiv CH$
C



D $CH_3CH_2CH_3$

8. Which of the following could *not* be an alkane?

A C_3H_8
B C_4H_8
C $CH_3CH_2CH_2CH_2CH_3$
D C_6H_{14}

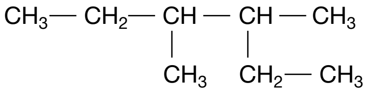
9. An alkyl group is:

A a strongly basic group in an organic molecule
B a group of organic compounds related to the alkanes
C a group equivalent to an alkane molecule less one hydrogen atom
D a group of organic compounds with the general formula C_nH_{2n+2}.

10. Which one or more of the following statements is *not* correct? A functional group:

A remains unchanged in most chemical reactions
B is common to all members of the particular homologous series
C may be regarded as being responsible for a characteristic set of chemical reactions for a given homologous series
D varies from member to member by $-CH_2$.

11. An organic compound has the formula:



$CH_3-CH_2-CH(CH_3)-CH(CH_2-CH_3)-CH_3$

The name of the compound above is:

A 2-ethyl-3-methylpentane
B octane
C 3,4-dimethylhexane
D 3-methyl-4-ethylpentane.

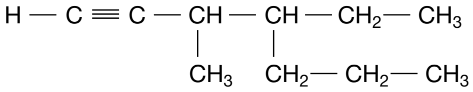
12. The group in which all the compounds are members of the same homologous series is:

A CH_4; C_2H_4; C_3H_4
B CH_3OH; CH_3CH_2OH; $CH_3CH_2CH_2OH$
C CH_4; CH_3OH; $HCOOH$
D CH_3OCH_3; CH_3CH_2OH; C_2H_5OH.

13. 2-methyl-2-butanol is an isomer of:

A butan-2-ol
B pentanoic acid
C pentan-1-ol
D 2,3-dimethyl-butan-2-ol.

14. An organic compound has the formula:

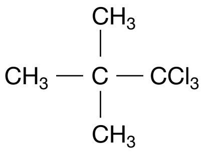


$H-C\equiv C-CH(CH_3)-CH(CH_2-CH_2-CH_3)-CH_2-CH_3$

The name of this compound is:

A 3-methyl-4-ethylhept-1-yne
B decyne
C 4-ethyl-3-methylhept-1-ene
D 3-methyl-4-propylhex-1-ene.

15. The formula of a compound is:



$CH_3-C(CH_3)_2-CCl_3$

The name of this compound is:

A trichloroisopentane
B 2,2-dimethyl-3,3,3-trichloropropane
C 1,1,1-trichloro-2,2,2-trimethylethane
D 1,1,1-trichloro-2,2-dimethylpropane.

16. The alkane in the following group with the highest boiling point is:

A	methane	**C**	propane
B	ethane	**D**	butane.

17. The organic compound in the following group with the highest boiling point is:

- **A** acetylene (ethyne)
- **B** ethane
- **C** 1,1,1-trichloroethane
- **D** ethan-2-ol.

18. When the ring on the Bunsen burner is closed during combustion, soot forms. Soot is:

A	carbon	**C**	carbon dioxide
B	carbon monoxide	**D**	unburnt propane.

Review questions

Hydrocarbons

1. (a) What is meant by the term 'hydrocarbon'?
   (b) Give the names and formulas of three hydrocarbons each containing three carbon atoms.

2. (a) Alkanes and alkenes are examples of hydrocarbons. In what way do they differ? Give an example of each.
   (b) Write an equation for the oxidation of the second member of the alkene series and the third member of the alkane series.

3. (a) How are alkanes 'produced'?
   (b) What are the major uses of alkanes?

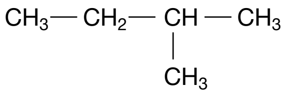
4. (a) How are alkenes 'produced'?
   (b) What are their uses?

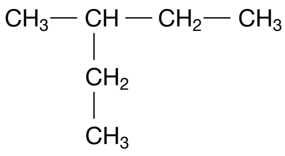
5. Define the term 'isomer'. Illustrate your answer with the isomers of pentane and butene.

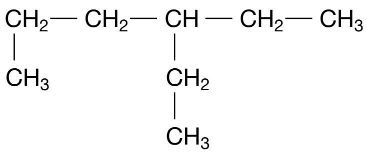
6. (a) Distinguish between saturated and unsaturated hydrocarbons.
   (b) Which is more reactive?

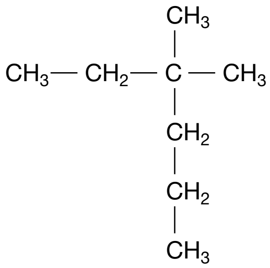
7. Write a formula for each of the following:
   (a) an alkane with 22 carbon atoms
   (b) an alkene with 17 carbon atoms
   (c) an alkyne with 13 carbon atoms.

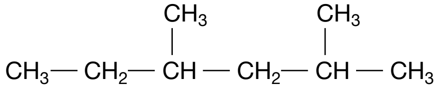
8. Name the following organic compounds. (*Note:* The longest continuous chain may not be written in a straight line.)

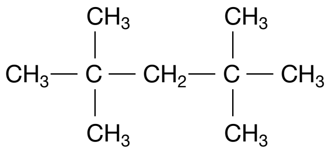
(a)

$$CH_3-CH_2-\underset{\begin{array}{c}|\\CH_3\end{array}}{CH}-CH_3$$

(b)

$$CH_3-\underset{\begin{array}{c}|\\CH_2\\|\\CH_3\end{array}}{CH}-CH_2-CH_3$$

(c)

$$\underset{\begin{array}{c}|\\CH_3\end{array}}{CH_2}-CH_2-\underset{\begin{array}{c}|\\CH_2\\|\\CH_3\end{array}}{CH}-CH_2-CH_3$$

(d)

$$CH_3-CH_2-\overset{\begin{array}{c}CH_3\\|\end{array}}{\underset{\begin{array}{c}|\\CH_2\\|\\CH_2\\|\\CH_3\end{array}}{C}}-CH_3$$

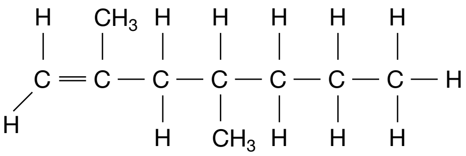
(e)

$$CH_3-CH_2-\overset{\begin{array}{c}CH_3\\|\end{array}}{CH}-CH_2-\overset{\begin{array}{c}CH_3\\|\end{array}}{CH}-CH_3$$

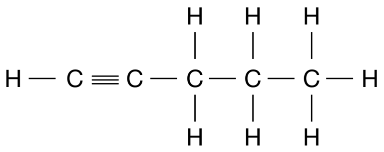
(f)

$$CH_3-\overset{\begin{array}{c}CH_3\\|\end{array}}{\underset{\begin{array}{c}|\\CH_3\end{array}}{C}}-CH_2-\overset{\begin{array}{c}CH_3\\|\end{array}}{\underset{\begin{array}{c}|\\CH_3\end{array}}{C}}-CH_3$$

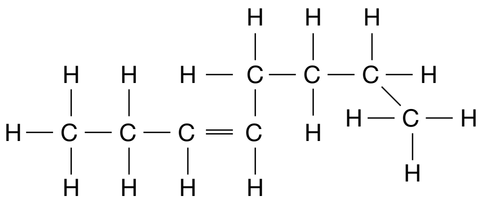
9. Draw structural formulas for the following compounds.
   (a) 3,3-dimethylpentane
   (b) 3-ethylhexane
   (c) 2,2,4-trimethylpentane
   (d) 2,3-dimethylhexane
   (e) 3-ethylheptane
   (f) 3-ethyl-2,4-dimethyloctane

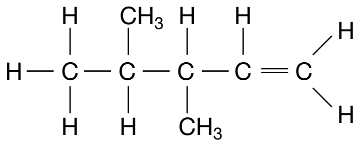
10. Draw structural formulas for:
   (a) 2-methylheptane
   (b) 2,3-dimethylbut-2-ene
   (c) pent-2-yne
   (d) 2,6-dimethylhept-3-ene
   (e) oct-3-ene
   (f) 3-methyl-4-ethylhex-2-ene
   (g) 3,4-diethyl-3-methylhexane
   (h) 3-ethyl-4,5-dipropyloctane.

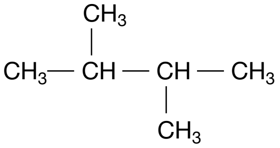
11. Name each of the following.

(a)


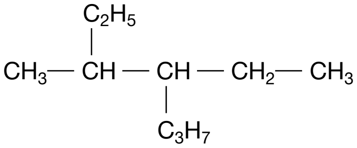
(b)

$$H-C\equiv C-\overset{\begin{array}{c}H\\|\end{array}}{\underset{\begin{array}{c}|\\H\end{array}}{C}}-\overset{\begin{array}{c}H\\|\end{array}}{\underset{\begin{array}{c}|\\H\end{array}}{C}}-\overset{\begin{array}{c}H\\|\end{array}}{\underset{\begin{array}{c}|\\H\end{array}}{C}}-H$$

(c)


(d) 

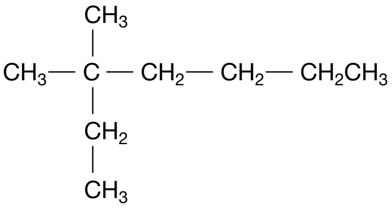
(e) 

$CH_3-\overset{\overset{CH_3}{|}}{CH}-\underset{\underset{CH_3}{|}}{CH}-CH_3$

(f) 

$CH_3-\overset{\overset{C_2H_5}{|}}{CH}-\underset{\underset{C_3H_7}{|}}{CH}-CH_2-CH_3$

12. Peter isolated the following compound and named it 2-ethyl-2,4-dimethylhexane. Sally said that the compound was incorrectly named. Explain why Sally is right and give the correct name for the compound.



$CH_3-\overset{\overset{CH_3}{|}}{\underset{\underset{\underset{\underset{CH_3}{|}}{CH_2}}{|}}{C}}-CH_2-CH_2-CH_2CH_3$

Functional groups

13. The alkanols form a homologous series. What does this mean?

14. Briefly explain the meanings of the following terms. Use examples in your answers.
  (a) Substitution reactions
  (b) Addition reactions
  (c) Functional group
  (d) Fractional distillation
  (e) Catalytic cracking

15. Draw structural formulas for the following compounds:
  (a) the paint remover dichloromethane
  (b) the water repellant used in Scotchgard, 1,1,1-trichloroethane.

16. Use a labelled diagram of a methanol molecule and a water molecule to explain why methanol is soluble in water.

17. (a) Draw the Lewis structures and structural and condensed formulas of ethane and ethanoic acid.
  (b) Outline the similarities and differences between these two compounds.

18. Volatile compounds readily vaporise.
  (a) Explain why esters are volatile compounds.
  (b) Would you expect esters to be soluble? Explain your answer.

19. Consider the following reaction:

$$\underset{\text{(i)}}{CH_3COOH} + \underset{\text{(ii)}}{CH_3CH_2OH} \rightarrow \underset{\text{(iii)}}{CH_3COOCH_2CH_3} + \underset{\text{(iv)}}{H_2O}$$

  (a) Draw the structural formulas represented by the equation above, and name the compounds (i) to (iv).
  (b) What type of reaction is this?
  (c) If four bottles contained, respectively, pure samples of (i), (ii), (iii) and (iv), how might you identify them?

20. Name the esters formed from the following reactions.
  (a) methanol and butanoic acid
  (b) 1-propanol and propanoic acid
  (c) ethanol and methanoic acid

21. Write formulas for the following esters.
  (a) ethyl butanoate
  (b) propyl ethanoate
  (c) butyl propanoate

22. A molecular compound contains four carbon atoms. Write a condensed formula for the compound if it is:
  (a) an alkene
  (b) an alcohol
  (c) a carboxylic acid
  (d) an ester.

Calculations in organic chemistry

Refer to chapter 5 to answer questions 23–27.

23. Find the percentage of carbon in:
  (a) ethane (b) ethene (c) ethyne.

24. A hydrocarbon contains 83.7% of carbon. Determine its empirical formula.

25. A 5.00 g sample of a hydrocarbon contains 0.908 g of hydrogen. Its molar mass is 44.0 g mol^{-1}.
  (a) Find its empirical formula.
  (b) Draw its structure.
  (c) State the homologous series to which it belongs.

26. An ester that is used in making nail polish remover is manufactured using ethanol. It has the following composition: 54.5% C, 9.15% H and 36.3%O. Its molar mass is 88.0 g mol^{-1}. Calculate the empirical formula and molecular formula of the ester and draw its structure.

27. An alcohol containing only the elements C, H and O was shown by analysis to contain 60.0% carbon and 13.3% hydrogen. Its molar mass is 60.0 g mol^{-1}. Identify the alcohol and draw possible structures.

studyon

Unit 1
AOS 2
Topic 3

Organic chemistry

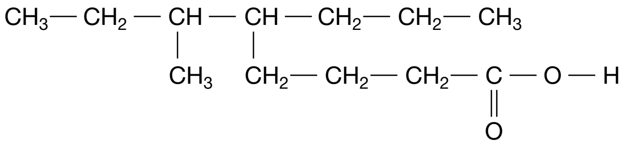
Sit topic test

Exam practice questions

In a chemistry examination you will be required to answer a number of short and extended response questions.

Multiple choice questions

1. An organic compound has the formula:



$$\begin{array}{l} CH_3-CH_2-\underset{\displaystyle CH_3}{\underset{|}{CH}}-\underset{|}{CH}-CH_2-CH_2-CH_3 \\ \qquad\qquad\qquad\quad CH_2-CH_2-CH_2-\underset{\displaystyle O}{\underset{\|}{C}}-O-H \end{array}$$

The name of this compound is:

A 3-methyl-4-butyloctanoic acid

B 4-propyl-5-methyloctanoic acid

C 5-propyl-6-methyloctanoic acid

D 3-methyl-4-propyloctanoic acid. **1 mark**

2. The boiling points of the homologous series of alkanes:

A decrease with increase in molecular size

B decrease with decrease in molecular mass

C are all approximately the same

D are all significantly higher than the corresponding alkenes containing the same number of carbon atoms. **1 mark**

Extended response questions

3. In terms of bonding, explain why heavier hydrocarbons such as diesel oil (those with a high number of carbon atoms) have higher boiling temperatures than lighter hydrocarbons such as propane (those with fewer carbon atoms). **2 marks**

4. If a mixture of lubricating oil, diesel oil, petrol, paraffin and kerosene were placed in a fractional distillation apparatus:

(a) which fraction would be collected first

(b) which would have the lowest boiling temperature

(c) which would have the highest boiling temperature

(d) what could be done if the fractions collected were found to still be mixtures? **4 marks**

CHAPTER

9 Polymers

In this chapter, you will study the polymerisation processes by which complex organic molecules, called polymers, may be formed from simpler molecules. You will examine various types of polymers and how their properties can be altered by different processes and additives. The advantages and disadvantages of the use of polymers are discussed, as well as which plastics can be recycled to form other useful materials.

YOU WILL EXAMINE:

- the relationship between the structure and bonding of polymer materials and their properties and uses
- thermoplastic and thermosetting polymers, and how their different structures affect their elasticity
- addition and condensation polymerisation
- the structures of high-density polyethene (HDPE), and low-density polyethene (LDPE), their properties and recycling
- the general structural model for polymer materials and the forces existing between the chains
- advantages and disadvantages of using polymers
- recycling plastics.

> Reuse, reduce, recycle, repair.
>
> *Anonymous*

Originally, toothbrushes were made from the neck hairs of wild hogs, and these hairs were attached to either a bone or bamboo handle. But, thankfully, with the advent of modern plastics, getting a new toothbrush usually requires only a short trip to the local supermarket. Toothbrushes today are constructed with nylon bristles fixed to a polypropene handle. This was one of the first commercial uses of nylon. Each of these materials has specific properties that make it suitable for this purpose. Unlike hog's hair, which is a natural polymer, both nylon and polypropene are synthetic polymers. What are polymers, and how are they customised and produced?

Polymers

It is difficult to imagine, now that we have grown so used to plastics, what life was like without them. Many of the items we use every day, from objects as insignificant as pens and zippers to larger items such as furniture, refrigerators and the interiors of cars, are made from plastics. Even our clothes are made from synthetic fibres. Many modern products are produced from a selection of plastics, each one chosen for its specific properties.

The first polymers

Polymers are large molecules made by joining smaller molecules, called monomers, together. Polymers, including plastics, are a wide range of substances, both natural and synthetic.

In this chapter you will discover how the widely used compounds called ‘plastic’ are manufactured. There are many different kinds of plastic, and collectively these are called **polymers**. There are natural and synthetic polymers. The polymer industry is mainly concerned with the development and manufacture of various kinds of synthetic polymer called **plastics**.

The development of plastics began in the middle of the nineteenth century, following a competition set up by a manufacturer of billiard balls. In those days, billiard balls were made of ivory and the manufacturer, Phelan and Collander, offered a substantial prize for the discovery of a satisfactory alternative. One of the entrants in the competition was John Wesley Hyatt, who developed a substance called celluloid. He did not win the prize, as celluloid is volatile, and the billiard balls would have exploded on impact. However, celluloid was the first **thermoplastic** — a substance moulded using heat and pressure that retains its shape once cooled — and celluloid became a widely used polymer. Its applications included dental plates and men’s collars, but it is most identified with its role in the early photographic and motion picture industries, where it was used to make film.

In 1907 the first completely synthetic polymer, called Bakelite, was produced by a Belgian–American chemist Leo H. Baekeland (1863–1944). Bakelite, unlike celluloid-based plastics, is a thermosetting polymer and cannot be softened by heat once it has set. It is also resistant to common acids and solvents. These properties made it virtually indestructible, and Bakelite was a landmark in the history of plastics. Its discovery started a large plastics industry that produced items such as telephones, billiard balls and insulators for electrical devices. The age of plastics had begun. Chemists went on to invent nylon, PVC, teflon, polyethene, velcro and many other new materials with important new properties.

The casings and dials of early telephones, radios and many other appliances were frequently produced from Bakelite. Modern phones are fabricated using polyethene or PVC, which are less brittle and can be manufactured in a range of different colours. (Bakelite usually had a brown appearance due to the oxidation of phenol in air.)

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Polymers

What is a polymer?

The name polymer comes from the Greek *polymeres* ('of many parts'). A polymer molecule is made up of thousands of units strung together into very long chains. The simple molecules that are strung together are called **monomers**. Monomers link together to form a polymer chain in a process called **polymerisation**. The monomers used to make the polymer can be the same or different. **Copolymers** are formed when two or more different monomers are used. Most polymerisation reactions require a catalyst. Polymers may be either natural or synthetic.

Wool, cotton, linen, hair, skin, nails, rubber and flesh are all naturally occurring polymers. Most natural polymers are made of proteins or cellulose. Cotton is nearly all cellulose and hair is a protein polymer.

Synthetic polymers are commonly referred to as plastics. 'Plastic' means pliable or able to be moulded, and this is true of these synthetics, at least during their production.

Thermosoftening polymers can be heated repeatedly, softened, reshaped and hardened by cooling. Thermosetting polymers do not soften when heated but char instead. They cannot be reshaped.

Thermoplastic and thermosetting polymers

Polymers may be classified on the basis of their thermal behaviour. Plastics that may be repeatedly melted, reshaped and hardened by cooling are called **thermosoftening** plastics or thermoplastics. An example of a thermoplastic polymer is polystyrene. Plastics that do not melt but char when heated are called **thermosetting** plastics. These plastics must be moulded or shaped during their manufacture. Bakelite is an example of a thermosetting plastic material. Further examples of each of these plastics and their uses are described in table 9.1.

TABLE 9.1 Common polymers and their uses

Thermal behaviour	Polymer	Use
thermoplastic	polyvinyl chloride (PVC)	water piping, upholstery covering, toys, flexible laboratory tubing, suitcases, packaged meats, juice bottles, credit cards, shampoo bottles
	nylon	hosiery, carpets, clothing, bearings, fishing lines, tennis racquet strings, toothbrush bristles, mascara brush bristles
	polystyrene	foam for insulation, plates, cups, trays, packaging, surfboards, beanbag filling, refrigerator and washing machine components
	polytetrafluoroethene (PTFE, Teflon®)	washers, coating on frying pans, razor blades, skis
	polymethyl methacrylate (Perspex®)	advertising signs, telephones, light fittings, aeroplane windows
thermosetting	urea formaldehyde or phenol formaldehyde (Bakelite®)	backs of television sets, ash trays, vacuum jugs
	expoxy resin (Araldite®)	handles of screwdrivers
	melamine formaldehyde (Formica®)	kitchen cupboards and benches
	polyurethane	foam rubber, pillow filling, packaging

For addition polymerisation to occur, the monomer must possess a carbon–carbon double bond. In the joining process, one of the bonds in the double bond breaks and its electrons are used to form new bonds that join the monomers together.

Polymer manufacture

A range of processes are used in polymer manufacture. There are two main ways to produce polymers from monomers by polymerisation. These are **addition polymerisation** and **condensation polymerisation**. In each of these processes, different **functional groups** can be used. As seen in chapter 8, a functional group is a group of atoms, a bond or an atom that gives a molecule its specific properties. The focus in this chapter is on addition polymerisation.

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Unit 1
AOS 2
Topic 4
Concept 1

Addition polymerisation
Summary screen and practice questions

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Making polymers
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Addition polymerisation

Addition polymers are widely used in the packaging industry, as well as featuring in many other products. The uses of the addition polymers polyvinyl chloride (PVC), polytetrafluoroethene (Teflon®) and polystyrene were shown in table 9.1. The uses of two other common addition polymers are shown in table 9.2. The manufacture and uses of polymers made from the monomer ethene are discussed in further detail below.

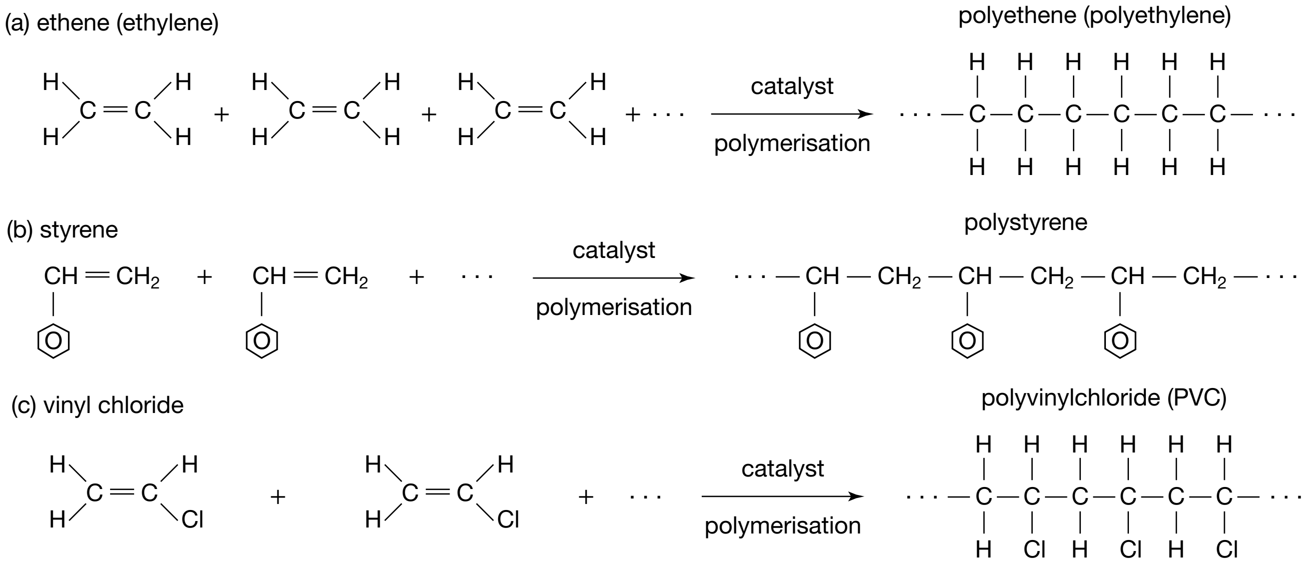
TABLE 9.2 Addition polymers and their uses

Polymer	Monomer	Use
polyethene	ethene	cling film, squeeze bottles, milk crates
polypropene (polypropylene)	propene	mixing bowls, ice-cream containers, moulded chairs

When ethene is subjected to high pressure it changes from a gas to a liquid. If the liquid ethene, still under pressure, is heated in the presence of a catalyst (a small quantity of oxygen), an **addition reaction** takes place in which the ethene molecules join together and form a long chain (**polyethene** or polyethylene). The length of these chains can vary from 100 to 1000 carbon atoms, depending on the temperature and pressure used.

For addition polymerisation to occur, the monomer must have a double bond between two carbon atoms. This double bond breaks to allow the long chains to form.

Modifying ethene by substituting different functional groups for hydrogen atoms produces other monomers that can be polymerised to make polymers with different properties, as shown in the figure below.



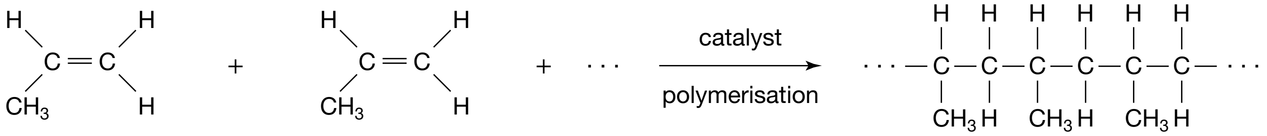
The formation of the polymers (a) polyethene, (b) polystyrene and (c) polyvinylchloride from their monomers ethene, styrene and vinyl chloride. The double bond between the two carbon atoms in the monomer breaks, allowing long chains to form. The resulting materials have an extraordinary array of applications, among them plastic toys and pipes used for plumbing made from polyvinyl chloride (PVC). More commonly known as vinyl, PVC was discovered in the 1920s by Waldo Semon, who was attempting to bind rubber to metal. The new material was inexpensive, durable, fire resistant and easily moulded, and is still widely used today.

Not all polymers are formed by addition polymerisation. Some, such as the nylon used to produce mascara brushes, are formed by condensation polymerisation.

Sample problem 9.1

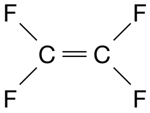
Use structural formulas to show the polymerisation of polypropene (poly-propylene) from the monomer propene.

Solution:

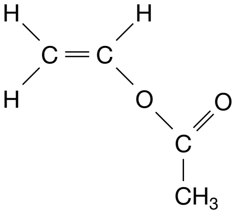


Revision questions

1. The monomer tetrafluoroethene has the structural formula shown at right.
   (a) Use expanded structural formulas to show how you would expect this monomer to polymerise.
   (b) Name the polymer formed.
   (c) Find out the common name for this polymer, and identify its main use.



2. The monomer vinyl acetate has the structural formula shown at right.
   (a) Use expanded structural formulas to show how you would expect this monomer to polymerise.
   (b) Name the polymer formed.
   (c) Find out one use of the polymer formed.



Changing the properties of linear polymers

The versatility of polymers is due to their many different properties. They can be hard or soft, flexible or rigid, transparent or coloured, brittle or able to be stretched. These properties of polymers are affected by the:

- extent of branching
- arrangement of side branches

- crosslinking of polymer chains
- degree of crystallinity
- length of the polymer chains
- addition of plasticisers
- additives.

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Unit 1 **AOS 2** **Topic 4** **Concept 4**

Effects of branching and chain length Summary screen and practice questions

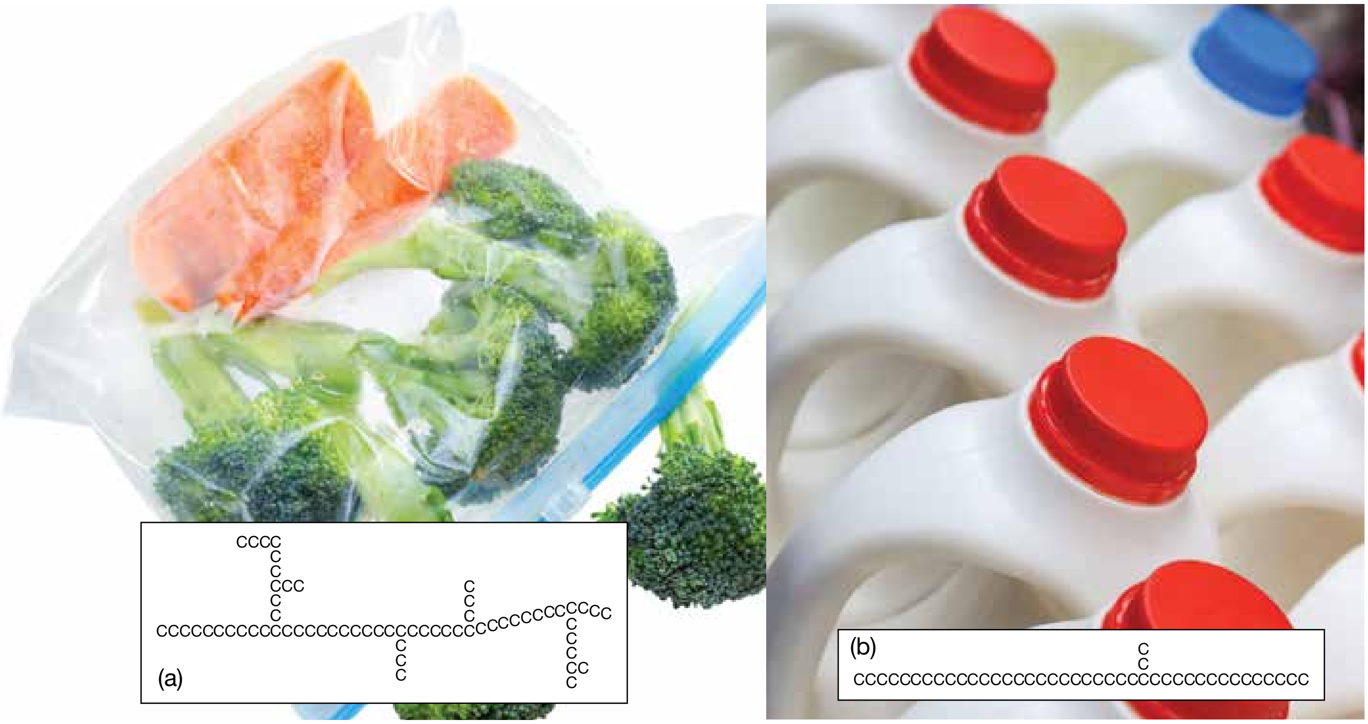
Extent of branching

Ethene can be polymerised to produce both low- and high-density polyethene. Polyethene was discovered in 1933. Scientists had been experimenting with the effect of heat and very high pressures on ethene. It was not until 1939 that full commercial production of polyethene began. The first polyethene produced was called low-density polyethene (LDPE), because its polymer chains support a large number of long side branches, producing a low-density substance (see the figure below). Since the only forces causing these polymer chains to attract each other are dispersion forces, the effect of the branches is to keep the chains apart. Because the attraction becomes weaker as the chains are further apart, the density of the resultant compound is low, and LDPE is soft, flexible and translucent, with a waxy surface that repels water.

In the early 1950s, ethene was polymerised using lower pressure and lower temperatures. The result was a polymer of ethene with very few branches, and any branches that developed were short. Dispersion forces act more effectively on these chains, and this type of polyethene is more rigid, stronger and more opaque than LDPE. It is slightly flexible and also has a waxy surface that repels water. It is called high-density polyethene (HDPE).

Most ethene is extracted from natural gas or naphtha by high-temperature cracking. Cracking is the process where large molecules are broken down into smaller ones.

Polyethene is produced in two main forms. These are low-density polyethene (LDPE) and high-density polyethene (HDPE). Different reaction conditions are used to produce these two variations. The differences in properties are due to the number of branches in the polymer chains and the resulting differences in how the chains can pack together.



The carbon backbone in a portion of a polymer chain in (a) LDPE and (b) HDPE. The extent of the side branching determines the properties of the two forms. LDPE, the first polyethene produced, supports large numbers of long branches to produce a low density substance. Among its uses are plastic bags. HDPE was later developed, with very few and short branches. It is a stronger, more rigid material and can be used to make bottles.

The properties of a polymer can be affected by physical factors such as how well its long chains can pack together. If the chains can pack closely together, the weak dispersion forces between them can act more effectively. Such polymers are usually higher in density and hence tougher and harder than polymers where the chains cannot get as close together.

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Unit 1 **AOS 2** **Topic 4** **Concept 2**

Effects of cross-linking
Summary screen and practice questions

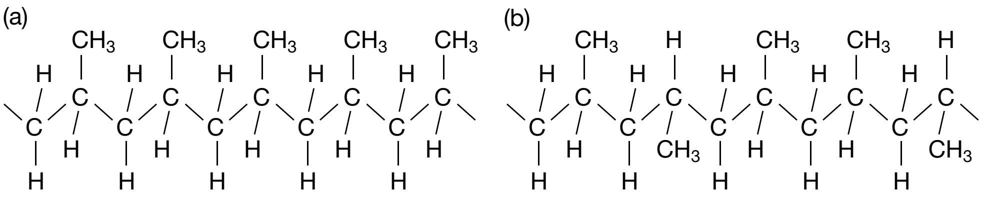
Arrangement of side branches in linear polymers

Linear polymers are those in which the main backbone is unbranched. When one of the hydrogen atoms on ethene, a linear polymer, is replaced by a —CH_3 group the molecule is propene. When propene is polymerised it becomes polypropene; the — CH_3 group is then called a side branch. The way in which side branches are arranged on linear polymers, such as polypropene, shown in the next diagram, can affect the properties of the polymer.

When side branches are arranged on the same side of a linear polymer, the polymer is **isotactic**. Due to the greater effect of dispersion forces, such polymers can pack together closely, producing a substance that has a high density and is rigid and tough with a high softening temperature. Such polymers have many uses. For example, buckets, milk crates, carpets and toilet seats are all made from isotactic polypropene.

When the branches occur at irregular points on both sides of the chain of a linear polymer, the polymer is **atactic**. In these polymers the chains of molecules cannot get as close together and a low-density substance is formed. Atactic polypropene, for example, is soft and waxy with little potential use.

Industrially, increased branching is brought about by the use of higher temperatures and pressures.



Isotactic polypropene (a) has branches on the same side of its chain, while atactic polypropene (b) has branches randomly placed above and below the chain. The different molecular structures produce two substances with very different properties.

Cross-linking of linear polymer chains

Cross-links are covalent bonds that can form between polymer chains. If the number of cross-links is small, an elastomer results. If the number is large, a hard inflexible thermosetting polymer is produced.

Thermoplastics are made from linear polymers where the forces between the chains are mainly weak dispersion forces. If a small amount of **cross-linking** (i.e. bonding) between the chains is introduced, an **elastomer** such as vulcanised rubber is produced. When a large amount of cross-linking is introduced, a rigid polymer results as the atoms are strongly bonded in all three dimensions. These polymers are *thermosetting* and do not melt when heated, but can char at high temperatures.

To make a thermosetting polymer, the long linear chains are produced first. The cross-linking is then brought about either by using heat or by adding a chemical to react between the lateral functional groups linking the chains together.

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Digital document
Experiment 9.1 Cross-linking an addition polymer to make slime
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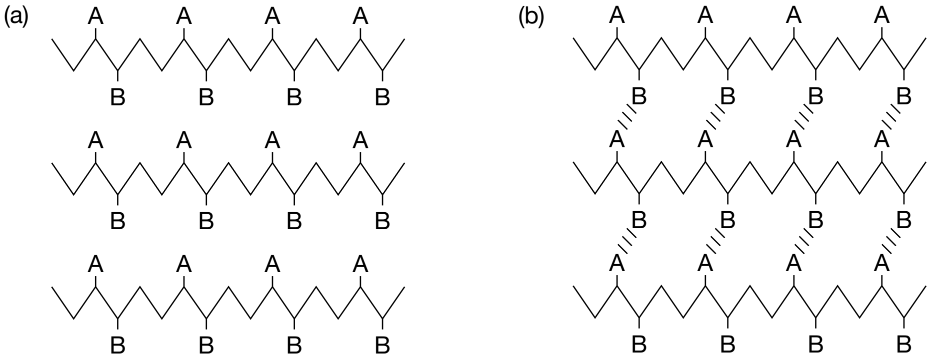


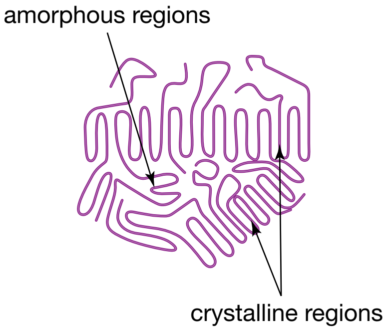
A collection of linear polymers with lateral functional groups A and B is shown at (a). When the A and B groups are linked, a rigid cross-linked polymer (b) results.

For example, a particular article may be made by placing the preliminary linear polymer into a mould and then heating it. Cross-links form and it is a rigid thermosetting article that is removed from the mould when it is opened. Note that thermosetting polymers may still be classed as plastics. This is because, in the early stages of their manufacture, before the cross-linking is introduced, they can still be moulded.

A familiar example of the second method of producing cross-links is when two-part glues are mixed. One tube contains the preliminary linear polymer with the lateral functional groups. The other tube contains the 'hardener'. When mixed, this chemical reacts with the lateral functional groups, bonding them together and linking the chains together with strong covalent cross-links.

Degree of crystallinity

Polymer chains can be arranged in two ways. They can be crystalline, in that they are regularly organised into lines, or randomly packed with no particular order. Polymers can be partially crystalline or almost totally amorphous. The percentage of regularly ordered chains usually ranges between 10% and 80%. The more crystalline a structure is the greater its hardness, tensile strength and opacity. Amorphous polymers are more easily deformed and often transparent. The factors that affect crystallinity include chain length, branching and interchain bonding, as seen previously for LDPE and HDPE. HDPE is more likely to conform to a crystalline structure because of its unbranched carbon chains.



The strength of a polymer is increased if there are more crystalline areas.

Length of polymer chains

As chain length increases, the strength, melting and boiling temperatures also increase. Long chains tend to get tangled and are less likely to slide over each other than smaller chains.

Addition of plasticisers

Plasticisers are major components added to polymers such as PVC during production. They are small molecules which cause the polymer chains to move slightly further apart resulting in a softer and more flexible polymer. There are many different plasticisers but the most commonly used are phthalates.

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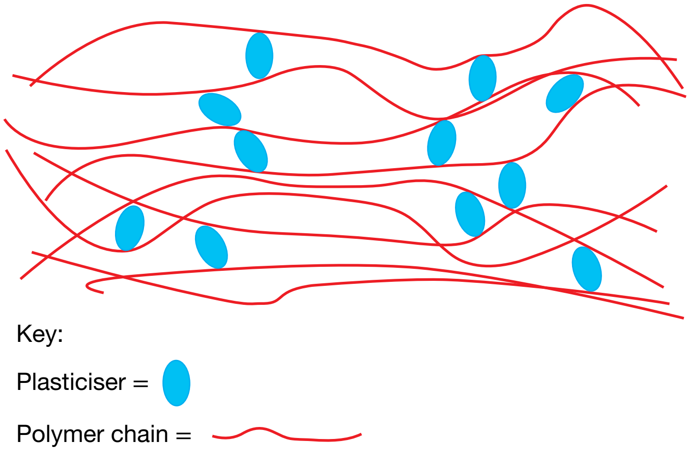
Unit 1
AOS 2
Topic 4
Concept 5

Customised polymers — copolymers, crystalline areas and plasticisers
Summary screen and practice questions



Plasticisers are added to polymers to increase flexibility.

Additives

Additives in polymers can have many different uses. Examples include:

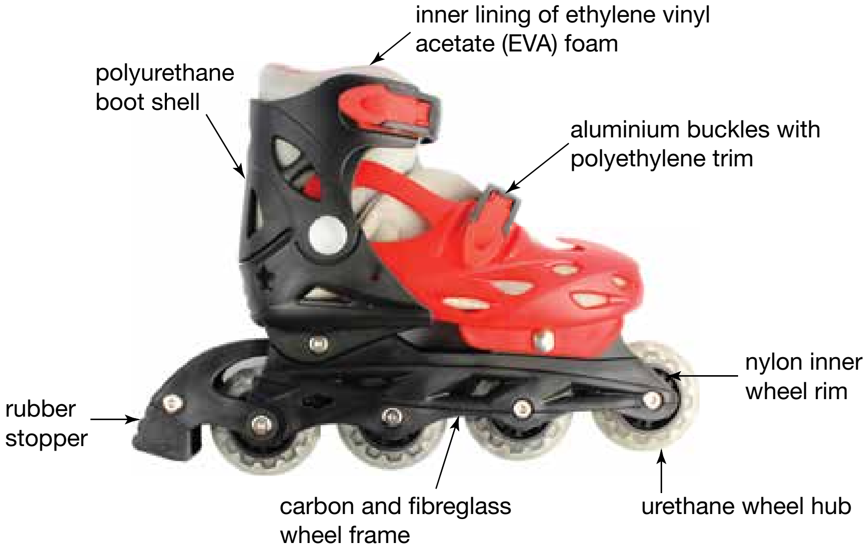
- UV stabilisers, which absorb UV rays to prevent the polymer breaking down
- flame retardants to reduce the tendency of the polymer to burn
- dyes to add colour or provide patterns.

Revision question

3. Explain the difference between an isotactic and an atactic polymer with respect to:
  - **(a) side branching**
  - **(b) intermolecular bonding**
  - **(c) formation**
  - **(d) uses.**

Commonly used polymers

The process of addition polymerisation has been widely applied in industry to produce a host of materials with very different properties. For example, many of the parts of the inline skate below are polymers, each one chosen for its specific properties.



Many items, such as this in-line skate, are fabricated from a diverse range of plastic components. Each of these is chosen for its particular properties.

Addition polymers

Addition polymers are widely used in the packaging industry, as well as featuring in many other products. The uses of the addition polymers polyvinyl chloride (PVC), polytetrafluoroethene (Teflon®) and polystyrene were shown in table 9.1. The uses of two other common addition polymers were shown in table 9.2.

Rubber

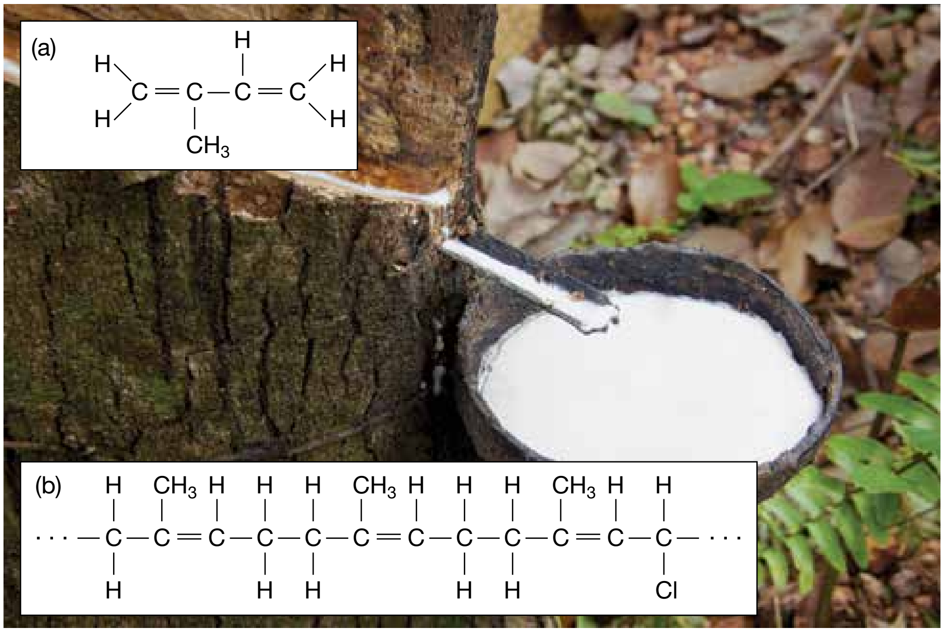
Rubber is an addition polymer that occurs naturally. It is a completely amorphous polymer. The monomer in natural **rubber** is **isoprene** or

2-methylbuta-1,3-diene. Isoprene polymerises to form long chains, as shown in the figure below, and the molecular formula is written as $(C_5H_8)_n$.

This chain is similar to that of polyethene (polyethylene) but with an important difference: rubber still contains double bonds that can be attacked by oxygen and, unlike polyethene, rubber can perish. Another disadvantage of natural rubber is that it is not elastic. When stretched, the long tangled chains straighten out and remain this way as there is little tendency for them to return to their original shape. Natural rubber is also susceptible to temperature changes, becoming very brittle when cold and sticky when hot.

An American inventor, Charles Goodyear, experimented for many years to find a way of countering these tendencies. In 1839, Goodyear accidentally dropped a piece of rubber that had been treated with sulfur onto a hot stove. He later found that this sample had improved elasticity and greater resistance to temperature change. His discovery formed the basis of the process of **vulcanisation**, which is still widely used in the rubber manufacturing industry today to improve the durability and elasticity of natural rubber.

Rubber is an example of a naturally occurring addition polymer. However, it needs to be vulcanised to make it useful. This involves heating the natural rubber with varying amounts of sulfur. The sulfur introduces cross-links between the polymer chains and makes the rubber more elastic and durable.

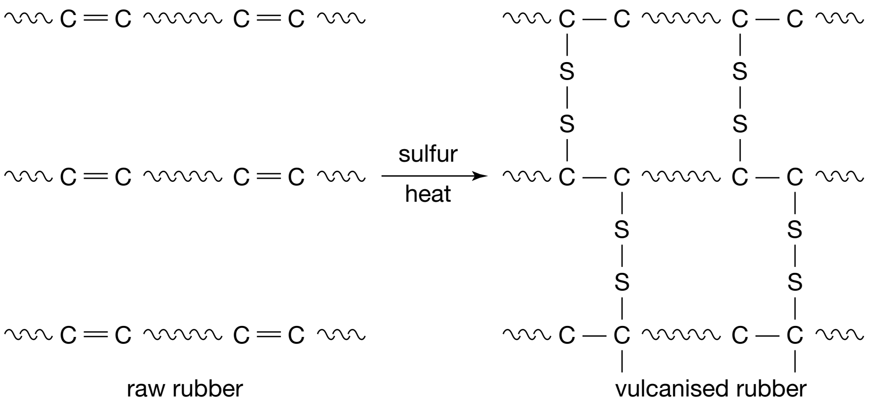


Structural formulas of (a) the monomer isoprene and (b) the polymer for raw (natural) rubber. Natural rubber is produced from latex, a milky white substance that can be harvested from rubber trees by making a cut in the bark and collecting the sap as it runs out. A well-managed plant can yield approximately 1.8 kg of dry crude rubber per year.

Natural rubber is vulcanised in an industrial process where it is mixed with sulfur and heated. The sulfur atoms form cross-links between chains of rubber molecules, reducing the number of double bonds, as shown in the figure on the next page.

When vulcanised rubber is stretched, the sulfur linkages stop the chains slipping past one another and the rubber returns to its original shape when the stretching force is removed. Vulcanised rubber is an example of a cross-linked polymer where the sulfur atoms link straight chains together. Rubber is used for tyres, carpet backing, tyre tubing and the soles of sports shoes.

The linear chains of natural rubber can be cross-linked, using heat and sulfur, to produce a more elastic and durable product.



Polymer selection

A huge range of polymers exist today and are used for many different applications. Their versatility has made them one of the most useful classes of substances that we rely on in today's society. This versatility can be attributed to the many different ways that they can be modified.

In 1963, the Nobel Prize winning chemist Giulio Natta said 'A chemist setting out to build a giant molecule is in the same position as an architect designing a building. He has a number of building blocks of certain shapes and sizes, and his task is to put them together in a structure to serve a particular purpose.' Today we can take this statement to mean that, due to their versatility, polymers can be produced for almost any imagined purpose.

Already in this chapter we have seen a number of ways that this can be done. We have seen that small chemical changes can be made to monomers, resulting in polymers with different features. Monomers such as ethene, vinyl chloride, propene, tetrafluoroethene and styrene illustrate this point. We have also seen that, even with the same monomer, a change in how the monomers join together results in significant differences in the polymer produced. The two main forms of polyethene (LDPE and HDPE) are examples of this. With cross-linked polymers, the degree of cross-linking can be varied to produce polymers with varying degrees of elasticity and, when the number of cross-links is large, polymers that are firm and rigid.

Additional ways also exist to modify polymers. Substances may be added to further modify the way a polymer behaves (these are often called 'fillers'). Vinyl is an immensely versatile polymer because of this. Such chemicals may also include filters and stabilisers to slow the environmental degradation of articles made from such polymers.

Finally, a range of polymers exist based on silicon rather than carbon. These are called 'silicones'. They are generally much more expensive and are used in specialised applications.

The most common families of polymers can be produced through controlled polymerisation from natural ingredients that are relatively easily obtained, such as wood, cotton, petrochemicals, limestone, water, salt, fluorspar and sulfur. Each family boasts several members, and the proliferation of synthetic materials has resulted in at least 50 commercial varieties.

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Waste and management

Advantages and disadvantages of using polymers

A quick look around the room will show the versatility of polymers and how dependent our society is on them. Plastics contribute to transport, construction, entertainment, packaging, clothing and other everyday items such as furniture, toys, bags and pens. There are many advantages of using plastics, but advantages such as durability are also a disadvantage when it comes to recyclability.

Advantages include:

- can be moulded into any shape
- don't corrode
- are cheaper raw materials than alternative materials
- are less dense so lighter
- are good insulators of heat (home insulation)
- are good insulators of electricity (wire covers)
- are water resistant
- some are recyclable
- most are chemically inert
- are unbreakable
- are relatively cheap to produce
- are useful for surfaces for construction purposes
- reduce the need to use limited natural materials such as timber.

Disadvantages include:

- use up a nonrenewable/finite resource
- take a long time to decompose
- take up space in landfill
- cause pollution in oceans
- produce toxic gases on combustion
- some cause health problems.

Recycling

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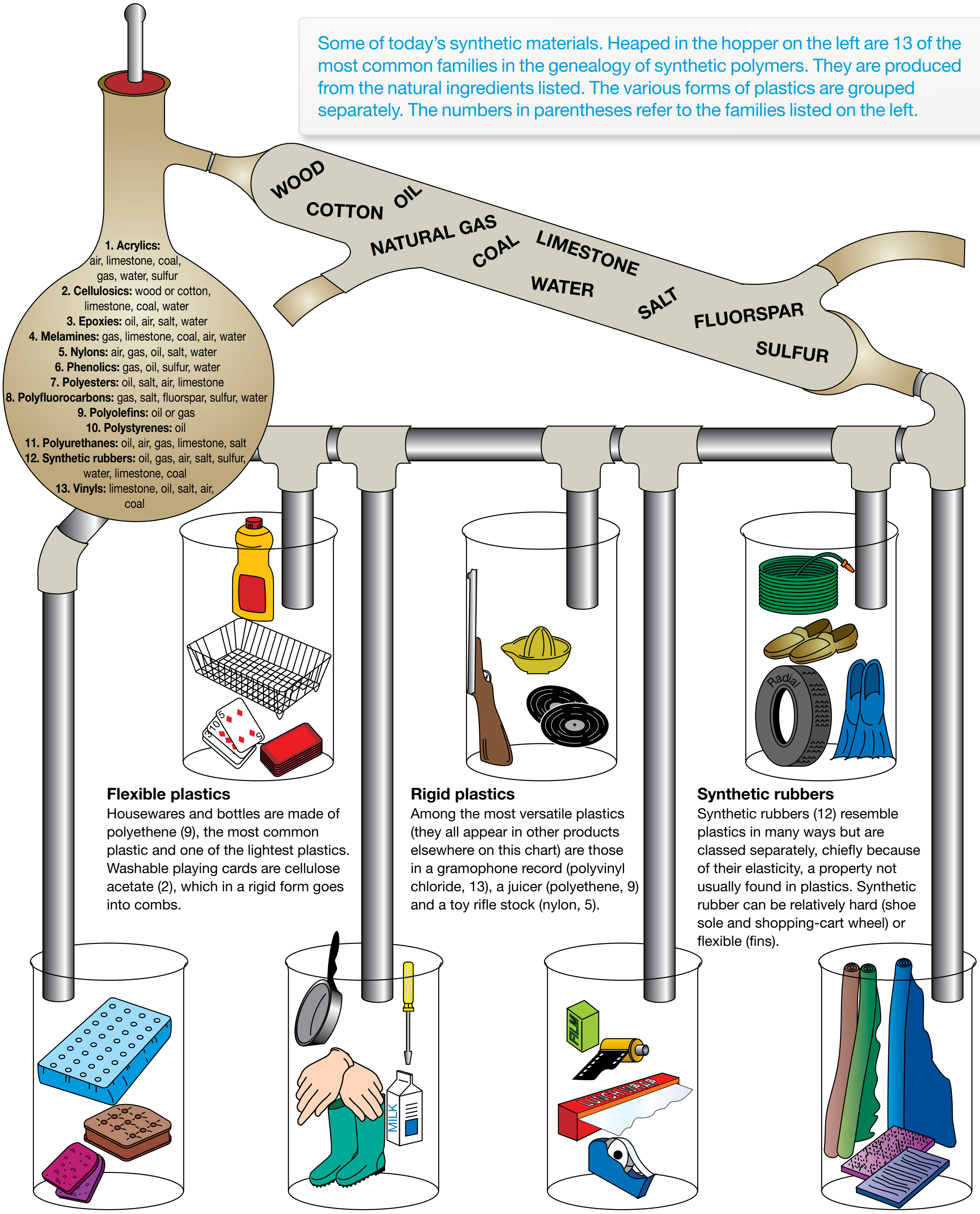
Unit 1
AOS 2
Topic 4
Concept 3

Recycling polymers
Summary screen and practice questions

Most plastics are produced from crude oil, coal or gas. The production of plastics therefore contributes to the depletion of finite resources. Their bulk adds to the cost of waste collection and disposal and, because many of them are not biodegradable, they have become a visible part of our environmental litter. For these reasons the recycling of plastics is important.

Amazingly durable and inexpensive to produce, plastics changed the way products, particularly food and beverages, were packaged. They offered superior hygiene, and were shatterproof, lightweight and cost-effective to transport. The explosion in disposable plastic products, however, has led to a significant litter problem. The very qualities that make plastics so useful can also make them an eyesore.

Some of today's synthetic materials. Heaped in the hopper on the left are 13 of the most common families in the genealogy of synthetic polymers. They are produced from the natural ingredients listed. The various forms of plastics are grouped separately. The numbers in parentheses refer to the families listed on the left.



1. Acrylics: air, limestone, coal, gas, water, sulfur
2. Cellulosics: wood or cotton, limestone, coal, water
3. Epoxies: oil, air, salt, water
4. Melamines: gas, limestone, coal, air, water
5. Nylons: air, gas, oil, salt, water
6. Phenolics: gas, oil, sulfur, water
7. Polyesters: oil, salt, air, limestone
8. Polyfluorocarbons: gas, salt, fluorspar, sulfur, water
9. Polyolefins: oil or gas
10. Polystyrenes: oil
11. Polyurethanes: oil, air, gas, limestone, salt
12. Synthetic rubbers: oil, gas, air, salt, sulfur, water, limestone, coal
13. Vinyls: limestone, oil, salt, air, coal

Flexible plastics

Housewares and bottles are made of polyethene (9), the most common plastic and one of the lightest plastics. Washable playing cards are cellulose acetate (2), which in a rigid form goes into combs.

Rigid plastics

Among the most versatile plastics (they all appear in other products elsewhere on this chart) are those in a gramophone record (polyvinyl chloride, 13), a juicer (polyethene, 9) and a toy rifle stock (nylon, 5).

Synthetic rubbers

Synthetic rubbers (12) resemble plastics in many ways but are classed separately, chiefly because of their elasticity, a property not usually found in plastics. Synthetic rubber can be relatively hard (shoe sole and shopping-cart wheel) or flexible (fins).

Foams

The mattress is a synthetic rubber foam, and the other items are plastic foams, made by allowing gas to bubble through the liquid plastic during its forming. The sponges are polyurethane (11); the seat cushions are vinyl (13).

Coatings

Synthetics are widely used for protective covering. Paint bases and glove coatings are synthetic rubber (12); other items are protected by plastics from various families: (9), boots, carton; (13), tool handle; (8), non-stick frying pan.

Plastic films

Various vinyls (13) are used for wrappings. Other wrappings include polyethene (9) and cellulose acetate (2), used for camera film and tape.

Synthetic fibres

The fibres used in synthetic textiles are not thought of as plastics, although they are made the same way, by polymerisation of simple materials.

With the joint aim of reducing the impact of the use of plastics on the environment, the plastics industry has implemented a coding system to enable plastics to be identified and sorted into their various types for recycling. Plastic products are now coded with one of seven symbols, as shown in table 9.3.

Many plastics, including LDPE and HDPE, can be recycled. Once the plastics have been separated, they are shredded into flakes, and washed to remove contaminants. The flakes are melted and extruded then chopped again into flakes, which can then be remoulded into new products.

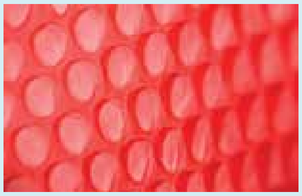
Recycled HDPE is used for compost bins, detergent bottles, pipes, plumbing fittings, household bags and irrigation pipes. Recycled LDPE is used for carry bags, packaging film and household bags.

Plastics can be recycled but first the type of plastic must be identified according to the codes shown. Once sorted, the plastic groups are shredded into flakes so that they can be washed, melted and remoulded.

TABLE 9.3 Recyclable plastics

Symbol and name	Examples
♳ polyethene terephthalate PETE	softdrink and water bottles 
♴ high-density polyethene HDPE	milk, juice and water bottles 

(continued)

Symbol and name	Examples
♵ PVC V	juice bottles, detergent bottles, credit cards, PVC piping 
♶ low-density polyethene LDPE	frozen food bags, bin liners, squeezable bottles, flexible container lids, cling film, bubble wrap 
♷ polypropene PP	reusable microwave containers, margarine tubs, kitchen ware, yoghurt containers, disposable cups and plates 
♸ polystyrene PS	egg cartons, packing 'peanuts', disposable cups, plates, trays and cutlery, disposable takeaway containers 
♹ other OTHER	beverage bottles, baby milk bottles, electronic casing 

Chapter review

studyon
Unit 1
AOS 2
Topic 4
Polymers
Summary screens and practice questions

Summary

- A polymer is a long chain of units called monomers that are linked together in a process called polymerisation. A copolymer is formed when two different monomers are used to make the polymer.
- Natural polymers include wool, cotton and hair. Synthetic polymers are commonly called plastics and have two forms.
  - Thermoplastic (or thermosoftening) materials soften when heated and can be remoulded. An early example is celluloid.
  - Thermosetting materials do not melt when heated. An early example is the first synthetic polymer, Bakelite.
- Polymers are usually made by either addition or condensation polymerisation.
  - Addition polymers are formed when an addition reaction causes monomers containing carbon–carbon double bonds to link together.
- Ethene polymerises to form polyethene (also known as polyethylene).
  - Ethene can be modified by substituting different functional groups for hydrogen atoms to produce other monomers.
  - Monomers are styrene, vinyl chloride, tetrafluorene and propene. These can be polymerised to produce polymers with different properties and many uses.
- The structure of a linear polymer determines its properties.
  - Extent of branching: Low-density polyethene (LDPE) and high-density polyethene (HDPE) are two important polymers produced from ethene. The two substances have very different properties due to the different structures of their polymer chains.
    - LDPE chains support a large number of long side branches, which keep the chains apart. As the attraction by dispersion forces is weaker when the chains are further apart, LDPE is a soft, flexible and translucent material with a waxy surface that can be used for packaging, flexible industrial piping, and wire and cable insulation sheathing.
    - HDPE is produced using lower pressure and temperatures than LDPE and the resulting chains have very few and short side branches, which are held closer together by dispersion forces, making the material less flexible, stronger and opaque. HDPE is used to make tough plastic bags, kitchen containers, car parts, rigid pipe and fibres.
  - Arrangement of side branches: The arrangement of side branches can result in either an isotactic polymer (branches on the same side) or an atactic polymer (branches on both sides), producing a high- and low-density substance respectively. This is because the presence of branches pushes the chains of polymers away from each other and acts against the dispersion forces that hold them together.
  - Cross-linking of polymer chains: The presence of cross-linking, or bonding between the chains, affects the elasticity and rigidity of the polymer. A small amount of cross-linking produces an elastomer, which is relatively elastic, and a large amount produces a rigid polymer, because atoms are strongly bonded in all three dimensions.
  - Degree of crystallinity: More regions of regular arrangement increase strength.
  - Length of the polymer chains: The longer the chain, the stronger the polymer.
  - Addition of plasticisers: Plasticisers are small molecules in polymers that make them more flexible.
  - Additives: These contribute to the appearance and stability of the polymer.
- Rubber is a natural addition polymer formed from the monomer isoprene. It can be treated with sulfur and heat, in a process known as vulcanisation, to introduce cross-links and make an elastomer that is more elastic and resistant to cold and heat.
- There are many advantages of using plastics but some of the advantages, such as their durability, are also a disadvantage when it comes to recyclability.
- Many plastics can be recycled. Discarded plastics can be sorted according to standard recycling codes. The different groups of plastics can then be shredded, washed and melted down to produce chips to be remoulded.

Multiple choice questions

1. Which one of the following is a natural polymer?
A Cellulose
B Polypropene
C Teflon
D Kevlar

2. Large molecules can be built up by the combination of smaller molecules. These smaller molecules are called:
A polymers
B isomers
C monomers
D allotropes.

3. Addition polymerisation involves monomers that:
- **A** are polar
- **B** contain a multiple bond
- **C** decompose easily
- **D** contain hydrogen and a functional group.

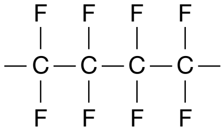
4. Which of the following best describes polymerisation?
- **A** Small molecules decompose to form a new substance.
- **B** Two or more chemicals react together.
- **C** Small molecules react to form very long molecules.
- **D** Small molecules react to form a thicker substance.

5. Which one of the following is *not* a characteristic of thermoplastic?
- **A** No cross-links between chains
- **B** Can be moulded
- **C** Chars when heated
- **D** Has weak dispersion forces between chains

6. Rubber is a polymer made from the monomer:
- **A** styrene
- **B** isoprene
- **C** vulcene
- **D** urea.

7. What structural characteristics are present in the monomers used to produce addition polymers?
- **A** They must contain only carbon and hydrogen.
- **B** They must all be the same.
- **C** They must contain a carbon–carbon double bond.
- **D** They must have a low molecular mass.

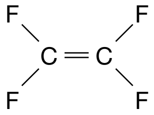
8. Teflon is made from the polymer polytetrafluoroethylene (PTFE). It has the structure:



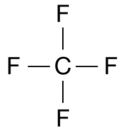
$$-\underset{F}{\overset{F}{C}}-\underset{F}{\overset{F}{C}}-\underset{F}{\overset{F}{C}}-\underset{F}{\overset{F}{C}}-$$

The monomer for this polymer is:

A

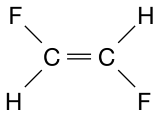


C

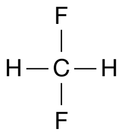


$$F-\underset{F}{\overset{F}{C}}-F$$

B

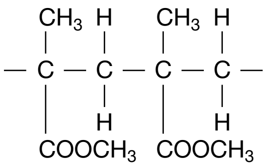


D



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

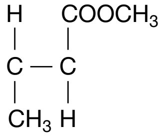
9. Polymerisation of methyl methacrylate produces 'perspex'. The structure of perspex is:



$$-\underset{COOCH_3}{\overset{CH_3}{C}}-\underset{H}{\overset{H}{C}}-\underset{COOCH_3}{\overset{CH_3}{C}}-\underset{H}{\overset{H}{C}}-$$

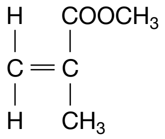
The structure of methyl methacrylate is:

A



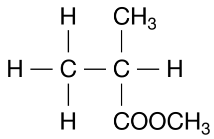
$$\underset{CH_3}{\overset{H}{C}}-\underset{H}{\overset{COOCH_3}{C}}$$

C



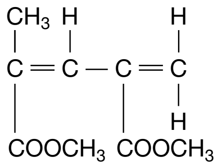
$$\underset{H}{\overset{H}{C}}=\underset{CH_3}{\overset{COOCH_3}{C}}$$

B



$$H-\underset{H}{\overset{H}{C}}-\underset{COOCH_3}{\overset{CH_3}{C}}-H$$

D

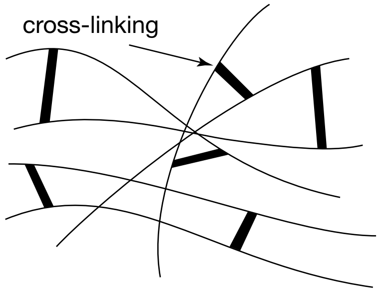


$$\underset{COOCH_3}{\overset{CH_3}{C}}=\overset{H}{C}-\underset{COOCH_3}{C}=\underset{H}{\overset{H}{C}}$$

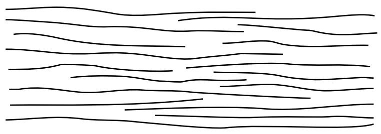
10. When the chain length of a polymer increases:
- **A** the melting temperature is lowered
- **B** the substance becomes more viscous
- **C** solubility in non-polar solvents increases
- **D** density decreases.

11. Which of the following diagrams depicts a thermosetting polymer?

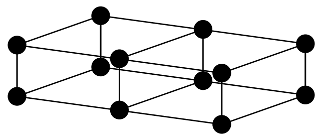
A



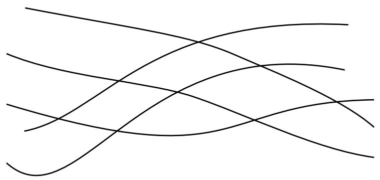
B



C



D



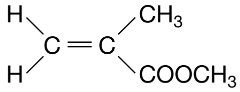
12. The difference in strength between low-density polyethene (LDPE) and high-density polyethene (HDPE) is mostly due to:
- **A** the degree of cross-linking
- **B** the degree of side branching
- **C** a difference between the two in relative molecular mass
- **D** the orientation of hydrogen atoms along the carbon backbone.

13. The vulcanisation of natural rubber:
   - **A** involves a condensation reaction
   - **B** increases side chain branches in the polymer
   - **C** involves cross-linking adjacent polymer chains
   - **D** increases flexibility.

Review questions

Addition polymerisation

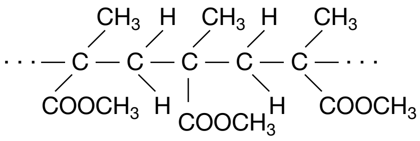
1. What are monomers and polymers?
2. Distinguish between thermosetting and thermoplastic polymers and give three examples of each.
3. Perspex is an addition polymer that has the appearance of glass. It is made from the monomer:



   (a) Draw part of the polymer of perspex.
   (b) In what situations would the use of perspex be superior to the use of glass?
   (c) In what situations would the use of glass be superior to the use of perspex?
4. Give the structural formula for natural rubber and the monomer from which it is made.
5. What is vulcanised rubber?
6. Explain why vulcanised rubber is an elastomer, whereas natural rubber is not.
7. Briefly explain how the rigidity of a polymer is related to the amount of cross-linking.

Using polymers

8. What are cross-linked polymers?
9. A polymer used in the manufacture of artificial eyes is shown in the figure below.



   (a) Draw the structure of the monomer from which the polymer is derived.
   (b) Describe the forces that would hold these chains together.
10. Identify the two different forms of polypropene and describe how the degree of branching in each polymer affects its final use.
11. Transparent film for packaging can be made from LDPE but would not be made from HDPE, which is more opaque. Explain this difference in terms of the molecular structure of the two forms of polyethene.
12. LDPE is soft and flexible whereas HDPE is stronger and more rigid. Explain.
13. Describe how the vulcanisation of rubber makes a more usable material.
14. List three advantages and three disadvantages of using plastics.

Bonding types

15. Polymers are sometimes called giant molecules because they contain so many atoms. As polymer chains are twisted together they form fibres. Diamond and graphite are also classed as giant molecules. A list of the types of bonds studied in this course follows. These bonds may exist either within a chain or molecule (intramolecular bonds) or between chains or molecules (intermolecular bonds). Copy and complete the table that follows, using the bond types.
Bond types: covalent, dispersion forces, hydrogen bonding, disulfide

	Bond type	
Polymer	**Intramolecular**	**Intermolecular**
diamond		
graphite		
polyethene		
natural rubber		
vulcanised rubber		

Recycling

16. Explain why the recycling of plastics is desirable.
17. Describe why the coding system used for plastics is beneficial.

studyon
Unit 1
AOS 2
Topic 4
Polymers
Sit topic test

Exam practice questions

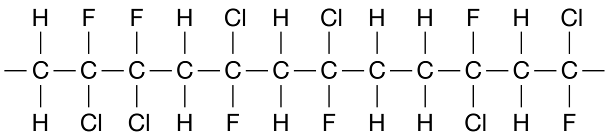
In a chemistry examination you will be required to answer a number of short and extended response questions.

Extended response questions

1. (a) Draw the full structural formula of the unsaturated molecule represented by the semi-structural formula CHCl CHF. **1 mark**

(b) Draw a section of the polymer produced when the monomer from part (a) undergoes polymerisation. **2 marks**

(c) A section of an addition polymer is shown below.



$$\begin{array}{ccccccccccccccccccccccccc}
 & \text{H} & & \text{F} & & \text{F} & & \text{H} & & \text{Cl} & & \text{H} & & \text{Cl} & & \text{H} & & \text{H} & & \text{F} & & \text{H} & & \text{Cl} & \\
 & | & & | & & | & & | & & | & & | & & | & & | & & | & & | & & | & & | & \\
- & \text{C} & - & \text{C} & - & \text{C} & - & \text{C} & - & \text{C} & - & \text{C} & - & \text{C} & - & \text{C} & - & \text{C} & - & \text{C} & - & \text{C} & - & \text{C} & - \\
 & | & & | & & | & & | & & | & & | & & | & & | & & | & & | & & | & & | & \\
 & \text{H} & & \text{Cl} & & \text{Cl} & & \text{H} & & \text{F} & & \text{H} & & \text{F} & & \text{H} & & \text{H} & & \text{Cl} & & \text{H} & & \text{F} &
\end{array}$$

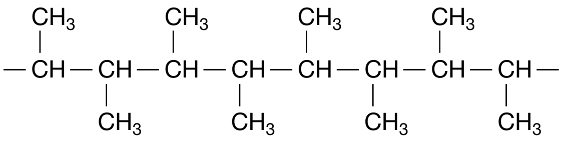
Draw the structural formula of the monomer that it is produced from. **2 marks**

2. (a) Draw the structural formulas of the following molecules.

(i) but-2-ene

(ii) but-1-ene **2 marks**

(b) Examine the section of polymer that is shown below.



$$\begin{array}{ccccccccccccccccc}
 & CH_3 & & & & CH_3 & & & & CH_3 & & & & CH_3 & & & \\
 & | & & & & | & & & & | & & & & | & & & \\
- & CH & - & CH & - & CH & - & CH & - & CH & - & CH & - & CH & - & CH & - \\
 & & & | & & & & | & & & & | & & & & | & \\
 & & & CH_3 & & & & CH_3 & & & & CH_3 & & & & CH_3 &
\end{array}$$

From which of the substances in part (a) could this polymer have been made? **1 mark**

(c) Draw a section of the polymer chain that could be made from the other molecule (i.e. the one *not* used for part (b)). **2 marks**

CHAPTER

10 Investigating and reporting

YOU WILL EXAMINE:

- guidelines for undertaking a research investigation and presentation of findings
- designing and performing practical exercises
- safety requirements
- recording information and reporting
- how to use significant figures and scientific notation
- strategies needed for effective study and exam preparation.

> All truths are easy to understand once they are discovered; the point is to discover them.
>
> *Galileo Galilei*

Performing practical investigations and writing clear and concise reports are essential skills in chemistry. In this chapter, you will find guidelines for carrying out practical tasks safely and preparing thorough reports using appropriate scientific conventions.

Research investigations

A research investigation is an exploration of a scientific question or questions and presentation of findings. In outcome 3 you have the opportunity to find out more about an aspect of the chemistry of materials that you have studied in Unit 1. This will involve researching for information, evaluating it, interpreting it and presenting it in a cohesive, accurate and detailed report.

Planning

1. Decide whether you are going to record your information electronically or in a notebook.
2. This is a great opportunity to investigate a topic that interests you. You can also select from the suggested questions provided in each of the options. Alternatively, after a discussion with your teacher, you may prefer to respond to questions of your own choosing or have a combination of both. Typical questions could begin with how, why, what, when, who or where. Record the questions that you are going to investigate. These questions may change as you start researching.
3. Once the key questions have been decided, it is time to commence the research using the abundance of available sources. Switching on the computer and looking on the internet is one way of doing research; other sources include books, newspapers, scientific magazines, podcasts, videos, people with expertise in the relevant field and the many organisations that can provide valuable material. It is wise to use a variety of resources so that the information is not biased or limited by a narrow perspective. Your teacher and librarians at school and public libraries can point you in the right direction.
4. When doing web searches, use key words. Make use of an advanced search to find a key phrase, recent document or a particular file type such as ppt or doc. Assess websites carefully and identify the point of origin. Government websites, and those from educational institutions and established organisations are usually reliable; when using other websites, it is a good idea to look for authors and dates to see if the information is verifiable and current. Evaluate your information by comparing it with other websites or other sources.
5. Record your information in an organised way. Prepare a separate page for each question and set up a table with headings such as:
   - Date accessed
   - Information (to be written *in your own words*). Dot points are adequate at this stage.
   - Reference: List all of the sources of your information; use correct bibliographical procedure and record this as you are writing down information.
   - Comment: For example, useful/not useful, detailed/little information.

   Once you have recorded information from a few sources, read your notes carefully and write it as if you were explaining it to another person, using your own words and sentence structure.
6. Analyse and interpret your information. Remember to note any gaps in the information or the presence of competing opinions, which is also useful. You can organise the information using highlighters to collate different groups of ideas or sections. Write a rough plan showing the flow of the report.
7. Decide how you will present your report. Do not plagiarise the material; this means that it is not acceptable to just change a few words or reorganise the sentence. If you are quoting another person, make this clear by using quotation marks, and remember to cite the author.

Presentation of findings

There are various ways of presenting a report, including written format, poster, slideshow, video, animation and oral presentation. Remember to include your name on the front page on a written report. Your report needs a title that is not too general but gives a good idea of what it is about. For example, 'The relationship between structure and properties of metals' is preferable to 'Metals'. The title and introductory paragraph can be written at the end of the drafting process because the final content will be established at this point. The introduction describes the focus of your investigations, the structure of the report and possibly the reasons for your explorations. Some definitions and the scope of the research may be useful early in your report.

Tables and images can provide added interest; however, they need to be relevant and referenced. Some images are copyright; a search through Creative Commons or using 'usage rights' in an advanced Google search engine will help you access copyright-free images.

Use correct and relevant terminology and explain all scientific terms; look up definitions while you are taking notes.

Divide the report into paragraphs and, where necessary, use subheadings to separate ideas. Take care with punctuation, spelling and grammar.

Finally, the conclusion should summarise and tie together the findings of the report and reflect back on the introductory paragraph. It might provide the opportunity to suggest areas for further investigation, but it should not include any new information. A bibliography, written in the correct format, is the way to end the report. There are many helpful online sites for writing a bibliography correctly.

Final checklist

Have you:

- understood your topic sufficiently
- used your own words
- included an introduction with an outline of what the topic is about
- communicated the information clearly, logically and accurately
- used correct chemical units and terminology
- explained important scientific terms
- referred to relevant chemical theories or models
- checked whether your information is reliable and trustworthy
- considered related ethical, environmental and social impacts where necessary
- referenced all diagrams and tables
- presented the report in an appropriate and succinct form that would be understood by a fellow student at this level
- written a concluding paragraph that relates back to the introduction
- provided a detailed bibliography?

Practical investigations

In Unit 2, outcome 3 you can use your developing experimental skills to investigate a question related to water quality. You will be doing a practical investigation that uses laboratory or fieldwork to respond to a question. A practical investigation involves considerable planning, expertise in working scientifically to design a method, record results, analyse findings and communicate conclusions. You will be familiar with the usual format of a practical report (see table 10.1).

TABLE 10.1 Format of a practical report

Title	short heading related to the investigation
Aim	what you are investigating and why
Hypothesis	what you think will happen; often an 'if/then' statement is used
Method	outline of the steps in your procedure
Materials	list of chemicals and equipment used
Results	includes the data collected, usually in a table
Discussion	describes whether your data supports the hypothesis
Conclusion	relates to the aim and summarises the findings of the investigation

Procedure

1. Start a logbook; scraps of paper can be easily lost. This could be an electronic logbook. Record everything in your logbook, including dates, questions, changes in method, observations and results.
2. In consultation with your teacher, decide on a question to investigate.
3. Consider ethical and environmental implications of your question.
4. Write a short and relevant title. This may change as the investigation progresses.
5. Record your hypothesis, which is an educated guess about what you think will happen.
6. Review the theory related to your topic.
7. Plan your method. Which variables will you keep constant? What quantity will you be measuring? Will it be a fair test? Can you complete it in the time given? Are the chemicals and materials accessible?
8. List the steps in the method. These need to be explicit enough so that someone else could repeat your experiment.
9. Consider safety implications. Refer to the section on safety on the next page.
10. Prepare a materials list. Be specific about quantities and equipment required, including size of beakers, measuring cylinders etc.
11. Do an initial rough trial, and modify the method if necessary.
12. Prepare a results table with appropriate headings and units.
13. Do your experiment, and repeat it if possible to reduce random errors.
14. Record detailed observations and results as accurately as possible. Photos can be useful here.
15. Complete calculations to the correct number of significant figures. Use scientific notation as required.

Writing your report

1. Choose how you will present your report, as this will affect which of the following points are relevant.
2. Include your name.
3. Check with your teacher whether you need an abstract, which is a *brief* summary of the question examined, method, general results and findings.
4. Record your final title, hypothesis, materials, method and safety assessment.
5. Include your table of results, photos or diagrams. Include fully labelled graphs where applicable.
6. The discussion is particularly important. You do not need to repeat your results. Identify any trends and relate them to the theory. Comment on any unexpected results. Suggest any sources of error and possible modifications to the experiment. Discuss difficulties and suggest future investigations. Check that you have used correct scientific terminology and conventions.
7. In the conclusion, state whether the hypothesis is supported. The conclusion should relate to the aim and include no new information.
8. Check punctuation, spelling and grammar.

9. Include any references used and acknowledge assistance received from external sources.
10. Refer to the criteria sheet to ensure that you have addressed every point thoroughly.

General safety rules

Some general safety precautions will help to ensure that you and others will not be injured in the laboratory.

- Wear protective clothing. This might include laboratory coat, safety glasses and gloves.
- Be aware of the position of safety equipment such as fire blanket, fire extinguisher and safety shower.
- Ask if you are unsure of how to operate equipment or how to use apparatus.
- Read labels carefully to confirm contents and concentration of chemicals.
- Clean and return all equipment to the correct places.
- Read instructions carefully before commencing an experiment.
- Prepare a risk assessment for required chemicals and equipment.

Risk assessments

Risk assessment is a procedure for identifying hazardous chemicals, what the risks are and how to work safely with them. Table 10.2 lists the usual requirements for a written risk assessment.

TABLE 10.2 Requirements for a written risk assessment

Summary of method	brief list of steps indicating how the chemicals and equipment will be used
Chemicals used	list of chemicals used and their concentrations
Chemical hazards	details of the hazards caused by the chemicals. These can be found by looking at the SDS (safety data sheets), which can be accessed online.
Risk control methods	precautions taken to limit risks. This will depend on the chemicals involved.
Equipment	list of materials used in the experiment
Equipment risk	list of hazards associated with the equipment
Risk control measures	how the risks associated with the equipment will be minimised

TABLE 10.3 Sample risk assessment

Risk assessment record for chemistry experiments					
Assessor (name)	Bonny Bunsen	Date of activity	10/02/2016	Class/Year level	11
Title and outline of experiment	Flame test A wire inoculating loop is placed in HCl, and then dipped into the chemical to be tested. The loop is then held in the Bunsen flame.				

Chemical	Amount and concentration	Hazards — risk phrases	Controls — safety phrases
HCl, hydrochloric acid	20 mL, 2M	Harmful if swallowed Toxic by inhalation Causes severe burns Risk of serious damage to eyes Skin contact may produce health damage.	Do not breathe gas/fumes/vapour. Avoid contact with eyes. Wear suitable protective clothing.

(continued)

Chemical	Amount and concentration	Hazards — risk phrases	Controls — safety phrases
$SrNO_3$, strontium nitrate	solid	Risk of explosion by shock, friction, fire or other sources of ignition Contact with combustible material may cause fire. Irritating to skin and eyes Moderately toxic	Keep away from combustible material. Do not breathe dust. Avoid contact with skin. Wear eye/face protection. Wear gloves.
$CuCl_2$, copper chloride	solid	Toxic if swallowed Skin and eye irritant Risk of serious damage to eyes	Do not breathe dust. Avoid contact with eyes. Wear suitable protective clothing. Wear gloves.
$BaCl_2$, barium chloride	solid	Harmful by inhalation Toxic if swallowed	Do not breathe dust. Wear suitable protective clothing. In case of insufficient ventilation, wear suitable respiratory equipment.
NaCl, sodium chloride	solid	Not hazardous	
Risk control for all chemical experiments	• Small quantities used • Safety goggles, gloves, wash hands after use, laboratory coats and closed shoes • Chemical labels, instructions for mixing and disposing of chemicals must be followed. • Obey Science Safety Rules.		
Disposal	Very little waste produced Contaminated acid can be disposed of via sink with plenty of water.		
Other risk controls			
Assessor (signed)	Bonny Bunsen	Date of record	10/2/2016

Revision questions

1. **Draw a plan of your laboratory and label the positions of safety equipment.**
2. **Describe the equipment hazards involved in heating a beaker of water.**
3. **(a) Identify and record the hazards involved when using 2.0M sodium hydroxide solution.**
   (b) Suggest appropriate safety precautions.

Working with figures

Scientific notation

When working with the sizes of atoms, we are dealing with very, very small numbers and yet, when working with numbers of atoms, we need to use very, very large numbers. For example, depending on how it is calculated, there are 33 500 000 000 000 000 000 water molecules in a drop of water. It is really not convenient to write these figures with large numbers of zeroes before or after the decimal point. Using scientific notation this number is written as 3.35×10^{19} water molecules.

There are two parts to scientific notation; the first part is the first digit of the number, and the second part is a decimal point followed by the remaining digits. This is multiplied by 10 raised to the power that returns the number to its original value.

Sample problem 10.1

Write 37 769 using scientific notation.

Solution: 37 769 = 3.7769 × 10 000 = 3.7769×10^4

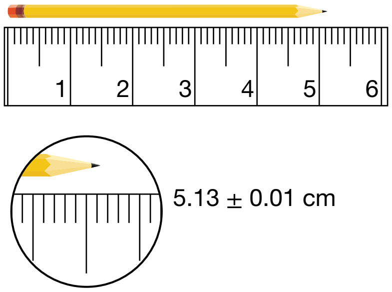
Remember that the power of 10 is positive if the number is greater than 1 and negative if the number is less than 1. On a calculator, the ^ key is used to raise 10 to a particular power. Scientific notation is also called standard form.

Revision questions

4. **Express the following measurements using scientific notation:**
 (a) **A red blood cell is about 0.000 008 m across.**
 (b) **A flea is about 0.0013 m long.**
 (c) **The moon is 384 400 000 metres from the Earth.**

5. **Convert each of the following numbers to scientific notation.**
 (a) **6942** **(b)** **248** **(c)** **0.000 49** **(d)** **3.17** **(e)** **0.0082** **(f)** **64 500**

6. **Convert each of the following numbers from scientific notation to an ordinary number.**
 (a) 7.21×10^{-4}
 (b) 5.739×10^{5}
 (c) 1.552×10^{1}
 (d) 7.112×10^{0}
 (e) 4.71×10^{2}
 (f) 2.577×10^{-6}
 (g) 9.179×10^{-5}



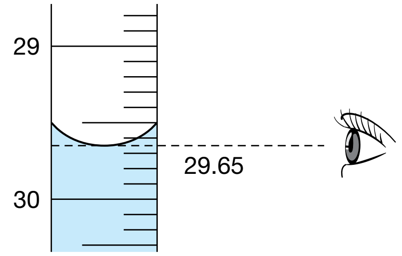
Significant figures

Every measurement that you make has an inherent level of uncertainty. This is due in part to the limitations of the measuring equipment itself and in part to user interpretation. Significant figures are one means of recognising these limitations. We say that significant figures indicate the level of precision in a measurement. When reading a scale, read it to $\frac{1}{10}$ of the smallest division in the scale.

Read the scale to $\frac{1}{10}$ of the smallest division on the scale.

Sample problem 10.2

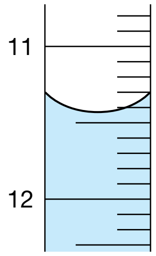
What is the reading on this burette?



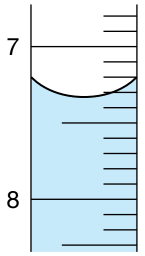
Solution: The reading at the bottom of the meniscus is 29.65 mL.

Revision questions

7. **Record the reading on the burette in this diagram.**



8. **Explain why a student reading this burette as 8.67 mL is incorrect.**



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How do we count significant figures?

When calculations are to be made from different measurements, it is important that we are able to assess the number of significant figures in each measurement. We can then quote the answer to an appropriate level of precision.

In assessing the number of significant figures in a measurement, we use two rules:

- All non-zero digits are counted.
- Zeros are counted if they come *after* a non-zero digit.

Examples

- A burette reading of 21.37 mL has 4 significant figures.
- A pipette reading of 20.00 mL has 4 significant figures.
- A measurement of 100 mL from a gas syringe has 3 significant figures.
- A mass of 1.003 g has 4 significant figures.
- A mass of 0.003 g has 1 significant figure.
- A mass of 0.300 g has 3 significant figures.

Note: Significant figures are not the same as decimal places. If a measurement is expressed in standard form, significant figures are counted from the decimal part only. For example, the Avogadro constant, when expressed as 6.02×10^{23}, contains 3 significant figures. The digits in the index are not counted.

study on

Unit 2
AOS 2
Topic 2
Concept 9

Significant figures
Summary screen and practice questions

Significant figures and calculations

The number of significant figures allowed in any calculated answer is governed by the number of significant figures in the least precise value from which that answer is calculated.

If a calculation involves defined conversion constants, these are not counted. Remember that it is only the measurements that are considered.

Working with significant figures

- When multiplying or dividing, the answer is written to the least number of significant digits.
- When adding or subtracting, the answer is written to the least number of decimal places.

Sample problem 10.3

In determining the density of a particular liquid, a student measured the volume of a sample as 8.3 mL. She then weighed the same sample and obtained a mass of 7.2136 g. Calculate the density to the correct level of precision.

Solution: $\text{density} = \dfrac{\text{mass}}{\text{volume}}$

$$= \frac{7.2136}{8.3} \quad \begin{matrix}\text{(5 significant figures)}\\ \text{(2 significant figures)}\end{matrix}$$

$$= 0.8691 \text{ (This needs to be rounded to 2 significant figures.)}$$

$$= 0.87 \text{ g mL}^{-1}$$

Revision questions

9. **How many significant figures are there in each of the following measurements?**
   - **(a) 256 litres**
   - **(b) 43 001 grams**
   - **(c) 0.230 °C**
   - **(d) 21.0 mL**
10. **Calculate the following to the correct number of significant figures.**
    - **(a) 4.25 + 9**
    - **(b) 0.04 + 3.7**
    - **(c) 5.640 + 70.435**
    - **(d) 0.80 – 0.3**
    - **(e) 840 – 627.03**
    - **(f) 12.01 + 6.7**
11. **Solve the following problems using the correct number of significant figures in the answers.**
    - **(a) density = 27.8 grams/1.2 mL =**
    - **(b) mass = 23.45 grams + 5.332 grams =**
    - **(c) moles = 1.221M × 2.6 L =**
    - **(d)** $m(\text{Cu}) = 0.45 \text{ mol} \times 63.5 \text{ g mol}^{-1} =$

Errors and uncertainties

The ideas of errors and uncertainties are related to the topic of significant figures.

Errors cause the value of a measurement to change. The errors associated with taking a measurement cause its value to be different from the true value. Errors (no matter how small) are present in every measurement we make. There are two types of errors:

- Random errors cause measurements to be sometimes higher than their true values and sometimes lower. The effect of random errors can therefore be minimised by doing a large number of measurements and averaging them.
- Systematic errors are always biased the same way. In other words, they cause values to be always too high or always too low. The effect of systematic errors cannot be minimised by multiple observations and averaging.

Uncertainty is an extension of significant figures. A plus/minus notation is used to indicate the possible range of values for the last significant figure. For example, when using an electronic balance, a student might record the result as:

$$35.964 \pm 0.002 \text{ g}$$

As discussed above, the use of 35.964 indicates that the '4' is the digit that we are least certain of. The '4' is our 'best estimate'. The uncertainty of ± 0.002 indicates how uncertain we are. In effect, we are stating that the measured mass could reasonably be expected to lie between 35.962 g and 35.966 g.

Effective study techniques to improve your results

Setting up a study routine

- Find a quiet place to study, whether it is at home or at a library.
- Prepare a realistic study timetable that takes into account sporting, work and other commitments.
- Study for about an hour and then take a short break. Use a timer on your computer to assist you.
- Limit your time spent on individual subjects; you will get more done if you realise that time is short.
- Put away all sources of distraction, including your mobile phone, and do not access any social networking sites.
- Reward yourself for persisting with your timetable.

Study skills

- Review each lesson as soon as practical after the class, highlighting/noting main points and learning terminology.
- Complete homework on time.
- Read the textbook regularly.
- Study the concepts and practise calculations. If you have any questions, put a sticky note on the page. Ask a friend first; if you are unable to get a satisfactory answer, ask your teacher.
- Make sure that the information is actively processed; reading is not an adequate study technique. Methods for processing information include using concept maps, preparing flash cards, writing a summary, telling a story, drawing a labelled diagram, making a video, creating a slide show, writing a song, creating a mnemonic, using a Venn diagram, recording a podcast or brainstorming using a large sheet of paper. Don't fall into the trap of spending so long preparing these materials that you don't have time to study them.
- A very effective way of learning is to teach someone else.
- Studying with a friend or small group can be very useful.
- Retaining information will be maximised with regular review and repetition.
- There are numerous excellent websites and YouTube videos that explain various concepts.
- Do past exams, preferably to time; analyse your mistakes and relearn the concept, rather than just writing the correct answers. Ask yourself if the mistakes are due to mathematical errors, not reading the question properly, lack of knowledge or lack of understanding. Learn from your mistakes.

Revision questions

12. **In your opinion, what are the five most important study hints?**
13. **Record four strategies used by other students in the class to review concepts.**

Exam preparation

- Eat healthily.
- Exercise regularly.
- Learn to relax; practising meditation could be beneficial.
- Have a good night's sleep before the actual exam.
- Organise the necessary materials the night before.
- Arrive in good time.
- Be positive.

The exam

Using reading time effectively

- Read the whole paper, particularly the last pages.
- Decide which questions you will do first: that is, the ones that you think you can do easily. You will gain confidence and have more time for the other questions.
- Consider the units given in a question and the units required in the answer.
- Check multiple choice questions for qualifying words such as *not*, *always* and *all*.
- Plan time for each section and stick to it.

Getting the best results in writing time

- Read questions carefully and highlight relevant data.
- Remember that multiple choice questions are worth only 1 mark so move on if you are unable to answer a question. Answer all multiple choice questions, eliminate incorrect alternatives and guess if you have to.
- Use a diagram for analysis questions.
- Balance chemical equations and include states.
- Think about the terminology required in a question and use it appropriately.
- Try to answer all questions. It is easier to get half of the whole exam correct than all of half the exam correct!
- Be aware of the difference between *a* and *the*; for example, 'write *a* structure (equation) for . . .' versus 'write *the* structure (equation) for . . .'
- When drawing structures, make sure that each atom has the correct number of bonds.
- If you find a question particularly difficult, read it carefully, highlight relevant information and come back to it later.
- If you are feeling stressed, take a minute to have a few deep breaths and relax.

Calculations

- Set work out clearly and show all working. Make your answer obvious.
- Give full information; for example, use 'n(NaCl) =' not just 'n ='.
- Check formulas and molar masses.
- Make sure you use the correct number of significant figures throughout the whole question.
- Consequential marks apply in chemistry so, even if you know you have an incorrect answer but do not know why, continue with the calculations.

UNIT 2

AREA OF STUDY 1

AREA OF STUDY 2

CHAPTER

11 Connecting chemical reactions and equations

YOU WILL EXAMINE:

- the difference between chemical and physical changes
- how chemical equations can be written to summarise chemical reactions
- how chemical equations are balanced to reflect the law of conservation of mass
- how patterns in chemical reactions can be used to predict the products of a chemical reaction.

> Measure what is measurable, and make measurable what is not so.
>
> *Galileo Galilei*

Sparkling wines have a characteristic 'pop and froth' effect due to the release of carbon dioxide gas that was bottled under pressure. Is this a physical change or a chemical change?

Chemical changes

Our environment is subject to countless **chemical changes**, most of them the result of natural processes that have evolved over millions of years. However, the actions of humans (especially since the Industrial Revolution) have added to this and placed much more stress on the environment than ever before. To develop new chemical processes that are sympathetic to the environment, a thorough knowledge of chemical changes is essential.

In a chemical change, one or more new substances are formed. A physical change does not produce any new substances. It simply involves a change of state in the substances already present.

A chemical change is a reaction in which one or more substances are transformed into one or more new substances. **Physical changes**, such as ice melting to form water or dry ice subliming to produce a theatrical 'fog', do not produce *new* substances. Physical changes are changes of state.

When chemical changes are taking place, we may observe one or more of the following:

- A substance disappears.
- A gas is given off. (The production of a gas as a result of a chemical change is called effervescence.)
- A solid is precipitated. (Precipitation is the formation of a new substance that is only slightly soluble and settles as a solid.)
- A colour change takes place.
- The temperature changes.
- A new odour is released.
- Light is given out.

Natural gas is used as a fuel in this flame thrower. Methane, CH_4, is the principal component of natural gas. In the fire, methane and oxygen in the air react and are transformed into carbon dioxide and water vapour or steam. Methane and oxygen are called the reactants — chemicals that are present before a chemical reaction takes place. Carbon dioxide and water are the products — the new chemicals that result from the chemical reaction.

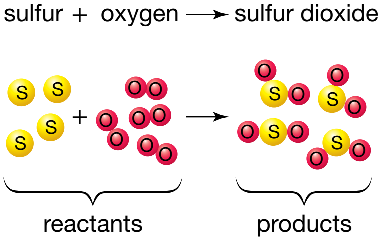
A **chemical equation** is a symbolic way of representing a chemical reaction, showing **reactants** and **products** and their physical states. It is also possible to indicate, above and below the arrow between reactants and products, the presence of catalysts or other agents needed to activate the reaction.

A chemical equation is a way of summarising a chemical change. It shows the formulas of the original reactants on the left and the formulas of the new substances (products) on the right. It also shows the physical state of each substance involved, and it may indicate conditions necessary for the reaction to occur.

Chemical equations

In a chemical reaction, the way in which the atoms are joined together is changed. Bonds are broken and new ones are formed as the reactants are changed into products. The same atoms (both number and type) are present before and after the reaction — they are just arranged differently.

$$\text{reactants} \longrightarrow \text{products}$$



Sulfur and oxygen gas combine to form sulfur dioxide gas in a chemical reaction. Note the rearrangement of atoms that takes place during the reaction.

For example, when sodium is heated and plunged into green chlorine gas, the white ionic solid sodium chloride is formed. This can be represented in a chemical equation using the symbol for each element and formula of the compound.

$$\mathrm{Na} + \mathrm{Cl_2} \longrightarrow \mathrm{NaCl}$$

$$\underbrace{\text{sodium} + \text{chlorine gas}}_{\textit{reactants}} \longrightarrow \text{sodium chloride}$$

$$\longrightarrow \textit{products}$$

Such equations, which show just the formulas of the reactants and products, are called ‘skeleton’ equations.

The physical states of the substances should be added to the equation. These are shown in brackets after each symbol or formula: (s) for a solid, (l) for a liquid, (g) for a gas, and (aq) for a solution in water.

$$2\mathrm{Na(s)} + \mathrm{Cl_2(g)} \longrightarrow 2\mathrm{NaCl(s)}$$

The states of the reactants and products should be indicated in the equation by (s) for solid, (l) for liquid, (g) for gas and (aq) for a solution in water.

In some chemical reactions, a substance called a **catalyst** is used. A catalyst speeds up a reaction without being used up. Because a catalyst is not used up, it is neither a reactant nor a product and is written above the arrow in the equation. For example, manganese dioxide is a catalyst that speeds up the decomposition of a solution of hydrogen peroxide, H_2O_2, to water and oxygen.

$$2\mathrm{H_2O_2(aq)} \xrightarrow{\mathrm{MnO_2}} 2\mathrm{H_2O(l)} + \mathrm{O_2(g)}$$

A similar notation is used to denote dissolving and heating. When a substance such as copper sulfate, $CuSO_4$, is dissolved, water merely causes the ions in the lattice to come apart. Since the water is not being changed, it is written above the arrow as shown below.

$$\mathrm{CuSO_4(s)} \xrightarrow{\mathrm{H_2O}} \mathrm{Cu^{2+}(aq)} + \mathrm{SO_4{}^{2-}(aq)}$$

Also, when a reaction requires heat, the word ‘heat’ is written above the arrow to denote that heat has been applied. For example, the melting of ice to form water requires heat and may be described by the following chemical equation:

$$\mathrm{H_2O(s)} \xrightarrow{\text{heat}} \mathrm{H_2O(l)}$$

Table 11.1 shows many of the symbols that are used in writing chemical equations.

TABLE 11.1 Symbols used in equations

Symbol	Meaning
(s)	a reactant or product in the solid state
(l)	a reactant or product in the liquid state
(aq)	an aqueous solution (the substance is dissolved in water)
(g)	a reactant or product in the gaseous state
$\xrightarrow{\text{heat}}$	indicates that heat is supplied to the reaction
$\xrightarrow{\mathrm{Pt}}$	a formula written above the arrow indicates its use as a catalyst (in this case, platinum)

Sample problem 11.1

When solid mercury(II) sulfide is heated in air, molten mercury and gaseous sulfur dioxide are produced. Write an equation for this chemical change.

Solution: The reactants are mercury(II) sulfide and oxygen, so they are written on the left-hand side of the reaction arrow. The products, mercury and sulfur dioxide, are written on the right-hand side of the reaction arrow. The states of all reactants and products are included.

$$\mathrm{HgS(s)} + \mathrm{O_2(g)} \longrightarrow \mathrm{Hg(l)} + \mathrm{SO_2(g)}$$

Sample problem 11.2

Write a sentence that completely describes the following reaction:

$$N_2O_3(g) + H_2O(l) \longrightarrow 2HNO_2(aq)$$

Solution: Gaseous dinitrogen trioxide reacts with water to produce a solution of nitrous acid.

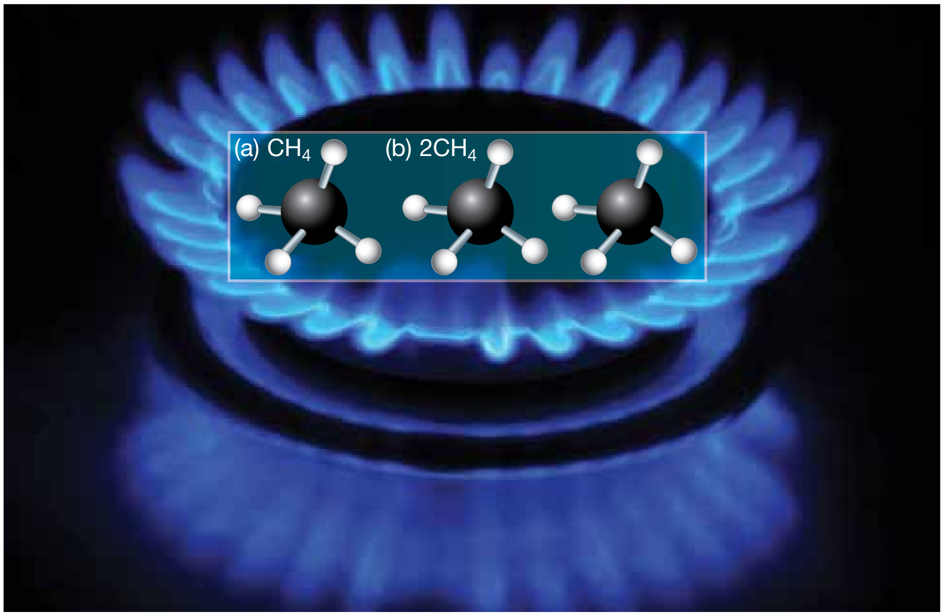
Revision questions

1. **Write a skeleton equation for each of these chemical changes.**
   - **(a) Aluminium reacts with oxygen in the air to form aluminium oxide powder.**
   - **(b) Oxygen gas can be made by heating potassium chlorate in the presence of the catalyst manganese dioxide. Solid potassium chloride is produced.**
2. **Write sentences that completely describe the following chemical changes.**
   - **(a)** $Pb(NO_3)_2(aq) + H_2SO_4(aq) \longrightarrow PbSO_4(s) + 2HNO_3(aq)$
   - **(b)** $4NH_3(g) + 5O_2(g) \xrightarrow{Pt} 4NO(g) + 6H_2O(g)$

Balancing chemical equations

A balanced chemical equation has the same number of each type of atom on each side of the equation. Balancing is achieved by placing numbers (called coefficients) in front of each formula.

To represent chemical reactions correctly, equations must be *balanced*. This means the number of atoms of each element must be the same on both sides of the equation. According to the **law of conservation of mass**, the total mass of the reactants in a chemical reaction is equal to the total mass of the products. Atoms are not created or destroyed, but are rearranged to form new substances. In order to balance an equation, numbers called **coefficients** are placed *in front* of the whole formulas. Remember that we cannot change the formula of any substance in the equation, as these are found by experiment. In the example on the previous page, where two formula units of NaCl are produced, 2NaCl is used, rather than Na_2Cl_2.



A coefficient in a balanced equation applies to the whole formula. (a) CH_4 means one molecule of methane consisting of one C atom and four H atoms. (b) Here, the coefficient '2' is used to indicate the presence of two molecules: $2CH_4$ means two molecules of methane, each consisting of one C atom and four H atoms.

The vast range of possible chemical reactions (and associated chemical equations) makes it difficult to produce a list of hard and fast rules for balancing equations. However, the following rules should be considered a starting point.

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Balancing chemical equations

Rules for balancing equations

1. Write the reactants and products using the correct formula and state for each substance.
2. Count the number of atoms of each element on the left-hand side of the equation. Do the same for the right-hand side and compare for each element. If any of these numbers do not match, the equation is not balanced and you need to proceed to the following steps.
3. Balance by placing coefficients in front of the formulas. *Do not change the actual formula.* If any substance is present as an element, leave the balancing of it to last.
4. Check all atoms or ions to ensure that they are balanced.
5. Make sure that the coefficients are in their lowest possible ratio.

Sample problem 11.3

At 1000 °C, ammonia gas, NH_3, reacts with oxygen gas to form gaseous nitric oxide, NO, and water vapour. Write a chemical equation for this reaction.

Solution:

STEP 1

Write the reactants and products including states.

$$NH_3(g) + O_2(g) \longrightarrow NO(g) + H_2O(g)$$

STEP 2

This is an unbalanced equation because there are three atoms of hydrogen on the left-hand side of the equation and only two atoms of hydrogen on the right-hand side.

STEP 3

Begin to balance the equation by correcting for hydrogen atoms. A coefficient of 2 for NH_3 and a coefficient of 3 for H_2O gives six atoms of hydrogen on both sides:

$$2NH_3(g) + O_2(g) \longrightarrow NO(g) + 3H_2O(g)$$

Now, an imbalance in nitrogen atoms has occurred.

Balance the nitrogen atoms by introducing a coefficient of 2 for NO.

$$2NH_3(g) + O_2(g) \longrightarrow 2NO(g) + 3H_2O(g)$$

STEP 4

Now there are two atoms of oxygen on the left and five on the right. Balance the oxygen atoms. This can be done by introducing a coefficient of $2\frac{1}{2}$ for O_2.

$$2NH_3(g) + 2\tfrac{1}{2}\,O_2(g) \longrightarrow 2NO(g) + 3H_2O(g)$$

STEP 5

Check the equation. The same numbers of N, H and O atoms are on both sides of the equation, and states have been included. The equation is balanced. But the ratios are not yet in the simplest ratio.

STEP 6

Since the usual practice is to have whole-number coefficients, multiply the entire equation by 2.

$$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(g)$$

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Revision questions

3. **Balance the following equations.**
   - **(a)** $N_2(g) + H_2(g) \longrightarrow NH_3(g)$
   - **(b)** $SiO_2(s) + C(s) \longrightarrow Si(s) + CO(g)$
   - **(c)** $FeO(s) + O_2(g) \longrightarrow Fe_2O_3(s)$
   - **(d)** $Cr(s) + S_8(s) \longrightarrow Cr_2S_3(s)$
   - **(e)** $NaHCO_3(s) \longrightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$
4. **When silicon tetrachloride, a liquid, reacts with solid magnesium, two solid products are formed: silicon and magnesium chloride. Write a balanced equation for this reaction.**
5. **Although each of the equations listed below is 'balanced', the equations are all incorrect. Identify the errors in each equation, then write a correct equation for each reaction.**
   - **(a)** $CH_4 + O_2 \longrightarrow CO_2 + H_4$
   - **(b)** $I_2 + Cl_2 \longrightarrow I_2Cl_2$
   - **(c)** $K + O_2 \longrightarrow KO_2$
   - **(d)** $Cl_2 + NaBr \longrightarrow NaCl_2 + Br$

Types of chemical reaction

Magnesium combining with oxygen to form MgO(s)

Patterns can be found in the vast number of chemical reactions taking place around us. Chemists have determined several main groups of chemical reactions that help us predict the products of these reactions. We should keep in mind, however, that when we write equations they should be based on experimental data for complete certainty.

The main types of chemical reaction discussed in this chapter are:

- precipitation
- acid–base
- combustion.

Precipitation reactions

Precipitation occurs when ions in solution combine to form a new compound of low solubility in water. This low-solubility compound forms as solid particles that eventually settle. It is called a **precipitate**. In order to predict whether a precipitate will form, we must know which substances are soluble in water and which substances are insoluble.

From table 11.2 we can predict that, when a solution of sodium chloride is mixed with a solution of silver nitrate, a precipitate of silver chloride will form:

$$NaCl(aq) + AgNO_3(aq) \longrightarrow NaNO_3(aq) + AgCl(s)$$

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Precipitation reactions

In a precipitation reaction, ions from the first reactant swap with ions from the second reactant to form new combinations, one of which is insoluble.

That is, the Ag^+ ions combine with the Cl^- ions to form solid AgCl while the Na^+ and NO_3^- ions remain in solution.

(LEFT) Silver chloride, a white precipitate, is formed when colourless solutions of silver nitrate and sodium chloride are mixed. Sodium nitrate is the other product in this reaction; it is soluble.

(BELOW) The formation of a precipitate of silver chloride from the mixing of solutions of sodium chloride and silver nitrate

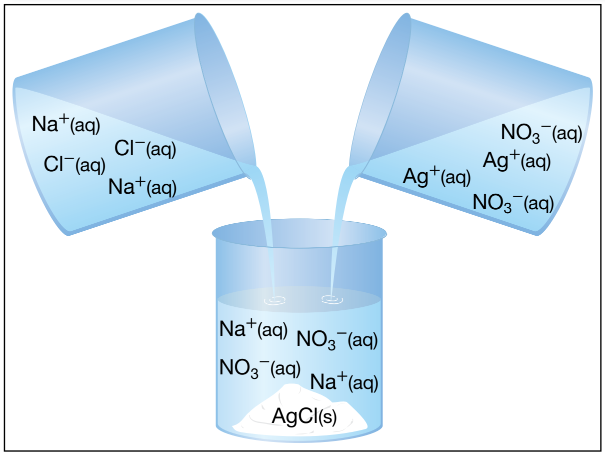

TABLE 11.2 Solubility in water of compounds of common ions

Name of ion	Symbol	Soluble compounds of ion	Insoluble compounds of ion
group I ions	Li^+, Na^+, K^+, Rb^+, Cs^+, Fr^+	all	none
ammonium	NH_4^+		
hydrogen	H^+		
nitrate	NO_3^-		
nitrite	NO_2^-		
chlorides	Cl^-	most	Ag^+, Pb^{2+}, Hg^{2+} ($PbCl_2$ is moderately soluble in hot water.)
bromides	Br^-		
iodides	I^-		
sulfates	SO_4^{2-}		Ba^{2+}, Pb^{2+} (Ag_2SO_4 and $CaSO_4$ are slightly soluble.)
carbonates	CO_3^{2-}	Na^+, K^+, NH_4^+	most
phosphates	PO_4^{3-}		
sulfides	S^{2-}	Na^+, K^+	most (MgS, CaS, BaS, Al_2S_3 and Fe_2S_3 decompose in water.)
hydroxides	OH^-	Na^+, K^+, Ba^{2+} (NH_4OH and $(NH_4)_2O$ do not exist as solids.)	most ($Ca(OH)_2$ is slightly soluble.)
oxides	O^{2-}		

Note: Most of the ionic compounds not listed here may be assumed to be insoluble in water.

Precipitation reactions are dealt with further in chapter 15.

Sample problem 11.4

Predict the products and write a balanced equation for the reaction occurring between sodium carbonate and copper(II) sulfate.

Solution: **STEP 1**

Write the reactants.

$$Na_2CO_3 + CuSO_4 \longrightarrow$$

STEP 2

Look at the solubility table to determine which of the ions are soluble. Sodium carbonate is aqueous (since all group I ions are soluble) and copper(II) sulfate is aqueous (since most sulfates are soluble). The products are sodium sulfate and copper(II) carbonate. Sodium sulfate is aqueous (since all group I compounds are soluble), but copper(II) carbonate is an insoluble compound.

STEP 3

Write the balanced equation for the reaction.

$$Na_2CO_3(aq) + CuSO_4(aq) \longrightarrow Na_2SO_4(aq) + CuCO_3(s)$$

Revision questions

6. **Determine whether each of the following ionic compounds is soluble:**
   - (a) **$BaSO_4$**
   - (b) **$Pb(NO_3)_2$**
   - (c) **PbI_2**
   - (d) **sodium sulfide.**
7. **Use table 11.2 to determine:**
   - (a) **which sodium compounds are soluble**
   - (b) **which potassium compounds are insoluble**
   - (c) **which silver compounds are soluble.**
8. **Predict products and write balanced equations for:**
   - (a) **$KOH(aq) + Ca(NO_3)_2(aq) \longrightarrow$**
   - (b) **$Na_2S(aq) + (CH_3COO)_2Pb(aq) \longrightarrow$**
   - (c) **$(NH_4)_3PO_4(aq) + CaCl_2(aq) \longrightarrow$**

Acid–base and neutralisation reactions

When an acid reacts with a base to form a salt and water, a **neutralisation reaction** occurs:

$$\text{acid} + \text{base} \longrightarrow \text{salt} + \text{water}$$

For example, hydrochloric acid reacts with the base sodium hydroxide to form sodium chloride (common table salt) and water:

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

Other common reactions involving acids are:

1. acid + metal ⟶ salt + hydrogen

   This reaction does not occur with Cu, Hg, or Ag.

2. acid + metal carbonate ⟶ salt + water + carbon dioxide

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

3. acid + metal oxide ⟶ salt + water

   $$2HCl(aq) + CuO(s) \longrightarrow CuCl_2(aq) + H_2O(l)$$

4. acid + metal hydroxide ⟶ salt + water

   $$H_2SO_4(aq) + 2NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

Since a neutralisation reaction is often difficult to detect, a chemical **indicator** can be used. Indicators are compounds that change colour when an acid or a base has completely reacted.

Acid–base reactions are dealt with further in chapter 17.

Sample problem 11.5

Write a balanced chemical equation for the reaction between sulfuric acid and magnesium carbonate.

Solution: The reaction of an acid (sulfuric acid) and a metal carbonate (magnesium carbonate) produces a salt (magnesium sulfate), water and carbon dioxide. Magnesium sulfate dissolves in water. The equation for the reaction is:

$$H_2SO_4(aq) + MgCO_3(s) \longrightarrow MgSO_4(aq) + H_2O(l) + CO_2(g)$$

Revision questions

9. **Predict the products and write balanced equations for:**
  - **(a)** $LiOH(aq) + HCl(aq) \longrightarrow$
  - **(b)** $NaOH(aq) + HNO_3(aq) \longrightarrow$
  - **(c)** $KOH(aq) + HCl(aq) \longrightarrow$

10. **Write balanced equations for:**
  - **(a) nitric acid reacting with lithium hydroxide**
  - **(b) sulfuric acid reacting with sodium carbonate**
  - **(c) sulfuric acid reacting with copper(II) oxide**
  - **(d) hydrochloric acid reacting with magnesium metal.**

Combustion reactions

When hydrocarbons burn in a plentiful supply of oxygen, they give off heat to their surroundings and produce carbon dioxide and water.

When a limited air supply is available for a **combustion reaction**, carbon monoxide may be formed in preference to carbon dioxide. For example, when octane, C_8H_{18}, is burned in air, the combustion reaction may be represented as:

$$2C_8H_{18}(l) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(g)$$

However, when octane is burned in an engine, where the supply of oxygen is limited, carbon monoxide gas is one of the products:

$$2C_8H_{18}(l) + 17O_2(g) \longrightarrow 16CO(g) + 18H_2O(g)$$

The products of hydrocarbon combustion are in the gaseous state.

A hydrocarbon is a compound that consists of only hydrogen and carbon atoms. When a hydrocarbon burns in air, carbon dioxide and water are always formed as the products.

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Weblink
Combustion reactions

When candles burn in oxygen, carbon dioxide and water are produced. When balancing combustion equations, balance the carbon atoms first, then the hydrogen atoms, followed by the oxygen atoms.

Sample problem 11.6

Candle wax, $C_{25}H_{52}$, burns in air. Write a chemical equation for the reaction.

Solution: **STEP 1**

Carbon dioxide and water are the products of combustion in air:

$$C_{25}H_{52}(s) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

STEP 2

Balance the equation, first for carbon, then hydrogen, and then oxygen.

$$C_{25}H_{52}(s) + 38O_2(g) \longrightarrow 25CO_2(g) + 26H_2O(g)$$

For hydrocarbon combustion, an algebraic chemical equation can be written. By knowing the values of just x and y, the combustion equation can be written quickly. This equation is:

$$C_xH_y + (x + \frac{y}{4})O_2 \longrightarrow xCO_2 + \frac{y}{2}H_2O$$

Revision questions

11. **Write balanced equations for the following reactions:**
- **(a)** **propane gas burning in oxygen ⟶**
- **(b)** **ethene gas burning in oxygen ⟶**
- **(c)** **ethane gas burning in a limited supply of oxygen ⟶**

12. **Butane, C_4H_{10}, in cigarette lighters burns with a sooty flame. The soot that forms is actually carbon. Water vapour is also produced in this reaction.**
- **(a)** **Explain why carbon, rather than carbon dioxide, forms when a cigarette lighter is lit.**
- **(b)** **Write a balanced molecular equation for the cigarette lighter reaction, showing the physical states of all reactants and products.**
- **(c)** **Write a balanced molecular equation for the complete combustion of butane in air, showing the physical states of all reactants and products.**

13. **Predict whether a reaction will take place between the following substances and, if so, write a balanced equation for it.**
(a) **$Cu + Pb(NO_3)_2$** **(b)** **$Mg + ZnCl_2$** **(c)** **$Fe + AgNO_3$**

Chemical reactions — by patterns

Many people find it useful to predict chemical reactions by using certain patterns — many of which have been alluded to in the previous section under their particular reaction type. However, it still must be remembered that it is only predictions that are being made. Exceptions can always occur and all such predictions should really be checked by experiment. Shown below is a list of some of the more common patterns.

- acid + metal hydroxide (base) ⟶ salt + water
- acid + basic oxide ⟶ salt + water
- acidic oxide + base ⟶ salt + water
- acid + metal ⟶ salt + hydrogen
  (no reaction for Ag, Cu, Pt or Au)
- acid + metal carbonate ⟶ salt + carbon dioxide + water
- acid + metal hydrogen carbonate ⟶
  salt + carbon dioxide + water
- metal carbonate $\xrightarrow{\text{heat}}$ metal oxide + carbon dioxide
- hydrocarbon + (plentiful) oxygen ⟶
  carbon dioxide + water

Crystals of vitamin C, ascorbic acid, are shown in a reflected light micrograph. The ascorbic acid ions start off dissolved in an aqueous solution, and then the solvent is slowly evaporated to form the crystals. This solidification reaction can be written as an ionic equation that is the reverse of the dissociation ionic equation.

Chapter review

Summary

- A chemical reaction involves a chemical change in which one or more kinds of matter (reactants) are transformed into one or more new kinds of matter (products). In this way it differs from a physical change, in which no new substances are produced.
- A chemical change may be indicated by:
  - disappearance of a substance
  - evolution of a gas
  - formation of a precipitate (solid)
  - colour change
  - temperature change
  - odour change.
- A chemical reaction may be expressed symbolically by a chemical equation. Principles for writing chemical equations include:
  - Indicate the physical states of the reactants and products in parentheses after each chemical formula.
  - Above the reaction arrow, show the presence of any catalysts (substances that are used to alter the rate of a chemical reaction but are not themselves used up in the reaction).
  - According to the law of conservation of mass, atoms are neither created nor destroyed in chemical reactions but rearranged to form new substances. Therefore all atoms in the equation must be *balanced*. This is done by placing numbers, called coefficients, in front of the whole formulas.
- Three main types of chemical reaction are:
  - A precipitation reaction produces a compound (precipitate) of low solubility in water. A solubility table may be used to predict the formation of a precipitate when solutions are mixed.
  - A neutralisation reaction occurs when an acid reacts with a base to produce a salt and water. These reactions can be detected using a chemical indicator. Common reactions of acids include:

    acid + metal hydroxide ⟶ salt + water

    acid + metal carbonate ⟶ salt + water + carbon dioxide

    acid + metal oxide ⟶ salt + water

    acid + metal ⟶ salt + hydrogen (not Cu, Hg or Ag).
  - A combustion reaction involves the burning of an organic compound in oxygen to produce carbon dioxide or carbon monoxide as well as water, generating heat.

    hydrocarbon + oxygen (plentiful supply) ⟶ carbon dioxide + water

    hydrocarbon + oxygen (limited supply) ⟶ carbon monoxide + water

Multiple choice questions

1. Which one of the following processes is *not* an example of a chemical change?

A Ice melts.
B Iron rusts.
C Wood burns.
D Leaves turn brown in autumn.

2. During any chemical reaction, there is *never* a change to the:

A total number of molecules in the system
B heat produced by the system
C heat absorbed by the system
D total mass of the system.

3. Consider the following equation:

$$a\mathrm{NH_3}(g) + b\mathrm{O_2}(g) \longrightarrow c\mathrm{N_2}(g) + d\mathrm{H_2O}(g)$$

The coefficients that balance the equation are represented by *a, b, c* and *d*. These are, respectively:

A 2, 1, 1, 3
B 2, 2, 1, 3
C 4, 3, 2, 6
D 4, 3, 2, 4.

4. Consider the following equation:

$$a\mathrm{C_6H_{14}}(g) + b\mathrm{O_2}(g) \longrightarrow c\mathrm{H_2O}(g) + d\mathrm{CO}(g)$$

The coefficients that balance the equation are represented by *a, b, c* and *d*. These are, respectively:

A 1, 13, 7, 6
B 1, 19, 7, 6
C 2, 13, 14, 12
D 2, 19, 14, 12.

5. Hydrogen gas burns readily in oxygen to produce water. Which of the following is the correctly balanced equation for this process?

A $\mathrm{H_2}(g) + \mathrm{O_2}(g) \longrightarrow \mathrm{H_2O}(l)$
B $2\mathrm{H_2}(g) + \mathrm{O_2}(g) \longrightarrow 2\mathrm{H_2O}(l)$
C $\mathrm{H_4}(g) + \mathrm{O_2}(g) \longrightarrow \mathrm{H_4O_2}(l)$
D $\mathrm{H}(g) + \mathrm{O}(g) \longrightarrow \mathrm{H_2O}(l)$

6. Which of the following is a correctly balanced equation?

A $2\mathrm{CuO}(s) + 2\mathrm{C}(s) \longrightarrow \mathrm{Cu}(s) + 4\mathrm{CO_2}(g)$
B $\mathrm{SO_2}(g) + \mathrm{O_2}(g) \longrightarrow \mathrm{SO_3}(g)$
C $\mathrm{H_2S}(g) + \mathrm{O_2}(g) \longrightarrow \mathrm{H_2O}(l) + \mathrm{SO_2}(g)$
D $\mathrm{P_4O_{10}}(g) + 10\mathrm{C}(s) \longrightarrow \mathrm{P_4}(g) + 10\mathrm{CO}(g)$

7. In which of the following cases would you expect a chemical reaction to *not* occur?

- **A** Magnesium is added to dilute sulfuric acid.
- **B** Copper is added to dilute hydrochloric acid.
- **C** Magnesium is added to dilute nitric acid.
- **D** Zinc is added to ethanoic acid.

8. Carbon dioxide gas is identified using the 'limewater test'. When the gas is bubbled through a solution of limewater (calcium hydroxide), the solution turns milky due to the formation of insoluble calcium carbonate. The equation for this reaction is:

$$Ca(OH)_2(aq) + CO_2(g) \longrightarrow CaCO_3(s) + H_2O(l)$$

Which of the following hydroxide solutions would produce a similar observation?

- **A** NH_4OH
- **B** $NaOH$
- **C** KOH
- **D** $Ba(OH)_2$

Review questions

Identifying chemical reactions

1. Explain how physical changes:
  (a) are similar to chemical changes
  (b) differ from chemical changes.

2. Describe four observations that provide evidence for a chemical reaction having taken place, illustrating your answer with examples from everyday life.

3. What is a chemical equation? Why does it have to be balanced?

4. Explain how the law of conservation of mass is related to the balancing of a chemical equation.

5. Consider the following balanced equation:

$$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$$

  (a) What information does this give about the reaction?
  (b) What information is not given?

Balancing and completing chemical equations

6. Balance the following equations.
  (a) $N_2(g) + H_2(g) \longrightarrow NH_3(g)$
  (b) $HCl(aq) + Al(s) \longrightarrow AlCl_3(aq) + H_2(g)$
  (c) $NH_3(g) + O_2(g) \longrightarrow NO(g) + H_2O(g)$
  (d) $NO(g) + O_2(g) \longrightarrow NO_2(g)$
  (e) $Na(s) + P_4(s) \longrightarrow Na_3P(s)$

7. Write fully balanced equations, including the physical states of all reactants and products, for each of the following word equations.
  (a) When solid sodium is added to water, aqueous sodium hydroxide and hydrogen gas are formed.
  (b) Hydrogen chloride gas and ammonia gas react to form ammonium chloride.
  (c) Nitrogen monoxide gas reacts with oxygen gas to form nitrogen dioxide.
  (d) Solid zinc sulfide reacts with oxygen gas to form solid zinc oxide and sulfur dioxide gas.

8. Balance the following equations.
  (a) $Cu + O_2 \longrightarrow CuO$
  (b) $HgO \longrightarrow Hg + O_2$
  (c) $AsCl_3 + H_2S \longrightarrow As_2S_3 + HCl$
  (d) $Fe_2O_3 + H_2 \longrightarrow Fe + H_2O$
  (e) $NaCl \longrightarrow Na + Cl_2$
  (f) $Al + H_2SO_4 \longrightarrow H_2 + Al_2(SO_4)_3$
  (g) $C_8H_{18} + O_2 \longrightarrow CO_2 + H_2O$

9. Complete and balance the following equations.
  (a) $HNO_3(aq) + Zn(s) \longrightarrow$
  (b) $HCl(aq) + MgO(s) \longrightarrow$
  (c) $HCl(aq) + Ca(OH)_2(s) \longrightarrow$
  (d) $H_2SO_4(aq) + Na_2CO_3(s) \longrightarrow$
  (e) $Ca(s) + O_2(g) \longrightarrow$

10. Write balanced formula equations for the following word equations.
  (a) granulated zinc + sulfuric acid ⟶
  (b) metallic barium + oxygen gas ⟶
  (c) hydrochloric acid + magnesium hydroxide ⟶
  (d) calcium oxide + water ⟶

11. Bubbling chlorine gas through a solution of potassium iodide produces elemental iodine and a solution of potassium chloride. Write a balanced chemical equation, showing the physical states of all reactants and products, for this reaction.

12. Write sentences to describe each of the reactions below.
  (a) $2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$
  (b) $FeO(s) + C(s) \xrightarrow{\text{heat}} Fe(s) + CO(g)$

13. Complete the following equations.
  (a) $Ca(OH)_2(s) + H_3PO_4(aq) \longrightarrow$
  (b) $H_2SO_4(aq) + Al(OH)_3(s) \longrightarrow$
  (c) $Fe(OH)_3(s) + HCl(aq) \longrightarrow$
  (d) $BaCl_2(aq) + (NH_4)_2CO_3(aq) \longrightarrow$
  (e) $AgNO_3(aq) + H_2S(g) \longrightarrow$
  (f) $K_2CrO_4(aq) + Pb(NO_3)_2(aq) \longrightarrow$

Combustion reactions

14. Write hydrocarbon combustion equations for ethene, propene and l-butene:
  (a) in a plentiful air supply
  (b) in a limited air supply.

15. Write hydrocarbon combustion reactions for pentane, 2-pentene and 2-pentyne:
  (a) in a plentiful air supply
  (b) in a limited air supply.

16. Octane, C_8H_{18}, is the major hydrocarbon present in petrol. Write a balanced molecular equation, including physical states, for the combustion of octane:
  (a) in a plentiful air supply
  (b) in a limited air supply.

17. Write an equation for the complete combustion of:
  (a) glycerol, $C_3H_8O_3$
  (b) sucrose, $C_{12}H_{22}O_{11}$
  (c) ethanoic acid, CH_3COOH.

Exam practice questions

In a chemistry examination you will be required to answer a number of short and extended response questions.

Extended response questions

1. (a) From the following list of substances, choose a precipitate that would form in an aqueous solution: $CuSO_4$, $Ba(OH)_2$, KOH, $BaSO_4$. **1 mark**

(b) Write a chemical equation for the reaction between two reactants of your choice that would produce the precipitate you chose in part (a). (Remember that the reactants must be soluble in water.) **2 marks**

2. Two common fuels for outdoor barbeques are natural gas and propane gas.

(a) Write the equation for the combustion of natural gas (assuming it to be methane) in a plentiful supply of oxygen. **2 marks**

(b) Write the equation for the combustion of propane in a plentiful supply of oxygen. **2 marks**

These barbeques often display a warning for them not to be operated in confined spaces due to the possibility that the oxygen supply might start to become limited.

(c) Write the equation for the combustion of propane in a limited supply of oxygen. **2 marks**

(d) Using your answer to part (c), suggest why it would be dangerous to operate an outdoor barbecue in a confined space. **1 mark**

(e) If the same number of molecules of methane and propane are burned, which fuel would produce more molecules of carbon dioxide? Explain. **2 marks**

Another commonly available fuel for barbeques is butane. The equation for the combustion of butane is:

$$2C_4H_{10}(g) + 13O_2(g) \longrightarrow 8CO_2(g) + 10H_2O(g)$$

(f) If the same number of molecules of propane and butane are burned, which fuel would produce more molecules of water? **1 mark**

CHAPTER

12 Properties of water

YOU WILL EXAMINE:

- the special properties of water and their explanations in terms of hydrogen bonding
- why the special properties of water are important to life
- the importance of the solvent properties of water to aquatic and marine life
- the mechanisms by which substances dissolve in water
- precipitation reactions and their representation by both full and ionic equations.

> We never know the worth of water till the well is dry.
>
> *Thomas Fuller*

Water is essential to the life of these birds. Flamingos rely on the salts dissolved in the water in which they live to create ideal conditions for the production of their food. The water also supports the growth of micro-organisms that produce alpha and beta carotene; these micro-organisms form part of the flamingos' diet, and the carotenes give them their vibrant colouring.

Water exists naturally on Earth in all three states. It is continuously being recycled due to the water cycle and has many important properties essential for maintaining life.

Water — the life force

More than 70% of the Earth's surface is covered by water. Through the water cycle, water is recycled in the ocean and the atmosphere, and on the landmass (in rivers, lakes and sub-surface water). The human body also consists largely of water. What is so special about water?

Water is essential to life on Earth. Plants and all living creatures need it to survive.

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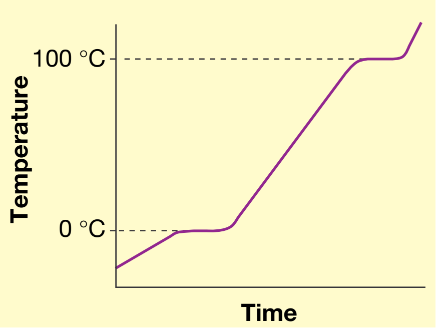
Unit 2	**Latent heat of vaporisation and evaporative cooling**
AOS 1	Summary screen and practice questions
Topic 1	
Concept 3	

The physical properties of water

Water is a polar, discrete molecular compound. At 25 °C, water is a colourless, odourless liquid. It freezes at 0 °C and boils at 100 °C. It is found in all forms — solid, liquid and gas.

Water has many properties that have resulted in the development of life on Earth in the form that we know it. For example, it needs a great deal of energy to heat up (having both a high melting temperature (0 °C) and a high boiling temperature (100 °C)), and it holds heat longer than most other substances.

Latent heat



The latent heat of water. Energy is needed to break the bonds between the water molecules. This causes the temperature to remain constant even though the water is being heated.

When a substance changes phase (for example, when ice melts and becomes water, or water becomes water vapour), energy is needed. This energy comes from the surrounding atmosphere. The energy needed to change the phase of a substance at its melting or boiling temperature is called the latent heat of that substance. The symbol for latent heat is *l*. Common units are kJ mol^{-1} and J kg^{-1}.

The **latent heat of fusion** of water is the amount of energy needed to change a fixed amount of water from a solid to liquid phase at 0 °C. The bonds between the ice molecules must be broken so that the water molecules can move around more freely in a liquid state. The latent heat of fusion of water is 6.02 kJ mol^{-1}, so 6.02 kJ of energy must be supplied to change each mole of water from solid to liquid.

Latent heat measures the amount of energy required to change the state of a given amount of substance. The latent heat of fusion measures the change from solid to liquid, while the latent heat of vaporisation measures the change from liquid to gas. Water has a relatively high latent heat for both these processes.

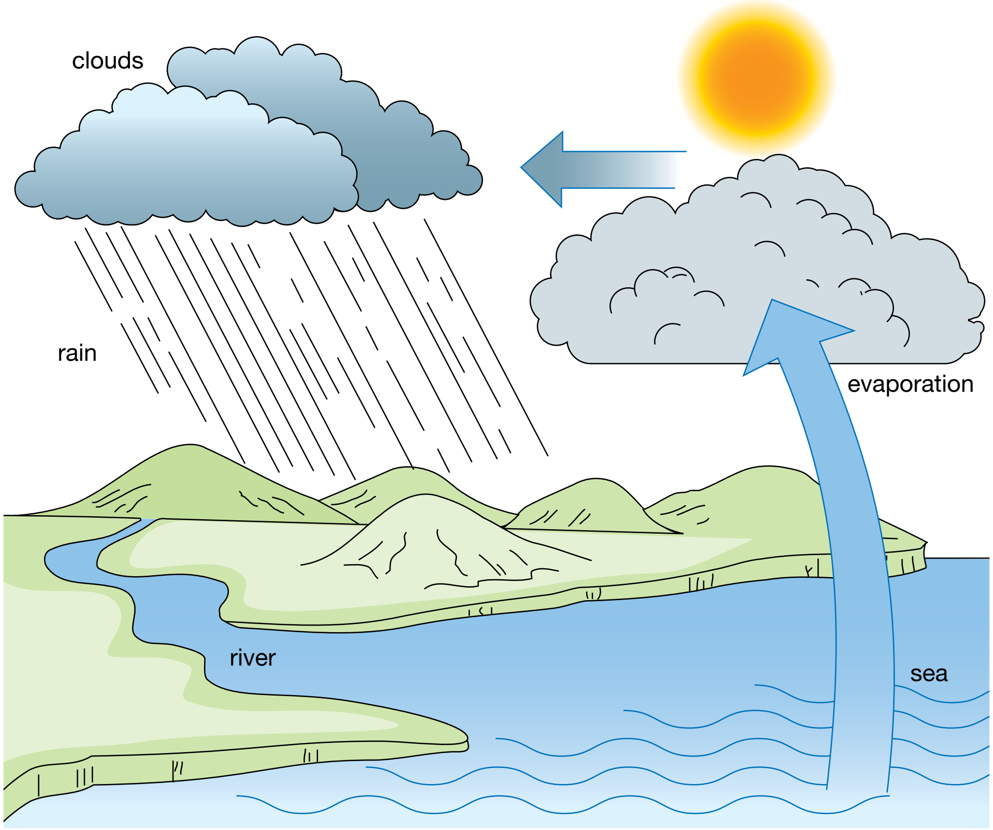
The **latent heat of vaporisation** of water is the amount of energy needed to change a fixed amount of water from a liquid to a gas at 100 °C. The remaining forces holding the water molecules together must be overcome so that the molecules may move around freely as a gas.

The higher the latent heat of a substance, the greater the amount of energy it needs when it changes phase. Water has much higher latent heat values than other substances of a similar size owing to the stronger intermolecular forces within its solid and liquid states.

The latent heat of vaporisation of water is 40.7 kJ mol^{-1}, which means that 40.7 kJ of energy must be supplied to vaporise one mole of water.

The relatively high latent heat of vaporisation of water is very useful in keeping organisms cool through perspiration. In humans, when a person perspires, droplets of liquid water on the skin evaporate by absorbing heat energy from the body. As seen above, this change from liquid water to water vapour requires a large amount of heat energy, making perspiration an efficient mechanism for removing unwanted heat from the body. It is interesting to note that, if there is already a lot of water vapour in the air (as is the case when humidity is high), this process is slow with the result that people often feel hot and uncomfortable on such days.

Another consequence of water's relatively high latent heat of vaporisation is that water evaporates more slowly than many other liquids. Given that we store large amounts of water in dams and lakes, this means that losses due to evaporation from their surfaces are lower than they would otherwise be.



The water cycle. The relatively high latent heat of vaporisation of water, 40.7 kJ mol^{-1}, means that a large amount of energy from the sun and surroundings is needed for evaporation.

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Unit 2
AOS 1
Topic 1
Concept 2

Specific heat capacity of water
Summary screen and practice questions

Specific heat capacity

The **specific heat capacity** of a substance is the amount of energy needed to raise the temperature of one gram of the substance by one degree Celsius. The formula $Q = mc\Delta T$ follows from this definition, where Q is the energy required (in J) to raise m g of a substance of specific heat capacity c by ΔT degrees Celsius.

Sample problem 12.1

How much energy is needed to raise 250 mL of water from 20 °C to 100 °C?

Solution:

$$
\begin{aligned}
Q &= mc\Delta T \\
&= 250 \times 4.15 \times (100 - 20) \\
&= 83 \text{ kJ}
\end{aligned}
$$

Beaches are popular places on hot days because of water's cooling properties. On hot days, people go to the beach because the shore is cooler than inland areas. The water can absorb heat from its warmer surroundings so that the air temperature by the seashore is lowered.

The higher the specific heat capacity of a substance, the more energy it must absorb to raise its temperature. The specific heat capacity of water is almost five times higher than that of soil particles or rocks. This makes it an excellent insulator of heat. Water also retains heat five times more effectively than land, but may take five times longer than land to heat up.

The high specific heat capacity of water is also important for life in the oceans and for climate moderation. It means that water effectively acts as a buffer against large fluctuations in temperature. This is due to its ability to absorb large amounts of heat energy with only a minimal increase in temperature. The temperature of marine environments is thus made more stable, so that organisms can live in optimal conditions. On a global scale, extremes in temperature are evened out due to oceanic water absorbing heat from the atmosphere in the tropics and releasing it when ocean currents take this water to the polar regions.

The specific heat capacity of a substance is the amount of energy required to raise one gram of it by 1 °C. It is therefore a measure of how easy or how difficult it is to change the temperature of a substance. Water has a high specific heat capacity (4.15 J g^{-1} C^{-1}), which means that its temperature is difficult to alter.

TABLE 12.1 Specific heat capacity of some common substances

Substance	Specific heat capacity (c) (J g^{-1} °C^{-1})
water	4.15
ethanol	2.5
vegetable oil	2.0
soil (wet)	1.5
aluminium	0.90
concrete	0.88
glass	0.84
sand (dry)	0.80
soil (dry)	0.80
copper	0.39
gold	0.13
lead	0.13

Note that the symbol for specific heat capacity is *c*.

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Phase changes
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Density

The density of a substance is a measure of the amount of mass that is contained in a certain volume of that substance. It is calculated by using the formula $d = m/v$ (where m is the mass of the substance and v is its volume). Typical units are g cm^{-3} and kg m^{-3}. The official symbol for density is ρ, although d is still commonly used.

The density of water varies across its three states due to expansion and contraction and this depends on its temperature. Table 12.2 shows the density of ice and water at selected temperatures.

Water has an unusual property: its solid form, ice, is less dense than its liquid form. The molecules in the solid form are arranged in a more structured hexagonal way. This means that ice floats on water and that water freezes from the surface down.

TABLE 12.2 Density of ice and water at selected temperatures

Ice		Liquid water	
Temperature (°C)	Density (ρ) (g cm^{-3})	Temperature (°C)	Density (ρ) (g cm^{-3})
–50	0.922	0	0.999 84
–40	0.921	2	0.999 94
–30	0.920	4*	0.999 98
–20	0.919	6	0.999 94
–10	0.919	8	0.999 85
0	0.916	10	0.999 7
		20	0.998 2
		25	0.997 1
		30	0.995 7

*Water has its highest density at 3.98 °C.

Solid ice floating on top of a lake supports the weight of this team of dogs and its sled.

An inspection of table 12.2 reveals two interesting facts about water. The first is that it has its maximum density not at its freezing point but at 4 °C. This means that water that is approaching 0 °C, and is therefore close to freezing, floats to the top as it is less dense than the surrounding water. The second is that ice has a density that is significantly less than water at the temperatures shown. This is the reason that ice floats on water.

The fact that water freezes with a layer of ice across the top is attributable to both these properties. The water that is about to freeze is at the surface and, when it does turn into ice, its lower density keeps it there. The layer of ice that forms then acts as an insulator, preventing the water below from freezing. This means that aquatic life can survive, even in sub-zero conditions. If ice were denser than water, some bodies of water would freeze solid during winter and the aquatic life would die.

This unusual property of ice also explains why bottles full of water crack if they are allowed to freeze in a refrigerator. The lower density of ice compared with water means that a given mass of water expands when it freezes. This resulting pressure breaks the bottle. The same principle can be an important weathering phenomenon, where water freezes after seeping into cracks in rocks. As it freezes, enough pressure can build up to crack the rocks and contribute to their eventual erosion.

Icebergs form when sections of the frozen icecap break off and float in the sea. When ice forms on the surface of water it acts as an insulator, preventing the water below from freezing. This means that aquatic life can survive, even in sub-zero conditions. If ice were denser than water, some bodies of water would freeze solid during winter and the aquatic life would die.

Electrical conductivity

A substance conducts electricity only if it contains charged particles that are able to move. The covalent bonding in water, which results in the formation of neutral molecules, means that it should not conduct electricity — even though these molecules move freely in the liquid state. However, pure water does conduct electricity very slightly. The amount is so tiny that it is still often described as a non-conductor. The reason for this is the ability of water molecules to react slightly with themselves to form a small number of ions. This is called the self-ionisation of water and is discussed more fully in chapter 14. The equation for this reaction is:

$$H_2O(l) + H_2O(l) \longrightarrow H_3O^+(aq) + OH^-(aq)$$

The small extent of this reaction means that only a few ions are formed and the resulting conductivity is virtually zero.

This fact may surprise a number of people. We are all accustomed to warnings that water and electricity do not mix and have seen such warnings on appliances such as hair dryers and bathroom heaters. These warnings are necessary because the water that we use every day is never pure, but contains other substances dissolved within it. Often these substances are ionic or react to form ions in the water. Once such ions are present, the water conducts freely as the charged particles are able to move.

Hydrogen bonding and the properties of water

Water molecules are not only polar, but also contain hydrogen–oxygen bonds, creating hydrogen bonding between the water molecules. This bonding is the reason for the important properties of water.

Water is a small molecule that contains only three atoms. It would therefore be expected that the dispersion forces between its molecules are quite weak. While this is true, there is also another type of force between water molecules. This is hydrogen bonding and, although it is certainly a lot weaker than the covalent bonds that hold the molecule together, it is nevertheless much stronger than the dispersion forces mentioned above.

When a molecule consists of atoms with different electronegativities, the bonding electrons are attracted toward the atom with the stronger electronegativity. The atom gaining a greater share of the bonding electrons becomes slightly more negatively charged than the atom that had the lower electronegativity. That atom then becomes slightly positively charged.

Differences in charges at either end of a molecule produce a polar molecule. These slight negative and positive regions cause the molecules to attract other polar molecules quite strongly. The stronger the attraction, the closer the molecules can get, and the stronger the resulting intermolecular bond.

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eLesson
Hydrogen bonding
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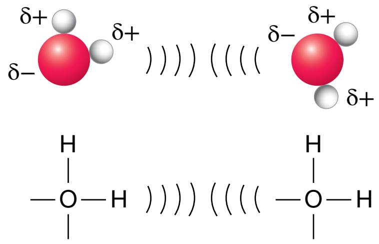
A special case exists when the atoms making up the molecule consist of hydrogen bonded to oxygen, fluorine or nitrogen. The high electronegativity difference causes the hydrogen's electron to move toward the non-metal, resulting in a polar molecule. When this molecule approaches another polar molecule, the slightly positive hydrogen end moves slightly toward the negative end of the other molecule. Hydrogen has only one electron, and when this electron has moved toward the non-metal making up the molecule, there are no other electrons left to shield the hydrogen nucleus from the negative region it is approaching. Such molecules can therefore attract other molecules very closely, resulting in a stronger bond than would be the case if hydrogen were not involved. This extra-strong intermolecular force is called a **hydrogen bond**.



The hydrogen bond attraction between two water molecules. These diagrams are different representations.

In water the effects of this hydrogen bonding can be seen clearly when we compare melting and boiling temperatures for other hydrides in group 16. These changes of state both involve overcoming intermolecular forces that exist between the appropriate molecules. Melting and boiling temperatures can therefore indicate the comparative strengths of these forces.

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Unit 2
AOS 1
Topic 1
Concept 1

Melting and boiling points of water
Summary screen and practice questions

TABLE 12.3 Melting and boiling temperatures of group 16 hydrides

Hydride	Formula	Molar mass (g mol^{-1})	Melting temperature (°C)	Boiling temperature (°C)
water	H_2O	18	0	100
hydrogen sulfide	H_2S	34	−86	−61
hydrogen selenide	H_2Se	81	−65	−42
hydrogen telluride	H_2Te	129.6	−51	−2

If either melting or boiling temperatures are compared, it seems that water is an anomaly. If water is ignored, there is an obvious trend of increasing temperatures as one goes down the group. The larger the molecule, the higher are its melting and boiling temperatures. This is exactly what would be expected if dispersion forces were the only significant intermolecular force present. The values for water, however, indicate that the intermolecular forces are much stronger than expected for just dispersion forces alone. This is because hydrogen bonding is also present as an additional, much stronger, force between the molecules. Water contains hydrogen–oxygen bonds, which, as mentioned previously, allow for the formation of comparatively strong permanent dipoles that are called hydrogen bonds.

Effects of hydrogen bonding on the properties of water

Hydrogen bonding significantly affects the physical properties of a substance. For example, the existence of hydrogen bonding between water molecules results in:

- *the relatively high melting and boiling points of water compared with similarly sized molecular substances*. More heat is required to enable the molecules to gain sufficient kinetic energy to break free of the hydrogen bonds, which are stronger than dispersion forces alone.
- *the expansion of water upon freezing*. The open, hexagonal crystalline lattice of ice places the water molecules further apart than occurs in the liquid state. Because water expands on freezing, it is less dense as a solid. Ice therefore floats on water.

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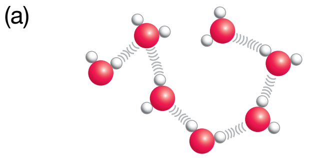
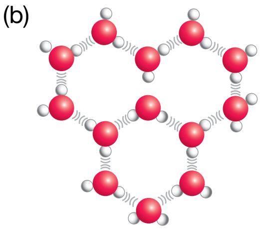
Water and ice structure



Water molecules move too quickly when in liquid form to form regular structures; however, they still remain attached.



When in solid form (ice), water molecules join together to form hexagonal structures.

Hydrogen bonding in water

- *the relatively high latent heat values for fusion and evaporation of water*. Because hydrogen bonding is a significant intermolecular force in water, more energy is required for the necessary changes in arrangement between molecules when either melting or evaporation occurs.
- *the relatively high specific heat capacity of liquid water*. This can also be explained by the higher than usual forces of attraction between molecules caused by hydrogen bonding. Temperature is related to the average kinetic

Unusual properties of water include its relatively high melting and boiling temperatures and its expansion upon freezing. Its ability to form hydrogen bonds also explains why it is such a good solvent for other polar substances as well as many ionic substances.

energy of the particles involved and, for particles that are all the same (i.e. water molecules), it is in turn related to their velocities. The higher than usual attractions between water molecules mean that more energy is required to speed them up and thus increase the temperature.

- *the ability of water to dissolve other substances*. Water's ability to form hydrogen bonds gives it a very strong attraction to molecules such as sugar and alcohol. Although these compounds are discrete molecules, they dissolve very easily in water because they also form hydrogen bonds with the water molecules. In fact, sugar is so soluble in water that it continues to dissolve until eventually syrup is formed.

Revision questions

1. **The kinetic theory of matter describes strong forces of attraction pulling molecules close together as they slow down upon cooling, with the substance contracting. However, water expands upon freezing. Explain why water behaves differently.**
2. **Explain why perspiring on a hot day is an effective mechanism for cooling warm bodies.**

Water as a solvent

Tap water contains varying amounts of dissolved minerals and gases. Chemically pure water never exists in nature because water dissolves so many substances.

Water is sometimes called '**the universal solvent**', because it dissolves so many different substances — in particular, salts. The issue raises a number of questions:

- Why does water dissolve salts?
- Why does oil not dissolve in water?
- What substances dissolve in oil?

To answer these, we must look at the forces that hold the molecules of water and each of the other substances together.

Water is a molecule containing an atom of oxygen bonded to two hydrogen atoms with covalent bonds. Each molecule of water has a slightly negative region (represented by the symbol $\delta-$, 'delta negative') where the oxygen atom is, and a slightly positive region (represented by the symbol $\delta+$, 'delta positive') where the hydrogen atoms are. In other words, water molecules are polar.

Solubility

Solubility measures the extent to which a solute dissolves in a solvent. Whether a solute dissolves can be predicted by the 'like dissolves like' rule. A polar solvent generally dissolves a polar solute. Non-polar substances usually dissolve in non-polar solvents.

Dissolving occurs when one substance is pulled apart by another, or moves into the spaces within another. The **solubility** (that is, the extent to which a **solute** can dissolve in a **solvent**) of one substance in another is determined to a large extent by the polar or non-polar nature of the solute and the solvent. This situation demonstrates the '**like dissolves like**' rule used by chemists, namely that, if both the solvent and solute particles contain charged regions, it is likely that the solute will dissolve. For example, if the solute (the substance being dissolved) is polar, it is more likely to dissolve in a solvent (the substance doing the dissolving) that is also polar. It is very unlikely that a non-polar solute will dissolve in a polar solvent because there is nothing for the polar molecules to attract. Conversely, a non-polar solute dissolves in a non-polar solvent because neither has strong forces holding their molecules together. It dissolves because its molecules become more disordered.

Ionic crystals (salts) are made up of positive and negative ions. They are more likely to dissolve in polar solvents than non-polar ones. Water is an excellent solvent for many ionic and polar substances. When a salt is placed in water and dissolves, the positive and negative ions become surrounded by water molecules, which pull them apart. When the individual ions are surrounded by water, they are said to be hydrated ions.

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Unit 2
AOS 1
Topic 1
Concept 4

Dissolving ionic substances
Summary screen and practice questions

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Unit 2
AOS 1
Topic 1
Concept 5

Dissolving molecular substances
Summary screen and practice questions

Summarising the behaviour of substances in water:

- They may remain undissolved. These substances are insoluble.
- Ionic substances (e.g. sodium chloride) may dissociate. This means that water pulls the ions from the solid structure and they then move apart and spread throughout the water to form a solution. These substances are soluble. Ions present become surrounded by water molecules. The resulting solution conducts electricity.
- Polar molecules may react with water to form ions. This is called ionisation, and is a feature of many acids and bases. Substances that ionise in water conduct electricity.
- Other polar molecules, although they do not react with water as above, may still form hydrogen bonds with water and undergo dissociation. These substances are also soluble. Sugar and ethanol are in this category. These solutions do not conduct electricity.

Solubility of gases in water

Polar gases dissolve easily in water. For example, both ammonia gas, NH_3, and hydrogen chloride gas, HCl, dissolve in water. Both of these gases are composed of polar molecules.

When the pressure is released on a bottle of carbonated drink, the dissolved carbon dioxide gas can no longer remain dissolved, and so it bubbles out. Eventually, the drink becomes ‘flat’.

A gas such as oxygen or carbon dioxide is made up of molecules that are non-polar. These **non-polar gases** do not dissolve as easily. Only a very small amount of oxygen gas dissolves in water, but this is enough for cellular respiration in plants and fish. Likewise, the small amount of dissolved carbon dioxide in water enables photosynthesis to take place in aquatic plants. An increase in pressure on the gas–water mixture causes more of the gas to dissolve in the water. Carbon dioxide is dissolved in water under pressure to make ‘carbonated’ drinks. When the top is removed from a bottle of carbonated drink, the pressure is released and the gas starts coming out of solution.

Temperature also influences the amount of gas that dissolves in water. An increase in temperature generally decreases the solubility of gases in liquids.

Air bubbles can be trapped in glacial ice for thousands of years.

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Unit 2
AOS 1
Topic 1
Concept 8

Aspects of water as a solvent
Summary screen and practice questions

Using water as a solvent

At the start of this chapter we saw that water has a number of special properties. These properties alone would make it a very useful and important chemical. In addition, as we have just seen, it is a very good solvent. This fact makes it even more useful in a wide range of contexts. Some of these include:

- industrial applications. Many chemicals are more easily handled and transported as solutions than in their pure states. Water facilitates many reactions by providing a convenient medium to bring reactants together. In electrolysis reactions, solutions in water often provide a lower temperature alternative to using molten materials to achieve conductivity. From a safety point of view, chemicals can be diluted so that they pose less of a risk — larger volumes of more dilute solutions can be used instead of smaller volumes of more concentrated and potentially dangerous solutions. Water can be used to conveniently capture and dissolve by-products of industrial reactions, allowing them to be turned into useful products or to be processed further before release into the environment.
- biological applications. The bodies of most organisms are made up of a large percentage of water. Humans, for example are about 60% water. Blood plasma and cells contain dissolved substances such as amino acids, proteins, mineral ions, various types of sugars, digestion products and hormones to name just a few. There are also waste products such as urea and carbon dioxide present. In plants, water in the cells and sap has a wide range of dissolved materials including nutrients, amino acids and sugar produced from photosynthesis.
- domestic applications. In the garden, many pesticides and nutrients are dissolved in water before being applied. Vinegar is a solution of ethanoic acid. Beverages contain a number of solutes responsible for their flavours. Aerated drinks contain dissolved carbon dioxide. In the medicine cabinet are many aqueous solutions or products that are dissolved in water before being taken. Personal care products such as perfumes and mouthwashes can be added to this list.

A special mention should also be made of soaps and detergents. These products are part of a larger group called surfactants. These contain molecules that are polar at one end and non-polar at the other. This enables them to dissolve in both water and oil, making them ideally suited as cleaning agents where oil and grease need to be removed.

Mixing solutions

When solutions are mixed, chemical reactions can occur between their dissolved components. A major use for water is, in fact, dissolving substances so that reactions between them can be carried out conveniently and at low cost. Neutralisation and precipitation reactions are carried out in water. Electrolysis reactions (such as electroplating) require that a substance conducts electricity and they often involve ionic substances. Such substances can be made to conduct either by melting them (which requires high temperatures and cost) or by dissolving them in water. The latter, in many cases, provides a more convenient and lower cost alternative.

Neutralisation reactions involve acids and bases. For example, antacid remedies for heartburn and indigestion contain a base, often sodium hydrogen carbonate (also known as sodium bicarbonate). On its own, the antacid is stable. However, once it is combined with acid in the stomach, a reaction occurs that produces carbon dioxide (as well as water and a salt) and reduces acidity in the stomach. These types of reactions are discussed further in chapter 13.

When solutions are mixed, chemical reactions may occur. Two common such reactions are neutralisation reactions and precipitation reactions. A precipitate is formed when a new combination of ions produces an insoluble substance.

Precipitation reactions are used in many areas of everyday life. Water authorities often add chemicals such as alum to their water to deliberately bring about precipitation. As the newly formed precipitate settles, it also traps very fine suspended particles from the water, thus effectively removing them and making the water much clearer. A second common use is the removal of 'hardness' from water by adding hydrated sodium carbonate (also known as washing soda). Ions that cause hardness, such as Mg^{2+} and Ca^{2+}, react with the carbonate ions to form the corresponding insoluble carbonate. The ions are removed from the water and cannot react with soap to form a scum.

The ionic equations for these two reactions are:

$$Mg^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow MgCO_3(s)$$

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow CaCO_3(s)$$

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Unit 2	**Precipitation reactions** Summary screen and practice questions
AOS 1	
Topic 1	
Concept 6	

Precipitation reactions

Although water is an excellent solvent for many ionic substances, there are some that it cannot dissolve. For substances such as these, water is unable to dissociate the ionic lattice structure and they are therefore insoluble. As we have seen briefly in chapter 11, when different solutions containing dissolved ions are mixed, it is sometimes possible that they will form an insoluble combination. This produces solid particles that eventually settle out as a precipitate. When equations for reactions such as these are written, the substance that forms the precipitate is easily recognised by examining the symbols of state. For example, when a solution of sodium chloride is mixed with a silver nitrate solution, silver chloride and sodium nitrate are formed. By examining a list of solubility rules (on the next page) or tables such as table 11.2, it becomes apparent that the precipitate observed must be silver chloride. The equation can therefore be written:

$$NaCl(aq) + AgNO_3(aq) \longrightarrow NaNO_3(aq) + AgCl(s)$$

Sample problem 12.2

By reference to the solubility rules, write equations for the following reactions.

(a) Sodium carbonate and magnesium chloride solutions are mixed.
(b) Barium nitrate and sodium sulfate solutions are mixed.

Solution: (a)

STEP 1

Predict the products and identify whether one of these is a precipitate. The predicted products are sodium chloride and magnesium carbonate, and the precipitate is magnesium carbonate.

STEP 2

Write the unbalanced equation using appropriate formulas.

$$Na_2CO_3 + MgCl_2 \longrightarrow NaCl + MgCO_3$$

STEP 3

Balance the equation.

$$Na_2CO_3 + MgCl_2 \longrightarrow 2NaCl + MgCO_3$$

STEP 4

Add the symbols of state.

$$Na_2CO_3(aq) + MgCl_2(aq) \longrightarrow 2NaCl(aq) + MgCO_3(s)$$

(b)

STEP 1

Predicted products are sodium nitrate and barium sulfate. The precipitate is barium sulfate.

STEP 2

Write the unbalanced equation using appropriate formulas.

$$Na_2SO_4 + Ba(NO_3)_2 \longrightarrow NaNO_3 + BaSO_4$$

STEP 3

Balance the equation.

$$Na_2SO_4 + Ba(NO_3)_2 \longrightarrow 2NaNO_3 + BaSO_4$$

STEP 4

Add the symbols of state.

$$Na_2SO_4(aq) + Ba(NO_3)_2(aq) \longrightarrow 2NaNO_3(aq) + BaSO_4(s)$$

Rules for solubility

The following rules classify salts according to their solubility:

1. All salts of sodium, Na^+, potassium, K^+, and ammonium, NH_4^+, are soluble.
2. All nitrates, NO_3^-, are soluble.
3. All chlorides, Cl^-, bromides, Br^-, and iodides, I^-, are soluble except compounds formed with silver and lead.
4. All sulfates, SO_4^{2-}, are soluble except compounds formed with silver, lead, calcium, strontium and barium. $AgSO_4$ and $CaSO_4$ are borderline.
5. All phosphates and carbonates, PO_2^{3-} and CO_3^{2-}, are insoluble except compounds formed with group 1 elements and ammonium.
6. All hydroxides, OH^-, are insoluble except compounds formed with group 1 elements, ammonium and barium. Calcium hydroxide is borderline.
7. All sulfides are insoluble except compounds formed with group 1 and group 2 elements and with ammonium ions.

These rules imply that compounds are either soluble or insoluble. This is not the case. They are all soluble to a greater or lesser degree. These rules simply help us to predict generally which salts are more likely to dissolve in water than others.

The water in the Dead Sea contains a high concentration of dissolved salts that increases the density of the water. This means it is very easy to float in.

Revision questions

3. Use the rules for solubility to decide which of the following substances would be soluble in water.

(a) potassium nitrite	(d) calcium carbonate
(b) sodium chloride	(e) ammonium hydroxide
(c) lead sulfate	(f) copper(II) sulfide

4. Explain why water can act as a solvent for many salts.

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Experiment 12.1 Testing for solubility
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Balancing ionic equations

Ionic equations

In all the preceding equations, the elements and compounds have been written in their molecular or formula unit forms. However, in aqueous solutions, the reactions are best represented by simpler equations called ionic equations. Ionic equations are equations that show only the species that are formed or changed in a reaction. Any ions that remain unchanged in a reaction are not included in an ionic equation. Ions that are present in a reaction but do not react are called **spectator ions**.

The rules for writing ionic equations are as follows:

1. Write the balanced chemical equation.
2. Decide, from the solubility table, which substances are soluble and which will form precipitates.
3. Expand the equation by dissociating all the soluble compounds into their free ions.
4. Check for any molecular substances such as acids and certain bases that react with water to produce ions (hydrolyse). Replace the formulas of these substances by the ions that they form.
5. Cancel all free ions that are unchanged on both sides of the equation (the spectators).
6. Write the net ionic equation.

Note: The equation must be balanced in charge as well as in the number of atoms.

Precipitation and neutralisation reactions are often represented by ionic equations.

Yellow lead(II) iodide precipitates in a solution of sodium nitrate.

Writing ionic equations for precipitation reactions

When a salt is dissolved in water it breaks up, or dissociates, into its constituent ions:

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

The symbol (aq) indicates that each ion has become surrounded by water molecules.

If a second soluble salt, silver nitrate, is added, it also dissociates:

$$AgNO_3(s) \xrightarrow{H_2O} Ag^+(aq) + NO_3^-(aq)$$

When the ions of both solutions come into contact, a white precipitate of AgCl is formed.

$$Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \longrightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$$

Precipitates are insoluble salts and so are not dissociated in solution.

The net ionic equation does not include the 'spectator' ions, Na^+ and NO_3^-; these remain unchanged and dissolved in solution (as if sitting on the sidelines watching).

$$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$$

Sample problem 12.3

Write the ionic equation for the reaction between lead(II) nitrate solution and sodium iodide solution.

Solution:

STEP 1

$Pb(NO_3)_2(aq) + 2NaI(aq) \longrightarrow PbI_2(s) + 2NaNO_3(aq)$

STEP 2

From the solubility table, all nitrates are soluble, so the precipitate is PbI_2.

STEPs 3 and 4

Expand the equation and then cancel any spectator ions.

$Pb^{2+}(aq) + \cancel{2NO_3^-}(aq) + \cancel{2Na^+}(aq) + 2I^-(aq) \longrightarrow PbI_2(s) + \cancel{2Na^+}(aq) + \cancel{2NO_3^-}(aq)$

STEP 5

$Pb^{2+}(aq) + 2I^-(aq) \longrightarrow PbI_2(s)$

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Unit 2
AOS 1
Topic 1
Concept 7

Writing ionic equations for precipitation reactions
Summary screen and practice questions

Revision questions

5. **Write dissociation equations for (a) $NaOH(s)$ and (b) $Na_3PO_4(s)$.**
6. **Write ionic equations for the following reactions.**
   - **(a)** **barium chloride and magnesium sulfate solutions reacting to form barium sulfate and magnesium chloride**
   - **(b)** **calcium nitrate reacting with sodium carbonate in solution to form solid calcium carbonate and sodium nitrate solution**

Chapter review

studyon

Unit 2
AOS 1
Topic 1

Properties of water
Summary screens and practice questions

Summary

- Water is a polar, discrete molecular compound, covering about 71% of the Earth's surface, where it can be found in all three states — solid, liquid and gas.
- Water is essential for life, and the water cycle produces a continual supply.
- The properties of water include:
  - a high latent heat of fusion — the latent heat of fusion is the amount of energy needed to change a fixed amount of water from the solid to the liquid state.
  - a high latent heat of vaporisation — the latent heat of vaporisation is the amount of energy needed to change a fixed mass of water from the liquid to the gaseous state.
  - a high specific heat capacity — the specific heat capacity is the amount of energy needed to raise the temperature of one gram of water by 1 °C.
  - a melting point (of 0 °C) and a boiling point (of 100 °C).
  - a higher density in the liquid state than in the solid state.
  - non-conductor of electricity when pure. Impure water, however, can be an excellent conductor due to the presence of dissolved substances.
- The values for the above properties of water are higher than for other substances of similar molecular size due to water's strong intermolecular forces — especially hydrogen bonding. A hydrogen bond is formed only when the atoms making up the molecule consist of hydrogen bonded to oxygen, fluorine or nitrogen, which can attract each other very closely.
- Water expands upon freezing. Marine life can survive in winter because the density of ice is less than liquid water. The ice floats on water due to its lower density, providing an insulating layer that prevents the water below from freezing.
- Water is sometimes called 'the universal solvent'. This is because it is an excellent solvent for many ionic and polar substances. This is due to its polarity and ability to form hydrogen bonds. Water is therefore attracted to ions and polar substances. If these attractions are strong enough to overcome the attractions between the molecules or ions of the other substance, that substance dissolves (the 'like dissolves like' rule).
- A solute is a substance that is dissolved in a solvent to make a solution.
- Solubility is the extent to which a solute can dissolve in a solvent and this depends on the temperature.
- Any salts that contain any of the ions Na^+, K^+, NH_4^+ and NO_3^- are soluble.
- Polar gases such as ammonia dissolve readily in water.
- Non-polar gases such as oxygen and nitrogen have low solubility, but enough to sustain aquatic life.
- An increase in pressure generally increases the solubility of a gas in liquids.
- An increase in temperature generally decreases the solubility of gases in liquids.
- Dissociation is the process by which particles in the solid state are pulled apart by water and become able to move around. This results in the solid dissolving and a solution being formed.
- Precipitation reactions occur when solutions of soluble substances are mixed and an insoluble substance is formed.
- An ionic equation is one that shows only the ions that are formed or changed during a chemical reaction. Those that do not react are known as spectator ions and are not shown. Ionic equations are particularly useful for summarising precipitation reactions.

Multiple choice questions

1. Which of the following is not a physical property of water?
A It freezes at 0 °C.
B It has a high specific heat capacity.
C It readily dissolves salts.
D It is odourless.

2. Which of the following cannot be explained by the existence of hydrogen bonding in water?
A The relatively higher melting point
B The ability of water to dissolve other substances
C The expansion of water upon freezing
D The shape of the water molecule

3. Which one of the following types of bonding would not be present in a sample of water?
A Ionic
B Covalent
C Dispersion
D Hydrogen

4. The structure of ice is a lattice of hexagonal crystals. The bonding that holds the water molecules in the lattice together is best described as:
A hydrogen bonding only
B dispersion forces and hydrogen bonding
C ionic bonding
D covalent bonding.

5. In which of the following substances would you not expect to find hydrogen bonding?

A hydrogen fluoride
B water
C methane
D $CH_3CH_2NH_2$

6. The hydrogen bonding in water is an example of:

A intermolecular bonding
B intramolecular bonding
C covalent bonding
D ionic bonding.

7. Which of the following terms best describes a substance that dissolves when added to water?

A A solvent
B A solute
C Insoluble
D Saturated

8. Which of the following includes only water-soluble substances?

A ammonium chloride, magnesium nitrate, sodium hydroxide, sodium sulfide
B ammonium nitrate, magnesium hydroxide, ammonium hydroxide, sodium chloride
C magnesium chloride, sodium nitrate, silver chloride, potassium chloride
D potassium nitrate, magnesium nitrate, copper sulfide, copper nitrate

9. Which of the following compounds would appear as a precipitate in solution?

A ammonium sulfate
B barium nitrate
C barium sulfate
D ammonium hydroxide

10. Which of the following is a correctly balanced *ionic* equation?

A $Ba(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$
B $HSO_4^-(aq) + H_2O(l) \longrightarrow SO_4^{2-}(aq) + H_3O^+(aq)$
C $Na^+(aq) + 2Cl^-(aq) \longrightarrow NaCl_2(s)$
D $2H^+(aq) + SO_4^{2-}(aq) + 2K^+(aq) + 2OH^-(aq) \longrightarrow 2K^+(aq) + SO_4^{2-}(aq) + 2H_2O(l)$

11. In which one of the following instances does a chemical reaction *not* occur?

A Solutions of ammonium chloride and lead nitrate are mixed.
B Solutions of magnesium bromide and sodium hydroxide are mixed.
C A piece of silver is dropped into dilute hydrochloric acid.
D Pentane, C_5H_{12}, is ignited.

Review questions

1. Water is important for life. Keep a 24-hour record of how much water you use:

- during a school day
- during Saturday or Sunday.

Record the purpose and quantity of water used, then answer the following questions.

(a) Determine how much of the water, as a percentage, could be conserved.
(b) Does your water usage vary during a weekday as compared with a weekend?
(c) Would you expect water usage to vary between males and females or between young and old people? How?

Physical properties

2. (a) List the properties of water.
(b) The properties of water can be explained by the nature of the intermolecular forces between water molecules. Explain.

3. Explain why water is different from most liquids with respect to its density in the solid and liquid states.

4. (a) What is latent heat?
(b) What is specific heat capacity?
(c) Explain how water is unique with respect to these two physical properties.

5. If ice were denser than water, life as we know it would not exist. Explain this statement.

Solvent properties

6. Chemists often refer to the 'like dissolves like' rule when describing the solubility of various substances in different liquids. How does this rule explain that:

(a) potassium chloride dissolves in water
(b) ammonia dissolves in water
(c) methane is insoluble in water?

Ionic equations

7. Identify the substances from the following list that are soluble in water.

(a) CaI_2
(b) KOH
(c) AgCl
(d) AgI
(e) $BaSO_4$
(f) NH_4Cl
(g) Na_2S
(h) $MgCl_2$
(i) CuS
(j) $Al(OH)_3$

8. Predict whether a precipitate forms when solutions of the following are mixed.

(a) $AgNO_3$ and KCl
(b) H_2SO_4 and $BaCl_2$
(c) H_2SO_4 and NaCl
(d) $NaNO_3$ and KCl
(e) $(NH_4)_3PO_4$ and $CaCl_2$
(f) $(NH_4)_2S$ and $Pb(NO_3)_2$
(g) $(NH_4)_2S$ and $NaNO_3$
(h) $CaCl_2$ and NaBr

9. Write ionic equations for the reactions that form precipitates in question 8.

10. Write a molecular equation and an ionic equation for each of the following reactions (if no reaction occurs, write 'no reaction').
- (a) Sodium hydroxide reacts with nitric acid.
- (b) Solutions of potassium sulfate and calcium nitrate are mixed.
- (c) Lead(II) nitrate solution is added to potassium iodide solution.
- (d) Dilute hydrochloric acid is mixed with sodium hydrogen carbonate solution.
- (e) Magnesium chloride is added to sodium phosphate solution.
- (f) Solutions of potassium sulfide and zinc chloride are mixed.
- (g) Ammonium carbonate solution is added to dilute nitric acid.
- (h) Solutions of sodium sulfate and potassium bromide are mixed.

11. What cations could be present in a solution that:
- (a) gave a precipitate with sodium sulfate solution but not with sodium chloride
- (b) gave a precipitate with sodium sulfate and with sodium chloride
- (c) gave a precipitate with sodium sulfate solution but not with sodium hydroxide
- (d) gave a precipitate with sodium carbonate but not with sodium hydroxide?

12. Describe simple tests that would allow you to distinguish between solutions of:
- (a) silver nitrate and zinc nitrate
- (b) barium chloride and magnesium chloride
- (c) silver nitrate and lead nitrate
- (d) potassium sulfate and potassium chloride.

Exam practice questions

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Unit 2
AOS 1
Topic 1

Properties of water

Sit topic test

In a chemistry examination you will be required to answer a number of short and extended response questions.

Extended response questions

1. Some ammonium carbonate solution is poured into two separate test tubes. Some sodium chloride solution is then added to one of these test tubes, and some barium chloride solution is added to the other test tube.

(a) In which test tube would a precipitate be observed to form? **1 mark**

(b) Write the formula equation for the formation of the precipitate that you have identified from part (a). **2 marks**

(c) Hence, write the ionic equation for this precipitation reaction. **1 mark**

(d) What are spectator ions? In the example from part (c), list the spectator ions that are present. **2 marks**

(e) Suppose that the precipitate is filtered and the remaining liquid (the filtrate) is then heated to evaporate the remaining water. Describe what happens, naming any new substances that might be formed. **2 marks**

2. Table 12.1 (page 222) shows the specific heat capacities of a range of substances.

(a) Define the term 'specific heat capacity'. **1 mark**

(b) If an equal amount of heat energy is added to 10 g of water and to 10 g of sand, which substance would experience the higher increase in temperature? **1 mark**

(c) Calculate the energy required to raise 80 g of vegetable oil by 70 °C. **2 marks**

3. Solutions of sodium chloride and hydrochloric acid are both good conductors of electricity. This means that they both have mobile ions present.

(a) Define the term 'dissociation'. **1 mark**

(b) Define the term 'ionisation'. **1 mark**

(c) Explain why a solution of sodium chloride can conduct electricity. Use the appropriate term from parts (a) and (b) in your explanation. **2 marks**

(d) Explain why hydrochloric acid can conduct electricity. Use the appropriate term from parts (a) and (b) in your explanation. **2 marks**

CHAPTER

13 Reactions in water

Write your injuries in dust, your benefits in marble.

Benjamin Franklin

YOU WILL EXAMINE:

- the importance of acids and bases and their reactions in the household and the environment, and examples of common weak and strong acids and bases
- properties of acids and bases
- the Brønsted–Lowry theory of acid–base behaviour
- conjugate acid–base pairs
- polyprotic acids and amphiprotic substances
- pH as a measure of acidity of a solution
- balanced chemical equations for reactions involving acids and bases including their reactions wtih water
- pH calculations for strong acids and bases and for solutions that have been diluted
- the distinction between strong and weak acids and bases, and between concentrated and dilute solutions
- the use of pH measurements to determine relative strengths of acids (and bases)
- the environmental issues of acid rain and ocean acidification
- oxidation and reduction (redox) reactions
- the electron transfer theory for redox reactions
- the representation of redox reactions by half-equations involving conjugate redox pairs, and by combined overall ionic equations
- the reactivity series of metals and metal displacement reactions
- corrosion of metals and treatments to control it.

The burning, cramping stitch felt in muscles during periods of intense exercise is caused by a build-up of acidic hydrogen ions produced by chemicals in the muscles. The body does not take in enough oxygen to keep up with the demands of the increased activity so the natural acidity regulatory buffering system in the muscle is overcome and the pH drops, causing pain.

Two new groups of reactions

Chemists make sense out of the huge number of chemical reactions by classifying them into groups. Examples of such groupings are polymerisation reactions, which you have already studied in Unit 1, and precipitation reactions, which you met in chapters 11 and 12. Often these groups are so large that they are further divided into smaller groups, as is the case with the reactions involved in organic chemistry.

Two further such groups are acid–base reactions and redox reactions. Although both of these large groups of reactions can occur in situations without water, they nevertheless occur frequently in and with water. Any study of water chemistry would not be complete without their consideration.

Introducing acids and bases

Acids and bases are used widely in industry and in the home. Their properties, which complement each other, have been known since ancient times.

What can send a shiver down your spine, bring tears to your eyes and make your mouth pucker? Simply the taste of a freshly cut lemon or unripe pineapple. The vinegar in salad dressing, though less stringent, has the same effect. Even a cola or a glass of sparkling wine may leave a slightly sour taste in your mouth. All of these foods have a similar effect because of the presence of compounds called **acids**.

Citrus fruits contain citric acid. This gives lemons and limes their characteristic sour taste.

Hydrochloric acid is found in our stomachs, where it is used to help break down food. It is also used in industry, where it is sometimes called 'spirit of salts', to clean bricks and to clean off the coating of oxide on corroded iron or steel before plating the metal with a protective layer of zinc or tin. Table 13.1 (see page 239) lists common acids and their uses.

Bases are compounds with properties that in some ways complement those of acids. This means that the characteristic properties of an acid can be reduced by adding a base. If enough base is added, the acidic properties completely disappear. When this occurs, we say that the acid has been neutralised by the base. Bases taste bitter, have a soapy feel and are commonly found in cleaning agents. Soap, detergent, oven and bathroom cleaners, toothpaste and shampoo all contain water-soluble bases in varying degrees of strength. Table 13.2 (see page 241) lists some common bases and their uses.

Given the corrosive nature of some of these acids and bases it is important to carefully read the labels on common supermarket cleaning products.

Safety with acids and bases

Always take care when using acids, especially if they are concentrated. To dilute a concentrated acid, always add the acid to the water, never the other way around. Bases are also dangerous and need to be handled with care.

Even so-called weak acids can be dangerous. A concentrated sample of ethanoic acid (acetic acid or vinegar), which is regarded as weak, can be extremely harmful to eyes and mucous membranes. Strong acids such as nitric acid, sulfuric acid and hydrochloric acid are extremely dangerous in concentrated form and should be handled with care even in their more dilute forms. *When diluting a concentrated acid, always add acid to water.* Adding water directly to concentrated acids may cause them to splatter, and severe burns may result.

Strong bases such as sodium hydroxide and potassium hydroxide are also very dangerous, particularly in concentrated form and when hot. Acid and base spills on skin should be rinsed with large quantities of running water.

TABLE 13.1 Common acids and their uses

Acid	Formula	Occurrence or use
nitric	HNO_3	used to etch copper and manufacture fertilisers
hydrochloric	HCl	produced in the stomach; used to clean bricks and concrete
sulfuric	H_2SO_4	used in car batteries, plastics, insecticides, detergents and pharmaceuticals
ethanoic	CH_3COOH	found in vinegar (used in pickling and fermentation of foods)
citric	$C_6H_8O_7$	found in lemon juice
carbonic	H_2CO_3	found in soft drinks
lactic	$C_3H_6O_3$	makes milk sour; used in wrinkle-smoothing face cream
ascorbic	$C_3H_8O_6$	found in citrus fruits

The properties of acids

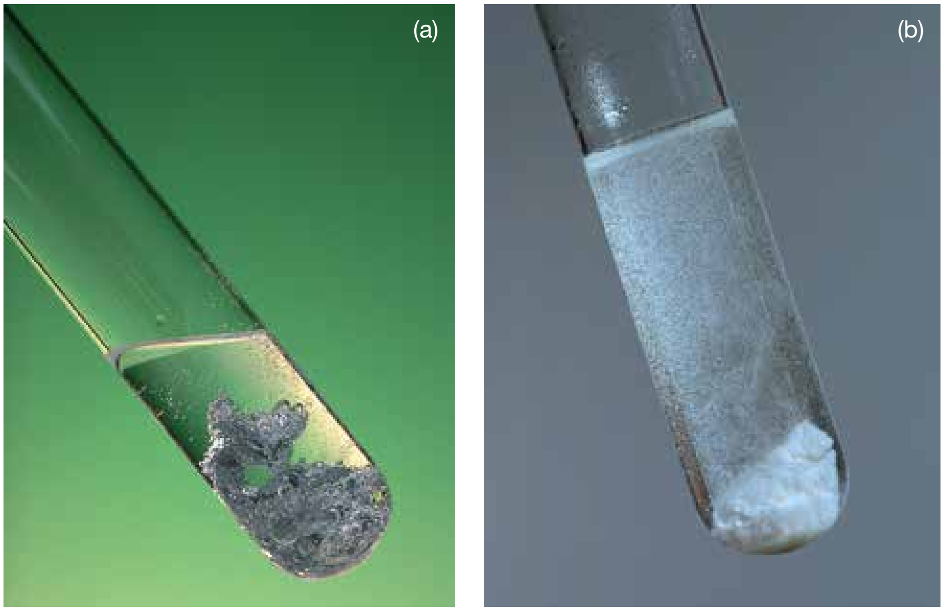
Acids have many common properties. They:

- usually taste sour
- are corrosive
- are molecular in structure and dissolve in water to produce an **electrolyte** (substance that conducts electricity)
- affect the colour of certain natural and synthetic dyes (they turn litmus, a plant dye, from blue to red)
- are neutralised by bases.

Some common acids found in the laboratory include hydrochloric acid, HCl, sulfuric acid, H_2SO_4, and nitric acid, HNO_3. Two common bases are sodium hydroxide, NaOH, and potassium hydroxide, KOH.

Reactions of acids

Some common reactions involving acids and their equations were introduced in chapter 11, and a brief review of these is provided here. Any of the acids H_2SO_4, HCl and HNO_3 can be substituted in the general equations that follow. For each general word equation, an example is given, showing first the molecular and then the ionic equation for the reaction.



(a) Zinc reacts with hydrochloric acid to liberate hydrogen gas.
(b) Limestone, calcium carbonate, reacts with hydrochloric acid to liberate carbon dioxide gas.

The reactions of acids are often predictable. This makes it easy to write equations for many reactions in which acids are involved.

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Unit 2
AOS 1
Topic 2
Concept 6

Common reactions of acids
Summary screen and practice questions

studyon

Unit 2
AOS 1
Topic 2
Concept 7

Writing ionic equations for acid–base reactions
Summary screen and practice questions

1. acid + metal (not Cu, Hg or Ag) ⟶ salt + hydrogen gas

   $H_2SO_4(aq) + Zn(s) \longrightarrow ZnSO_4(aq) + H_2(g)$

   $2H^+(aq) + Zn(s) \longrightarrow Zn^{2+}(aq) + H_2(g)$

2. acid + metal carbonate ⟶ salt + carbon dioxide gas + water

   $H_2SO_4(aq) + CaCO_3(s) \longrightarrow CaSO_4(aq) + CO_2(g) + H_2O(l)$

   $2H^+(aq) + CaCO_3(s) \longrightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(l)$

3. acid + metal hydrogen carbonate ⟶ salt + carbon dioxide gas + water

   $HCl(aq) + NaHCO_3(s) \longrightarrow NaCl(aq) + CO_2(g) + H_2O(l)$

   $H^+(aq) + NaHCO_3(s) \longrightarrow Na^+(aq) + CO_2(g) + H_2O(l)$

4. acid + metal sulfite ⟶ salt + sulfur dioxide gas + water

   $2HCl(aq) + FeSO_3(aq) \longrightarrow FeCl_2(aq) + SO_2(g) + H_2O(l)$

   $2H^+(aq) + SO_3^{2-}(aq) \longrightarrow SO_2(g) + H_2O(l)$

5. acid + metal sulfide ⟶ salt + hydrogen sulfide gas

   $2HCl(aq) + FeS(s) \longrightarrow FeCl_2(aq) + H_2S(g)$

   $2H^+(aq) + FeS(s) \longrightarrow Fe^{2+}(aq) + H_2S(g)$

6. acid + metal oxide (basic oxide) ⟶ salt + water

   $2HNO_3(aq) + MgO(s) \longrightarrow Mg(NO_3)_2(aq) + H_2O(l)$

   $2H^+(aq) + MgO(s) \longrightarrow Mg^{2+}(aq) + H_2O(l)$

7. acid + base (metal hydroxide) ⟶ salt + water

   $H_2SO_4(aq) + 2NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$

   $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

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Experiment 13.1 The reactions of acids
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Revision questions

1. **Write both full and ionic equations for the reactions that result when zinc and HCl are mixed.**
2. **Write both full and ionic equations for the reactions that result when the following substances are mixed. Assume all reactions go to completion.**
   - (a) **magnesium and sulfuric acid**
   - (b) **aluminium and sulfuric acid**
   - (c) **copper(II) carbonate and nitric acid**
   - (d) **potassium hydrogen carbonate and sulfuric acid**
   - (e) **sodium sulfite and nitric acid**
   - (f) **lead(II) sulfide and phosphoric acid, H_3PO_4**
   - (g) **copper(II) oxide and hydrochloric acid**
   - (h) **potassium hydroxide and sulfuric acid**

The properties of bases

Bases have many common properties. They:

- usually taste bitter
- feel slippery (they react with the natural oils in the skin and produce soap, which gives the characteristic ‘slippery feel’)
- turn litmus from red to blue
- are electrolytes (substances that conduct electricity)

- may be corrosive (for example, NaOH, which breaks down organic substances such as fat, hair and vegetable matter) and are good drain cleaners
- are generally ionic substances
- are oxides or hydroxides of metals (ammonium hydroxide is a base)
- are usually insoluble in water. A base that is soluble in water is called an **alkali**.

TABLE 13.2 Common bases and their uses

Base	Formula	Use
ammonia	NH_3	fertilisers and detergents
sodium hydroxide (caustic soda)	NaOH	soaps and detergents
sodium carbonate	Na_2CO_3	manufacture of glass; washing powder and detergents
calcium oxide (quicklime)	CaO	bricklayers' mortar
lead(II) oxide	PbO	house paint (now phased out)
calcium hydroxide (slaked lime)	$Ca(OH)_2$	garden lime, plaster and cement
ammonium hydroxide	NH_4OH	cleaning agents
magnesium hydroxide	$Mg(OH)_2$	milk of magnesia (for treatment of indigestion)

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eLesson
The Brønsted–Lowry theory
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The Brønsted–Lowry theory of acids and bases

An acid–base theory suggested independently by Lowry and Brønsted in 1923 is still used today. This theory is summarised by the following two statements:

- An acid is a proton donor.
- A base is a proton acceptor.

This means that a reaction between an acid and a base involves a proton transfer from the acid to the base. This occurs only when both an acid and a base are present.

The proton described by Brønsted and Lowry is simply a hydrogen ion, H^+. The hydrogen ion is transferred from one substance to another. The substance that loses an H^+ ion is the acid and the one that accepts it is the base.

$$\overbrace{HNO_3(l) + H_2O(l)}^{H^+} \longrightarrow H_3O^+(aq) + NO_3^-(aq)$$

acid loses a proton	base gains a proton

According to the Brønsted–Lowry theory, the water in this equation is acting as a base since it accepts a proton.

The Brønsted–Lowry theory defines acids and bases as follows:

- An acid is a substance that donates a proton (H^+ ion) to another substance.
- A base is a substance that accepts a proton (H^+ ion) from another substance.

Since a hydrogen atom is simply a proton and an electron, removing the electron leaves a proton, H^+.

study on

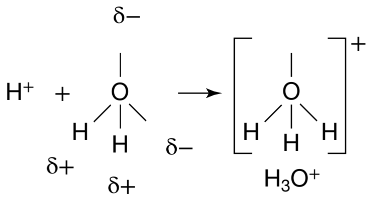
Unit 2	**Acid–base definitions**
AOS 1	Summary screen and practice questions
Topic 2	
Concept 1	

Revision question

3. **Which of the following reactions can be classified as proton-transfer reactions?**
   (a) $HNO_3(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + NO_3^-(aq)$
   (b) $NH_3(aq) + HNO_3(aq) \longrightarrow NH_4^+(aq) + NO_3^-(aq)$
   (c) $Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$

(d) $H_3PO_4(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + H_2PO_4^-(aq)$
(e) $O^{2-}(aq) + H_2O(l) \longrightarrow 2OH^-(aq)$
(f) $Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$
(g) $CO_3^{2-}(aq) + H_2O(l) \longrightarrow HCO_3^-(aq) + OH^-(aq)$

Ionisation and the production of the hydronium ion



Formation of a hydronium ion. Experimental evidence indicates that each proton is associated with a particular water molecule for about 10–11 seconds at room temperature.

The following equation shows hydrochloric acid acting as an acid and donating a proton to water. The reaction is called **ionisation**.

$$HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

Ionisation is a reaction in which a substance reacts with water to produce ions.

Along with the chloride ion, a **hydronium ion** is produced. A proton, H^+, cannot exist by itself as it is attracted to the negative end of the polar water molecule to form a hydronium ion, H_3O^+.

Hydrolysis

When an ionic substance dissolves in water, the resulting solution is often not neutral. This is because the ions can act as acids or bases when reacting with water and produce solutions that are either acidic or basic; this is known as **hydrolysis**. A hydrolysis reaction is one in which a substance reacts with water to form OH^- or H_3O^+ ions. Note that hydrolysis reactions can also be called ionisation reactions.

Ionisation occurs when a substance reacts to form ions. A common example is when acids and bases react with water to form H_3O^+ or OH^- ions. Reactions such as these may also be termed hydrolysis reactions.

Not all cations and anions hydrolyse. For example, if sodium chloride, NaCl, is dissolved in water, the solution is neutral as neither Na^+ nor Cl^- hydrolyses to any significant extent.

However, if ammonium chloride, NH_4Cl, is dissolved in water, ammonium and chloride ions are produced. The ammonium ions react as follows:

$$NH_4^+(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + NH_3(aq)$$

This solution is acidic due to the formation of $H_3O^+(aq)$.

If sodium carbonate, Na_2CO_3, is dissolved in water, sodium ions and carbonate ions are produced. The carbonate ions react as follows:

$$CO_3^{2-}(aq) + H_2O(l) \longrightarrow HCO_3^-(aq) + OH^-(aq)$$

This solution is therefore alkaline owing to the formation of OH^- ions.

Dissociation of bases

When ionic bases dissolve in water, they **dissociate** or separate into their constituent ions. They do not ionise since they do not actually react with the water to produce ions as acids do.

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

Molecular bases such as NH_3 do not contain ions, but react with water instead. For example, ammonia ionises according to:

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

$\rightleftharpoons$ *Note:* The arrows drawn in this manner indicate that the reaction proceeds partially to the right. This means that, at any given time, there are more reactant species than product species present.

Neutralisation

When an acid and base react, they usually produce a salt and water. This is called **neutralisation**. For example:

$$\underset{\text{acid}}{HCl(aq)} + \underset{\text{base}}{NaOH(aq)} \longrightarrow \underset{\text{salt}}{NaCl(aq)} + \underset{\text{water}}{H_2O(l)}$$

The ionic equation of this reaction is:

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \longrightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$$

Since Na^+ and Cl^- are *spectator ions* and do not take part in the reaction, neutralisation may be shown simply as:

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

This equation is called the **ionic equation of neutralisation**. The equation can also be written as follows:

$$H_3O^+(aq) + OH^-(aq) \longrightarrow 2H_2O(l)$$

since we have already seen that the hydrogen ion, H^+, does not exist by itself in water.

Neutralisation reactions are one way of producing pure samples of salts. For example, if hydrochloric acid and calcium hydroxide are mixed together, an aqueous solution of calcium chloride results.

$$2HCl(aq) + Ca(OH)_2(aq) \longrightarrow CaCl_2(aq) + 2H_2O(l)$$

The water can be removed by evaporation. Table 13.3 lists some common salts and some of their uses.

One application of neutralisation reactions is in antacid tablets. Although the hydrochloric acid in the gastric juices in our stomachs helps break down the food we eat, if too much hydrochloric acid is produced, indigestion may result. Antacid tablets contain a base such as magnesium hydroxide, aluminium hydroxide or sodium bicarbonate that neutralises the acid in the stomach.

Neutralisation is the process of an acid reacting with a base. The acid and base properties can be cancelled out if the exact amounts of both are used. A neutralisation reaction produces a salt and water as its products. All neutralisation reactions can be represented by the following ionic equation:

$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

or, alternatively,

$H_3O^+(aq) + OH^-(aq) \longrightarrow 2H_2O(l)$

The neutralisation of a solution of the base ammonia by hydrochloric acid. The indicator changes colour from yellow to red.

TABLE 13.3 Some common salts and their applications

Name	Formula	Application
ammonium sulfate	$(NH_4)_2SO_4$	fertiliser
barium sulfate	$BaSO_4$	X-ray of digestive tract
calcium sulfate dihydrate	$CaSO_4{\cdot}2H_2O$	plasterboard
copper sulfate pentahydrate	$CuSO_4{\cdot}5H_2O$	dyeing; fungicides
sodium chloride	$NaCl$	table salt
potassium chloride	KCl	sodium-free salt substitute

Revision question

4. **Use the following acids and bases to write neutralisation equations to produce the salts in table 13.3. You may assume that the water in the hydrated salts remains in solution.**
   Acids: **H_2SO_4, HCl**
   Bases: **NaOH, KOH, NH_4OH, $Cu(OH)_2$, $Ca(OH)_2$, $Ba(OH)_2$**

Acid–base terms

Conjugate acid–base pairs

In an acid–base reaction, the substance acting as the acid gives away a proton and forms a conjugate base. The substance acting as a base, after accepting a proton, forms its conjugate acid. An acid–base reaction therefore forms two conjugate pairs, with the formulas of each pair member differing by the H^+ ion that was transferred.

Conjugate base = acid – H^+

Conjugate acid = base + H^+

When an acid and a base react, a **conjugate** acid and base are formed. In an equation where an acid donates a proton to a base, the conjugate pairs are represented as shown in the following two examples:

$$\begin{array}{ccccccc} HNO_3(aq) & + & H_2O(l) & \longrightarrow & H_3O^+(aq) & + & NO_3^-(aq) \\ \text{acid}_1 & & \text{base}_2 & & \text{acid}_2 & & \text{base}_1 \end{array}$$

In this case, the conjugate pairs are (HNO_3/NO_3^-) and (H_3O^+/H_2O).

$$\begin{array}{ccccccc} NH_3(g) & + & H_2O(l) & \longrightarrow & NH_4^+(aq) & + & OH^-(aq) \\ \text{base}_1 & & \text{acid}_2 & & \text{acid}_1 & & \text{base}_2 \end{array}$$

The conjugate pairs in this reaction are NH_4^+/NH_3 and H_2O/OH^-.

Conjugate acid–base pairs differ by a proton, H^+. To find the conjugate base of an acid, we subtract one H^+; to find the conjugate acid of a base we add one H^+.

$$\text{conjugate acid} \underset{+H^+}{\overset{-H^+}{\rightleftharpoons}} \text{conjugate base}$$

For example:

$$\begin{array}{ccc} \text{conjugate} & & \text{conjugate} \\ \text{acid} & & \text{base} \\ H_3PO_4 & \underset{+H^+}{\overset{-H^+}{\rightleftharpoons}} & H_2PO_4^- \end{array}$$

Note that a conjugate acid and a conjugate base can re-form the original acid and base by once again transferring a proton in a 'backwards' reaction according to our definitions.

Polyprotic acids

A polyprotic acid can donate more than one proton. Acids can be classified as monoprotic, diprotic or triprotic depending on the actual number of protons that can be donated.

A **polyprotic** acid can donate more than one proton per molecule of acid. An acid such as sulfuric acid is called a **diprotic** acid, as it can donate two protons. This occurs in two stages, as shown in the following equations:

$$\text{Stage 1}\quad H_2SO_4(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + HSO_4^-(aq)$$

$$\text{Stage 2}\quad HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_4^{2-}(aq)$$

The first stage is more extensive than the second.

Hydrochloric acid, HCl, and ethanoic acid, CH_3COOH, are examples of **monoprotic** acids, which can donate only one proton. Phosphoric acid, H_3PO_4, is an example of a **triprotic** acid as it has three protons that can donate.

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Unit 2	**Polyprotic acids and amphiprotic substances**
AOS 1	Summary screen and practice questions
Topic 2	
Concept 2	

Amphiprotic substances and ampholytes

Some substances can act as acids or bases, according to their chemical environment. This means that they can donate or accept protons. Such substances are described as **amphiprotic**.

Water can act as an acid or base depending on whether it is reacting with a stronger acid or base. This is illustrated in the examples in the section on

Amphiprotic substances and ampholytes are substances that can act as either acids or bases. The way they react depends on the relative strengths of the acids and bases they are reacting with. Water is a common example of an amphiprotic substance. Ampholytes are electrolytes of ionic substances such as HSO_4^- and HPO_4^{2-}.

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Unit 2	**pH and the ionic product of water**
AOS 1	Summary screen and practice questions
Topic 2	
Concept 4	

conjugate pairs. Other examples are HSO_4^-, HCO_3^- and HS^-. These ions have both a proton that they can donate and a negative ion to attract an H^+. Ampholytes are ionic amphiprotic substances.

$$\underset{\text{acid}}{HCO_3^-(aq)} + \underset{\text{base}}{H_2O(l)} \longrightarrow \underset{\text{acid}}{H_3O^+(aq)} + \underset{\text{base}}{CO_3^{2-}(aq)}$$

$$\underset{\text{base}}{HCO_3^-(aq)} + \underset{\text{acid}}{NH_4^+(aq)} \longrightarrow \underset{\text{acid}}{H_2CO_3(aq)} + \underset{\text{base}}{NH_3(aq)}$$

Self ionisation of water

Because water can act as both an acid and a base, it can also react with itself according to the equation:

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

Careful measurements have shown that this does indeed occur and that the concentration of these ions in neutral water is 10^{-7} M (at 25 °C). Such values indicate that this reaction occurs only to an extremely small extent. Furthermore, measurements have shown that, even when dissolved acids and bases are present, the mathematical product of these two ions is always 10^{-14}. If the concentration of H_3O^+ ions increases due to the presence of a dissolved acid, the concentration of OH^- ions decreases such that the product is still 10^{-14}. The reverse applies if a source of OH^- ions (a dissolved base) is present. Therefore, for any acidic, basic or neutral solution, we may write:

$$K_w = [H_3O^+][OH^-] = 10^{-14}$$

where K_w is termed the *self-ionisation constant* for water, and $[H_3O^+][OH^-]$ is called the *ionic product*.

It follows that:

- if $[H_3O^+] > [OH^-]$, the solution is acidic
- if $[OH^-] > [H_3O^+]$, the solution is basic
- if $[H_3O^+] = [OH^-]$, the solution is neutral.

However, in all cases, the product $[H_3O^+][OH^-]$ equals 10^{-14}.

Revision questions

5. Write the formulas of the conjugate bases of the following acids.

(a) H_2SO_4
(b) H_2S
(c) HS^-
(d) H_2O
(e) NH_4^+

6. Write the formulas of the conjugate acids of the following bases.

(a) OH^-
(b) HCO_3^-
(c) S^{2-}
(d) H_2O
(e) CN^-

7. Consider the following equation and select the substances that are acting as acids in this reaction.

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

(a) $NH_3(aq)$ and $H_2O(l)$
(b) $NH_3(aq)$ and $OH^-(aq)$
(c) $H_2O(l)$ and $NH_4^+(aq)$
(d) $H_2O(l)$ and $OH^-(aq)$

8. (a) Write equations to illustrate H_3PO_4 acting as a triprotic acid in water.
(b) Identify the conjugate acid–base pairs for the equations you wrote in part (a).

9. Write equations to illustrate the amphiprotic nature of the following species.

(a) HS^-
(b) HSO_4^-

Strengths of acids and bases

The strength of an acid or base is related to the ease with which it donates or accepts a proton.

A strong acid donates protons readily.

A strong base accepts protons readily.

Weak acids or bases do not donate or accept protons readily.

Other properties of an acid, such as its conductivity, are related to its strength. For example, hydrochloric acid is a strong acid and its reaction in water is nearly complete, producing many ions. The solution therefore is a good electrolyte.

$$\underset{\text{strong acid}}{HCl(g)} + H_2O(l) \xrightarrow{\substack{\text{virtually}\\ \text{complete reaction}}} H_3O^+(aq) + Cl^-(aq)$$

However, ethanoic acid is weak, does not react to completion and its solution is a poor conductor of electricity because not many ions are formed.

$$\underset{\text{weak acid}}{CH_3COOH(l)} + H_2O(l) \underset{\text{partial reaction}}{\rightleftharpoons} H_3O^+(aq) + CH_3OO^-(aq)$$

Strangely shaped pillars create an eerie atmosphere in moist limestone caves. They are formed by a long process that begins with rainwater dissolving carbon dioxide in the atmosphere to form carbonic acid. When this acidic rain drips through limestone rocks, an acid–base reaction occurs. The calcium carbonate in the rocks reacts with the weak carbonic acid to form calcium hydrogen carbonate, which trickles down into the caves. As the droplets evaporate slowly, the calcium hydrogen carbonate decomposes to re-form carbon dioxide and solid calcium carbonate, creating stalactites and stalagmites.

Acid strength is an indication of how easily an acid donates protons. A strong acid donates protons readily and forms a large number of ions in water.

Strong acids form solutions that are good conductors of electricity. Weak acids do not donate protons readily. As a result, only a few ions are formed by reaction with water and the resulting solutions are poor electrical conductors.

The strength of a base also affects its conductivity. A base is said to be strong if it produces many hydroxide ions in solution because the hydroxide ion readily accepts protons. For example, sodium hydroxide is a strong base because it dissociates completely, producing many hydroxide ions. Sodium hydroxide is therefore a good electrolyte.

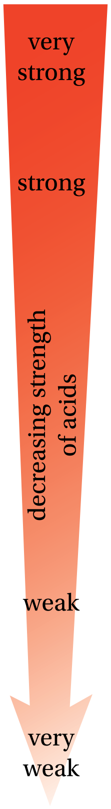
$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

However, a weak base such as ammonia does not readily ionise and accept protons. It forms very few ions in solution. Ammonia is therefore a poor conductor of electricity.

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

As this conductivity probe experiment shows, solutions of hydrochloric acid (a strong acid), sodium hydroxide and sodium chloride are strong electrolytes as they produce many ions in solution. Water is a weak electrolyte as it produces very few ions and is therefore able to conduct only a limited amount of electricity.

TABLE 13.4 Relative strengths of common acids and their conjugate bases



Name of acid	Formula	Conjugate base	Name of base
hydrochloric	HCl	Cl^-	chloride ion
nitric	HNO_3	NO_3^-	nitrate ion
sulfuric	H_2SO_4	HSO_4^-	hydrogen sulfite ion
hydronium ion	H_3O^+	H_2O	water
phosphoric	H_3PO_4	$H_2PO_4^{2-}$	dihydrogen phosphate ion
hydrofluoric	HF	F^-	fluoride ion
ethanoic (acetic)	CH_3COOH	CH_3COO^-	ethanoate (acetate) ion
carbonic	H_2CO_3	HCO_3^-	hydrogen carbonate ion
hydrogen sulfide	H_2S	HS^-	hydrogen sulfide ion
ammonium ion	NH_4^+	NH_3	ammonia
hydrogen carbonate ion	HCO_3^-	CO_3^{2-}	carbonate ion
hydrogen sulfide ion	HS^-	S^{2-}	sulfide ion
water	H_2O	OH^-	hydroxide ion
hydroxide ion	OH^-	O^{2-}	oxide ion
hydrogen	H_2	H^-	hydride ion

Notes

1. All strong acids react *completely* with water and their anions have essentially no ability to attract protons.
2. The hydronium ion is listed so that water can appear as a base on the table.
3. Oxide ions do not exist in water as they always revert to hydroxide ions.

Revision questions

10. **Using HNO_3 and HS^- as examples, write equations, with appropriate arrows, to illustrate the difference between a strong acid and a weak acid. Explain the difference in strength in terms of the conductivity of the resulting solution.**

11. **Dimethylamine, $HN(CH_3)_2$, is a weak base that reacts in water like ammonia, NH_3. Write an equation to illustrate the reaction of dimethylamine with water.**

12. **Acids, such as nitric and hydrochloric, are used by artists to make engravings on metal plates made from copper, zinc or iron.**
   - **(a)** **Use table 13.4 to find the conjugate pairs of the acids used by the artists for their etchings.**
   - **(b)** **What safety precautions would you recommend an artist working with strong acids should take?**

Strength versus concentration

The strength of an acid is different from the concentration of an acid. The strength of a solution is determined by the number of ions present. A strong acid is completely ionised in solution (producing many ions). A weak acid does not ionise to any great extent and so contains a larger number of molecules compared with the number of ions produced in solution. Concentration, however, refers to the amount of an acid or base that is dissolved in a given volume of water. A large amount always produces a concentrated solution whereas a small amount in the same volume of water produces a dilute solution. It is possible to have a weak, concentrated acid or a dilute solution of a strong acid.

'Strong' and 'concentrated' have different meanings when applied to acids and bases. Likewise, the terms 'dilute' and 'weak' also have different meanings.

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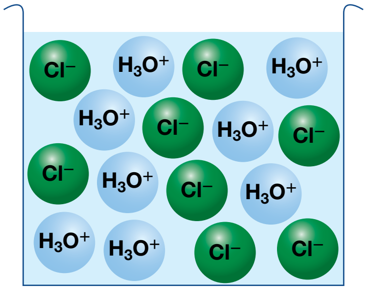
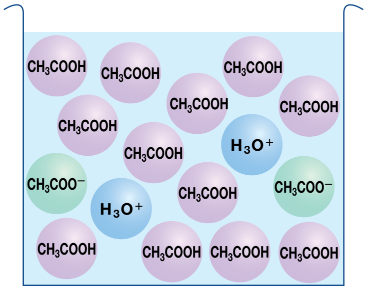
Unit 2
AOS 1
Topic 2
Concept 3

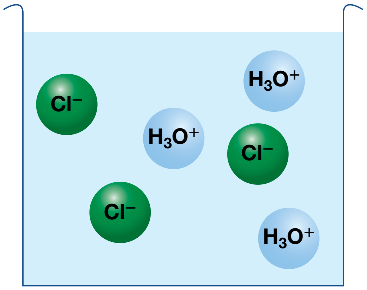
Strong and weak acids versus concentrated and dilute acids
Summary screen and practice questions



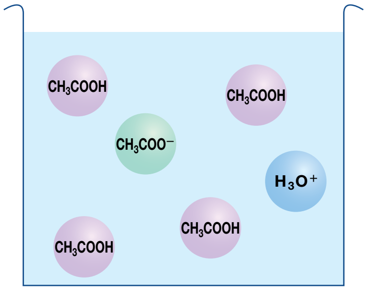
concentrated strong acid



concentrated weak acid



dilute solution of strong acid



dilute solution of weak acid

Both the concentration and strength of an acidic or basic solution determine the number of ions present in the solution.

The pH scale

The **pH scale** is usually applied over a range from 1 to 14. Using this scale, a neutral solution has a pH of 7. Values lower than 7 indicate an acidic solution. The stronger the acid the lower the pH value. For example, a solution of pH 3 is stronger than a solution of pH 4.

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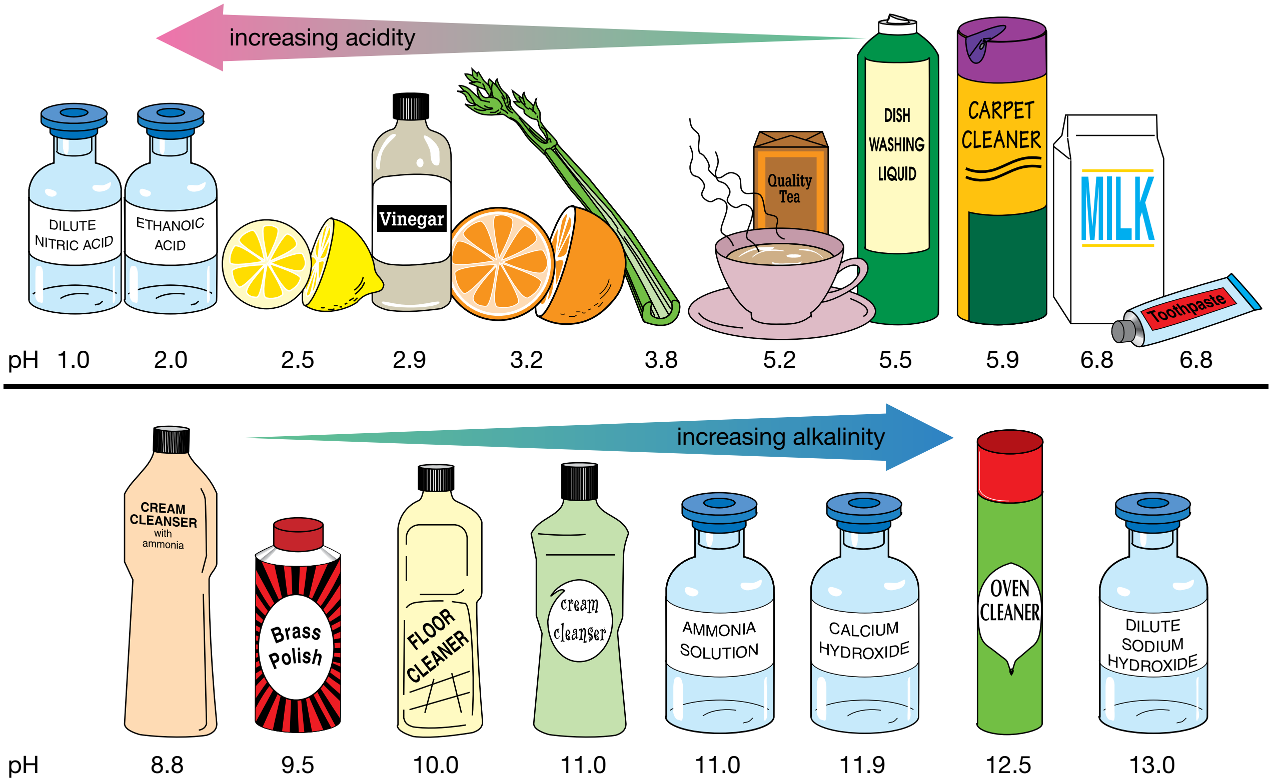
Acid strength

Any pH value greater than 7 indicates a basic solution. In this case, the higher the value, the stronger the base. The pH scale of some common acids and bases is shown below.



The pH scale for common acids and bases

Have you ever wondered why the pink hydrangeas in your garden can bloom a different shade from year to year? The colour of the hydrangea bloom depends on the acidity of the soil. If the soil becomes too acidic, pink hydrangea blooms turn blue. To correct this, calcium oxide, CaO, generally known as ‘lime’ or ‘quicklime’ to the gardener, is added. Calcium oxide is a base that dissolves in water to form calcium hydroxide, $Ca(OH)_2$, or ‘slaked lime’. This base changes the colour of the blossoms back to pink.

Definition of pH

pH is defined according to the formula:

$$pH = -\log_{10}[H_3O^+]$$

This formula enables pH to be calculated for any known H_3O^+ or OH^- concentration by using a calculator with a log function.

Square brackets are used to denote concentrations measured in M.

When $[H_3O^+]$ is written as a power of 10, the pH is equal to the power multiplied by –1. This formula can also be written as:

$$[H_3O^+] = 10^{-pH}$$

Sample problem 13.1

A solution is found to have an H_3O^+ concentration of 0.0001 M. Calculate its pH.

Solution:

STEP 1

$[H_3O^+] = 0.0001 = 10^{-4}$ M (by inspection)

STEP 2

pH = 4 (multiply the power above by –1)

Sample problem 13.2

A solution is found to have an H_3O^+ concentration of 3.5×10^{-5} M. Calculate its pH.

Solution:

STEP 1

$[H_3O^+] = 3.5 \times 10^{-5} = 10^{-4.5}$ M (using log function on calculator)

STEP 2

pH = 4.5

Sample problem 13.3

A solution is found to have an OH^- concentration of 2.8×10^{-4} M. Calculate its pH.

Solution:

STEP 1

$[OH^-] = 2.8 \times 10^{-4} = 10^{-3.6}$ M (using calculator)

STEP 2

Using the ionic product for water:

$[H_3O^+][10^{-3.6}] = 10^{-14}$

$[H_3O^+] = 10^{-10.4}$ M

STEP 3

pH = 10.4

pH values of strong acids and bases

As we have seen, strong acids undergo virtually complete ionisation with water. For a monoprotic acid such as hydrochloric acid, HCl, this means that the concentration of acid molecules originally dissolved equals the concentration of H_3O^+ ions produced when these react with water according to:

$$HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

The pH of such solutions can then be predicted using calculations such as those above.

The pH of strong bases can also be calculated in exactly the same way, using the ionic product of water as an extra step.

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Unit 2
AOS 1
Topic 2
Concept 5

pH and values of the acidity constant
Summary screen and practice questions

Sample problem 13.4

The label on a bottle of nitric acid states that its concentration is 0.005 M. Calculate its pH.

Solution: **STEP 1**

As nitric acid is a strong acid, the concentration of H_3O^+ ions is also 0.005 M, due to complete ionisation.

STEP 2

$[H_3O^+] = 0.005 = 10^{-2.3}$ M (using calculator)

STEP 3

$pH = 2.3$

Sample problem 13.5

The label on a bottle of potassium hydroxide states that its concentration is 0.02 M. Calculate its pH.

Solution: **STEP 1**

As potassium hydroxide is a strong base, the concentration of OH^- ions is also 0.02 M.

STEP 2

$[OH^-] = 0.02 = 10^{-1.7}$ M (using calculator)

STEP 3

Using the ionic product for water, $[H_3O^+] = 10^{-12.3}$ M

STEP 4

$pH = 12.3$

Revision questions

13. **What is pH and how can it be calculated?**
14. **What is the difference between a dilute and a concentrated solution of an acid or base?**
15. **Determine the pH of the following given that $[H_3O^+]$ is:**
    (a) **10^{-5} M** (b) **0.1 M.**
16. **Determine the pH of the strong acid $HClO_4$ whose concentration is 0.001 M.**
17. **Determine the pH of the following given that $[OH^-]$ is:**
    (a) **10^{-4} M**
    (b) **0.1 M.**
18. **90 mL of water is added to 10 mL of HNO_3 with pH 2. What is the new pH?**
19. **Determine the pH of a solution where:**
    (a) **$[H_3O^+] = 9.2 \times 10^{-4}$ M**
    (b) **$[OH^-] = 1.8 \times 10^{-5}$ M.**
20. **Using the formula for pH, explain why a neutral solution has a pH of 7 (at 25 °C).**

Dilution and pH

Dilution is the process of adding water to a solution. For solutions that are acidic or basic, this affects the concentration of H_3O^+ ions that are present and, hence, the pH. Because pH is on a log scale, the factor by which H_3O^+ changes is not the same as that by which pH changes. For example, if a solution whose H_3O^+ concentration is 0.01 (10^{-2}) M is diluted by a factor of 100 to 0.0001 (10^{-4}) M, the pH changes from 2 to 4.

Sample problem 13.6

A solution of hydrochloric acid of concentration 0.050 M is diluted by adding an equal volume of water. This halves the concentration to 0.025 M. Calculate the change in pH.

Solution: **STEP 1**

Hydrochloric acid is a strong acid, so all acid concentrations are the same as the H_3O^+ concentrations, due to complete ionisation.

Therefore, we can write initial $[H_3O^+] = 0.050 = 10^{-1.3}$ M and initial pH = 1.3.

STEP 2

After dilution, $[H_3O^+] = 0.025 = 10^{-1.6}$ M.

Therefore, pH after dilution = 1.6.

Therefore, pH changes from 1.3 to 1.6.

Note the common-sense check here. We would expect the solution to be less acidic after dilution and therefore have a higher pH.

Relative strengths of acids and bases and pH

The strength of an acid or base is measured by its tendency to donate (in the case of acids) or accept (in the case of bases) protons. The strength of an acid has to be taken into consideration when calculating the pH, as it determines the acid's ability to ionise.

It has been found experimentally that not all acids of 0.1 M concentration have the same pH, and do not conduct electricity to the same extent. Remember that the conductivity of a solution is directly related to the number of ions in solution. Consider two acids, HCl and CH_3COOH, with the same concentration of 0.1 M at 25 °C. Experimental results confirm that hydrochloric acid has a pH of 1; however, ethanoic acid has a pH of 3. This is explained by the differing abilities of the two acids to ionise. Almost every HCl molecule donates protons to water. This is consistent with the fact that HCl is a strong acid and ionises virtually completely to produce a good electrolyte. CH_3COOH, on the other hand, is a weak acid. Only one in every hundred CH_3COOH molecules loses a proton, and therefore a solution of ethanoic acid is a poor electrolyte and hence a poor conductor of electricity.

The pH of a basic solution is also affected by the ability of the base to dissolve or dissociate in water. For example, a 0.1 M solution of NaOH, a strong base, has a higher pH than a 0.1 M solution of NH_3, a weak base. The sodium hydroxide solution is therefore a stronger base than the ammonia solution.

This dependence of pH on acid strength can be used to compare the relative strengths of a range of acids. Provided that the concentrations of the acids are all the same, the lower the pH the stronger the acid is. This is illustrated in table 13.5. Note that nitrous acid is the strongest acid of those listed (lowest pH) while hypochlorous acid (highest pH) is the weakest. It does not matter which concentration is used to make this comparison, as long as it is the same for all the acids that you are comparing.

TABLE 13.5 pH values for solutions of a range of acids

Acid	Formula	pH (0.01 M solution)	pH (0.001 M solution)
benzoic acid	C_6H_5COOH	3.1	3.6
ethanoic acid	CH_3COOH	3.4	3.9
hypochlorous acid	HOCl	4.8	5.3
lactic acid	$HC_3H_5O_3$	2.9	3.4
methanoic acid	HCOOH	2.9	3.4
nitrous acid	HNO_2	2.6	3.1

The growth of algae and bacteria is controlled in swimming pools through the use of chlorine compounds. Sodium hypochlorite, NaOCl, and calcium hypochlorite, $Ca(OCl)_2$ are used, undergoing hydrolysis to produce the weak acid hypochlorous acid, HOCl, according to the equation:

$$OCl^-(aq) + H_2O(l) \longrightarrow HOCl(aq) + OH^-(aq)$$

However, if the pH is too high, the concentration and effectiveness of HOCl are reduced. To lower the pH, an acid solution (such as hydrochloric acid or solid sodium hydrogen sulfate) can be used to neutralise the excess hydroxide ions. If the pH is too low, the acid content of the water can cause eye and skin irritation and damage to pool fittings. The pH can be raised by neutralising with sodium carbonate or sodium hydrogen carbonate.

Measuring pH

Indicators

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Experiment 13.2 Indicators and pH
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Indicators are dyes that usually are themselves weak acids or bases. They have the important characteristic that they show a different colour in an acidic solution from that in a basic solution. They can be used to determine whether a substance is acidic or basic and also how acidic or basic it is. Litmus, phenolphthalein, methyl orange and universal indicator are the most commonly used indicators. **Universal indicator** is a mixture of several indicators. It changes colour gradually, from red to green to violet as the solution changes from acidic to neutral to basic.

Indicators are useful tools, although they have certain limitations. The indicator colours found in charts are quoted for 25 °C; at other temperatures an indicator may change colour at a different pH. If a solution being tested has a colour of its own,

Universal indicator is a mixture of several indicators and can be used to estimate the pH of a solution. The colour assumed by the indicator after mixing with the solution is compared with a standard colour chart.

the colour of the indicator may be distorted. As the naked eye is used to make the comparison between the indicator colour and the pH chart, this can be a problem, especially for people who are colourblind.

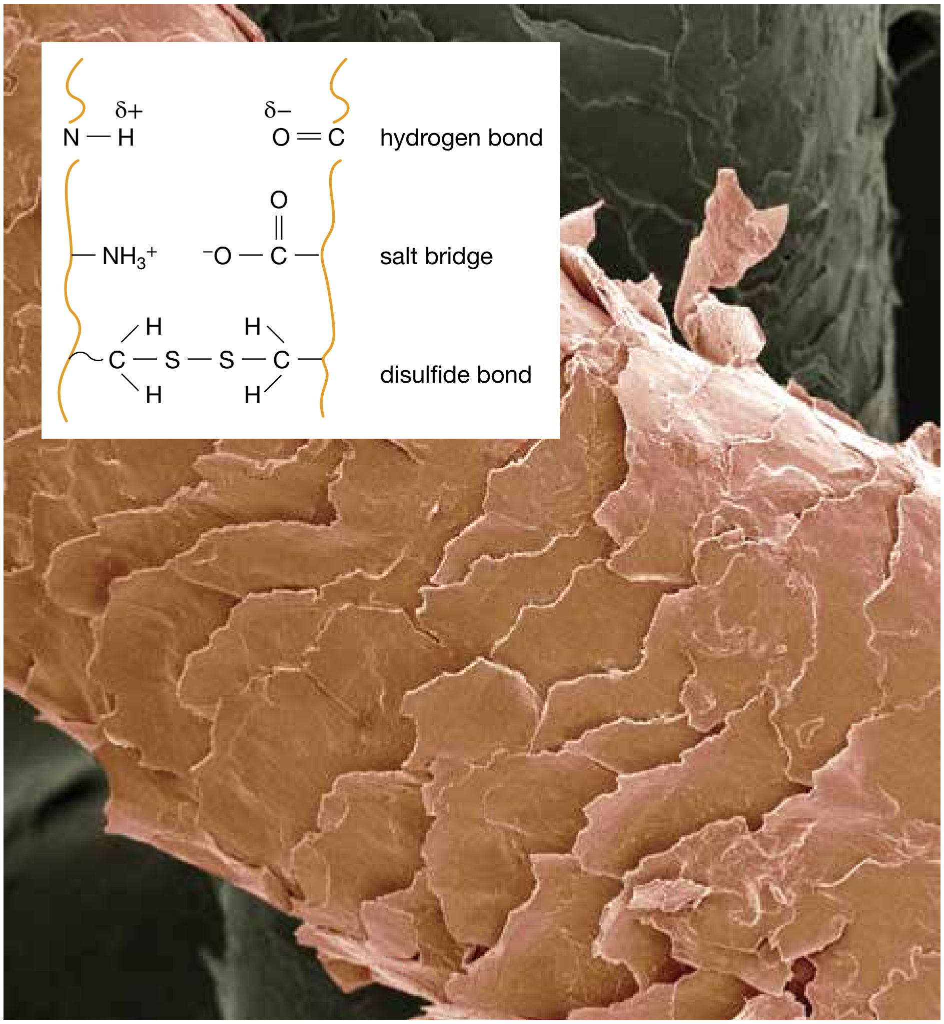
The pH of a solution can be measured by adding universal indicator and matching the colour shade obtained against a special chart. More accurate readings can be obtained through the use of specially designed pH meters.

The pH meter

Most laboratories in industry have a pH meter that is used to make rapid, accurate measurements of pH. It can be connected to a computer to monitor pH changes continuously.

The voltage of the electrode changes with the [H_3O^+] in the solution into which it is dipped. Values of pH obtained by a pH meter are accurate to within 0.01 units of the true pH and are not affected by the colour and cloudiness of the unknown solution. Hospitals use pH meters to find small but significant changes of pH in blood and other body fluids. Soils, ponds, rivers, sewage and industrial effluents are easily monitored with a pH meter.

Pure hydrogen chloride, HCl, is a gas; it can be dissolved in water to produce a concentrated solution. Both concentrated HCl and ammonium hydroxide, NH_4OH, produce fumes. These fumes react to produce a fine white solid, ammonium chloride, NH_4Cl. This is an example of an acid–base reaction that does not involve water.



Amino acids polymerise to make hair, a natural polymer whose chains are held together by hydrogen bonding; salt bridges between the acid and base groups; and disulfide bonds. Hair is normally acidic (pH 4 to 5), and for this reason alkaline shampoos are followed by an 'acidic rinse' or conditioner. In strongly acidic solutions (pH 1 to 2), the H–bonding and salt bridges of the hair are broken. Slightly alkaline (pH 8.5) solutions break some of the disulfide bonds, causing the outer surface of hair strands to become ruffled and appear dull, as light is unevenly reflected from their surfaces.

study on

Unit 2
AOS 1
Topic 2
Concept 8

Environmental effects — acid rain and ocean acidification
Summary screen and practice questions

When acids cause trouble

The presence of pollutant gases in the atmosphere, together with increasing levels of carbon dioxide has, and is, leading to some environmental issues that may have serious implications, on both local and global scales.

Acid rain

Normal rain is slightly acidic due to dissolution of carbon dioxide in water according to the equation:

$$CO_2(g) + H_2O(g) \longrightarrow H_2CO_3(aq)$$

The product is the weak acid carbonic acid, which reacts slightly with water to produce a few hydronium ions as per the equation:

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

Many industrial processes involve the burning of coal, oil or some other fossil fuel. Many of these fuels contain sulfur in varying amounts. When sulfur is burned in air, it forms sulfur dioxide, SO_2. This gas is often released into the air in vast quantities.

It must be noted that sulfur dioxide is released in many natural processes as well. In particular, active volcanoes release a large amount of sulfur dioxide into the air. The atmosphere can cope with large quantities of sulfur dioxide if it is given time to disperse the gas and break it down. When a large number of industries are all producing the gas over a small area, it cannot disperse in the air fast enough and becomes too concentrated to be safe.

Sulfur dioxide (SO_2) can react with water to produce sulfurous acid (a weak acid). It can also react with oxygen in the air to produce sulfur trioxide, SO_3, which then reacts with water to produce the strong acid sulfuric acid. The equations for these three processes are:

$$SO_2(g) + H_2O(l) \longrightarrow H_2SO_3(aq)$$

$$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$$

$$SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)$$

Other gases contribute to the **acid rain** problem and many of these are produced by industry and by every car that uses the internal combustion engine.

In a car engine, the temperature produced by the spark plug is sufficient to cause the nitrogen and oxygen in the air to combine and form nitrogen monoxide, NO. Nitrogen monoxide combines spontaneously with oxygen in the air to form nitrogen dioxide, NO_2. This nitrogen dioxide reacts with rainwater to form nitric acid, another contributor to the problem known as acid rain.

The equation for this reaction is:

$$3NO_2(g) + H_2O(g) \longrightarrow 2HNO_3(aq) + NO(g)$$

It must be noted, however, that carbonic acid and the oxides of sulfur are largely responsible for acid rain.

Impacts and control

Acid rain has been responsible for the defoliation of significant amounts of forest in both Europe and North America. In Australia it is one of the principal causes of the 'lunar' landscape that occurred around Queenstown in Tasmania where large amounts of sulfur dioxide were produced as a result of copper ore smelting during the last century.

Where acid rain has run off into lakes, there have been instances where the pH of such lakes has dropped to the point where they have been unable to sustain life.

A range of responses has now evolved to reduce the input of gases into the atmosphere that cause acid rain. These include:

- catalytic converters in car exhausts to change nitrogen oxides into nitrogen and oxygen
- the use of natural gas as a fuel (natural gas contains lower levels of sulfur impurity) in place of other fuels such as oil and coal
- switching to coal with a lower sulfur content
- taking measures to assist the dispersal of SO_2, such as by using taller chimney stacks
- *scrubbing*. This refers to treating exhaust gases to remove undesirable emissions. An increasing number of processes are now removing SO_2 from their emissions and concentrating it so that it can be used as a feedstock to produce sulfuric acid. This has an extra advantage because sulfuric acid is a valuable industrial acid and can be sold to offset the costs of the original process.

Ocean acidification

As we have seen previously, carbon dioxide dissolves in water to produce carbonic acid. The oceans are slightly alkaline, but research has shown that, since the start of the Industrial Revolution, their pH has dropped from 8.2 to 8.1. While there are many localised variations in these figures, such a change is significant. This is directly attributable to the increased level of carbon dioxide in the atmosphere. Although the effect of this is the subject of ongoing research, problems that might arise as a result of this acidification include:

- coral bleaching
- interference with marine organisms, particularly those at the lower end of the food chain
- reproductive disorders in certain marine organisms
- interference with shell building in shellfish and other similar organisms. The shells of many marine species have calcium carbonate as an important constituent. Others, including corals, form skeletons using calcium carbonate. Although calcium carbonate is virtually insoluble, there are Ca^{2+} and CO_3^{2-} ions in sea water at concentrations that are low but nevertheless important. However, carbonate ions can be removed by a reaction between themselves, water and carbon dioxide according to:

$$CO_2(g) + H_2O(l) + CO_3^{2-}(aq) \longrightarrow 2HCO_3^{-}(aq)$$

If the carbonate ion concentration is lowered too much, it is possible that the calcium carbonate in the shells of marine organisms may dissolve in an attempt to replace the removed carbonate ions.

$$CaCO_3(s) \longrightarrow Ca^{2+}(aq) + CO_3^{2-}(aq)$$

Introducing redox reactions

Redox reactions are very common. The term ‘oxidation’ was used to describe combining with oygen, but this definition has been expanded.

Oxidation–reduction or **redox** reactions are one of the largest groups of chemical reactions. Redox chemistry is an important aspect of everyday life.

Our bodies work by redox reactions — the food we eat is oxidised to enable us to obtain the energy we need to live. If we break a bone, electric currents generated around the damaged area are involved in the healing process. Redox chemistry is involved in the conduction of impulses in nerves. The batteries we use in radios, clocks, cars and calculators rely on spontaneous redox reactions. Many metals are extracted by reduction reactions. Municipal water supplies and swimming pools are treated with oxidants that act as bactericides and algicides. Bushfires are uncontrolled redox reactions on a large scale, and explosions are very fast redox reactions. Redox reactions occur all around us.

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Unit 2
AOS 1
Topic 3
Concept 1

Redox terms
Summary screen and practice questions

An explosion is a very fast redox reaction.

What is oxidation–reduction?

Early chemists knew that a gas called oxygen was essential for the survival of living things. This prompted them to study the reactions of oxygen with other substances in an effort to find out more about its behaviour. They found that many substances combined with oxygen, so the term **oxidation** was used to describe these reactions. The combustion reactions of hydrocarbons such as the burning of propane:

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$$

and the burning of iron in air:

$$4Fe(s) + 3O_2(g) \longrightarrow 2Fe_2O_3(s)$$

were described as oxidation reactions because a reactant, or some part of it, combined with oxygen.

Steel wool burns as it reacts with oxygen in air to form iron(III) oxide.

Conversely, reactions that involved the decomposition of a compound, with the loss of oxygen, were called **reduction** reactions because the compound seemed to be ‘reduced’ to something simpler. For example, copper(II) oxide may be reduced to copper by hydrogen:

$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(g)$$

Similarly, iron(III) oxide is reduced to iron by carbon monoxide in a blast furnace:

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g)$$

As one reactant is reduced, the other reactant is oxidised. This can be seen when magnesium burns in steam to form magnesium oxide and hydrogen:

$$Mg(s) + H_2O(g) \longrightarrow MgO(s) + H_2(g)$$

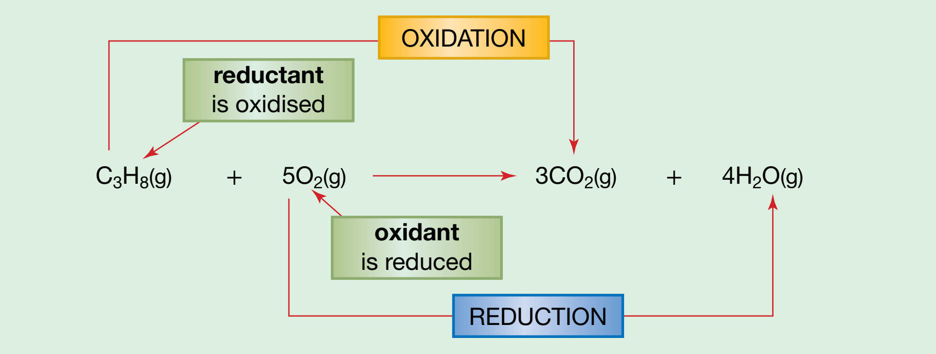
Oxidants and reductants

The oxidant is the species that is reduced.

The reductant is the species that is oxidised.

In any oxidation–reduction reaction:

- The **oxidant** is the species that *causes* oxidation of *another substance* and is *itself* reduced.
- The **reductant** is the species that *causes* reduction of *another substance* and is *itself* oxidised.



The definitions of oxidation and reduction on the one hand, and oxidants and reductants on the other, can seem confusing at first. It is useful to remember that oxidation and reduction are *processes*. Oxidants and reductants, however, are *substances* that these processes happen to.

An electron transfer view

Note: You can remember oxidation–reduction by the acronym ‘OIL RIG’:
Oxidation
Is
Loss of electrons
Reduction
Is
Gain of electrons

The modern definition of oxidation and reduction processes describes redox reactions in terms of electron transfer.

For example, the combustion of magnesium in oxygen to form magnesium oxide involves the oxidation of magnesium.

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

A similar reaction occurs when magnesium chloride is formed by the combustion of magnesium in chlorine, but in this case no oxygen is involved.

$$Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$$

Redox reactions are electron transfer reactions.

Analysis of these two reactions in terms of a transfer of electrons reveals that they are essentially the same process. Both products are ionic substances and so contain Mg^{2+} ions. In the process of oxidation, each atom of magnesium has lost two electrons.

$$Mg \longrightarrow Mg^{2+} + 2e^-$$

A substance that is oxidised is one that loses electrons.

In general, a substance that is *oxidised* is one that *loses electrons* and is therefore an electron donor.

When magnesium reacts in this way with either oxygen or chlorine, these atoms accept these electrons and become O^{2-} and Cl^- anions in the ionic lattice of the products.

$$O_2 + 4e^- \longrightarrow 2O^{2-}$$

$$Cl_2 + 2e^- \longrightarrow 2Cl^-$$

A substance that is reduced is one that gains electrons.

In general, a substance that is *reduced* is one that *gains electrons* and is therefore an electron acceptor.

Oxidation can be defined as a loss of electrons to another substance. Reduction can be defined as an acceptance of electrons from another substance. Redox reactions, therefore, are those in which electrons are transferred from one reactant to another.

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Redox-electron transfer
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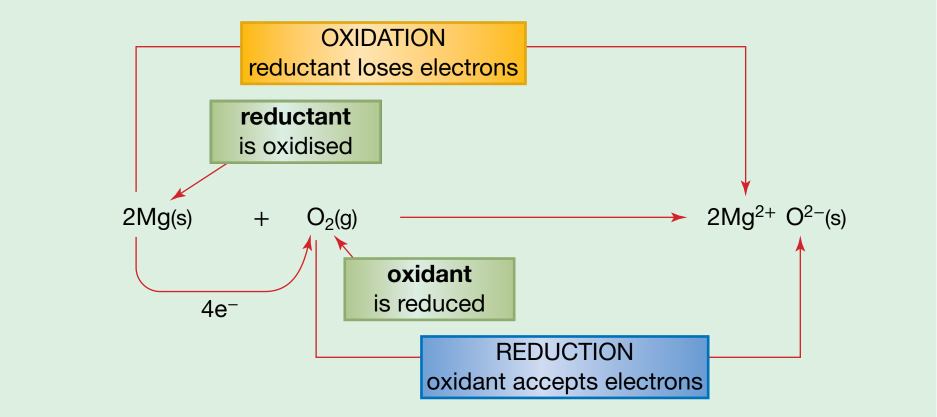
Oxidation must be accompanied by reduction; that is, oxidation cannot occur unless reduction occurs simultaneously. When writing equations for redox reactions, we do not show electrons as all electrons given off during oxidation are taken in during reduction. However, if we are considering oxidation or reduction reactions separately, it is appropriate (and necessary) to write reactions that do show the electrons, as has been done in the examples on the previous page. Such equations are called *partial equations* or *half-equations*. It should be noted that, if you multiply the two half-equations so that there are the same number of electrons on each side, and then add them together, you do, in fact, get the overall ionic equation.

The reductant loses electrons.

The reactant that is losing electrons (undergoing oxidation) is called the reductant or reducing agent, because it transfers electrons to another substance and causes it to be reduced.

The oxidant gains electrons.

The reactant that is gaining electrons (undergoing reduction) is called the oxidant or oxidising agent, as it accepts the electrons from the reductant and causes it to be oxidised. This can be illustrated by the following diagram.



When steel wool burns in chlorine, iron in the wool is oxidised to form iron(III) ions, Fe^{3+}. At the same time, chlorine is reduced to form chloride ions, Cl^-.

We can summarise these ideas in another way.

Reactant	Reactant undergoes	Reactant is called
Mg	oxidation (loses e^-)	reductant
O_2	reduction (gains e^-)	oxidant

Consider the reaction of steel wool burning in chlorine gas:

$$2Fe(s) + 3Cl_2(g) \longrightarrow 2FeCl_3(s)\ (Fe^{3+}, 3Cl^-)$$

Fe	Cl_2
donates electrons	accepts electrons
is oxidised	is reduced
is the reductant	is the oxidant

Revision question

21. **Most of the following equations represent redox reactions. Use a diagram similar to that above to demonstrate which reactant is being oxidised and which reactant is being reduced. Identify the reactions that are not redox, and explain why.**

(a) $Mg(s) + S(s) \longrightarrow MgS(s)$
(b) $Fe(s) + 2Ag^+(aq) \longrightarrow 2Ag(s) + Fe^{2+}(aq)$
(c) $Cl_2(aq) + 2Br^-(aq) \longrightarrow Br_2(aq) + 2Cl^-(aq)$
(d) $Fe^{3+}(aq) + 3OH^-(aq) \longrightarrow Fe(OH)_3(s)$

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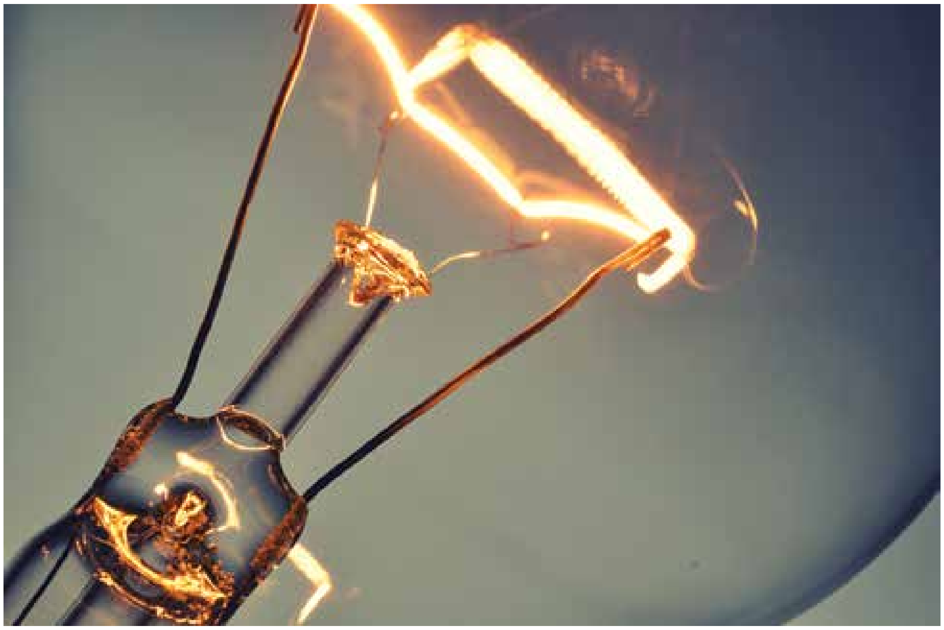
Unit 2 **Writing half-equations for redox reactions**

AOS 1 Summary screen and practice questions

Topic 3

Concept 3



The filaments of some light bulbs are often made from tungsten metal, W. If hot tungsten is exposed to air it oxidises to form tungsten oxide. This is why inert argon gas is used to fill the inside of light globes.

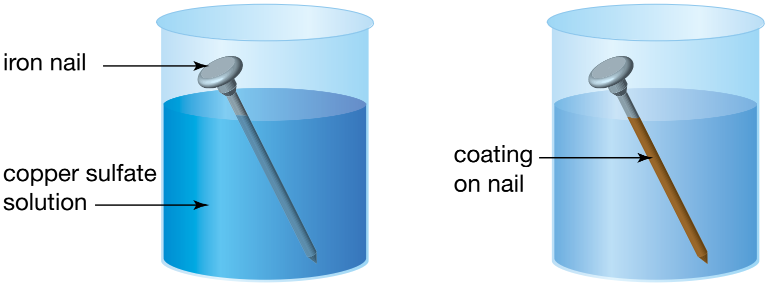
Half-equations

Redox equations can be written as two half-equations.

Half-equations are a useful way of understanding the processes involved in a redox reaction. The following example describes how half-equations are written.

When an iron nail is placed in a blue copper sulfate solution, the nail becomes coated with metallic copper and the blue colour of the solution fades. A redox reaction has taken place as electrons have been transferred from the iron nail to the copper ions in the solution, allowing solid copper to form. The full equation for the reaction is as follows:

$$Fe(s) + CuSO_4(aq) \longrightarrow FeSO_4(aq) + Cu(s)$$



Since copper sulfate and iron sulfate are in solution, we can write an ionic equation for this reaction, eliminating the sulfate spectator ions.

$$Fe(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow Fe^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$$

The ionic equation becomes:

$$Fe(s) + Cu^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Cu(s)$$

A conjugate redox pair is made up of two species that differ by a certain number of electrons.

Although the oxidation and reduction reactions occur simultaneously, it is possible to consider the two reactions separately. To do this we separate the conjugate pair of oxidant and reductant. A **conjugate redox pair** is made up of two species that differ by a certain number of electrons. Each has its own half-equation.

So for the oxidation conjugate pair we can write:

$$Fe(s) \longrightarrow Fe^{2+}(aq)$$

and for the reduction conjugate pair we can write:

$$Cu^{2+}(aq) \longrightarrow Cu(s)$$

Another way to remember oxidation and reduction is to look at the half-equations. Electrons on the left mean that it is reduction (*red* is on the left of the term 'redox'). Electrons on the right mean that it is oxidation (*ox* is on the right of the term 'redox').

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Experiment 13.3 Simple redox equations
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In order to balance the conjugate pair to produce proper half-equations, electrons need to be shown. So, an iron atom must have lost two electrons during the reaction to go from iron metal to iron ions. So the half-equation for the oxidation part of the reaction becomes:

$$Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^-$$

Similarly, each copper ion must have gained two electrons in order to become a copper atom with no charge. So the half-equation for the reduction part of the reaction becomes:

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

These half-equations are balanced with respect to both atoms and charge. Combining these two half-equations will yield the ionic equation, shown on the previous page, for the reaction as a whole.

Revision questions

22. **The balanced ionic equation for the displacement of silver from an aqueous silver nitrate solution by metallic lead is:**

$$2Ag^+(aq) + Pb(s) \longrightarrow 2Ag(s) + Pb^{2+}(aq)$$

- **(a) Write balanced oxidation and reduction half-equations for the reaction.**
- **(b) Which reactant accepts electrons?**
- **(c) Which reactant is oxidised? What is it oxidised to?**
- **(d) Which reactant is the reductant?**
- **(e) From which reactant are electrons taken?**

23. **The reaction of aluminium with hydrogen ions in a dilute solution of hydrochloric acid can be represented by the following half-equations:**

$$Al(s) \longrightarrow Al^{3+}(aq) + 3e^-$$
$$2H^+(aq) + 2e^- \longrightarrow H_2(g)$$

- **(a) Which reactant accepts electrons?**
- **(b) Which reactant is reduced? What is it reduced to?**
- **(c) Which reactant is the reductant?**
- **(d) From which reactant are electrons taken?**
- **(e) Which reactant is oxidised? What is it oxidised to?**
- **(f) Which reactant is the oxidant?**

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Writing balanced half-equations for ions in aqueous solution

At first glance, the statement that MnO_4^- is *reduced* to Mn^{2+} seems incorrect because, according to the charges on the ions, there appears to be a loss, rather than gain, of electrons:

$$MnO_4^-(aq) \longrightarrow Mn^{2+}(aq)$$

It is only when the entire half-equation for the change is written that its true nature as reduction becomes obvious, with five electrons being accepted by each permanganate ion:

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \longrightarrow Mn^{2+}(aq) + 4H_2O(l)$$

The steps involved in obtaining such equations, using as an example the oxidation of chloride ions by permanganate ions in acidified solution are given below. The products are molecular chlorine, $Cl_2(aq)$, and manganese(II) ions, $Mn^{2+}(aq)$. Do not attempt to balance a single equation.

To balance half-equations: balance the elements, add water molecules, balance by adding hydrogen ions, and balance the difference in charge with electrons.

STEP 1 Write the half-equations for oxidation and reduction showing conjugate pairs. If you are having trouble identifying the conjugate pairs, use oxidation numbers to identify which substance has been oxidised and which has been reduced.

$$\text{reduction } MnO_4^-(aq) \longrightarrow Mn^{2+}(aq)$$

$$\text{oxidation } Cl^-(aq) \longrightarrow Cl_2(aq)$$

STEP 2 Balance all elements except hydrogen and oxygen, which will be balanced later.

$$MnO_4^-(aq) \longrightarrow Mn^{2+}(aq)$$

$$2Cl^-(aq) \longrightarrow Cl_2(aq)$$

STEP 3 Balance oxygen atoms, where needed, by adding water.

$$MnO_4^-(aq) \longrightarrow Mn^{2+}(aq) + 4H_2O(l)$$

$$2Cl^-(aq) \longrightarrow Cl_2(aq)$$

STEP 4 Balance hydrogen atoms, where needed, by adding H^+.

$$MnO_4^-(aq) + 8H^+(aq) \longrightarrow Mn^{2+}(aq) + 4H_2O(l)$$

$$2Cl^-(aq) \longrightarrow Cl_2(aq)$$

STEP 5 Balance the charge of each half-equation by adding electrons.

$$MnO_4^-(aq) + 8H^+(aq) \longrightarrow Mn^{2+}(aq) + 4H_2O(l)$$

charges as the equation stands	−1	+8	+2	0
net charge	+7		+2	

Since the net charge of the left-hand side is greater than that of the right-hand side, it is necessary to add five electrons to the left-hand side to make the net charge on both sides equal at +2. The reduction half-equation becomes:

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \longrightarrow Mn^{2+}(aq) + 4H_2O(l)$$

$$2Cl^-(aq) \longrightarrow Cl_2(aq)$$

charges as the equation stands	2 × −1	0
net charge	−2	0

Since the net charge of the left-hand side is less than that of the right-hand side, it is necessary to add two electrons to the right-hand side to make the net charge on both sides equal at −2. The oxidation half-equation becomes:

$$2Cl^-(aq) \longrightarrow Cl_2(aq) + 2e^-$$

Each half-equation is now balanced. The net charge on each side of the reduction equation is +2 and the net charge on each side of the oxidation equation is −2.

STEP 6 The previous step has resulted in an uneven number of electrons, which will not cancel if the half-equations are added. To overcome this, multiply each half-equation by factors that will lead to the same number of electrons in each half-equation. In the case of the reduction half-equation, multiply by 2, and for the oxidation half-equation multiply by 5. The reduction half-equation becomes:

$$2MnO_4^-(aq) + 16H^+(aq) + 10e^- \longrightarrow 2Mn^{2+}(aq) + 8H_2O(l)$$

The oxidation half-equation becomes:

$$10Cl^-(aq) \longrightarrow 5Cl_2(aq) + 10e^-$$

STEP 7 Add the half-equations.

$$2MnO_4^-(aq) + 10Cl^-(aq) + 16H^+(aq) + 10e^- \longrightarrow 2Mn^{2+}(aq) + 5Cl_2(aq) + 8H_2O(l) + 10e^-$$

Then cancel the electrons.

$$2MnO_4^-(aq) + 10Cl^-(aq) + 16H^+(aq) \longrightarrow 2Mn^{2+}(aq) + 5Cl_2(aq) + 8H_2O(l)$$

This method of balancing equations is called the half-reaction or half-equation method.

Summary of steps for balancing half-equations

1. Write the half-equations for oxidation and reduction showing conjugate pairs.
2. Balance all elements except hydrogen and oxygen.
3. Balance oxygen atoms, where needed, by adding water.
4. Balance hydrogen atoms, where needed, by adding H^+.
5. Balance the charge by adding electrons.
6. Multiply each half-equation by factors that will lead to the same number of electrons in each half-equation.
7. Add the half-equations and omit the electrons.

Sample problem 13.7

When copper metal is placed in concentrated nitric acid, the brown gas nitrogen dioxide is formed and the solution turns green. Nitrogen dioxide is often seen as a pollutant in industrial areas. Write a balanced equation for this redox reaction.

Solution: **STEP 1**

Write the half-equations for oxidation and reduction showing conjugate pairs.

reduction $NO_3^-(aq) \longrightarrow NO_2(g)$

oxidation $Cu(s) \longrightarrow Cu^{2+}(aq)$

STEP 2

It is not necessary in this example to balance elements other than hydrogen and oxygen as the main elements are already balanced.

STEP 3

Balance the oxygen atoms (where needed) by adding water.

$$NO_3^-(aq) \longrightarrow NO_2(g) + H_2O(l)$$

STEP 4

Balance hydrogen by adding H^+ where needed.

$$NO_3^-(aq) + 2H^+(aq) \longrightarrow NO_2(g) + H_2O(l)$$

STEP 5

Balance the charges by adding electrons so that the net charge on each side is equal to zero.

$$NO_3^-(aq) + 2H^+(aq) + e^- \longrightarrow NO_2(g) + H_2O(l)$$

$$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$$

STEP 6

Balance the electrons by multiplying the first equation by 2. The two half-equations are

$$2NO_3^-(aq) + 4H^+(aq) + 2e^- \longrightarrow 2NO_2(g) + 2H_2O(l)$$

$$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$$

STEP 7

Add the two half-equations:

$$Cu(s) + 2NO_3^-(aq) + 4H^+(aq) + 2e^- \longrightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l) + 2e^-$$

Cancel the electrons

$$Cu(s) + 2NO_3^-(aq) + 4H^+(aq) \longrightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$$

Revision question

24. Balance the following equations using the half-equation method.

(a) $I_2(s) + H_2S(g) \longrightarrow I^-(aq) + S(s)$

(b) $NO_3^-(aq) + H_2S(g) \longrightarrow NO(g) + S(s)$

(c) $Cr_2O_7^{2-}(aq) + H_2S(g) \longrightarrow Cr^{3+}(aq) + S(s)$

(d) $SO_3^{2-}(aq) + MnO_4^-(aq) \longrightarrow SO_4^{2-}(aq) + Mn^{2+}(aq)$

(e) $Cu(s) + NO_3^-(aq) \longrightarrow NO(g) + Cu^{2+}(aq)$

Potassium metal reacts very vigorously with water in a redox reaction to produce hydrogen gas and potassium hydroxide. The heat given off during this reaction is sufficient to ignite the hydrogen gas and allows it to burn in air with a bright flame. Write the two redox equations for the reaction shown in the photograph.

The reactivity of metals

When zinc is placed in copper sulfate solution a reaction occurs, causing copper metal to form on the zinc, as shown below. This is a **displacement reaction**: zinc removes copper from the solution and, as a result, the deep blue colour of the solution pales. If you place a copper strip, however, in a solution of zinc sulfate, no reaction occurs. Therefore, zinc is more reactive than copper.

Different metals have varying abilities to react with other substances. Potassium is so reactive that it is found naturally as a compound rather than a pure metal. When extracted, it must be stored in oil to prevent exposure to moisture and oxygen. Gold, on the other hand, is very unreactive. It occurs in nature as a pure element and can be found by mining or panning. Ornamental gold is valuable because it remains untarnished for centuries.

Metals vary in their reactivity.

Potassium and gold are at opposite ends of the **reactivity series of metals**, shown in table 13.6 (see page 265). This series, sometimes called the activity series, lists the metals in order of how readily they react with oxygen, water, steam, dilute acids and salt solutions. Each of these reactions is a redox reaction, in which the metal is always the reductant. The reactivity series can be established experimentally. The most direct method to determine the relative reactivity of metals is to use displacement reactions.

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Unit 2	**Metal displacement reactions** Summary screen and practice questions
AOS 1	
Topic 3	
Concept 4	

To understand how displacement reactions may be used, consider our earlier example, where zinc displaced copper ions from solution to form copper metal and zinc ions according to the equation:

$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

Copper metal does not displace zinc from solution. This demonstrated difference in reactivity forms the basis of the displacement method of determining relative reactivity. Samples of each metal are immersed in separate solutions of metal cations. These cations accept electrons from any other more reactive solid metal atoms that are introduced, forming solid metal atoms themselves. The more reactive metal therefore loses electrons and forms ions in solution.

The order of reactivity can be determined by comparing the displacement reactions of each metal in different metal solutions.

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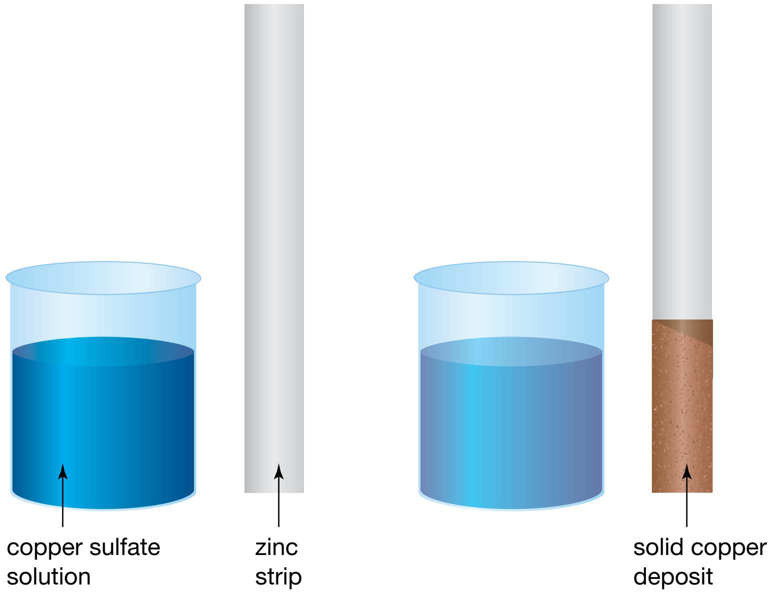
Metals in aqueous solutions



When a strip of zinc is placed in a blue copper sulfate solution, the zinc becomes coated with metallic copper and the blue colour of the solution fades.

TABLE 13.6 Reactivity series of metals

Element	Appearance of metal	Reaction with oxygen	Reaction with water or steam	Displacement of metals	Reaction with dilute acids
K	dull; stored under oil	oxidises in air at room temperature to give oxides	hydrogen formed from cold water	Any metal in this series displaces any other that occurs lower in the series (i.e. change the metal ion back to the metal). The further apart the metals are, the easier the displacement.	hydrogen formed violently
Na					
Ca	generally dull				hydrogen formed with cold acid
Mg			hydrogen formed with steam		
Al					
Zn					
Fe					
Sn		oxidises when heated in air or oxygen to give oxides	no reaction at Bunsen burner temperatures		no visible reaction
Pb					
Cu	generally shiny		no reaction		no reaction
Hg					
Ag		no reaction			
Au					

Note: Aluminium metal often appears to be less reactive than the reactivity series indicates because it has a coating of aluminium oxide and this protects it from reacting further with oxygen. This coating can be thickened by a process called anodising.

Sample problem 13.8

Use the reactivity series to describe what happens when a piece of lead is dipped into a solution of silver nitrate. Write an equation for the reaction and determine the oxidant and reductant.

Solution: Since lead is above silver in the reactivity series, a displacement reaction occurs between lead metal and silver ions. Lead displaces silver from solution, producing silver metal according to the equation:

$$\mathrm{Pb(s) + 2Ag^{+}(aq) \longrightarrow Pb^{2+}(aq) + 2Ag(s)}$$

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Unit 2	**Using the electrochemical series to predict metal displacement reactions** Summary screen and practice questions
AOS 1	
Topic 3	
Concept 5	

From this equation we can deduce that the two half-equations are:

$$Pb(s) \longrightarrow Pb^{2+}(aq) + 2e^-$$

$$Ag^+(aq) + e^- \longrightarrow Ag(s)$$

Lead is, therefore, oxidised and is the reductant. Silver ions are reduced and are, therefore, the oxidant.

Revision questions

25. **Use the reactivity series to determine if a reaction occurs when the following are mixed. Write an equation for any reaction that occurs and identify the oxidant and the reductant.**
  - **(a)** $Mg(s) + Cu^{2+}(aq)$
  - **(b)** $Cu(s) + Fe^{2+}(aq)$
  - **(c)** $Al(s) + Pb^{2+}(aq)$
  - **(d)** $Ag(s) + Zn^{2+}(aq)$

26. **Use the reactivity series to predict what reaction, if any, occurs when the following reagents are mixed. Write half-equations and overall equations for any reaction that you predict.**
  - **(a)** $Pb^{2+}(aq)$ **and** $Fe(s)$
  - **(b)** $Al^{3+}(aq)$ **and** $Fe(s)$
  - **(c)** $Sn^{2+}(aq)$ **and** $Zn(s)$
  - **(d)** $Zn^{2+}(aq)$ **and** $Al(s)$

27. **A student used the reactivity series to predict that a strip of aluminium metal would react with copper(II) sulfate. When he tried the experiment he observed no reaction.**
  - **(a)** **Explain his observations.**
  - **(b)** **What steps would he need to take to prove that his prediction was correct?**

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Zinc/copper

Corrosion

There are two types of corrosion: wet and dry.

When a metal corrodes it suffers damage as a result of chemical change. **Corrosion** is the oxidation of metals by materials in their environment. There are two types of corrosion:

1. **Dry corrosion** occurs when a metal reacts directly with oxygen. Some metals are very reactive and must be stored under oil. Potassium and sodium combine readily with oxygen, so they undergo dry corrosion very quickly. Iron undergoes dry corrosion, but at normal temperatures its rate of dry corrosion is very slow.
2. **Wet corrosion.** This occurs when a metal reacts directly with oxygen in the presence of water. Water can speed up some corrosion reactions, such as that of iron, which corrodes much more quickly in a moist environment.

Corrosion returns metals to their natural states — the original ores. Except for gold and platinum, which are virtually unreactive, most metals are found in the ground as oxidised metals or ores. Since corroded metal often loses its structural purpose and attractiveness, this oxidation reaction has a great economic impact. In Australia, billions of dollars are spent each year preventing and treating corrosion.

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Corrosion

The wet corrosion of iron — an electrochemical process

Corrosion is an electrochemical process.

The wet corrosion of iron is commonly known as **rusting**. Iron rusts when a water drop containing dissolved impurities lands on it. Oxidation usually takes place at a ‘stress site’ on the iron, such as a dent in a car or a scratch on the surface. Iron is oxidised, releasing electrons.

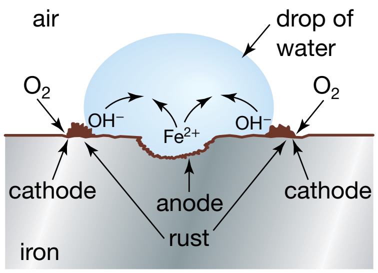
$$Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^-$$

These electrons travel through the metal and are accepted by oxygen, usually at the point where the edge of a water drop is in contact with the air. Oxygen can also come from dissolved oxygen in the water drop. Oxygen accepts electrons and is reduced.

$$O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$$

Fe^{2+} and OH^- ions produced migrate toward each other and react to produce an insoluble green iron(II) hydroxide precipitate. The overall equation is:

$$2Fe(s) + O_2(g) + 2H_2O(l) \longrightarrow 2Fe(OH)_2(s)$$



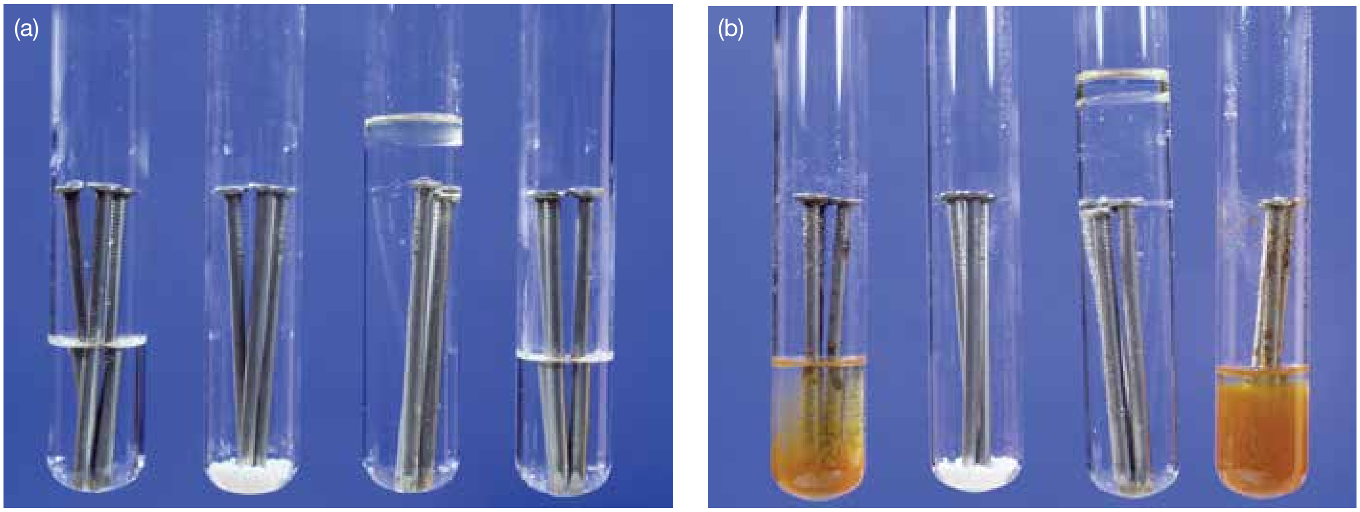
$$2Fe(s)+O_2(g)+2H_2O(l) \longrightarrow 2Fe(OH)_2(s)$$

The corrosion of iron

study on

Unit 2	**Corrosion in the environment**
AOS 1	Summary screen and practice questions
Topic 3	
Concept 6	

These nails are in brown agar gel in which an indicator has been dissolved. Corrosion forms on nails at weak points or tiny pits, shown here in blue. When rust forms, Fe^{2+} ions are formed, shown as blue spots. The pink spots are where phenolphthalein shows that hydroxide ions form and no rusting occurs.



(a) From the left, iron nails are placed in water and air, a drying agent, distilled water topped with oil, and salty water (far right). (b) After several weeks, rust (hydrated iron oxide) is seen in the test tube with water and air. The process has been accelerated in the test tube with salt water and air. The nails exposed to air but no water and to water but no oxygen have not rusted.

Iron(II) hydroxide is very unstable in air and quickly reacts with oxygen to produce brown iron(III) hydroxide.

$$4Fe(OH)_2(s) + O_2(g) + 2H_2O(l) \longrightarrow 4Fe(OH)_3(s)$$

Iron pipes exposed to the atmosphere rust very quickly. Even pipes buried beneath the ground rust. To prevent corrosion, blocks of a more reactive metal, such as magnesium, are bolted to aboveground pipes. For underground pipes, bags of magnesium are placed in holes in the ground and are connected to the pipe by a conductor. When the magnesium has corroded away, it is replaced.

This brown precipitate then partially dehydrates (loses water) to produce iron(III) oxide, $Fe_2O_3{\cdot}xH_2O$, where x is less than 3. The degree of hydration (the value of x) affects the colour of the rust.

This process is responsible for the corrosion of cars, tools, bridges, buildings, machinery and virtually anything made of iron and steel that is not protected in some way. Because it is soft and flaky, rust is easily dislodged, which accelerates further rusting. Water is vital in the process; without a 'salt bridge', iron does not rust. Cars last much longer in dry areas; however, the dissolved salts in the air in seaside areas accelerate rusting. The dissolved ions increase the conductivity of the moisture and therefore encourage corrosion. Other factors increase corrosion, including the presence of acids and pollutants such as sulfur dioxide and nitrogen dioxide, and imperfections in the metal surface.

Iron materials require rust prevention treatments.

Rusting adversely affects the properties of iron.

Adverse effects of corrosion

- Metal oxides, or rust, have less tensile strength and less elasticity than uncorroded metal. Corroded buildings, bridges and machinery may develop weaknesses that result in malfunction or breakage.
- A severely corroded metal cannot conduct electricity. Products of corrosion formed in a car battery, between the terminals and the leads, may cause the electrical system to malfunction.
- Corroded copper pipes and hot water tanks leak as the copper compounds dissolve in the water flowing through them.
- Products of corrosion are brittle and flake off, resulting in holes in car mufflers or iron roofing. In car radiators, iron oxide flakes may cause blockages.
- A corroded metal occupies a larger volume than the original metal. Corroded nuts and bolts may jam machinery.

Three main methods of corrosion prevention are surface protection, alloying and electrochemical protection.

Corrosion protection

It is important to protect iron and steel from corrosion. There are three main methods: surface protection (as described in chapter 4), **alloying** and **electrochemical protection**.

Metals can be protected with plastic, paint, grease or another metal.

Surface protection

If a metal surface can be prevented from coming into contact with oxygen or water then it will not corrode. An obvious way of achieving this is to coat the surface in some way. Commonly used surface coatings are plastic, paint, grease or oil, noble metals and sacrificial metals. Specific details of each of these methods may be found in chapter 4.

Alloying can be used to minimise corrosion.

Alloying

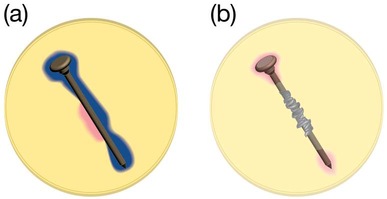
Iron can be alloyed with small quantities of metals, such as chromium, nickel, manganese and molybdenum, to produce stainless steel. A wide variety of grades of stainless steel is produced for specific purposes. The most widely used stainless steel contains approximately 18% chromium and 8% nickel. This is called 18-8 stainless steel and is used for kitchen sinks and cutlery. The formation of a stable film of chromium(III) oxide, Cr_2O_3, provides a very strong surface protection for this alloy, although the metal may still be attacked by chloride ions. It is important, therefore, to choose an appropriate alloy for any specific use. For example, stainless steel containing molybdenum is more resistant to chloride ion attack and is therefore recommended for use in sea water; it is used in valves on scuba diving equipment.

Electrochemical protection

Electrochemical protection involves placing a more reactive metal in electrical contact with the metal that needs protection. It involves 'sacrificing' one metal to save another.

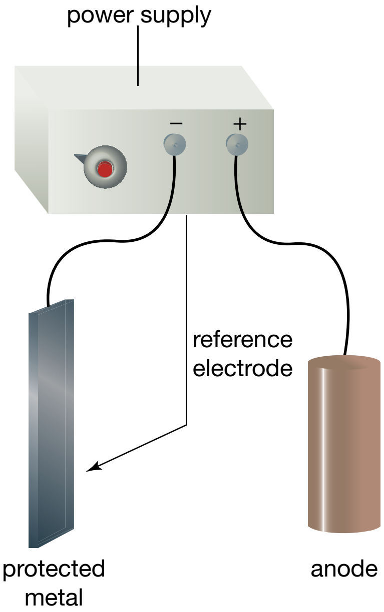
If a block of a metal higher in the reactivity series is connected to iron, then that metal is corroded while iron remains intact. Underground pipes may be protected by attaching bags of magnesium scraps at intervals and replacing them when they have been corroded away. The hulls of ships can be protected by attaching blocks of zinc, which are sacrificed to protect the iron. This is sometimes referred to as **sacrificial protection** or **cathodic protection**. It involves the selection of an appropriate sacrificial anode. In the diagram above left, magnesium wrapped around an iron nail protects the nail from corrosion.

Another method of cathodic protection is to make iron the cathode by connecting it to the negative terminal of a battery while a conductor such as graphite is connected to the positive terminal. The formation of Fe^{2+} ions is inhibited by the negative impressed current on the iron.



A demonstration of cathodic protection. Identical nails are introduced to two beakers of agar, each containing phenolphthalein indicator and potassium hexacyanoferrate, $K_4Fe(CN)_6$. Nail (a) is unprotected. Nail (b) has magnesium ribbon wrapped around its centre. The pink colouring indicates rusting.

Electrochemical protection involves placing a more reactive metal in electrical contact with the metal that needs protection.



Impressed-current cathodic protection supplies electrons to the protected metal by connecting it to the negative terminal. The current is monitored by a reference electrode that keeps the potential difference between the cathode and the anode high enough to ensure protection.

Car manufacturers treat newly formed car body shells so that they become resistant to corrosion and attractively coloured. This involves immersing each car body in a series of treatment and paint tanks containing zinc. The body then receives an electrocoat primer on the inside and outside surfaces to provide corrosion protection and prepare the surface for the coloured top-coat. Paint spraying is usually performed by automatic machines using electrostatic attraction. The car body is then heated to produce a hard, glossy finish.

studyon

Unit 2	**Oxidation numbers**
AOS 1	Summary screen and practice questions
Topic 3	
Concept 2	

Oxidation numbers

So far you have learned to identify oxidation and reduction in terms of production and consumption of electrons, respectively. From an overall ionic equation, it has been possible to identify conjugate redox pairs and then write the half-equations. From the positions of the electrons in these half-equations, it has then been possible to identify the oxidation and reduction processes. This procedure is excellent for consolidating an understanding of redox reactions but can be time consuming where reactions are complicated.

In order to simplify the identification of oxidation and reduction processes, chemists have developed *oxidation numbers*. These are assigned to each of the atoms in an equation using another easily learned set of rules.

Oxidation numbers will be introduced in Unit 3.

Revision questions

28. *Corrosion collage*

In a small group, collect samples of rusted metals to produce a display. Design and perform experiments that investigate the corrosion of the metal samples your group has collected. Label each sample with:
   - (a) **the name of the metal**
   - (b) **common uses of the metal**
   - (c) **the chemical reaction that describes the corrosion of the metal**
   - (d) **methods of corrosion control for the metal**
   - (e) **experimental investigation results relating to the metal's corrosion rate**
   - (f) **a description of the reactivity of the metal with respect to its position in the periodic table**
   - (g) **the cost to society of the corrosion of the metal.**

29. **Prepare a multimedia presentation that shows examples of metal corrosion in your home and wider environment. Your presentation should include the following:**
   - **chemical reactions for the examples of metal corrosion featured**
   - **a discussion of the effects of corrosion on the properties and thus the uses of your featured metals**
   - **the environmental factors that contribute to the corrosion**
   - **the cost of such corrosion to society.**

You should also describe the practices currently employed to control corrosion of the examples shown, and suggest other reasonable methods of corrosion minimisation. A digital camera could be used to document your examples.

30. **Find out from a local industry or small business how they deal with either corrosion prevention or replacement due to corrosion.**

Chapter review

studyon
Unit 2
AOS 1
Topics 2 & 3
Reactions in water
Summary screens and practice questions

Summary

- Common acids include HCl, H_2SO_4, HNO_3, H_3PO_4, CH_3COOH and H_2CO_3. These have similar properties. They taste sour, are corrosive and molecular in structure and turn litmus red.
- Chemical reactions of acids include:
  - acid + metal (not Cu, Hg or Ag) $\longrightarrow$ salt + hydrogen gas
  - acid + metal carbonate $\longrightarrow$ salt + carbon dioxide gas + water
  - acid + metal hydrogen carbonate $\longrightarrow$ salt + carbon dioxide gas + water
  - acid + metal sulfite $\longrightarrow$ salt + sulfur dioxide gas + water
  - acid + metal sulfide $\longrightarrow$ salt + hydrogen sulfide gas
  - acid + metal oxide (basic oxide) $\longrightarrow$ salt + water
  - acid + base (metal hydroxide) $\longrightarrow$ salt + water.
- Common bases include metal hydroxides (e.g. LiOH, NaOH, KOH) and molecular ammonia, NH_3. These have similar properties. They taste bitter, feel slippery, are corrosive and turn litmus blue.
- Alkalis are bases that dissolve in water (e.g. group 1 hydroxides).
- The Brønsted–Lowry theory defines acids and bases as follows:
  - An acid is a proton donor.
  - A base is a proton acceptor.

Acid–base reactions thus involve proton transfer, in which one hydrogen ion, H^+, is transferred from the acid to the base.

- Acids and bases may behave in water in the following ways:
  - Ionisation occurs when an acid donates a proton to water. The water acts as a base and forms the hydronium ion, H_3O^+.
  - Hydrolysis occurs when an anion reacts with water to produce OH^-, or a cation reacts with water to produce H_3O^+.
  - When an ionic base dissolves in water, it dissociates or separates into its constituent ions.
- When an acid and base react, the process is called neutralisation and a salt and water are usually produced. The ionic equation of neutralisation is:

$$H^+\,(aq) + OH^-\,(aq) \longrightarrow H_2O(l)$$

(where H^+ is written in place of H_3O^+).

- A conjugate acid–base pair is formed when an acid reacts with a base.
  - These pairs differ by a proton, H^+
  - (e.g. HCl/Cl^-, H_3O^+/H_2O).
  - To find a conjugate pair, subtract one H^+ from the acid or add it to the base.
- Acids can be classified according to their ability to donate protons:
  - Monoprotic acids can donate one proton (e.g. HCl).
  - Polyprotic acids can donate more than one proton, diprotic acids can donate two protons (e.g. H_2SO_4), and triprotic acids can donate three protons (e.g. H_3PO_4).
  - Amphiprotic substances can act as acids or bases depending on their chemical environment (e.g. H_2O, HS^-, $HSO_4{}^-$).
- The strengths of acids (or bases) relate to their ability to readily donate (or accept) protons. By contrast, concentration of acids (or bases) relates to the amount of the substance that is dissolved in a given volume of water. A large amount results in a concentrated solution while a small amount gives a dilute solution.
  - Strong acids include nitric, sulfuric and hydrochloric acids. Ethanoic acid is a weak acid.
  - Strong bases include sodium hydroxide and potassium hydroxide. Ammonia is a weak base.
- The pH scale is a convenient way of measuring the level of acidity or basicity in a solution.
  - In neutral solutions, pH = 7.
  - In acidic solutions, pH < 7.
  - In basic solutions, pH > 7.
- pH depends on H_3O^+ concentration and is defined by the formula:

$$pH = -\log_{10}[H_3O^+]$$

- In any aqueous solution at 25 °C:

$$[H_3O^+][OH^-] = 10^{-14}$$

- pH measurements can be used to compare the relative strengths of acids (and bases) providing the concentrations of the acids (or bases) are the same.
- Two significant environmental issues associated with acid–base chemistry are acid rain and ocean acidification.
- An indicator is a dye that is a weak acid or base. It has a different colour in acidic and basic solutions. Universal indicator is a commonly used indicator.
- Oxidation–reduction (or redox) reactions are those in which electrons are transferred from the reductant to the oxidant.
- Oxidation and reduction are complementary processes; one cannot occur unless the other occurs simultaneously.
- Oxidation can be defined as the loss of electrons.

- Reduction can be defined as the gain of electrons.
- An oxidant can be defined as a substance:
  - that accepts electrons
  - that undergoes reduction.
- A reductant can be defined as a substance:
  - that donates electrons
  - that undergoes oxidation.
- Oxidation–reduction reactions can be represented as the sum of two half-equations — one for oxidation and one for reduction. A conjugate redox pair is made up of two species that differ by a certain number of electrons, and each has its own half-equation.

  These must be balanced to show the relative amounts of oxidant and reductant such that the number of electrons taken from the reductant is the same as the number of electrons accepted by the oxidant.
- Metals can be ranked in an approximate order of reactivity, called the reactivity (or activity) series of metals, according to how readily each reacts with oxygen, water, steam, dilute acids and salt solutions. Displacement reactions, where a more reactive solid metal displaces the ions of another metal from solution, can be used to determine this ranking.
  - The order of some selected metals in the reactivity series is, from most reactive to least reactive: K, Na, Ca, Mg, Al, Zn, Fe, Sn, Pb, Cu, Hg, Ag, Au.
  - More reactive metals displace the cations of less reactive metals from solution.
- Corrosion is the oxidation of metals by materials in their environment. It has two types, wet and dry corrosion.
- Rusting is the corrosion of iron by oxygen in a moist environment.
  - Rust can be represented by the formula $Fe_2O_3{\cdot}xH_2O$ where x is a number from 1 to 3.
  - Rusting requires the presence of oxygen, to act as the oxidant, as well as water and dissolved salts, to act as the electrolyte.
  - Corrosion of iron occurs where the metal lattice has been deformed.
- Corrosion affects the properties of metals, which can be protected from corrosion by one of three methods:
  - surface protection, which may be achieved using plastic, paint, grease or oil, or through the application of metal coatings, such as noble coating and sacrificial coating
  - alloying
  - electrochemical protection, also known as sacrificial protection or cathodic protection, where a sacrificial anode of another more reactive metal (such as zinc or magnesium) is introduced.
- Common metal coatings discussed in this chapter include:
  - Galvanised iron has a sacrificial coating of zinc, which is more reactive than iron. At a break in the sacrificial coating, protection of the steel is maintained.
  - Stainless steel is produced by alloying and is corrosion resistant because of a protective layer of chromium(III) oxide.
  - Aluminium is corrosion resistant because of a tightly bound layer of aluminium oxide.

Multiple choice questions

1. Which of the following is the best definition of an acid?

A An acid is an electron acceptor.
B An acid is a substance that tastes sour.
C An acid is a substance that donates H^+.
D An acid is a substance that accepts H^+.

2. Consider the following equation and state the species that are acting as bases in this reaction.

$$HCN(aq) + H_2O(l) \longrightarrow H_3O^+ (aq) + CN^-(aq)$$

A $HCN(aq)$ and $H_2O(l)$
B $HCN(aq)$ and $H_3O^+(aq)$
C $HCN(aq)$ and $CN^-(aq)$
D $H_2O(l)$ and $CN^-(aq)$

3. The sourness of food is a reasonable guide to its acidity, since sourness increases with increasing acidity. If food A is more sour than food B, then it is likely that:

A A has a higher pH than B
B A has more $OH^-(aq)$ ions than B
C B is a stronger acid than A
D A has a higher concentration of $H_3O^+(aq)$ ions than B.

4. Hydrogen sulfide, H_2S, is a gas at room temperature and dissolves in water. When this solution is tested, it conducts electricity and turns litmus solution red.

The solution is acidic because:

A the H_2S has gained a proton
B the H_2O has gained a proton
C the H_2S and H_2O have gained protons
D the H_2S and H_2O have lost protons.

5. The conjugate base of the species $H_2PO_3{}^-(aq)$ is:

A $H_3PO_3(aq)$
B $HPO_4{}^{2-}(aq)$
C $H_2PO_4(aq)$
D $HPO_3{}^{2-}(aq)$.

6. Consider the following equations and select those that can represent acid–base reactions.

(i) $Zn(s) + 2H_3O^+(aq) \longrightarrow Zn^{2+}(aq) + H_2(g) + 2H_2O(l)$
(ii) $Cl^-(aq) + H_3O^+(aq) \longrightarrow HCl(aq) + H_2O(l)$
(iii) $H_3O^+(aq) + OH^-(aq) \longrightarrow 2H_2O(l)$

A (i) and (ii)
B (i) and (iii)
C (ii) and (iii)
D (i), (ii) and (iii)

7. Which of the following materials is the most acidic?

A Grapefruit juice of pH 3.0
B Detergent of pH 10.5
C Gastric juice of pH 1.0
D Sea water of pH 8.5

8. Ammonia solution is alkaline due to the presence of:

A NH_4^+ ions
B OH^- ions
C NH_3 molecules
D both NH_4^+ and OH^- ions.

9. When hydrogen chloride gas is dissolved in water, the resulting solution can conduct an electric current. The best explanation is that:

A ions in the hydrogen chloride gas are released when dissolved in water
B the water reacts with the hydrogen chloride to form ions
C the water molecules carry the electric current in one direction while the hydrogen chloride molecules carry it in the other direction
D the water contains impurities that react with the hydrogen chloride to form ions.

10. Which one of the following species acts as an amphiprotic species in aqueous solution?

A SO_4^{2-}
B HCO_3^-
C PO_4^{3-}
D Cl^-

11. The strength of an acid is determined by:

A the number of hydrogen ions present
B its concentration
C its degree of ionisation in aqueous solution
D its ability to change the colour of litmus.

12. The pH of four acids, labelled A, B, C and D, was measured at equal concentrations of each. The values obtained were 4.2, 2.3, 5.3 and 4.1 respectively. The relative strengths of these acids, from strongest to weakest is:

A C, B, D, A
B B, D, A, C
C B, C, D, A
D C, A, D, B.

13. The pH of a solution having an OH^- concentration of 0.000 01 M is:

A 5
B 7
C 9
D 14.

14. When dissolved in rain, which of the following gases does not cause it to be acidic?

A O_2
B CO_2
C NO_2
D SO_3

15. Oxidation is best described as a process that involves:

A the gain of protons
B the loss of oxygen
C the gain of electrons
D the loss of electrons.

16. Which of the following statements is *false*?

A The reactant from which electrons are taken in a redox reaction is reduced.
B Oxidation may involve the addition of oxygen.
C A reductant undergoes oxidation.
D Oxidation does not take place unless a reduction reaction can occur simultaneously.

17. The process of black and white film photography depends upon a reaction initiated by light and represented by the following equation:

$$2Ag^+(aq) + 2Br^-(aq) \longrightarrow 2Ag(s) + Br_2(aq)$$

Name the species that is oxidised in this reaction.

A $Ag^+(aq)$
B $Br^-(aq)$
C $Ag(s)$
D $Br_2(aq)$

18. For the reaction $Al \longrightarrow Al^{3+} + 3e^-$, which of the following is *true*?

A Aluminium loses electrons from its nucleus.
B Al is reduced to Al^{3+}.
C The electrons remain free in aqueous solution.
D This reaction can occur only if there is another reaction that is occurring simultaneously and using electrons.

19. Consider the following half-equation:

$$2Cl^-(aq) \longrightarrow Cl_2(aq) + 2e^-$$

This is an example of:

A an acid–base reaction
B a precipitation reaction
C a reduction reaction
D an oxidation reaction.

20. In the reaction:

$$FeI_2(aq) + Cl_2(g) \longrightarrow FeCl_2(aq) + I_2(aq)$$

the chlorine molecule:

A oxidises the $I^-(aq)$ ion
B oxidises the $Fe^{2+}(aq)$ ion
C reduces the $I^-(aq)$ ion
D reduces the $Fe^{2+}(aq)$ ion.

21. In which of the following reactions does $Fe^{2+}(aq)$ act as a reductant?

A $Fe^{2+}(aq) + S^{2-}(aq) \longrightarrow FeS(s)$
B $Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$
C $Fe^{2+}(aq) + Mg(s) \longrightarrow Fe(s) + Mg^{2+}(aq)$
D $2Fe^{2+}(aq) + Br_2(aq) \longrightarrow 2Fe^{3+}(aq) + 2Br^-(aq)$

22. Zinc powder reacts with sulfur and this reaction may be represented by the equation:

$$Zn(s) + S(s) \longrightarrow ZnS(s)$$

In this reaction:

A zinc is the oxidant
B sulfur is reduced
C sulfur is the reductant
D zinc is reduced.

23. The half-equation for the reduction of NO_3^- to N_2O can be represented as follows:

$$a\mathrm{NO_3^-(aq)} + b\mathrm{H^+(aq)} + 8e^- \longrightarrow c\mathrm{N_2O(g)} + d\mathrm{H_2O(l)}$$

where the values of *a* and *b* respectively are:

A 2 and 5
B 2 and 10
C 1 and 4
D 1 and 5.

24. The more reactive a metal:

A the less easily it is oxidised
B the more easily it gains electrons
C the more easily it is reduced
D the greater is its tendency to form positive ions.

25. A small piece of silver was placed in a solution containing both magnesium nitrate and copper(II) nitrate.
Which one of the following will occur?

A No reaction occurs.
B The silver dissolves and only copper is precipitated.
C The silver dissolves and only magnesium is precipitated.
D A mixture of copper and magnesium forms on the silver.

26. If a piece of shiny aluminium metal is placed in a $CuSO_4$ solution:

A sulfur dioxide gas is liberated
B no reaction occurs
C copper oxide is formed on the aluminium metal
D copper metal is deposited on the aluminium metal.

27. What would happen if you tried to store $FeSO_4$ solution in a container made of Cu metal?

A The $FeSO_4$ solution would be stored quite safely.
B The copper of the container would dissolve and Fe metal would form.
C The copper container would dissolve and Fe^{2+} ions would form.
D The copper container would dissolve and O_2 gas would form.

28. Which of the following metals could be used to make a container to store an aqueous copper sulfate solution?

A Sn
B Pb
C Ag
D Fe

29. Aluminium, although a fairly reactive metal, does not react with oxygen gas in air because:

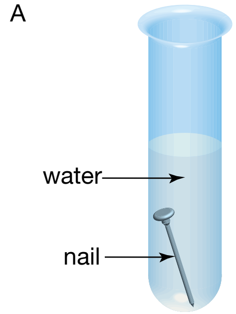
A it is protected by small amounts of its own impurities
B its oxide forms a hard protective layer on the metal
C nitrogen and carbon dioxide gas in the air prevent oxidation
D it forms a hard protective layer by a reaction with sulfur in the air.

30. The underground pipe used to deliver gas to some homes passes through damp marshland. Iron pipes are particularly susceptible to corrosion in the environment. The overall corrosion reaction involves:

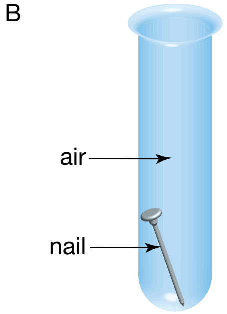
A the reduction of iron
B the oxidation of iron to Fe^{3+} ions
C the oxidation of oxygen to OH^- ions
D the oxidation of iron with acid from dissolved carbon dioxide.

31. The diagrams below represent four tests in an experiment designed to find the conditions necessary for a steel nail to rust.

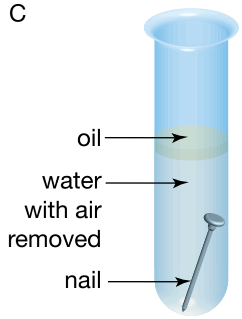
A



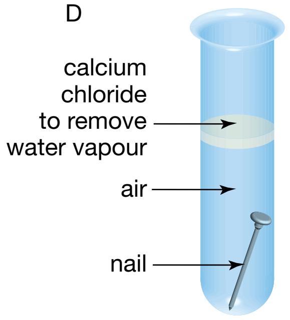
B



C



D



Which diagram represents the most favourable conditions for rust to appear on the nail?

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

32. Iron coated with a complete layer of tin does not rust because:

A the tin acts as a sacrificial metal, becoming oxidised to prevent the iron rusting
B the tin prevents the iron coming into contact with moisture and air
C the tin combines with the iron to form a new compound that does not rust
D particles of rust are unable to stick to the tin surface.

33. A large piece of magnesium buried in the ground and connected to an iron pipe prevents the corrosion of the iron pipe.
The best explanation for this fact is that:

A the Mg forms a protective coating of $Mg(OH)_2$ on the iron
B the Mg is sacrificially oxidised, and the electrons released flow to the Fe and prevent its corrosion

C a protective coating of Fe^{2+} forms on the iron as electrons flow from Fe to Mg

D Mg atoms fill the vacancies in the iron crystals left by the Fe atoms lost during corrosion.

34. Iron rubbish bins coated with a complete layer of zinc do not rust because:

A the zinc acts as a sacrificial metal, becoming oxidised to prevent the iron rusting

B iron is a more reactive metal than zinc

C the zinc combines with the iron to form a new compound that does not rust

D particles of rust are unable to stick to the zinc surface.

Review questions

Properties of acids and bases

1. (a) What are the characteristic properties of acids?
(b) Illustrate two of these properties with an appropriate equation.

2. (a) List three properties of bases.
(b) Give two uses of bases that illustrate these properties.

3. (a) Define a base and give two examples.
(b) Define an alkali and give two examples.

4. Classify the following list of properties and uses as: acidic; basic; both acidic and basic; or neutral.
(a) A sour taste
(b) The bubbles in champagne
(c) Corrosive
(d) Dissolves oxide coating on metals
(e) Dissolves marble statues
(f) Feels slippery
(g) Foods that taste bitter
(h) Breaks down vegetable matter
(i) Can be used to clean drains
(j) Produced in stomach
(k) Used for upset stomachs

Reactions of acids

5. Write both full and ionic equations for the reactions that result when the following substances are mixed. Assume all reactions go to completion.
(a) lead(II) sulfide and phosphoric acid (H_3PO_4)
(b) copper(II) oxide and hydrochloric acid
(c) potassium hydroxide and sulfuric acid
(d) aluminium and hydrochloric acid
(e) calcium and carbonic acid
(f) lithium carbonate and nitric acid

The Brønsted–Lowry theory of acids and bases

6. In terms of the Brønsted–Lowry theory, define an acid, a base and a proton-transfer reaction. You may use an equation in your explanations.

7. Distinguish between the terms ‘hydrolysis’, ‘ionisation’ and ‘dissociation’.

8. Explain, using equations, why a solution of hydrochloric acid conducts electricity but pure hydrogen chloride does not.

9. Give two common uses for sodium carbonate, Na_2CO_3.

10. Antacid tablets are used to neutralise the hydrochloric acid in the stomach. Use the three bases mentioned in the discussion of antacids on page 243 to write neutralisation equations for the reactions that may happen in our bodies.

11. A white cloud is formed when ammonia is neutralised by HCl. Write an equation for the formation of the white cloud in this reaction.

12. (a) What are conjugate acid–base pairs?
(b) State the formulas of the conjugate acid of the following bases:
(i) Cl^-
(ii) CO_3^{2-}
(iii) HSO_4^-
(iv) H_2O
(v) OH^-.
(c) State the formulas of the conjugate base of the following acids:
(i) HF
(ii) HCO_3^-
(iii) HSO_4^-
(iv) H_2O
(v) H_2S.

13. Identify the conjugate pairs in the following equations.
(a) $HS^-(aq) + NH_3(aq) \longrightarrow NH_4^+(aq) + S^{2-}(aq)$
(b) $NH_4^+(aq) + CH_3COO^-(aq) \longrightarrow NH_3(aq) + CH_3COOH(aq)$
(c) $HSO_4^-(aq) + H_2O(l) \longrightarrow SO_4^{2-}(aq) + H_3O^+(aq)$
(d) $HPO_4^{2-}(aq) + H_2O(l) \longrightarrow OH^-(aq) + H_2PO_4^-(aq)$
(e) $HNO_3(aq) + OH^-(aq) \longrightarrow H_2O(l) + NO_3^-(aq)$

14. Write two equations to illustrate the ampholyte HCO_3^-.

Strengths of acids and bases

15. Sulfuric acid is a strong, diprotic acid. Explain, using equations, what this term means.

16. (a) Explain why an aqueous solution of an acid conducts electricity.
(b) What is the term given for such a solution?
(c) Explain why a strong acid conducts electricity to a greater extent than a weak acid.

17. Andrew took the lid off a bottle of glacial ethanoic acid, CH_3COOH, and immediately noticed the very potent smell of vinegar. ‘This must be a very strong acid’, he exclaimed. ‘No’, said his sister Thalia, ‘It is merely a concentrated acid’. Using appropriate definitions, explain who is right.

18. Artists often use acids to etch into metals.
(a) What property of hydrochloric and nitric acids enables artists to etch into metals?
(b) Why would HCl ‘bite’ into an artist’s plate more deeply than HNO_3?

The pH scale

19. (a) Identify each of the following solutions as acidic, basic or neutral at 25 °C.

  (i) pH = 6.3
  (ii) pH = 7
  (iii) pH = 8.4
  (iv) pH = 1.2

(b) Which solution has the highest $[H_3O^+]$? Give a reason for your answer.

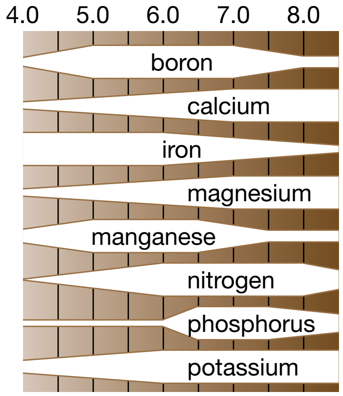
20. A person who ingests too much aspirin (acetyl salicylic acid) may die. Explain what happens to the pH and hydrogen ion concentration.

21. In a pH test on a swimming pool, the water is found to have a pH of 4.5.

(a) What would you add to the water to raise the pH?
(b) The pH should be about 7.5 for best conditions. Is this acidic or alkaline?

22. Soil contains a number of different elements. Plants need to take up these elements in order to grow and flower. If the pH of soil is too high or too low, plants cannot take up these elements.

The figure below shows how the pH of the soil influences the amount of an element that can be taken up by a plant. The narrower the bar, the harder it is for the plant to take up the element.



(a) Which elements can easily be absorbed at a pH below 4.5?
(b) Which elements can be absorbed at a pH of 8?
(c) Which elements cannot be taken up easily if the pH is 6?
(d) Which elements cannot be taken up easily if the soil pH is 8?
(e) Fuchsias require the soil to be acidic. Which elements would be most important to these plants?
(f) Azaleas require a pH of around 5.5. Which elements are not important to these plants?
(g) A gardener comes to you with a problem. He says that he cannot grow lettuce successfully any longer. It is not an insect problem or a water problem. Design an experiment you could do to find out the problem with his soil. (*Note:* Lettuces need a lot of iron.)

pH calculations

23. Calculate the pH of solutions that have the following H_3O^+ concentrations.

(a) 0.0010 M
(b) 0.00050 M
(c) 4.2×10^{-4} M
(d) 6.7×10^{-9} M

24. Explain why, when an acid is diluted by a factor of 2, its pH does not change by the same factor.

25. Calculate the pH of solutions that have the following OH^- ion concentrations.

(a) 0.0010 M
(b) 0.000500 M
(c) 4.20×10^{-4} M
(d) 6.7×10^{-9} M

Environmental issues

26. Choosing either acid rain or ocean acidification, produce a poster or a PowerPoint presentation.

Your poster or presentation must include a description of the problem and its scope, the chemistry involved, including relevant equations, the effect on living things and possible preventative action that could be taken to reduce the problem.

Oxidation and reduction

27. Define oxidation and reduction reactions in terms of the transfer of electrons.

28. What is an oxidant?

29. Use the electron transfer approach to identify the oxidant, reductant, direction of electron transfer and the direction of oxidation and reduction in the following equations. (The diagram on page 259 may be used as a guide.)

(a) $Sn^{2+}(aq) + 2Ce^{4+}(aq) \longrightarrow Sn^{4+}(aq) + 2Ce^{3+}(aq)$
(b) $Pb^{2+}(aq) + Cd(s) \longrightarrow Pb(s) + Cd^{2+}(aq)$
(c) $2Al(s) + 3Cl_2(g) \longrightarrow 2Al^{3+}(aq) + 6Cl^-(aq)$
(d) $2H^+(aq) + Mg(s) \longrightarrow H_2(g) + Mg^{2+}(aq)$
(e) $Na(s) + 2H_2O(l) \longrightarrow Na^+(aq) + 2OH^-(aq) + H_2(g)$

Redox equations

30. Identify the following half-equations as being either oxidation or reduction.

(a) $Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$
(b) $Al^{3+}(aq) + 3e^- \longrightarrow Al(s)$
(c) $O_2(aq) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$

31. Show, using oxidation half-equations and reduction half-equations, that the following reactions are redox reactions.

(a) $4Ag + O_2 \longrightarrow 2Ag_2O$
(b) $Fe + S \longrightarrow FeS$
(c) $Pb + Cl_2 \longrightarrow PbCl_2$
(d) $2Hg + O_2 \longrightarrow 2HgO$

Identify the oxidants and reductants.

32. Balance the following equations using half-equations.

(a) $Br^-(aq) + SO_4^{2-}(aq) \longrightarrow SO_2(g) + Br_2(l)$
(b) $Al(s) + Cl_2(g) \longrightarrow AlCl_3(s)$
(c) $I_2(s) + H_2S(g) \longrightarrow I^-(aq) + S(s)$
(d) $Cu(s) + HNO_3(aq) \longrightarrow Cu^{2+}(aq) + NO(g)$
(e) $Cu(s) + HNO_3(aq) \longrightarrow Cu^{2+}(aq) + NO_2(g)$
(f) $CuO(s) + NH_3(g) \longrightarrow N_2(g) + Cu(s)$
(g) $PbS(s) + H_2O_2(l) \longrightarrow PbSO_4(s) + H_2O(l)$
(h) $Cr_2O_7^{2-}(aq) + CH_3CH_2OH(g) \longrightarrow CH_3COOH(aq) + Cr^{3+}(aq)$

Reactivity series of metals

33. Explain why silver and gold are used for ornamental purposes more often than zinc and iron.

34. What is a displacement reaction? Explain using an example.

35. How does the reactivity of a metal relate to the reducing strength of that metal?

36. Where are the most reactive metals found in the periodic table?

37. Which of the following reactions is more likely to happen if the reactants are mixed?

$$Zn(s) + Pb^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Pb(s)$$
$$Pb(s) + Zn^{2+}(aq) \longrightarrow Pb^{2+}(aq) + Zn(s)$$

38. Develop overall redox equations by first writing the oxidation and reduction equations for the following reactions. Label the oxidant and reductant in each case.

(a) Zinc solid is placed in a copper sulfate solution.
(b) An iron nail is placed in a solution of tin(II) nitrate.
(c) A small piece of lead is placed in silver nitrate solution.
(d) A small piece of aluminium is reacted with zinc chloride solution.

39. Copper sulfate solutions cannot be stored in tin containers. Why is this so?

40. An iron nail is placed in the following aqueous solutions: $AgNO_3$, $Mg(NO_3)_2$, $CuSO_4$, $Pb(NO_3)_2$, $Zn(NO_3)_2$.

(a) In which of the solutions would you expect a coating of another metal on the iron nail?
(b) Write a redox equation for any reaction that may occur.

41. A metal *X* is placed in (a) a solution of NaCl or (b) a solution of $CuSO_4$. There is no reaction with sodium chloride, but *X* is covered with a thin film of copper in the second solution. What can you say about the relative reactivity of *X*, Na and Cu?

42. Tin displaces copper metal from $CuSO_4$ solution and cadmium displaces tin metal from $SnCl_2$ solution. The three metals and their cations are as follows: Sn/Sn^{2+}, Cu/Cu^{2+}, Cd/Cd^{2+}.

(a) Ignoring spectator ions, write an ionic equation for the two displacement reactions described above.
(b) Use the information given to deduce the relative order of metals from most reactive to least reactive.
(c) State whether you would expect the following reactions to occur.
    (i) $Sn^{2+}(aq) + Cu(s) \longrightarrow Sn(s) + Cu^{2+}(aq)$
    (ii) $Cd^{2+}(aq) + Cu(s) \longrightarrow Cd(s) + Cu^{2+}(aq)$
    (iii) $Sn(s) + Cd^{2+}(aq) \longrightarrow Sn^{2+}(aq) + Cd(s)$
    (iv) $Cd(s) + Cu^{2+}(aq) \longrightarrow Cd^{2+}(aq) + Cu(s)$

43. Four hypothetical metallic elements, *A*, *B*, *C* and *D*, form soluble nitrates with the formulas A:NO_3, B: $(NO_3)_2$, C:NO_3 and D:$(NO_3)_3$.
Strips of each of the four metals were immersed in solutions of the other metal nitrates and the following observations were recorded.

- Metal B underwent reaction in all solutions.
- Metal A reacted only with CNO_3.

(a) State the formulas of the metal cations in the soluble nitrates.
(b) Deduce the position of *D* in the series by listing the three metals *A*, *B* and *C* in decreasing order, with respect to their relative reactivities.
(c) List the four metals in order from most reactive to least reactive.
(d) Which metal cation is the weakest oxidant?

44. Six different strips of metal, *A* to *F*, were placed in solutions of their metal nitrates to see if there was any reaction. The results are shown in the table below.

Metal	*A* nitrate	*B* nitrate	*C* nitrate	*D* nitrate	*E* nitrate	*F* nitrate
A	—	✓	✓	✓	×	✓
B	×	—	×	×	×	×
C	×	✓	—	✓	×	×
D	×	✓	×	—	×	×
E	✓	✓	✓	✓	—	✓
F	×	✓	✓	✓	×	—

✓ = metal displaced × = no reaction

(a) Deduce the order of reactivity of the metals *A* to *F* and list them in order, beginning with the most reactive.
(b) Identify the six metals.

Corrosion

45. (a) What is corrosion?
(b) Why is corrosion sometimes referred as a 'cancer'?

46. Explain, using equations, why the corrosion of iron to form rust is a redox process.

47. The properties of metals are changed during corrosion. Explain, using three examples.

48. Salt is often spread on roads in cold countries to melt the ice. Explain why cars rust quickly in these countries.

Corrosion protection

49. List the methods of corrosion protection of steel that depend on:
- (a) prevention of water and air from coming into contact with the metal
- (b) providing electrons to the steel.

50. Explain what would happen if the following coatings of a steel object are damaged.
(a) Plastic (b) Paint (c) Tin (d) Zinc

51. (a) What does the term 'noble coating' mean?
- (b) Why is it undesirable to coat screws with tin as a protection from corrosion?

52. Food containers are tin-plated. Explain why:
- (a) galvanising steel tins is undesirable
- (b) it is dangerous to eat the food from dented tin-lined cans.

53. Which alloy is used for sinks and cutlery? Explain how it protects the steel.

54. It has been estimated that over 20% of the iron produced in the world is used to replace that lost by corrosion.
- (a) List the substances needed for the rusting of iron to occur.
- (b) Name three other substances that speed up the rusting of iron.
- (c) Explain why iron objects are less likely to rust in the following environments:
  - (i) in a desert
  - (ii) in outer space
  - (iii) on the ocean floor.

55. Which of the following methods would be successful in preventing iron from rusting?
- (a) Plating the iron with zinc
- (b) Plating the iron with copper
- (c) Connecting the iron with a wire to a block of zinc
- (d) Connecting the iron with a wire to a block of copper
- (e) Alloying the iron to produce stainless steel

Justify your choices.

56. An iron sculpture was produced to commemorate the anniversary of the founding of a small town. To stop it rusting, the mayor attached it by a wire to a block of zinc that was stored underground out of sight.
- (a) Explain how the mayor's action would prevent corrosion.
- (b) What would happen if a block of lead were to be attached to the iron rather than the block of zinc?
- (c) List two other ways in which corrosion of the statue could be prevented.

Exam practice questions

studyon

Unit 2
AOS 1
Topics 2 & 3

Reactions in water

Sit topic test

In a chemistry examination you will be required to write a number of short and extended response questions.

Extended response questions

1. Ethanoic (acetic) acid is a weak monoprotic acid.

(a) What is meant by the term 'monoprotic'? **1 mark**

(b) Write the equation for the ionisation of ethanoic acid in water. **1 mark**

(c) Write an equation for the neutralisation of ethanoic acid by sodium hydroxide. **1 mark**

(d) What is meant by the term 'weak'? Use your equation from part (b) to illustrate your answer. **2 marks**

(e) Draw the structure of an ethanoic acid molecule. Given that it is monoprotic, circle the hydrogen atom that is donated when it reacts with a base. **2 marks**

2. The hydrogen carbonate ion is amphiprotic.

(a) What is meant by the term 'amphiprotic?' **1 mark**

(b) Write two different equations for the reaction of the hydrogen carbonate ion with water that demonstrate its amphiprotic nature. **2 marks**

(c) Solutions of the hydrogen carbonate ion in water are slightly basic. Comment on this with respect to the two equations that you have written for part (b). **1 mark**

3. A clean piece of tin metal is placed into a solution of lead nitrate and left for a period of time. It is then observed that the tin has become coated with a layer of dull lead metal.

(a) Write the half-equation for the oxidation reaction that has occurred. **1 mark**

(b) Write the half-equation for the reduction reaction that has occurred. **1 mark**

(c) State the oxidant for this reaction. **1 mark**

(d) State the reductant for this reaction. **1 mark**

(e) If the tin is removed, cleaned of lead, and then placed into a solution of magnesium chloride, would a further redox reaction occur? Explain. **2 marks**

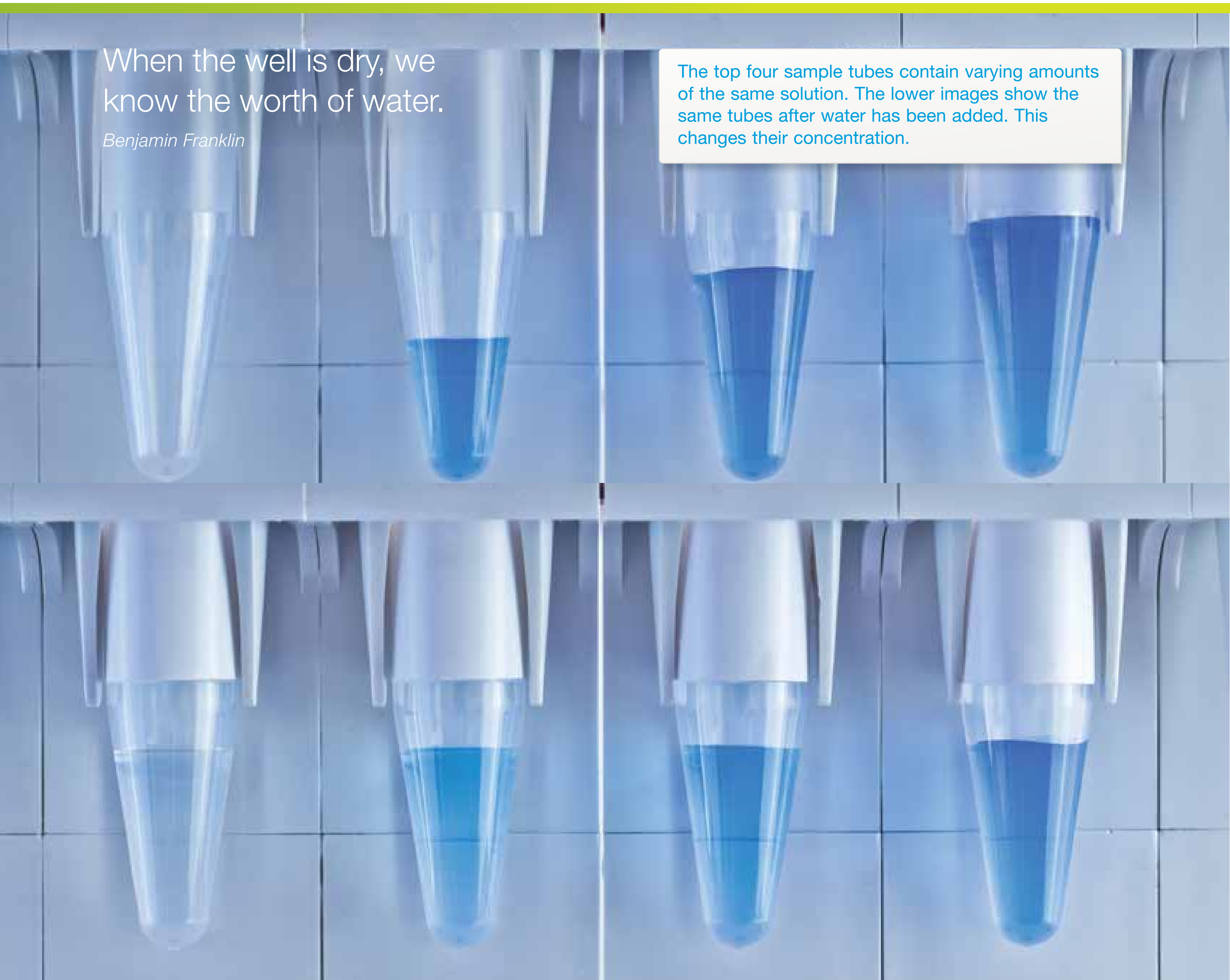
CHAPTER

14 Measuring solubility and concentration

YOU WILL EXAMINE:

- the distribution of water available on the Earth's surface
- how solutions may be described based on the level of solute they contain
- solubility curves as a means of summarising the relationship between solubility and temperature, and their use as predictive tools
- the range of units used to measure concentration
- environmental issues such as how the solubility of gases in the oceans may be affected by increasing temperature and eutrophication
- how supplies of drinking water may be increased by distillation and by reverse osmosis, and how new technologies may affect these processes.

> When the well is dry, we know the worth of water.
>
> *Benjamin Franklin*



The top four sample tubes contain varying amounts of the same solution. The lower images show the same tubes after water has been added. This changes their concentration.

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Unit 2
AOS 2
Topic 1
Concept 1

Availability of fresh water
Summary screen and practice questions

studyon

Unit 2
AOS 2
Topic 1
Concept 2

Water sampling protocols
Summary screen and practice questions

Despite the huge amount of water on the Earth, very little of it is both pure enough and easily sourced for drinking purposes.

Using water

Nearly all the water that is used around the world every day, including that used for drinking, is not pure. This is because, as we have seen in previous chapters, water is an excellent solvent. In the environment, water comes into contact with many substances that dissolve in it. As water falls through the atmosphere in the form of rain, gases from the atmosphere, both natural and pollutants, dissolve in it to some extent. When it travels over the Earth's surface, or soaks into the ground, naturally occurring salts also dissolve. As all water eventually flows into the oceans as part of the water cycle, this is precisely why the sea is so salty and contains a wide range of dissolved substances.

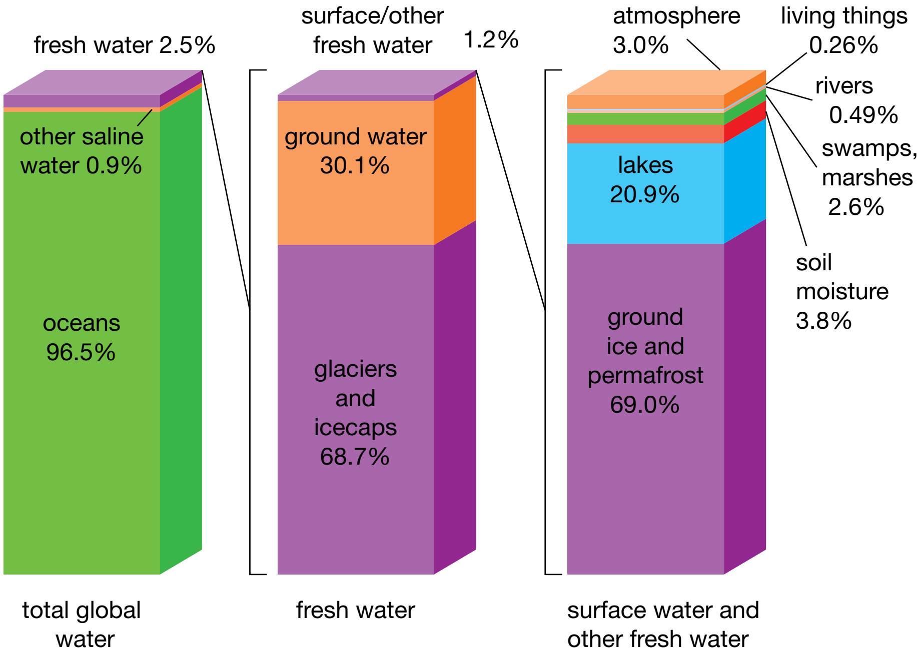

TABLE 14.1 Major ions in sea water

Ion	Typical concentration (mg L^{-1})
chloride, Cl^-	18 980
sodium, Na^+	10 556
sulfate, SO_4^{2-}	2 649
magnesium, Mg^{2+}	1 262
calcium, Ca^{2+}	400
potassium, K^+	380
bicarbonate, HCO_3^-	140
strontium, Sr^{2+}	13
bromide, Br^-	65
borate, BO_3^{3-}	26

The question that we therefore need to ask, from a realistic viewpoint, is whether the water we use is suitable for the purpose we have in mind. To answer such a question, we need to know not only *what* is dissolved in the water, but also *how much*. And we also need to know *how to measure* what the water contains.

Solubility of solids in water

What do cola, coffee, tea, lemonade, soup and cordial have in common? They are all mixtures and they all contain water. Some of them are heterogeneous mixtures (known simply as 'mixtures') and others are homogeneous mixtures called **solutions**. For example, when sodium chloride is dissolved in water, sodium chloride is the solute and water is the solvent. The end result is called the solution.

Depending on the amount of solute present, solutions may be described as saturated, unsaturated or supersaturated.

The solubility of a substance in water depends on temperature. A solubility curve is a graph of solubility (g per 100 g water) versus temperature. Most solids are more soluble as temperature increases.

A **saturated solution** is one that contains the maximum amount of solute for the volume of solution at a particular temperature. When less than the maximum amount of solute has been added to the solvent, the solution is called an **unsaturated solution**.

A **supersaturated solution** can be prepared by slowly cooling a saturated solution. When this is done, the solution contains more solute than it should at that temperature. The addition of a small crystal seed or even some dust causes the excess solute to crystallise.

The solvent properties of water make it an excellent medium for chemical reactions. Water samples containing dissolved substances are called **aqueous solutions**.

Most salts are soluble in water to some extent. Solubility is influenced by the temperature of the solution. Most salts become more soluble at high temperature, but a few become less soluble as the temperature increases. Salts with a solubility greater than 0.1 M of water are classed as soluble and salts with a solubility less than 0.1 M of water are classed as insoluble.

TABLE 14.2 The solubility of various salts

Salt	Formula	Solubility (g per 100 g water at 20 °C)
barium chloride	$BaCl_2$	36.0
barium sulfate	$BaSO_4$	0.000 24
calcium chloride	$CaCl_2$	74.0
calcium sulfate	$CaSO_4$	0.21
copper(II) sulfate	$CuSO_4$	20.5
copper(II) sulfide	CuS	0.000 03
lead(II) sulfate	$PbSO_4$	0.004
potassium chlorate	$KCIO_3$	7.3
potassium nitrite	KNO_2	300.0
silver chloride	$AgCl$	0.000 000 1
silver nitrate	$AgNO_3$	217.0
sodium chloride	$NaCl$	36.0
sodium nitrate	$NaNO_3$	87.0

Solubility tables

Table 11.2 on page 212 is an example of a solubility table. These allow us to predict whether a salt is soluble or insoluble. As we have already seen in chapters 11 and 12, such tables can be used to identify a likely precipitate from the mixing of two or more solutions. In many industries, knowledge of likely precipitates can be used to achieve a desired goal. In chapter 15, we will see how gravimetric analysis, an important analytical technique, relies on the ability to form suitable precipitates. In a popular method for large-scale water purification, aluminium hydroxide precipitate is formed *in situ* by mixing alum (potassium aluminium sulfate, $KAl(SO_4)_2$, and limewater (calcium hydroxide, $Ca(OH)_2$) together. Freshly prepared aluminium hydroxide is a gelatinous precipitate that adsorbs and traps particles of dirt and bacteria from the impure water as it settles. Subsequent filtering easily removes the precipitate and these undesired particles. A solubility table would also warn us that one of the other possible products, calcium sulfate, is only slightly soluble; this would need to be taken into account when calculating the amounts of alum and limewater that need to be added.

Solubility curves

The solubility of soluble substances can be easily measured by experiment. A convenient unit used for this purpose is g/100 g of water. When this is done across a range of temperatures, the result may be drawn as a graph of solubility versus temperature. Such graphs are called *solubility curves*. These find extensive use in medical, pharmaceutical and industrial applications where compounds need to be isolated and purified.

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Unit 2	**Solubility tables**
AOS 2	Summary screen
Topic 1	and practice
Concept 4	questions

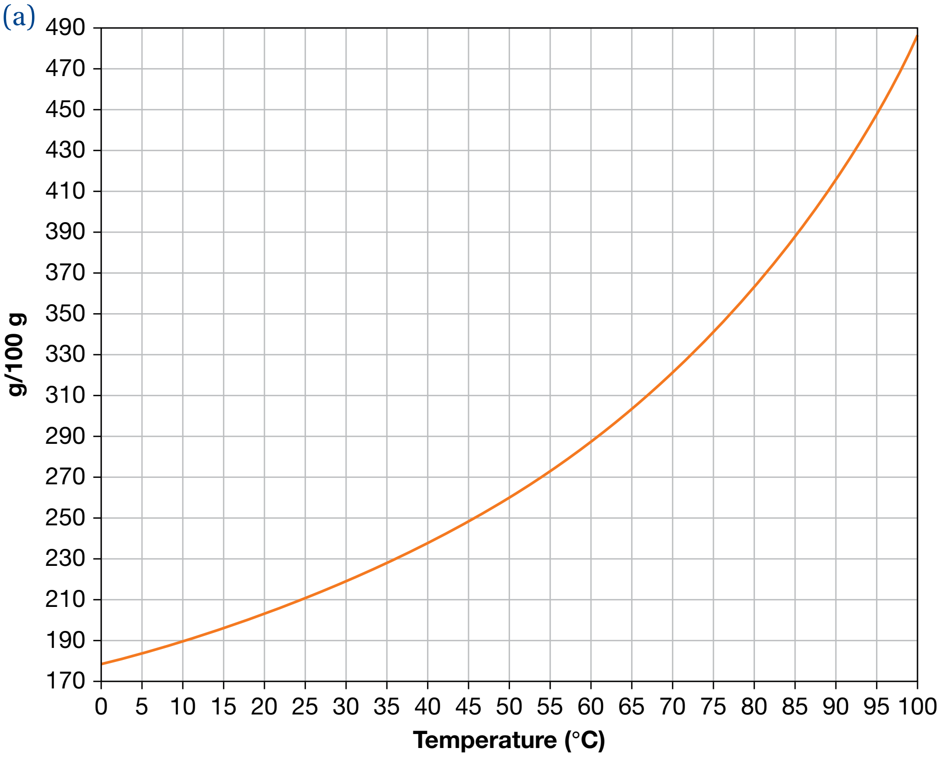
Sample problem 14.1

The solubility of a white crystalline substance known to be a type of sugar was determined over a range of temperatures. The results are shown in the following table.

Temperature (°C)	10	30	50	70	80	90
Solubility (g/100 g)	190	220	287	320	362	415

(a) Plot the solubility curve for this sugar.
(b) What mass of sugar dissolves in 100 g of water at 75 °C?
(c) What mass of sugar dissolves in 250 g of water at 20 °C?
(d) What mass of water is required to dissolve 110 g of sugar at 50 °C?

Solution: (a)



studyon

Unit 2	**Solubility curves**
AOS 2	Summary screen
Topic 1	and practice
Concept 5	questions

(b) At 75 °C, solubility = 340 g/100 g (from graph).

Therefore, 340 g sugar dissolves in 100 g of water.

(c) At 20 °C, solubility = 204 g/100 g.

Therefore, $204 \times \frac{250}{100} = 510$ g dissolves in 250 g of water.

(d) At 50 °C, solubility = 260 g/100 g.

Therefore, $100 \times \frac{110}{260} = 42.3$ g of water is required at this temperature.

Crystallisation

Crystallisation is a technique that is used to isolate substances and is based on the difference in solubility between two temperatures. As a hot solution is cooled, it eventually reaches a temperature at which it is saturated. Further cooling results in it becoming supersaturated — an unstable situation. Crystals then precipitate out of the solution and, the more it is cooled, the more crystals are formed. These crystals can then be collected easily by filtration. This technique can also be used to purify substances. Solubility curves are an important predictive tool for both these processes.

Sample problems 14.2 and 14.3 illustrate both of these techniques.

Sample problem 14.2

A solution of potassium chlorate contains 40 g dissolved in 100 g of water at 90 °C. It is cooled to 20 °C. What mass of crystals form?

(Use the solubility curve for potassium chlorate on page 285.)

Solution: At 90 °C, solubility = 46 g/100 g. Therefore, all the 40 g is dissolved — the solution is unsaturated.

At 20 °C, solubility = 6 g/100 g. Therefore, only 6 g can remain dissolved.

Therefore, the mass of crystals formed = 46 – 6 = 40 g.

Sample problem 14.3

A mixture is known to contain 46 g of potassium chloride and 12 g of potassium sulfate. The mixture is dissolved in 100 g of water at 80 °C.

(a) Do both these salts dissolve at this temperature?

(b) The solution is cooled to 20 °C. Describe what you would expect to happen.

(Use the solubility curves shown on the next page.)

(c) If we attempt to obtain a pure sample of potassium chloride, why would it be inadvisable to cool the solution below 20 °C?

Solution: (a) At 80 °C, solubility of potassium chloride = 46 g/100 g
and solubility of potassium sulfate = 21 g/100 g.

Therefore, both the salts dissolve.

(b) At 20 °C, solubility of potassium chloride = 32 g/100 g.

Therefore, 46 – 32 = 14 g of crystals forms.

Solubility of potassium sulfate = 12 g/100 g.

Therefore, all the potassium sulfate is still dissolved, and no crystals form.

We therefore predict that 12 g of pure potassium chloride crystals is formed.

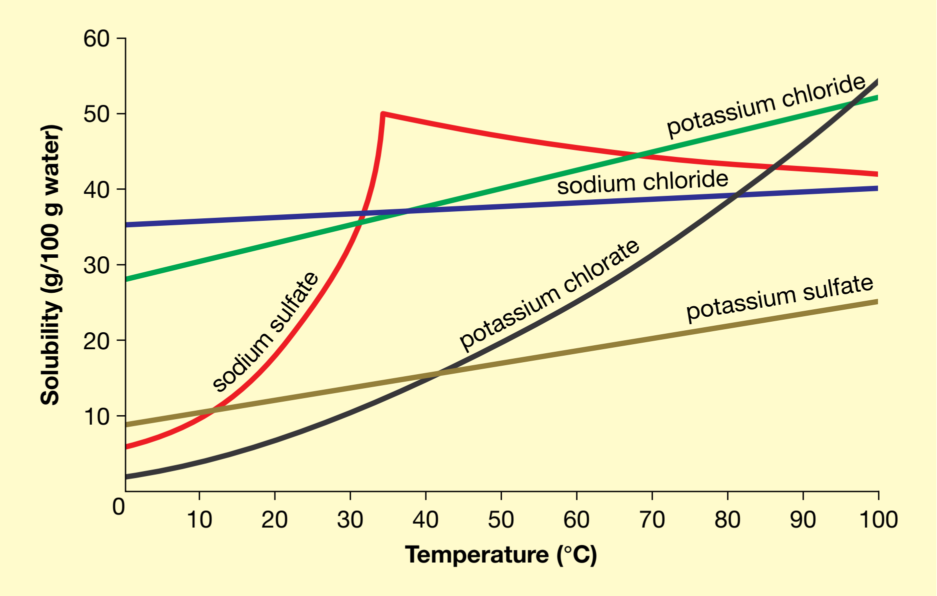
(c) We would not cool below 20 °C, because potassium sulfate crystals would then start to form as well, and the potassium chloride crystals would no longer be pure.

Reduced-pressure crystallisation

Crystallisation is used extensively in industry. However, as heating, especially on a large scale, is a major cost, a modification of the process described on the previous page is often employed. If pressure is reduced, water boils at a lower temperature. It also evaporates faster under such conditions. By reducing the pressure, the water may therefore be 'sucked' out of solution, making the remaining solution more and more concentrated. Once it becomes saturated, removing further water from the solution results in the formation of crystals. Using reduced pressure to achieve crystallisation, rather than heating and then cooling, is more cost effective in some situations.

Revision questions

1. **Use the following graph to determine the amount of potassium chlorate that dissolves in 100 g of water at 35 °C.**



2. **What mass of potassium chloride dissolves in 50 g of water at 80 °C?**
3. **Which salt is most soluble at 40 °C?**
4. **Use the information in table 14.3 to draw solubility curves for silver nitrate and sodium chloride on the same set of axes.**

TABLE 14.3

$AgNO_3$		NaCl	
Temperature (°C)	Solubility (g of salt per 100 g of water)	Temperature (°C)	Solubility (g of salt per 100 g of water)
0	122	0	35.7
10	170	10	35.8
20	222	20	36.0
30	300	30	36.3
40	376	40	36.6
50	455	50	37.0

Gases can also dissolve in water. As a general rule, the solubility of gases decreases as temperature increases.

Solubility of liquids and gases

As is the case with solids, the solubility of liquids and gases depends on the polarity of the solute. Liquids that dissolve in water (such as ethanol) are said to be miscible. Liquids that do not dissolve (such as oil and water) form layers if left to settle. Such liquids are termed 'immiscible'.

With gases, the situation is the same. It should also be noted that gases usually behave opposite to solids as far as temperature is concerned. That is to say, as the temperature increases, the solubility of a gas decreases.

Environmentally, this has the potential to become a significant problem for the Earth. As atmospheric temperatures increase due to higher levels of greenhouse gases (such as carbon dioxide), the upper layers of the oceans also increase in temperature. This leads to reduced amounts of gases such as oxygen and carbon dioxide being dissolved and significant amounts of these gases being returned to the atmosphere. This would further increase the level of carbon dioxide in the atmosphere and would seriously hamper respiration and photosynthesis for marine organisms due to the lower amounts of dissolved oxygen and carbon dioxide that remain.

The concentration of substances

The concentration of a solution is a measure of how much solute has been dissolved in a certain volume of solvent. It is often expressed as a ratio of the solvent amount unit to the volume unit. A number of different units therefore exist.

Concentration is usually defined in terms of the amount of solute per volume of solvent. The greater the quantity of solute added to the solvent, the more concentrated the solution becomes. (For example, the more sugar you dissolve in your cup of coffee, the higher the concentration of sugar in the coffee solution.) When there is only a very small amount of solute present, the solution may be called **dilute** (i.e. its concentration is relatively low). Often, we use the terms 'concentrated' and 'dilute' comparatively. A concentrated solution of hydrochloric acid, for example, contains 37 g of HCl in 100 g of solution. If the same volume of the solution contained only 10 g of HCl, it would be called a dilute solution of hydrochloric acid. The unit of concentration depends on the units for the quantity of solution and the quantity of solute.

A concentrated solution may be diluted by the addition of more solvent.

Each flask contains half the amount of solute of the flask on its left. The lightest coloured flask has $\frac{1}{16}$ the concentration of the darkest flask.

Ways of expressing concentration

There are many ways to express concentration units. A common form is the ratio of the units of solute to the units of volume (e.g. $g\ L^{-1}$, $mg\ L^{-1}$).

$$\begin{array}{c}\text{Concentration } (c) \\ \text{in grams per litre } (g\ L^{-1})\end{array} = \frac{\text{mass of solute in grams}}{\text{volume } (V) \text{ of solution in litres}}$$

Mass conversions

$1\ g = 10^6\ \mu g$

$1 \mu g = 10^{-6}\ g$

$1\ mg = 10^{-3}\ g$

When the concentration unit $mg\ L^{-1}$ is required, the mass in grams is converted to mass in milligrams by multiplying by 1000. For smaller solute amounts, $mg\ L^{-1}$ is usually more convenient to use. (Some useful conversions are listed in the margin.)

Volume conversions

$1\ m^3 = 1000\ L = 1\,000\,000\ mL$

$1\ m^3 = 10^3\ L = 10^6\ mL$

$1\ mL = 10^{-3}\ L$

Sample problem 14.4

A fence post preservative solution is prepared by dissolving 4.00 g of zinc chloride in enough water to make 2250 mL of solution. Find the concentration (c) of this solution in $g\ L^{-1}$ and then in $mg\ L^{-1}$.

Solution:

STEP 1

Convert 2250 mL to a volume in litres. Since 1000 mL = 1 L:

$$V = \frac{2250}{1000} = 2.250\ L$$

STEP 2

Use the equation for concentration in $g\ L^{-1}$.

$$\begin{aligned} c &= \frac{\text{mass of solute in grams}}{\text{volume of solution in litres}} \\ &= \frac{4.00}{2.250} \\ &= 1.78\ g\ L^{-1} \end{aligned}$$

STEP 3

To convert to $mg\ L^{-1}$, multiply 1.78 by 1000.

$$c = 1780\ mg\ L^{-1}$$

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Unit 2
AOS 2
Topic 1
Concept 6

Concentration units
Summary screen and practice questions

studyon

Unit 2
AOS 2
Topic 1
Concept 7

Calculations involving concentrations
Summary screen and practice questions

Sample problem 14.5

Find the mass of sodium bromide required to prepare 50 mL of a 0.40 $g\ L^{-1}$ solution.

Solution:

STEP 1

Convert mL to L.

$$V = 50\ mL = 0.050\ L$$

STEP 2

Manipulating the concentration equation, we get:

$$\text{mass of solute (g)} = \text{concentration } (g\ L^{-1}) \times \text{volume of solution (L)}$$

STEP 3

Calculate the mass of solute.

$$\begin{aligned} \text{mass of sodium bromide} &= 0.40 \times 0.050 \\ &= 0.020\ g \end{aligned}$$

Sample problem 14.6

What volume of solution is required to dissolve 125 mg of lithium chloride to give a concentration of 0.0500 g L^{-1}?

Solution: **STEP 1**

Convert mg to g.

$$\text{Mass of solute} = 125\text{ mg} = 0.125\text{ g}$$

STEP 2

Use the equation for volume of solution (L).

$$\begin{aligned} V &= \frac{\text{mass of solute (g)}}{\text{concentration (g L}^{-1})} \\ &= \frac{0.125}{0.0500} \\ &= 2.50\text{ L} \end{aligned}$$

Revision questions

5. **Calculate the concentration in g L^{-1} and mg L^{-1} for each of the following solutions.**
   (a) **0.425 g barium nitrate in 25.0 mL of solution**
   (b) **26.4 mg sodium bicarbonate in 55.0 mL of solution**
6. **How many grams of sodium chloride would you need to prepare 475 mL of a solution with a concentration of 0.200 g L^{-1}?**
7. **If you dissolve 200 mg of potassium chloride in sufficient solution to give a concentration of 50.0 g L^{-1}, what volume of solution is required?**

Common units of concentration include:
- grams per litre (g L^{-1})
- milligrams per litre (mg L^{-1})
- parts per million (ppm)
- micrograms per gram (μg g^{-1}).
- micrograms per litre (μg L^{-1})
- parts per billion (ppb).

A number of percentage measures are also used — %(m/m), %(m/v), %(v/v).

Micrograms per gram (ppm)

When very small quantities of solute are dissolved to form a solution, the concentration can be measured in **parts per million (ppm)**. This can be written as mg L^{-1} or μg g^{-1}. The concentration of chemicals in the environment and trace elements in the soil are often expressed in ppm. This is a method of comparing the amount of solute, measured in millionths of a gram (micrograms or μg), to the mass of solution, measured in grams. So this is really a mass per mass ratio.

$$\text{parts per million (ppm)} = \frac{\text{mass of solute in micrograms}}{\text{mass of solution in grams}}$$

Our drinking water is treated with chlorine to kill bacteria. Too much chlorine is dangerous, and too little does not kill the bacteria. A safe amount is about 1 part per million (ppm), or 1 microgram of chlorine in each gram of solution. (A microgram, μg, is 10^{-6} g, or 0.000 001 g.)

Parts per billion (ppb)

For even smaller concentrations, parts per billion is an appropriate unit to use. This is the same as μg kg^{-1} or μg L^{-1}.

A number of organic water pollutants, for example, may have their allowable levels measured in this unit. Dieldrin, a chemical previously used against soil pests and termites, has been banned for many years but may still persist in the environment and therefore pollute water. Australian guidelines for drinking water state that the allowable level for dieldrin is 0.3 ppb.

Sample problem 14.7

(a) If 10 g of chlorine gas is dissolved in every 2 500 000 L of water, express the concentration of the chlorine water in **(i)** g L^{-1} and **(ii)** ppm (μg g^{-1}).

(b) Would this water be fit to drink?

Solution: (a) (i)

STEP 1

Convert the volume unit to litres.

$$V = 2.5 \times 10^6 \text{ L}$$

STEP 2

Use the equation for concentration in g L^{-1}.

$$\begin{aligned} c &= \frac{\text{mass of solute in grams}}{\text{volume of solution in litres}} \\ &= \frac{10 \text{ g}}{2.5 \times 10^6} \\ &= 4 \times 10^{-6} \text{ g L}^{-1} \end{aligned}$$

(ii)

STEP 1

Use the equation for concentration in ppm to find the mass of the solute and the mass of the solution.

$$c = \frac{\text{mass of solute in micrograms}}{\text{mass of solution in grams}}$$

Mass of solute:

Convert g to μg; mass of solute = $10 \times 10^6 = 1 \times 10^7$ μg

Mass of solution:

Density of water = 1 g mL^{-1} (which means that 1 mL has a mass of 1 g).

So we have $2.5 \times 10^3 \times 10^6$ mL, which has a mass of 2.5×10^9 g.

STEP 2

Calculate the concentration in ppm.

$$\begin{aligned} c &= \frac{1 \times 10^7}{2.5 \times 10^9} \\ &= 4 \times 10^{-3} \text{ ppm} \end{aligned}$$

(b) The water is unfit to drink as there is insufficient chlorine to kill the bacteria. (*Note:* When the solvent is water, mg L^{-1} also equals ppm.)

Sample problem 14.8

If 5.00 m^3 of river water was found to contain 1.250 mg of an organic pollutant, calculate its concentration in ppb.

Solution: **STEP 1**

Convert mg to μg.

mass of solute = 1.250 mg = 1250 μg

STEP 2

Use the equation for concentration in μg L^{-1} (ppb).

$$
\begin{aligned}
c &= \frac{\text{mass of solute in micrograms}}{\text{volume of solution in litres}} \\
&= \frac{1250}{5.00} \\
&= 250\ \text{ppb}
\end{aligned}
$$

Revision questions

8. **Calculate the concentration in μg g^{-1} of a solution formed by dissolving 7.50 mg of aluminium sulfate in each 2.00 L of water processed by a water treatment plant (density of water = 1 g mL^{-1}).**

9. **A sample of pond water contained 88.00 μg g^{-1} of dissolved oxygen gas. Calculate the amount in grams of dissolved oxygen in 4.00 L of pond water.**

10. **A desirable level of dissolved oxygen in river water is 8 ppm. A sample of river water is found to have less than this, with 5.60 ppm oxygen. If 1.40 μg of dissolved oxygen was actually present in the sample, calculate the mass of water tested.**

Percentage by mass, %(m/m)

Solutions of concentrated acids and some household cleaners often carry a label with **%(m/m)** (for example, sulfuric acid, 98%(m/m). The unit %(m/m) expresses the concentration of the solute as a percentage of the mass (or weight) of solution. So, 98% sulfuric acid means there is 98 grams of sulfuric acid solute per 100 g of solution. We can calculate this mass percentage using the equation below.

$$\text{Concentration in }\%(\text{m/m}) = \frac{\text{mass of solute in g}}{\text{mass of solution in g}} \times 100\%$$

Sample problem 14.9

Find the concentration in %(m/m) obtained when 18.5 g of hydrochloric acid is dissolved in 50.0 g of solution.

Solution:

$$
\begin{aligned}
\%(\text{m/m}) &= \frac{\text{mass of solute in g}}{\text{mass of solution in g}} \times 100\% \\
&= \frac{18.5}{50.0} \times 100\% \\
&= 37.0\%
\end{aligned}
$$

Revision questions

11. **Calculate the concentration in %(m/m) for the following solutions.**
  (a) **38.4 g of NaCl in 150.0 g of solution**
  (b) **0.25 g benzalkonium chloride (active ingredient in an antibacterial kitchen cleaner) in 50.0 g of solution**
  (c) **144 mg of sugar in 200 g of solution**

12. **Calculate the mass of sodium hydrogen carbonate in 250 g of a 0.500%(m/m) solution.**

13. **Calculate the mass of solution required to prepare a 6.00%(m/m) tile and household cleaner solution containing 15.5 g of sodium phosphate.**

14. **Calculate the concentration in %(m/m) when 24.5 g of copper sulfate is dissolved in 50 g of solution.**

Percentage mass/volume, %(m/v)

An oven cleaner may have on its label 'active ingredient 1%(m/v) sodium hydroxide'. The expression **%(m/v)** expresses the concentration as mass of solute in grams per 100 mL of solution. We can use the following formula to perform calculations:

$$\text{Concentration }\%(\text{m/v}) = \frac{\text{mass of solute in g}}{\text{volume of solution in mL}} \times 100\%$$

Sample problem 14.10

Find the %(m/v) of a physiological saline solution consisting of 4.30 g of sodium chloride dissolved in 0.500 L of aqueous solution.

Solution: $V = 0.500\text{ L} = 500\text{ mL}$

$$\%(\text{m/v}) = \frac{\text{mass of solute in grams}}{\text{volume of solution in mL}} \times 100\% = \frac{4.30}{500} \times 100\%$$

$$= 0.860$$

Revision questions

15. **Calculate the %(m/v) of the following solutions.**
  - **(a)** **200 mg of potassium chloride in 50.0 mL of solution**
  - **(b)** **50 kg of sodium bromide in 250 L of solution**
  - **(c)** **0.025 g NaCl in 500 mL of solution**

16. **Calculate the mass of sodium hypochlorite required to prepare 375 mL of a 1.5%(m/v) solution of bathroom cleaner.**

17. **Calculate the volume of solution required to dissolve 100 mg of plant food to make a concentration of 0.50%(m/v).**

18. **Calculate the volume of solution required to prepare a 10%(m/v) solution of sodium hydroxide from a 10 mg sample of sodium hydroxide.**

Percentage by volume, %(v/v)

The percentage by volume, **%(v/v)**, expresses the concentration unit for liquid solutes as volume of liquid solute in mL per 100 mL of solution. For example, a one-litre bottle of liqueur is labelled 40% alc/vol. This means there is 40 mL of ethanol (alcohol) in each 100 mL of liqueur. Calculations can be performed using the equation:

$$\text{Concentration }\%(\text{v/v}) = \frac{\text{volume of solute in mL}}{\text{volume of solution in mL}} \times 100$$

Sample problem 14.11

If a standard glass (0.200 L) of a particular brand of beer contains 9.80 mL of ethanol, calculate the concentration of the beer in %(v/v).

Solution: Volume of solution = 200 mL

$$\%(\text{v/v}) = \frac{\text{volume of solute in mL}}{\text{volume of solution in mL}} \times 100\%$$

$$= \frac{9.80}{200} \times 100\%$$

$$= 4.90$$

Revision questions

19. **Glycerol is a syrupy, sweet tasting liquid used in cosmetics and candy. If 10.0 mL of glycerol is dissolved in 0.455 L of aqueous solution, calculate the concentration as %(v/v).**

20. **Ethanoic acid is dissolved in aqueous solution to make a 5.00%(v/v) solution of vinegar.**
  - **(a) What volume of ethanoic acid is in 0.750 L of vinegar?**
  - **(b) What volume of vinegar contains 25.0 mL of ethanoic acid?**

21. **A particular brand of champagne has an alcohol concentration of 11.5%(v/v). What is the alcohol content of a 750 mL bottle?**

The concentration of a solution is the number of moles that it contains in each litre. The concentration of a solution is calculated using the formula $c = \frac{n}{V}$. This can be transposed to find the number of moles, $n = cV$.

A special unit of concentration

Concentration can be expressed in a number of ways. Chemists most commonly use **molar concentration** or **molarity**, which is the amount of solute, in moles, present in each litre of solution. The symbol for concentration is *c*. The unit of measurement for *molar* concentration is moles per litre, which can be expressed as M or sometimes mol L^{-1}. Therefore a 1.5 M solution has a concentration of 1.5 mol L^{-1}.

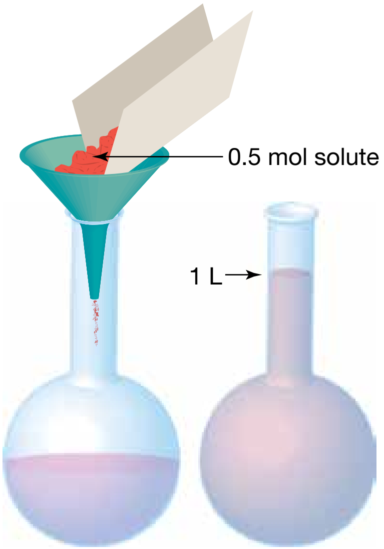
If 1 mol of solute is dissolved in a total volume of 1 L of water, the concentration of the solution is 1 M. If 0.5 mol of solute is dissolved in a total volume of 1 L of water, the concentration of the solution is 0.5 M. If, however, 0.5 mol of solute is dissolved in a total volume of 0.5 L of water, the concentration of the solution is 1 M.

Molar concentration is defined as:

$$\text{concentration } (c) = \frac{\text{quantity of solute (moles)}}{\text{volume of solution (L)}}$$

$$c = \frac{\text{moles}}{\text{volume}}$$

$$= \frac{n}{V} \text{ (in M)}$$



The preparation of a solution of known concentration. To make a 0.5 M solution, add 0.5 mole of solute to a 1.0 litre volumetric flask half filled with water. Swirl the flask carefully to make the solute dissolve. Finally, fill the flask with water exactly to the 1.0 L mark.

The formula for concentration can be manipulated to calculate the amount, and hence the mass, of the solute.

$n = c \times V$ (*Note:* Volume must be in litres.)

$m = n \times$ molar mass

In order to prepare a particular volume of solution of known concentration, the following five steps should be followed:

1. Calculate the number of moles of solute that are needed to obtain the correct concentration of solution for the volume of solvent to be used, according to the formula $n = cV$.
2. Calculate the mass of the solute needed, using the formula $m = nM$.
3. Partially fill a volumetric flask with water, and add the correct mass of solute.
4. Dissolve the solute.
5. Add water to the required volume.

Sample problem 14.12

Calculate the number of moles of sodium chloride needed to prepare 500 mL of a 0.0800 M salt solution. What mass of sodium chloride would be weighed out?

Solution: **STEP 1**

List the known information.

volume $(V) = 0.500$ L

concentration $(c) = 0.0800$ M

STEP 2

Calculate the number of moles (n) of NaCl needed.

$$\begin{aligned} n &= c \times V \\ &= 0.0800 \times 0.500 \\ &= 0.0400 \text{ mol} \end{aligned}$$

STEP 3

Calculate the mass represented by the number of moles.

$$\begin{aligned} m &= n \times \text{molar mass} \\ &= 0.0400 \times 58.5 \\ &= 2.34 \text{ g} \end{aligned}$$

Sodium bicarbonate tablet dissolving in water, releasing carbon dioxide bubbles. Many chemical reactions occur in water.

Revision question

22. Calculate the concentration (molarity) of copper(II) sulfate, $CuSO_4$, in:
- **(a) 1 L of solution containing 200 g $CuSO_4{\cdot}5H_2O$**
- **(b) 2.0 L of solution containing 250 g $CuSO_4{\cdot}5H_4O$**
- **(c) 700 mL of solution containing 750 g $CuSO_4{\cdot}5H_2O$**
- **(d) 125 mL of solution containing 50 g $CuSO_4{\cdot}5H_2O$.**

The concentration of water

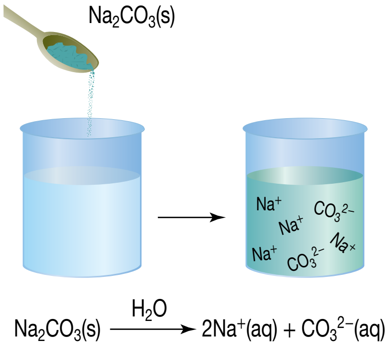
All pure liquids have a concentration, since concentration can be defined as the number of particles in a given volume. Given that it is possible to mix two liquids together and get a smaller volume, it follows that empty space exists between the liquid molecules. Although, to the naked eye, water appears to have no gaps, at the molecular level there is always empty space (vacuum) between the water molecules. So liquid water has a concentration.

The concentration of water can be expressed as grams per litre (density) or moles per litre.

Given that the density of water is 1 g mL^{-1}, it follows that 1000 mL of water has a mass of 1000 g. This can be converted to moles using the formula $n = \frac{m}{M}$. The molar mass of water (M) is 18 g mol^{-1}. Therefore:

$$\begin{aligned} n(H_2O) &= \frac{1000}{18} \\ &= 55.5 \text{ mol} \end{aligned}$$

One litre of pure water therefore contains 55.5 mol of water molecules.



Dissolving sodium carbonate brings about dissociation of the sodium and carbonate ions.

Calculating the concentration of ions in solution

Many ionic substances are soluble. The process of dissolving involves the ionic lattice breaking up and the anions and cations dissociating from each other.

The concentrations of the resulting ions can be calculated. These concentrations are called **ionic concentrations** and they are designated by square brackets. For example:

> $[Na^+]$ = 0.5 M means that the ionic concentration of Na^+ ions in solution is 0.5 M.

Sample problem 14.13

Sodium carbonate, Na_2CO_3, is often used as an ingredient in washing powders because it softens hard water. It dissolves readily in water. If 10.6 g of sodium carbonate is dissolved in 500 mL of water:

(a) what is the concentration of the solution

(b) what is the concentration of each ion in the solution?

Ionic concentrations are calculated by finding the concentration of the solution, and then multiplying the mole ratio of ions in the dissociation equation.

Solution: (a)

STEP 1

Calculate the number of moles.

$$n(Na_2CO_3) = \frac{m}{M} = \frac{10.6}{106}$$
$$= 0.100 \text{ mol}$$

STEP 2

Calculate the concentration.

$$c(Na_2CO_3) = \frac{n}{V} = \frac{0.100}{0.500}$$
$$= 0.200 \text{ M}$$

(b)

STEP 1

Write the dissociation equation:

$$Na_2CO_3(s) \xrightarrow{H_2O} 2Na^+(aq) + CO_3{}^{2-}(aq)$$

STEP 2

Write the mole ratios.

1 mol Na_2CO_3 produces 2 mol Na^+ and 1 mol $CO_3{}^{2-}$

STEP 3

Use mole ratios to determine ion concentration as the volume is constant:

$$c(Na^+) = 2 \times c(Na_2CO_3) = 2 \times 0.200 = 0.400 \text{ M}$$
$$c(CO_3{}^{2-}) = c(Na_2CO_3) = 0.200 \times 1 = 0.200 \text{ M}$$

Sample problem 14.14

In an $Al_2(SO_4)_3$ solution, the concentration of $SO_4{}^{2-}$ is 0.050 M. What is the concentration of solute?

Solution: **STEP 1**

Write the dissociation equation.

$$Al_2(SO_4)_3(s) \xrightarrow{H_2O} 2Al^{3+}(aq) + 3SO_4{}^{2-}(aq)$$

STEP 2

Compare the mole ratios. Since 1 mol $Al_2(SO_4)_3$ produces 3 mols of $SO_4{}^{2-}$:

$$\frac{n(Al_2(SO_4)_3)}{n(SO_4^{2-})} = \frac{1}{3}$$
$$n(Al_2(SO_4)_3) = \frac{1}{3} \times n(SO_4^{2-})$$

STEP 3

Since the volume is constant:

$$c(\mathrm{Al_2(SO_4)_3}) = \frac{1}{3} \times c(\mathrm{SO_4^{2-}}) = \frac{1}{3} \times 0.050\ \mathrm{M}$$
$$= 0.0166\ \mathrm{M}$$
$$= 0.017\ \mathrm{M}$$

Revision questions

23. **Caustic soda, NaOH, is an ingredient in household drain cleaners. If a 250 mL bottle of cleaner contains 20 g of caustic soda, calculate:**
  - **(a)** **the concentration of the solution**
  - **(b)** **the concentration of each ion in the solution.**

24. **Calculate the concentration of an $Mg(NO_3)_2$ solution in which the concentration of nitrate ions is 0.030 M.**

25. **Calculate the concentration of a solution of ammonium phosphate, $(NH_4)_3PO_4$, in which the concentration of ammonium ions is 1.25 M.**

Environmental issues

The excellent solvent properties of water mean that it is difficult to obtain pure water and that it is also easily polluted.

In many ways it is unfortunate that water is such a good solvent. Almost anything we put in the sea is eventually dissolved or broken down into constituent atoms, molecules or ions. For hundreds of years it has been a convenient and highly efficient water treatment plant. Salts have dissolved in rivers as water runs over rocks and soil. Effluent, too, has been poured into rivers, which eventually flow out to the sea.

Recently, we have been hearing of beaches becoming dangerous: sea water too dangerous to swim in; effluent being washed up on the sand. Why? What has gone wrong with the sea? Why won't it do its job anymore?

The volume and type of waste being poured into the sea is astounding. Oil is washed out of the holds of tankers. Factories discharge waste directly into the sea. Many cities pour their sewage directly into the sea without any pretreatment.

The more sophisticated our technology becomes, the more complex are the pollutants that we discharge. This means that it takes longer for the sea to degrade the waste. Some waste is **biodegradable**. This means that it is broken down by natural processes. However, an increasingly larger proportion of waste is **non-biodegradable**.

Pollution of the water from a sewer outfall can cause excessive growth in aquatic plants. This can upset the natural balance in the ecosystem.

studyon

Unit 2
AOS 2
Topic 1
Concept 3

Contamination of water
Summary screen and practice questions

Micro-organisms

Fresh water contains many microscopic organisms such as phytoplankton, bacteria, algae and protozoa. These organisms may be harmful or beneficial to other forms of life. For example, the biological decay of plant and animal matter often produces nitrogen-rich urea and ammonia, both of which are very soluble. Bacteria may then react with these compounds, ultimately converting the nitrogen in them into free atmospheric nitrogen. This is a critical step in the nitrogen cycle, so it is important to life. On the other hand, disease-causing (pathogenic) bacteria may be present, especially if run-off has been in contact with animal waste, or poorly treated sewage has entered the water. By testing water for the presence of various bacteria, scientists can determine whether water is contaminated. The presence of *Escherichia coli* (*E. coli*) indicates that the water contains pathogenic bacteria.

Eutrophication

When conditions are right, phytoplankton (and especially blue-green algae or cyanobacteria) may reproduce explosively to seriously affect water quality. These conditions usually include still water, sunlight and an excess of nutrients required for growth. Two important such nutrients are nitrates and phosphates, which may be present in excess as a result of human activity. Domestic waste water, agricultural practices and sewage are the main sources of these excess nutrients. This out-of-control growth is often noticed as an 'algal bloom' and, besides being unsightly, can have serious consequences. When these organisms die, their subsequent decomposition seriously depletes the level of dissolved oxygen in the water. As a result of this, animals, especially fish, and even plants may die due to a lack of oxygen required for respiration. Additionally, all this decay may produce biotoxins, which can be a serious health hazard to any organism that consumes this water.

This condition is called **eutrophication**.

Escherichia coli (commonly known as *E. coli*) is a micro-organism commonly found in the bowel. If it is present in waterways, however, it can indicate a serious health hazard.

Heavy metals

The dangers of some heavy metals were first exposed after a tragedy in Minemata Bay in Japan during the 1950s. Before the cause had been identified, at least 60 people had been seriously affected. The victims had been eating fish. These fish had accumulated a large quantity of methyl mercury in their bodies. The source of the mercury was a chemical company that had been using mercury as a catalyst in the production of organic chemicals and was discharging large quantities of mercury waste into the sea. Bacteria converted the waste to toxic methyl mercury.

Mercury, lead and cadmium are just some of the heavy metals that pose a major disposal problem. Inorganic and organic lead compounds are extremely poisonous and are difficult to dispose of.

Shellfish accumulate heavy metals and should not be eaten if they are taken from near a sewage outlet or a polluted river.

TABLE 14.4 Maximum permitted levels of some contaminants in drinking water

Contaminant	Maximum drinking water level (ppm)
mercury	0.002
lindane (an insecticide)	0.004
lead	0.05
2,4-D (a herbicide)	0.1
trihaloalkanes (solvents)	0.1
sodium	160.0

Water for drinking

Water quality is becoming a critical issue for many people around the world as the population increases and traditional fresh water supplies become harder to find and develop. Humans require high-quality supplies of fresh water for agriculture, washing, industry and, of course, for drinking. Experts agree that we are rapidly running out of such supplies. The fact that there is so much salt water present on the planet has led to much research and development of desalination — methods by which salt and other substances can be removed from sea water, leaving fresh water. Desalination technologies are currently used extensively in the Middle East, and are growing fast in areas such as North America, Australia and northern Africa. Small-scale desalination has been used for some time on board ships and submarines.

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Improving water quality

The production and treatment of water for drinking is subject to standards to ensure public health. These standards cover allowed levels of biological, chemical and radiological contaminants, as well as other issues such as pH and turbidity. Various acts describe what the standards are and how they are to be monitored. These acts include:

- *Safe Drinking Water Act, 2003* and *Safe Drinking Water Regulations, 2005*, which were set and administered by the Victorian Department of Health
- Australian Drinking Water Guidelines 2011, which were developed by the National Health and Medical Research Council.

Methods of desalination

Desalination is the process of removing salt and other unwanted substances from sea water and waste water. It is anticipated that this will become an important method for producing water that is fit for humans to use in the future. Two important methods for desalination are distillation and reverse osmosis.

At the moment there are two main methods of desalination used around the world. These are distillation and reverse osmosis. Both of these methods use a large amount of energy. This results in the fresh water produced being higher in cost than fresh water from traditional sources. However, with further research and attention to energy-saving measures, it is anticipated that these energy requirements will be reduced, resulting in desalination becoming an economically viable alternative.

Another problem is the disposal of the waste products. Removing salt and other substances from water means that a waste product is created in the form of residual water, in which is dissolved everything that was removed to make the fresh water. This waste has to be disposed of carefully. Although the residual water can be pumped back into the ocean, care must be taken to avoid causing localised areas of higher salinity and thus affecting the local marine flora and fauna. In inland areas, care must be taken to prevent such disposal contaminating existing fresh water supplies such as rivers, lakes and aquifers.

Piles of salt are mined from deep beneath the surface of the Earth. This salt was deposited here from ancient oceans that have since receded.

In addition to the two methods mentioned above, there are a number of new developments that use advances in membrane technology and nanotechnology. These are much more energy efficient than the traditional large-scale methods, with some of them also showing promise for small-scale, ‘point-of-use’ applications.

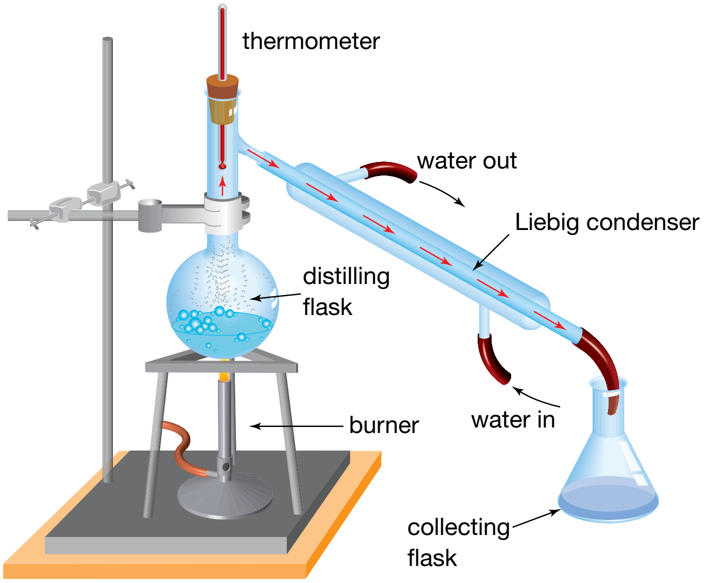
Principles of distillation

You may recall from Unit 1 that distillation is a process that can separate the components of crude oil. This was *fractional distillation*. In more general terms, Distillation is a process that can separate substances with different boiling temperatures. It is often used to recover a solvent from a solution or to separate mixtures of liquids.

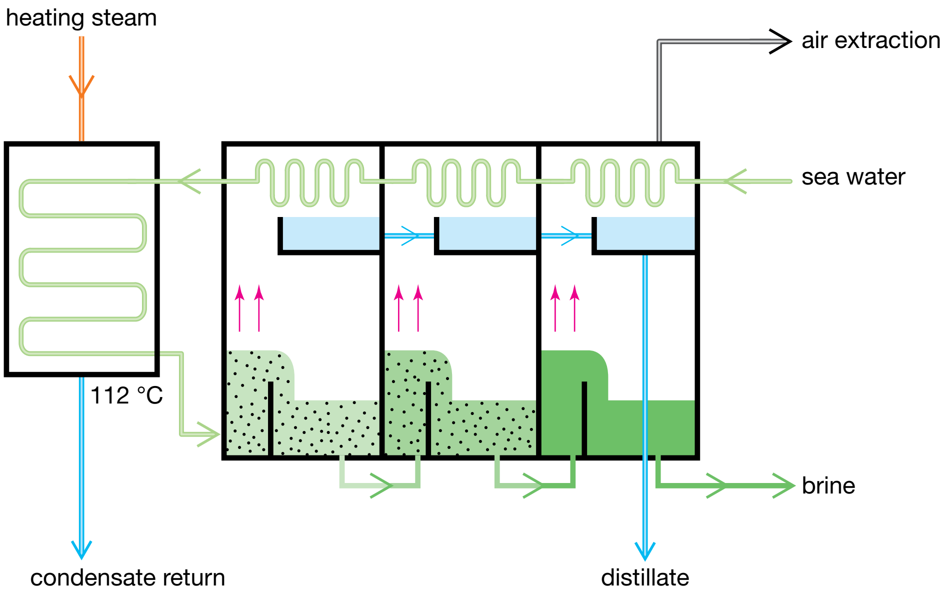
distillation can be used to separate a mixture of substances that have different boiling temperatures. If the boiling temperatures are similar, fractional distillation is required. However, if there is a large difference in boiling points, as is the case with ionic solutions, *simple distillation* will suffice.

For simple distillation in the laboratory, the impure mixture is placed in the distillation flask and heated. As the temperature rises and the solvent begins to boil, its vapours rise and are led into the condenser. This then cools the vapours and results in condensation back into droplets of the pure solvent. As these droplets coalesce, they run down the condenser and into the collecting flask. The impurities are left behind in the distilling flask.

A typical laboratory set-up for simple distillation is shown below.



Distillation is used on a large scale to desalinate water. To save energy, it is nearly always performed in a series of stages, in a process called *multi-stage flash distillation*, with successive stages being at lower and lower pressures. This allows the water to be boiled at lower and lower temperatures and results in a significant saving in heating costs. A typical arrangement for such a plant is shown below.

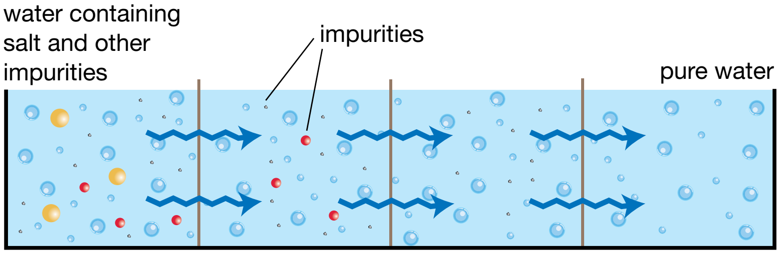


Multi-stage flash distillation saves energy by lowering the water pressure. This causes the water's boiling point to be reduced, meaning that less heating is required.

Notice that there are a number of stages, each one cooler and at a lower pressure than the one before it. The steam in each stage is cooled by pipes containing the cold, incoming sea water to form the distillate (pure water). The incoming sea water is, in turn, heated by the steam through which it passes.

Reverse osmosis

Osmosis is the diffusion of a solvent through a membrane towards an area of higher concentration of dissolved substances. This is important in many biological situations as a means of moving water, and is a spontaneous process that does not require energy. To purify water, however, we want the water to move away from a region of dissolved substances so that it can subsequently be collected as pure water. This process is called reverse osmosis, and it cannot occur without energy.



Reverse osmosis

When used to desalinate water, this process forces pressurised water through a series of membranes. These membranes contain tiny holes (or pores) that allow the smaller water molecules to pass through but prevent larger particles of contaminants. To prevent clogging of these membranes, they are often used in series, with each successive membrane becoming more selective. The diagram on the left shows a summary of this process.

New developments in desalination and water purification

Recent developments in *nanotechnology* have suggested many new and exciting possibilities for increasing the supply of drinkable (potable) water and lowering its cost.

Nanofiltration membranes can be made from materials such as carbon nanotubes and work in a similar way to reverse osmosis. However, they operate at a lower pressure, resulting in energy and cost savings. They also appear easier to clean after becoming fouled due to blockage of their pores. Nanomaterials that could be used in desalination and water purification include carbon nanotubes, graphene, nanoparticles of metals (especially silver and gold) and nanoparticles made from various metal oxides.

Another technique that is under development is the use of nanocomposite materials to enable distillation of sea water in the presence of sunlight. These materials convert sunlight into heat energy and, when added in small amounts to sea water, can lead to a significant increase in the temperature of the water when compared with solar heating alone. This results in higher rates of evaporation, where the water vapour produced can subsequently be condensed to produce pure water. Nanomaterials made from graphene and nanosilver, and from graphene and nanogold, appear to be very suitable for this application. An additional advantage is that nanosilver has been shown to have antibacterial properties.

Nanomaterials can also be 'functionalised' by adding various chemical groups to promote adsorption and sterilisation of substances in impure water. Their high surface area means that they are highly efficient at adsorption or when acting as catalysts in reactions to remove impurities. This means that small-scale devices that operate at the 'point of use' may soon become available. Such devices could be used following natural disasters, for example, when large-scale water treatment and purification plants may be rendered inactive. They could also help produce potable water from impure supplies in many developing countries. An example of such a device currently under development is a straw-like device called a 'waterstick'.

Chapter review

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Unit 2 **Measuring solubility and concentration**
AOS 2
Topic 1 Summary screens and practice questions

Summary

- Water is such a good solvent that it is rarely pure. Fresh water constitutes only 2.5% of the Earth's total water but may still contain a number of dissolved substances.
- A solute is a substance that is dissolved in a solvent to make a solution. A dilute solution contains a small amount of solute. Volumes of water containing dissolved substances are called aqueous solutions.
- Solubility is the extent to which a solute can dissolve in a solvent and this depends on the temperature.
  - A saturated solution contains the maximum amount of solute for the volume of solution at a given temperature.
  - An unsaturated solution contains less than the maximum amount of solute.
  - A supersaturated solution is prepared by slowly cooling a saturated solution so that it contains more solute than normal at that temperature.
- A solubility curve is a graph of temperature versus solubility. It can be used to predict how much solute dissolves at a given temperature.
- Increasing temperatures in the upper levels of the oceans reduces the amount of dissolved carbon dioxide and oxygen, which will have serious implications for life in these regions.
- Concentration is defined as the amount of solute per volume of solution. When more solute is added, a concentrated solution results. When more solvent is added, a dilute solution results.
- Concentration units can be expressed as g L^{-1}, mg L^{-1}, g μL^{-1}, parts per million (ppm), parts per billion (ppb), %(m/m), %(m/v) and %(v/v). Molar concentration (M) is an especially useful unit in chemistry.
- Many pollutants dissolve in water as it is such a good solvent. Some waste is biodegradable and can be broken down by natural processes, but much is non-biodegradable. This poses environmental problems where species may be endangered due to water pollution.
- Eutrophication is caused by excessive levels of nutrients in water, which leads to explosive growth of micro-organisms such as cyanobacteria. When these die, their subsequent decomposition leads to depleted levels of dissolved oxygen and the death of plants and fish.
- Desalination is the process of removing salt and other unwanted substances from water. Two common methods for doing this are distillation and reverse osmosis. New developments in membrane technology and in nanotechnology will offer further methods for desalination of water in the future.

Multiple choice questions

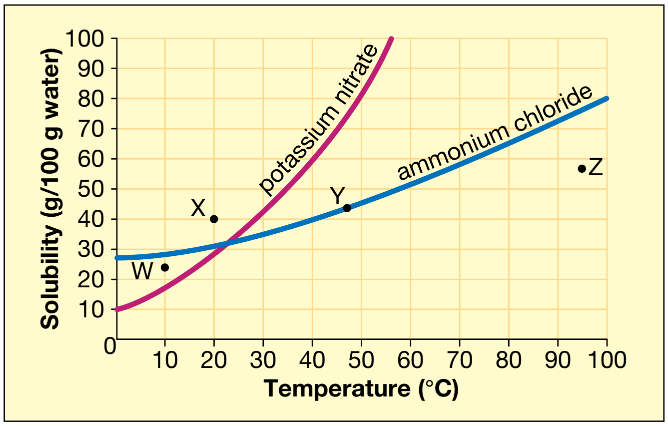
1. The percentage of fresh water on the Earth is closest to:

A 2.5%
B 30%
C 70%
D 97.5%.

2. A saturated solution is best described as a solution in which:

A there is a small amount of solute in a large amount of solvent
B crystals of the solute can stay without growing or dissolving
C the addition of further solvent produces a precipitate
D there is a very large amount of solute in a small amount of solvent.

Questions 3–5 refer to the information shown in the following graph.



3. To completely dissolve 35.0 g of potassium nitrate in 50.0 g of water, the minimum temperature would have to be:

A 25 °C
B 35 °C
C 45 °C
D 55 °C.

4. Various solutions of ammonium chloride are represented by the letters W, X, Y and Z. Which letter represents a supersaturated solution?

A W **B** X **C** Y **D** Z

5. A total of 80 g of ammonium chloride was dissolved in 200 g of boiling water. If the solution were cooled slowly, the first crystals of ammonium chloride would be expected to appear at:

A 100 °C
B 40 °C
C 80 °C
D 10 °C.

6. The concentration of sodium chloride in a sample of sea water is 28.5 g L^{-1}. The mass of salt remaining if 2.00 L of the sea water were evaporated to dryness would be:
A 14.3 g **C** 28.5 g
B 57.0 g **D** 42.8 g.

7. What volume of water is required to dissolve 0.15 g of copper sulfate to make a solution of 250 mg L^{-1}?
A 0.60 L **C** 6.0×10^{-4} L
B 1.7 L **D** 1.6×10^{3} L

8. A recommended safe level of the heavy metal cadmium in drinking water is 0.01 ppm. The mass of cadmium in a litre of drinking water would be (density of water = 1.0 g mL^{-1}):
A 10 mg **C** 0.010 g
B 0.010 μg **D** 0.010 mg.

9. What is the %(m/v) of a 1 L solution of commercial bleach that contains 50 g of sodium hypochlorite?
A 50%(m/v) **C** 5.0%(m/v)
B 2.0%(m/v) **D** 0.050%(m/v)

10. A solution contains a mixture of two salts: potassium nitrate and potassium sulfate. If the concentration of potassium ions is 1.40 M and the concentration of nitrate ions is 0.600 M, what is the concentration of the sulfate ions?
A 0.400 M
B 0.800 M
C 0.900 M
D 1.800 M

11. Warming of ocean waters is likely to cause the level of dissolved:
A oxygen and carbon dioxide to both increase
B oxygen to increase but carbon dioxide to decrease
C oxygen to decrease but carbon dioxide to increase
D oxygen and carbon dioxide to both decrease.

12. Which of the following would *not* be likely to contribute to eutrophication?
A Phosphates
B Sunlight
C Running water
D Nitrates

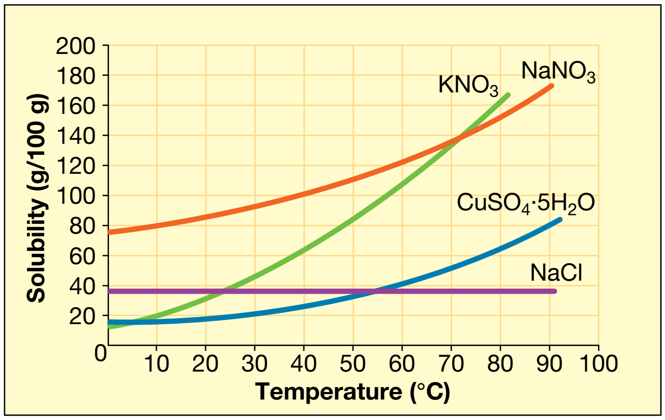
Review questions

1. Eutrophication can cause the stagnation of waterways. Find out what eutrophication is and explain how hydrogen bonding in water contributes to this process.

Solubility

2. An amount of 0.2 g of solute is dissolved in 500 L (500 kg) of aqueous solution. Express this as:
(a) g L^{-1} (b) ppm (μg g^{-1}).

3. Consider the following solubility curves and answer the questions that follow.



(a) At what temperature is the solubility of $CuSO_4.5H_2O$:
(i) 30 g/100 g
(ii) 40 g/100 g
(iii) 20 g/100 g?
(b) At what temperature does 60 g of KNO_3 dissolve in 50 g of water?
(c) What is the solubility of the following solids at 35 °C?
(i) NaCl
(ii) $NaNO_3$
(iii) KNO_3

4. Distinguish between the terms 'unsaturated', 'saturated' and 'supersaturated'.

Concentration

5. Calculate the percentage by mass, %(m/m), of the solute in each of the following.
(a) A 45 g solution containing 5 g of potassium nitrate
(b) A 24.5 g solution containing 250 mg of ammonium nitrate

6. A solution of potassium chloride is 8.5%(m/m). An experiment requires 3.4 g of KCl. How many grams of the solution do you need for this experiment?

7. The label on a bottle says that it contains a 7.5%(m/m) solution of ammonium nitrate. The bottle contains 125 g of this solution. How many grams of ammonium nitrate are in the bottle?

8. A chemist dissolves 3.5 mL of ethanoic acid in water to give a total of 120 mL. What is the concentration in %(v/v)?

9. What volume of ethanol is required to prepare the following alcoholic solutions?
(a) 4.0 L of 10%(v/v)
(b) 350 mL of 40%(v/v)

10. A saline solution contains 2.5%(m/v). How much salt would be present in 2.0 L of this solution?

Molar concentration

11. Calculate the mass of solute in each of the following solutions.
  (a) 300 mL of 1.5 M NaOH
  (b) 250 mL of 2.0 M H_2SO_4
  (c) 17.5 L of 1.5 M Na_2CO_3
  (d) 200 mL of 2.5 M $CuSO_4{\cdot}5H_2O$
  (e) 120 mL of 1.7 M $Mg(NO_3)_2$

12. Calculate the concentration (molarity) of solute in each of the following solutions.
  (a) 250 mL of solution containing 17 g sodium bromide, NaBr
  (b) 500 mL of solution containing 200 g magnesium sulfate, $MgSO_4$
  (c) 1.5 L of solution containing 1500 g magnesium nitrate, $Mg(NO_3)_2$
  (d) 100 mL of solution containing 1000 g sodium phosphate, Na_3PO_4
  (e) 2.00 L of solution containing 500 g potassium carbonate, K_2CO_3
  (f) 40.0 mL of solution containing 12.5 g aluminium chloride, $AlCl_3$

13. Calculate the concentrations of the following solutions.
  (a) 4.0 g of sodium hydroxide in 200 mL of solution
  (b) 12.6 g of sodium carbonate in 350 mL of solution
  (c) 5.35 g of magnesium carbonate in 500 mL of solution

14. Calculate the number of moles of solute needed to prepare:
  (a) 30 mL of 0.10 M $AgNO_3$
  (b) 300 mL of 1.5 M $Mg(NO_3)_2$
  (c) 230 mL of 0.40 M KCl
  (d) 2.5 L of 0.2 M KNO_3.

15. Calculate the mass of solute needed to make each of the following quantities of solution.
  (a) 250 mL of a 1.5 M sodium bromide, NaBr, solution
  (b) 500 mL of a 1.75 M magnesium sulfate, $MgSO_4$, solution
  (c) 1.50 L of a 0.575 M magnesium nitrate, $Mg(NO_3)_2$, solution
  (d) 100 mL of a 0.850 M sodium phosphate, Na_3PO_4, solution
  (e) 2.00 L of a 0.00500 M potassium carbonate, K_2CO_3, solution
  (f) 40.0 mL of a 2.30 M aluminium chloride, $AlCl_3$, solution

16. How many grams of CH_3COONa are obtained when 400.0 mL of a 0.500 M solution of CH_3COONa is evaporated to dryness?

17. What is the concentration of each of the following solutions?
  (a) 58.5 g of H_2SO_4 dissolved in enough water to produce 2.00 L of solution
  (b) 2.7 g of HCl dissolved in enough water to produce 500 mL of solution
  (c) 4.04 g of KNO_3 dissolved in enough water to produce 150.0 mL of solution
  (d) 234 g of sodium chloride dissolved in enough water to produce 6.00 L of solution.

Concentration of ions in solution

18. Most common fertilisers contain nitrogen compounds. Ammonium nitrate, NH_4NO_3, is soluble in water and hence it is quickly taken up by the plant's root system. If a 1.5 L bucket contains 150 g of ammonium nitrate calculate:
  (a) the concentration of the solution
  (b) the concentration of each ion in the solution.

19. Aluminium chloride, $AlCl_3$, is found in antiperspirants. A student wishes to do some tests on antiperspirant, so a jar of it is dissolved into a beaker containing 700 mL of water. The label on the jar states that it contains 200 g of aluminium chloride. Calculate:
  (a) the concentration of the solution
  (b) the concentration of each ion in the solution.

20. Calcium hydroxide, $Ca(OH)_2$, is found in plaster and cement, and is used in the treatment of drinking water. If the concentration of OH^- of a solution of calcium hydroxide is 0.050 M, calculate the concentration of the calcium hydroxide.

21. In an $Fe_2(SO_4)_3$ solution the concentration of SO_4^{2-} is 0.25 M. What is the concentration of solute?

Environmental issues

22. Excess fertiliser run-off can be a major contributing factor to eutrophication. Research and answer the following questions relating to this problem.
  (a) Describe in general terms how run-off becomes polluted with fertilisers.
  (b) What are the chemical names of the fertilisers that contribute to this problem?
  (c) Describe how this issue may be managed to minimise eutrophication.

Water treatment

23. Choose a method of desalination.
  (a) Draw an annotated diagram of the process.
  (b) Prepare an explanation of your chosen process that would be suitable for students at your year level who are not studying chemistry.
  (c) Find out some locations where your chosen process is used.

Exam practice questions

In a chemistry examination you will be required to answer a number of short and extended response questions.

studyon

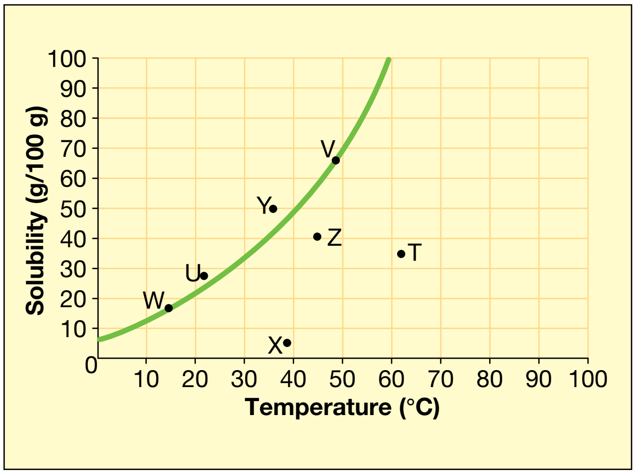
Unit 2
AOS 2
Topic 1

Measuring solubility and concentration

Sit topic test

Extended response questions

1. Various solutions of potassium nitrate are represented by the letters T, U, V, W, X, Y and Z on the solubility curve shown in the figure below. Which letters represent:
   (a) saturated solutions
   (b) unsaturated solutions
   (c) supersaturated solutions? **3 marks**



2. Calculate the concentration of the following aqueous solutions in:
   (i) g L^{-1}
   (ii) mg L^{-1} (ppm).
      (a) 3 g of solute in 100 mL solution
      (b) 0.004 g in 800 mL
      (c) 500 mg in 200 mL **6 marks**
3. It has been proposed that gold could be obtained from sea water. The concentration of gold varies from 0.1 to 2.0 mg per tonne.
   (a) Assuming that 1 tonne is equal to 1000 L of sea water, calculate the concentration of gold in parts per billion.
   (b) Assuming that there is 1.0 mg of gold per tonne of sea water, calculate the volume of sea water that would need to be processed to produce a one-kilogram gold ingot.
   (c) Why do you think this proposal would be uneconomical?

CHAPTER

15

Analysing water for salts

As we have already seen, water is an excellent solvent and is hardly ever found in the pure state. In this chapter, we will focus on dissolved salts. Other common dissolved substances, such as acids, bases and organic compounds, will be discussed in chapters 16 and 17.

YOU WILL EXAMINE:

- the importance of analysis and the difference between qualitative analysis and quantitative analysis
- protocols for obtaining water samples for analysis
- the sources of salts found in water
- a definition of a contaminant, and how contaminants may enter waterways
- how calculations can be made from chemical equations using stoichiometry
- a range of analytical techniques applicable to water samples including conductivity, colorimetry, UV–visible spectroscopy, atomic absorption spectroscopy and gravimetric analysis.

Wisdom begins in wonder.

Socrates

The stalactites and stalagmites in this cave were formed over thousands of years by the deposition of materials that found their way, dissolved in water, through the rocks and soil above the cave.

What is chemical analysis?

Chemical analysis is the process of determining what substances are in a sample. It also frequently involves determining the amounts of such substances. For example, it might be used to measure the masses of various elements in a compound so that its formula can be calculated. It can be used to test urine samples from athletes and racehorses to see whether they have taken any prohibited drugs. In the environment, analysis of samples can ascertain what is naturally in air and water and also identify possible contaminants. This enables statutory bodies such as the Environment Protection Agency (EPA) to set allowable limits and to impose penalties in order to protect the environment. It also allows water that might be unsuitable for one purpose to be safely used for another, thus making more efficient use of precious water resources.

Knowing what is in a water supply may permit its use in some situations but not in others. This allows a more efficient use of water.

A large number of methods are available to the analytical chemist. However, they can be divided into two broad groups: *qualitative analysis* and *quantitative analysis*.

Qualitative analysis

Qualitative analysis finds what substances are present. Quantitative analysis measures the quantity of substance present.

In **qualitative analysis**, the chemist is merely interested in *what* is present. In testing a urine sample from a racehorse, for example, the mere presence of a banned drug (or its metabolised products) is all that is necessary. In the confectionery industry, a sample of imported food dye might be tested to see whether it contains chemicals that are banned in this country.

Quantitative analysis

On the other hand, the question of *how much* may need to be answered — this is **quantitative analysis**. A brewer of a low alcohol beer, for example, needs to know whether its alcohol content is below a certain limit. Health authorities might need to know whether the level of mercury in samples of fish is below the allowed level.

A logical sequence often used in many analytical procedures is to perform a qualitative analysis first to find out what is present, and then to perform a quantitative analysis to find the various amounts of substances present.

Methods of chemical analysis

The analytical chemist uses methods that range from sophisticated to very simple. Techniques such as gravimetric analysis (analysing by mass) and volumetric analysis (analysing with accurate concentrations and volumes of solution) may be used. These techniques may already be familiar to you through your laboratory work. More likely, however, the analyst will use advanced instrumental techniques such as atomic absorption spectroscopy (AAS) and various types of chromatography. Due to advances in computer-chip technology, instruments for these techniques and many others have become cheaper, more powerful and more user friendly. For example, a common instrumental set-up today is to have a gas chromatograph (GC) connected to a mass spectrometer (MS) and a computer. Once the instrument has been configured and calibrated, a sample can be injected, its components identified and measured, and accurate results printed out quickly and reliably. Previously, chemists would have had to manually measure and interpret gas chromatographs from a paper printout one peak at a time, and then run a separate mass spectrum.

A technician injects a sample into a gas chromatograph. This sensitive instrument can analyse samples as small as one microlitre (10^{-6} L).

Deciding on an analytical method

When deciding on a method of analysis, the properties of the substance under investigation must be considered. A chemist therefore gives careful thought to physical properties such as melting and boiling temperatures, colour and solubility. The way the substance reacts chemically is also considered. Acid–base reactions are often important, while in other cases, redox reactions might be chosen as the basis for analysis. For example, to analyse a sample of oven cleaner containing the base sodium hydroxide, an acid might be used. However, a sample of bleach containing the oxidant sodium hypochlorite would most likely be analysed by reaction against a suitable reductant. In other cases, the most appropriate method for identifying and determining a substance may be by adding something that forms a precipitate, or alternatively, something that produces a colour that can be matched against a set of standards (colorimetric analysis).

Similarly, if instrumental analysis is being considered, the properties of the substance under test are also critical to the final choice. Many metals, for example, lend themselves to atomic absorption spectroscopy. The boiling temperature of a substance is important in distinguishing between gas chromatography and high-performance liquid chromatography if chromatographic instruments are being considered.

The type of analysis used will be determined by the properties of the substance to be analysed. For example, atomic absorption spectroscopy can be used for metal ions. Liquid chromatography is used to analyse many organic molecules.

Another important consideration for the analyst is the degree of accuracy required. This could mean the difference between choosing a tedious but accurate method in preference to a faster but less accurate one. Is a qualitative determination all that is needed, or is a quantitative one called for? An appreciation of the strengths and weaknesses of common analytical techniques is therefore necessary, so that the most appropriate method is chosen.

Standard tests to identify substances are also important. In the analysis of an unknown compound, for example, it would be important to establish that an evolved gas is carbon dioxide rather than oxygen or hydrogen.

All the considerations above are important in designing a method of analysis. If no particular standard method suits, the chemical analyst may have to display initiative and creativity in designing a new, and maybe innovative, method to get the job done.

Revision questions

1. **Classify the following analyses as either qualitative or quantitative.**
   - (a) **A pregnancy test in which a chemical is added to a sample of urine and a colour change is sought**
   - (b) **A chlorine test in which the colour of a chemical is compared with reference standards to estimate the chlorine level in a home swimming pool**
   - (c) **Placement of a detector in the exhaust pipe of a car during a tune-up procedure to measure the level of carbon monoxide emissions**
   - (d) **Testing for monosodium glutamate, MSG, in a sample of food claimed to be 'MSG free'**
2. **Oil tankers need to clean residual oil from their tanks before they load new cargo. The washings are supposed to be stored on board but unscrupulous captains sometimes dump these wastes at sea to reduce costs.**

   In Victoria, there have been numerous cases of such wastes washing ashore and polluting beaches. Sometimes this has occurred in the vicinity of Phillip Island and has affected its colony of fairy penguins.

   There have been instances where the analysis of such material has led to the prosecution of offending tanker captains.
   - (a) **Describe how qualitative analysis might assist in such cases.**
   - (b) **Given that such oil is usually a complicated mixture, describe how quantitative analysis could lead to the identification of the ship causing such pollution.**

Analysing water

Sources of salts in water

The salts found naturally in water come from a wide variety of sources. Many minerals in the ground dissolve as the water flows over the ground or percolates through layers of rock underneath it. Salt (sodium chloride) is the most common salt found in water, and the fact that the oceans contain so much sodium chloride is evidence of this.

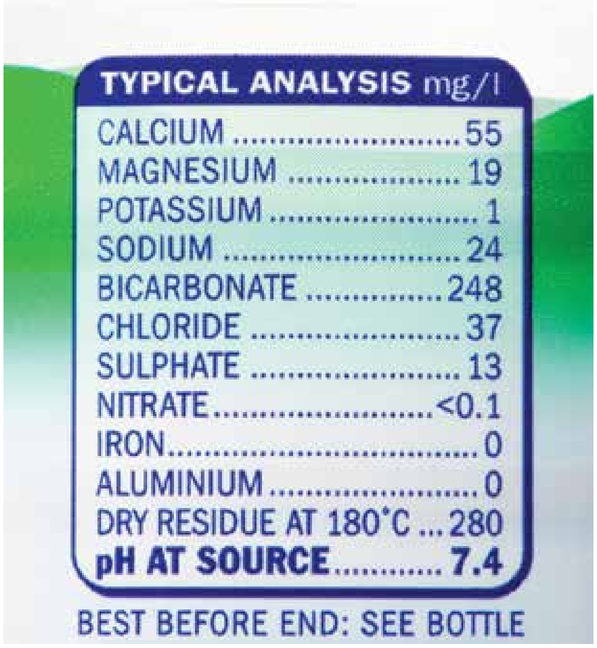
Soluble salts can hardly ever be mined because they dissolve so readily in water. They are usually found in regions with an extremely arid climate. One such region is the Atacama Desert in Chile, which has extensive deposits of Chile saltpetre (sodium nitrate, $NaNO_3$). This substance is rarely used these days, so many mining sites are now ghost towns, like this one in Humberstone, Chile.

It should be remembered, however, that salts do not need to dissolve in large amounts to have a significant effect on how water may be used. Good examples of this are salts of calcium such as calcium sulfate, $CaSO_4$, and calcium carbonate, $CaCO_3$. These may dissolve from minerals such as gypsum and dolomite, and rocks such as limestone and chalk, as water flows over or through them. Calcium sulfate has limited solubility in water (2.5 g L^{-1} at normal temperatures). Calcium carbonate is virtually insoluble but reaction with water and dissolved carbon dioxide produces calcium hydrogen carbonate, which is much more soluble. The equation for this reaction is:

$$CaCO_3(s) + CO_2(aq) + H_2O(l) \longrightarrow Ca(HCO_3)_2(aq)$$

Dissolved calcium ions, Ca^{2+}, along with magnesium ions, Mg^{2+}, are a major cause of hardness in water, which makes it difficult to lather and also makes it unsuitable for use in boilers and other high-temperature applications.

Human activity can also affect the levels of dissolved salts. A good example of this is the increasing salinity levels in the Murray River. Until 100 million years ago, the Murray–Darling Basin was covered by a shallow sea. After this sea retreated, the salt deposits left behind were buried underground. Since the advent of extensive irrigation, the extra water soaking into



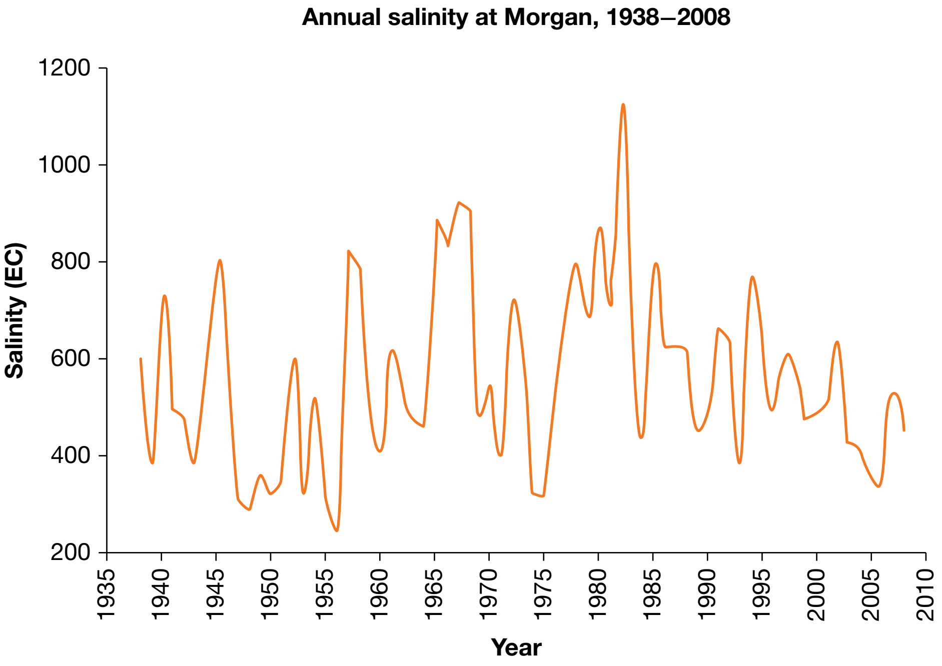
TYPICAL ANALYSIS	mg/l
CALCIUM	55
MAGNESIUM	19
POTASSIUM	1
SODIUM	24
BICARBONATE	248
CHLORIDE	37
SULPHATE	13
NITRATE	<0.1
IRON	0
ALUMINIUM	0
DRY RESIDUE AT 180°C	280
pH AT SOURCE	**7.4**

BEST BEFORE END: SEE BOTTLE

The label on this bottle of mineral water shows the many substances that are dissolved in it. These have come from the rocks and minerals at the source of the water.

Increased salinity levels along the Murray River are evidence of water's excellent solvent properties.

the ground has caused water tables to rise. This has brought large amounts of this previously undisturbed salt to, or near to, the surface from where it can enter rivers and streams, which eventually flow into the Murray River. The response to this problem has been the establishment of a number of salt interception schemes along the river. The graph below shows how salinity levels increased at Morgan in South Australia until 1982, but now appear to be decreasing, perhaps as a result of such schemes.



Salinity levels increased at Morgan in South Australia until 1982 but now appear to be decreasing.

Revision question

3. Explain why electrical conductivity provides a measurement of overall salt content rather than accurate values for particular salts.

Sources of contaminants in water

Dissolved salts in water can come from natural sources as well as the results of human activity.

A contaminant can be defined as an unwanted substance that makes water unsuitable for an intended use. Both inorganic and organic contaminants enter water in many ways. These include:

- *direct discharge* from factories and other sources. This can also happen during wet weather when some older sewerage systems overflow and when ships discharge ballast water inappropriately.
- *stormwater run-off*. As water flows over the surface, it may come into contact with surface contaminants and dissolve chemicals from them in the process. It may also flow through contaminated soil, dissolving soluble components and carrying soil particles along as a suspension.
- *contaminated air contact*. Components of contaminated air may either dissolve directly in water or react with it to produce a range of undesirable chemicals.

- *contact with contaminated ground water*. Ground water may become contaminated when surface water comes into contact with contaminated soil and then soaks into the ground. It may also occur when leaching through old buried deposits such as landfill sites. Sometimes, the material being leached out is organic in nature and largely insoluble in water. In cases such as this it may form 'pools' in underground reservoirs and become a source of long-term pollution as it dissolves very slowly over many years.

TABLE 15.1 Typical sources of some selected inorganic contaminants

Contaminant	Sources	Allowable level*
arsenic	manufacture of electrical components; wood preservatives and pesticides; gold mining; processing of ceramic materials	10 ppb
cadmium	contaminant in metals used to galvanise pipes and their subsequent erosion; metallurgical industries; improper waste disposal (e.g. nickel–cadmium batteries)	2 ppb
chromium	metallurgical processes such as alloying and electroplating; catalysts and oxidants; paint pigments	50 ppb
lead	plumbing systems; paint pigments; cable sheathing; ammunition; solder; batteries	10 ppm
mercury	mining; industrial wastes; fungicides; electrical equipment; batteries	1 ppb
nitrate	fertiliser run-off; farm animal wastes; sewage run-off	50 ppm (less than 3 months old) 100 ppm
sulfate	saltwater intrusion; mineral dissolution (e.g. $CaSO_4$); many industrial processes	250 ppm

**Source:* Australian Drinking Water Guidelines, 2011.

Organometallic contaminants

A number of heavy metals can form compounds with organic molecules. While inorganic heavy metals are often only sparingly soluble, organometallic heavy metal compounds are much more soluble and can therefore enter the food chain more easily. Methyl mercury is one such example. While dimethyl mercury is virtually insoluble (and a very dangerous poison), monomethyl mercury is essentially ionic and dissolves much more readily in water. Its structure is shown on the left.

$H_3C{-}Hg^{+}X^{-}$

The structure of monomethyl mercury, often called methyl mercury, where *X* is an appropriate anion

You will note from this structure that there is a covalent bond between the mercury atom and the methyl group. This is an example of a more complex type of bonding that does not fit the models that we considered in Unit 1.

A number of heavy metals can form organometallic compounds. This makes them much more environmentally dangerous.

The dangers of methyl mercury were first exposed in 1956 when residents of the Minamata Bay area in Japan became seriously ill after eating large amounts of fish. Many died and the effects still linger today in the form of mutations that occurred at birth. Estimates of the cumulative death toll vary but it is generally agreed to be many thousands. The source of the contamination was traced to a chemical company that had been releasing methyl mercury into the water. Methyl mercury was being used as a catalyst in the manufacture of

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Unit 2
AOS 2
Topic 1
Concept 2

Water sampling protocols
Summary screen and practice questions

acetaldehyde, a precursor to PVA (polyvinyl acetate). Other instances have occurred around the world since then. What makes this compound particularly dangerous is that it can also be *produced* when micro-organisms in water combine with organic residues that may be present in the water as well.

Water sampling protocols

To obtain an accurate analysis of a water sample, it is important that a number of steps are followed. These include:

- making sure the sample collected is a *representative* sample. This may include sampling from a variety of locations and depths or taking samples from a location where the water is considered to be well mixed. It would also involve taking it at a time that is representative of the 'average conditions' for the location. It would not be appropriate, for example, to take samples during a flood unless that was the specific reason for the analysis.
- use of *proper techniques* in the field to prevent cross contamination
- use of *equipment and procedures for storage of the samples* that prevent pre- and post-collection contamination. These include the types of containers to be used, whether they are to be nearly filled or completely filled, whether they need to be rinsed before collection and even what types of lids can be used.
- collection of *duplicate* samples
- *preservation* of samples. This may involve cooling or freezing, or adding chemicals to preserve the sample. If chemicals are added, the preparation of 'blanks' is important; these contain everything the real sample does but are made up with distilled water rather than the sampled water. Once collected it is important that samples are analysed within certain prescribed times.
- *security during transport* to avoid contaminants such as exhaust gases and dust. This is important as the expected levels of contaminants in the sample may be in parts per million or even parts per billion.
- *choice of appropriate analytical technique.* Prescriptions exist for analysis of particular contaminants. Additionally, such tests are often required to be carried out only by suitably qualified laboratories.

Strict protocols exist for the collection of water samples in order to produce accurate analytical results. These apply before, during and after sampling takes place.

The level of sophistication required can be illustrated by the collection of samples for Fe^{2+} analysis. Many metals can be collected in glass, polyethene or polypropene containers provided they are first washed with nitric acid to provide a pH between 1 and 2. However, for Fe^{2+}, hydrochloric acid must be used instead, as strong nitric acid is an oxidant and may convert some of the Fe^{2+} into Fe^{3+}. For the same reason, it is specified that the containers must be completely filled to exclude air, as oxygen gas is also an oxidant.

There are also a number of tests that are done 'in the field'. These include measurements of temperature, pH, dissolved oxygen, some anions such as fluoride and sulfide, and turbidity (cloudiness). Portable instruments are used for these tests, and these must be regularly calibrated against laboratory standards to maintain their accuracy.

Revision question

4. **Explain why it is necessary to have strict protocols for the collection of water samples for analysis.**

A closer look at analytical methods

As mentioned earlier in this chapter, a range of different methods exist for analysing the chemicals in a water sample. We will now take a more detailed look at some of these.

Electrical conductivity provides a quick and inexpensive way of measuring the combined level of dissolved salts. Measurements are usually made in electrical conductivity units (ECs).

Electrical conductivity

Salts are ionic compounds; when they dissolve in water, they dissociate to produce mobile ions. This means that they can conduct electricity to an extent that depends on how many ions are present. Electrical conductivity is therefore a quick and cheap method that can be used where the *total* salt content of a water sample is to be measured; it does not determine which particular salts are present. It is measured in electrical conductivity units (EC). The graph on page 310 showing salinity levels at Morgan shows salinity levels in EC units.

study on

Unit 2
AOS 2
Topic 2
Concept 1

Electrical conductivity
Summary screen and practice questions

The water in this river contains many dissolved salts. The total amount of these salts can be measured easily using electrical conductivity.

Colorimetry

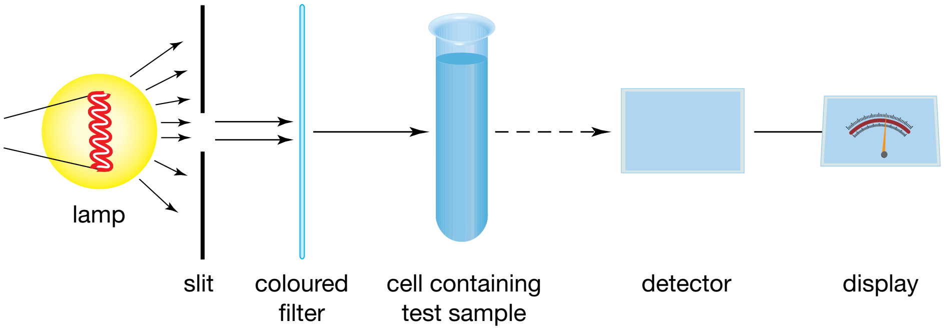
Colorimetry is a technique in which colour levels are compared with a set of standards; where the intensity of the colour depends on the component being analysed.

If a solution is naturally coloured, or can be coloured by adding certain chemicals, colorimetric analysis may be performed. Two common examples of this are the use of universal indicator to measure pH, where the colour produced is compared by eye with a standard chart, and estimating the level of chlorine in swimming pool water, where the sample is coloured by adding chemicals and the results compared with a chart.

Instrumental colorimetry is a quick method suitable for some coloured solutions. It involves shining light of a complementary colour through the solution and measuring the amount absorbed.

More accurate results can be obtained by using an instrument called an *instrumental colorimeter*. This compares the colour in the test sample with the colours produced in samples of known concentration (standards) that have been treated identically to the test sample. From the readings produced, a graph of absorbance versus concentration (called a calibration curve) may be produced and the concentration in the sample may be read directly from this graph.

The basic design of a colorimeter is shown below.



The essential features of the colorimeter

A colour's complement is the colour left after the original colour is removed from white light.

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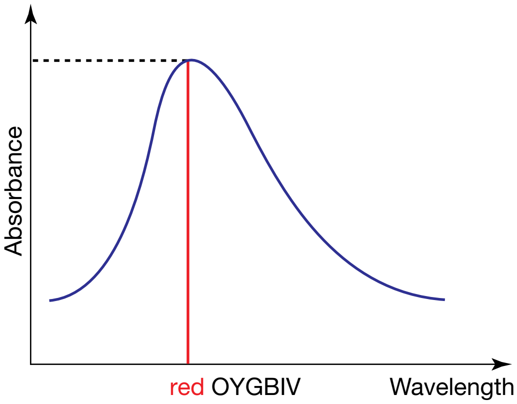
Unit 2	**Measuring phosphates by colorimetry**
AOS 2	Summary screen and practice questions
Topic 2	
Concept 2	

In a colorimeter, the detector measures the amount of light that passes through the test cell and converts this into a reading that is displayed. While normal white light can be used, more accurate results are obtained if light that is *complementary* to the colour in the test cell is used. When light of a particular colour is removed from white light, the remaining light produces a different colour. The second colour is said to be complementary to the original colour. For example, if copper sulfate levels are being measured, the best colour to shine through such samples is red light.

TABLE 15.2 Observed colours and their complementary colours

Colour observed	Colour absorbed	Frequency of absorbed colour ($\times 10^{14}$ Hz)
green	deep red	3.85–4.41
blue–green	red	4.41–4.83
blue	orange	4.83–5.17
blue–violet	yellow	5.17–5.45
violet	yellow–green	5.45–5.76
deep red	green	5.76–6.00
red	blue–green	6.00–6.38
orange	blue	6.38–6.81
yellow	blue–violet	6.81–7.14
yellow–green	violet	7.14–7.89



A blue substance strongly absorbs red light, so this wavelength is chosen for the analysis.

Colorimetry is relatively cheap and is obviously useful for measuring the concentration of a coloured species (or a colourless species from which a coloured derivative can be easily prepared). Its major drawback is that care must be taken to ensure that no other coloured species are present in a sample that could also absorb the particular colour of light being used. This method can produce results to within 1–2% accuracy.

UV–visible spectrometry and atomic absorption spectrometry are two further methods that can be used if more accurate results are required.

Determining phosphate levels — an example of colorimetry in use

Phosphates are important nutrients for plants but can become an environmental problem when they enter water systems in large amounts. As we have already seen, this can lead to eutrophication if certain other conditions are present as well. Sources of phosphate include fertilisers such as ammonium phosphates (typically $(NH_4)H_2PO_4$ and/or $(NH_4)_2HPO_4$) and superphosphate (a mixture of calcium dihydrogen phosphate, $Ca(H_2PO_4)_2$, and calcium sulfate, $CaSO_4$), which are applied in large amounts as part of current agricultural practices. Until recently, laundry detergents were another source of this nutrient; they were added to remove hardness and buffer the wash water to maintain a slightly alkaline pH. Environmental awareness by consumers, however, has now led to phosphate removal from virtually all such detergents.

Testing for phosphate levels is a relatively easy application of colorimetry. The process involves the addition of ammonium molybdate, a chemical that forms a blue compound if phosphate is present. The more phosphate, the more intense the blue colour of the solution. If a set of standards containing known phosphate levels is produced in the same way as the unknown, the level in the

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Unit 2	**Standard solutions**
AOS 2	Summary screen and practice questions
Topic 2	
Concept 3	

unknown sample may then be determined. This can be done roughly by eye using colour matching, or by determining a calibration curve produced from an instrumental colorimeter.

UV–visible spectroscopy

UV–visible spectroscopy measures the absorption of light at a particular frequency. It is suitable for coloured solutions, solutions that can be made coloured, and solutions that absorb in the ultraviolet region.

In many ways, **UV–visible spectroscopy** is a more sophisticated development of instrumental colorimetry. The basic idea is the same — the amount of absorption is related to the concentration of the substance being tested. However, this technique is far more selective and therefore less likely to suffer interference from similarly coloured compounds. This is because light of a specific frequency is used. For example, although two compounds in a sample to be tested appear blue, one of these might absorb strongly at a frequency that the other does not. This frequency can then be used in the subsequent analysis to distinguish between them.

Just as many substances absorb light from the visible section of the electromagnetic spectrum (and therefore appear coloured), there are also substances that absorb radiation from the ultraviolet region of the spectrum. As our eyes are not able to detect this radiation, substances absorbing in this region may not necessarily appear coloured. UV–visible spectroscopy is therefore suitable for many colourless substances as well as coloured ones.

The UV–visible spectrophotometer can be used both qualitatively and quantitatively.

For qualitative analysis, the sample to be analysed is dissolved using a suitable solvent. A spectrum is obtained by measuring the absorbance against a range of frequencies. This spectrum can then be compared to known spectra for the suspected substances in the sample. In practice, UV–visible spectroscopy is used only as supporting evidence in an identification, not as a qualitative analysis method in its own right.

In quantitative analysis, a pure sample of the substance to be measured would first have its spectrum determined as described above. From this spectrum, a frequency would be chosen at which strong absorption occurs. The absorbance of the test sample at this frequency would then be compared to the absorbance of a set of standards of known concentration at this chosen frequency. For example, to measure the level of glucose in a sample of urine, the spectrum of pure glucose would be obtained and a suitable frequency chosen. Glucose samples of known concentration would then be tested at this chosen frequency and their absorbances noted. Finally, the sample of urine would be tested and the absorbance compared to that of the standards to obtain the glucose concentration. Note that in choosing a suitable frequency for such an analysis, care must be taken to choose a frequency at which the glucose in the urine is the only substance that absorbs.

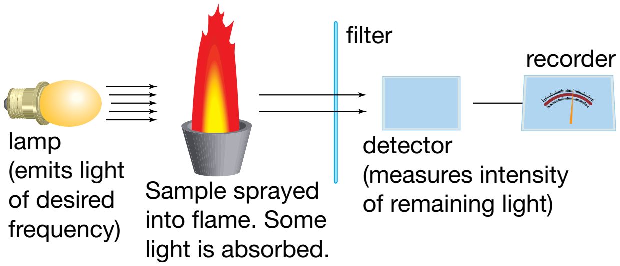
Atomic absorption spectroscopy (AAS)

Atomic absorption spectroscopy is suitable for the detection of many metals and metalloids. Atoms of a particular element, when energised in a flame, absorb light from an emission lamp containing that same element.

Atomic absorption spectroscopy, which was developed in Australia by the CSIRO, is useful for detecting the presence of metal ions.

The instrument used is called an atomic absorption spectrometer. It uses the absorption of light to measure concentrations of metal ions. It works on the principle that atoms absorb light if the frequency (and therefore the energy) of this light is of the correct value to promote an electron from its ground state energy level to a higher energy level.

Solutions of known concentration are analysed first, followed by the solution being tested. These solutions are drawn into the flame. A lamp is used to shine light of a very specific wavelength through the flame towards a detector. The amount of light absorbed by the flame is measured and can be used to determine the concentration of the substance being analysed.



The atomic absorption spectrometer

The flame is usually an air/acetylene, C_2H_2, mixture. However, for some analyses a hotter flame is required. In such cases nitrous oxide, N_2O, may be added to the gas mixture.

This technique is both very sensitive as well as being very selective. Concentrations of parts per million are easily measured, and, for some ions, parts per billion. The selectivity of the instrument allows a particular component of a mixture to be analysed without having to separate it from other components. For example, the proportions of sodium and potassium in a salt substitute can be measured without one ion affecting the other. This selectivity is due to the frequency of the light from the lamp. If, for example, a lamp is used that emits light at a frequency found in the sodium spectrum but not the spectrum of potassium, only the sodium atoms in the flame absorb the light. The sodium content can thus be measured with no interference from the potassium.

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Interactivity
Atomic absorption spectroscopy (AAS)
int-6356

Allan Walsh — the Australian inventor of the atomic absorption spectrometer

study on

Unit 2	**Atomic absorption spectroscopy (AAS) and calibration curves**
AOS 2	
Topic 2	
Concept 4	Summary screen and practice questions

Sample problem 15.1

A sample of drinking water was collected for analysis of its sodium (ion) content.

Immediately prior to analysis by AAS, it was diluted by the addition of an equal volume of deionised water.

Use the results to determine the level of sodium in the original sample.

Solution: A set of sodium standards were analysed, followed by the diluted sample. The results are shown in table 15.3.

TABLE 15.3 Absorbance versus concentration of sodium

Solution concentration (ppm)	Absorbance reading
5	0.367
10	0.712
15	1.110
20	1.470
test sample	0.989

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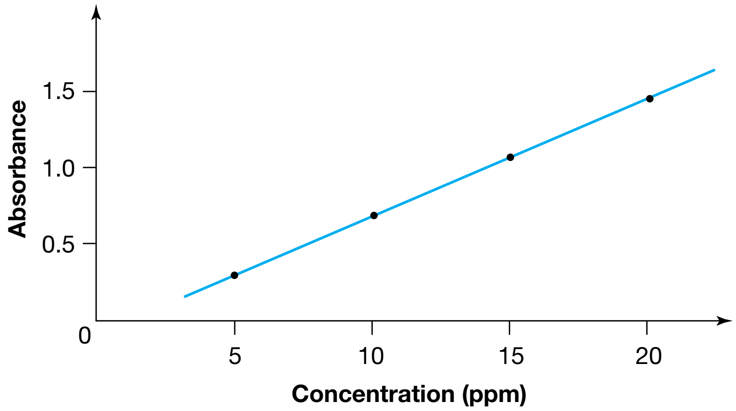
Digital document
Experiment 15.1 Zinc content of cornflakes using AAS
doc-16015

A calibration curve was produced from these results, as shown below. Reading the absorbance value off this graph revealed that the test sample had a concentration of 13.5 ppm.

For quantitative use, both instrumental colorimetry and atomic absorption spectroscopy require calibration by taking readings from a set of known standards.



Calibration curve produced from the data in table 15.3

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Unit 2
AOS 2
Topic 2
Concept 5

Calculation of metal concentrations using AAS
Summary screen and practice questions

Therefore, from the calibration curve:

concentration (diluted sample) = 13.5 ppm

The original sample was diluted by a factor of 2, so:

concentration (undiluted sample) = 13.5 × 2 = 27 ppm

Note: 27 ppm is the same as 27 mg L^{-1}.

Such a level may need to be taken into account for people on low-sodium diets.

Revision question

5. **A second sample of water suspected to contain a higher level of sodium was analysed on the same instrument as in sample problem 15.1, using the same standards. Due to its suspected higher level of sodium, it was diluted by adding 90 mL of deionised water to a 10 mL sample of the sampled water.**
   AAS produced a reading of 0.790 from the diluted sample.
   What is the concentration of sodium (in mg L^{-1}) of this sample?

Gravimetric analysis

Gravimetric analysis is analysis by weight. Among other steps, it involves collecting and drying precipitates, and then weighing them. In order to understand the calculations involved, it is important to know how chemical equations can be used to calculate masses of reactants and products involved in chemical reactions.

We will now consider this important topic before returning to a more detailed look at gravimetric analysis.

Stoichiometry—calculations using balanced equations

How is chemistry involved in the operation of an automotive airbag? Before being inflated, airbags contain a small amount of solid sodium azide, NaN_3. On impact with another vehicle or object, an electric spark is released, acting as a trigger. Very quickly, the sodium azide in the airbag decomposes to form solid sodium, Na, and nitrogen gas, N_2, according to the chemical equation:

$$2NaN_3(s) \longrightarrow 2Na(s) + 3N_2(g)$$

The sodium then reacts with water vapour in air. The nitrogen gas produced in the reaction inflates the airbag in time to absorb much of the impact energy of the crash, which may otherwise have caused injury to the driver. The sodium azide reaction must produce just the right amount of nitrogen gas to inflate the bag to the correct pressure. How do designers know how much sodium azide to use so that the airbag is correctly inflated? They use stoichiometric calculations.

Stoichiometry is the term used for calculations based on chemical equations. Stoichiometry predicts the relative amounts of substances involved in chemical reactions.

The study of relative amounts — ratios — of substances involved in chemical reactions is known as **stoichiometry**. The word 'stoichiometry' comes from two Greek words meaning 'element' and 'measure'. Knowledge of stoichiometry is essential whenever quantitative information about a chemical reaction is required.

Stoichiometry is used every day in the home and in industry. Baking a cake in the kitchen requires a recipe to ensure proper quantities of all ingredients are used. Imagine how a cake would taste if a chef added three times the required amount of flour! Many cooks have disasters in the kitchen by not using ingredients in the correct ratios. On a larger scale, an industrial chemist reducing iron in a blast furnace needs to know exactly how much iron ore, Fe_2O_3, coke, C, and limestone (calcium carbonate), $CaCO_3$, to feed into the top of the furnace in order to produce a certain amount of iron. A disaster in industry can be life threatening because incorrect ratios of chemicals can result in uncontrolled, violent explosions.

Solving stoichiometric problems is particularly important in fields such as food chemistry, drug chemistry, forensic science and in any industry where chemicals are manufactured. Stoichiometry may be thought of as the means of obtaining a recipe for a chemical reaction.

What a balanced chemical equation tells us

Chemical equations show the relative number of molecules participating in a chemical reaction. They can also show the relative number of moles, the mole ratio, of each substance involved.

Chemically speaking, if we wish to say, 'One molecule of nitrogen gas reacts with three molecules of hydrogen gas to form two molecules of ammonia gas', we simply write:

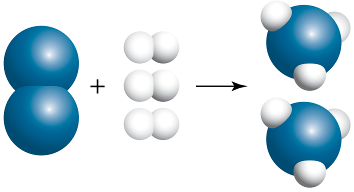
$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

However, it would be impossible and impractical to make ammonia molecule by molecule. Chemists work with larger quantities called moles (symbol mol). The coefficients in a balanced chemical equation tell us the relative numbers of moles of reactants and products taking part in the reaction.

For example, 1 mole of nitrogen gas reacts with 3 moles of hydrogen gas to produce 2 moles of ammonia gas.

The equation actually indicates the mole ratio of reactants and products. The beauty of this method is that the ratios never change. If 1 mole of nitrogen reacts with 3 moles of hydrogen, then 2 moles of nitrogen reacts with 6 moles of hydrogen, 0.5 mole of nitrogen reacts with 1.5 moles of hydrogen, and so on.

Assuming that there is plenty of nitrogen, how much ammonia is produced by 3 moles of hydrogen? The equation tells us: 2 moles of ammonia. Similarly, 6 moles of hydrogen produces 4 moles of ammonia.



The balanced chemical equation for the formation of ammonia can be interpreted in several ways.

A chemical equation does not indicate the rate (how fast) or extent (how far) of a reaction.

An equation also tells us about the ratio between other quantities of the reactants and the products including the number of particles and the masses of the reactants and products.

So, when 1 mole of N_2 gas reacts with 3 moles of H_2 gas to produce 2 moles of NH_3 gas, we know that this can be restated as: 28 g of N_2 gas (or 6.02×10^{23} N_2 molecules) reacts with 6 g of H_2 gas (or 1.806×10^{24} H_2 molecules) to produce 34 g of NH_3 gas (or 1.204×10^{24} NH_3 molecules).

What a chemical equation *does not* tell us

An equation conveys no information about the rate of a reaction. It may be fast, as in the explosive oxidation of hydrogen to form water, or it may be slow, as in the oxidation or rusting of iron.

An equation does not tell us whether a reaction requires heat or gives off heat, or what temperature or pressure is needed.

Finally, an equation gives no details as to how the individual atoms or molecules are transformed from reactants to products. This information is essential for full understanding of how a chemical reaction takes place.

Different fuels are suited for different purposes. High-performance racing cars may run on diesel, ethanol, methanol or petrol. All of these fuels react in oxygen to produce carbon dioxide and water vapour.

Sample problem 15.2

Methanol, CH_3OH, is used as a fuel in some racing cars. The equation for the combustion of methanol in the car's engine is:

$$2CH_3OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(g)$$

(a) How many moles of oxygen are needed to react with 4 moles of methanol?

(b) How many moles of water are formed when 1 mole of oxygen reacts completely with methanol?

Solution: (a) The equation tells us that 2 moles of methanol reacts with exactly 3 moles of oxygen. Use this ratio (2 : 3) to calculate that 4 moles of methanol reacts with 6 moles of oxygen exactly.

(b) The equation tells us that 3 moles of oxygen produces 4 moles of water. Use the ratio (3 : 4) to conclude that 1 mole of oxygen produces $1\frac{1}{3}$ moles of water.

Revision questions

6. **Silver tarnishes partly because of the presence of small amounts of hydrogen sulfide, H_2S (a gas that originates from the decay of food and smells like rotten eggs), according to the reaction:**

$$4Ag(s) + 2H_2S(g) + O_2(g) \longrightarrow 2Ag_2S(s) + 2H_2O(l)$$

(a) How many moles of silver sulfide form from the complete reaction of 1 mole of silver?
(b) How many moles of hydrogen sulfide react with 1 mole of silver?
(c) How many moles of silver sulfide form from 3.5 moles of hydrogen sulfide?

7. **Methane is the main component of the natural gas that we use as a fuel. When methane burns in air, the following reaction takes place:**

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

(a) How many moles of methane are needed to react with 1 mole of oxygen gas?
(b) How many moles of oxygen are needed to react with 0.1 moles of methane?
(c) How many moles of carbon dioxide are produced from 0.1 moles of methane?
(d) How many moles of water are produced from 0.1 moles of methane?
(e) How many moles of carbon dioxide are produced by 0.1 moles of oxygen gas?
(f) How many moles of oxygen gas react completely with 0.25 moles of methane?
(g) How many moles of water are produced from 8 moles of oxygen gas?

Mass–mass stoichiometry

Mass–mass stoichiometry involves solving a problem in which the mass of a reactant or product is given. You are then asked to calculate the mass of another reactant or product. Mass–mass stoichiometry requires the conversion of masses of substances to moles, or moles of substances to masses. Such conversions make use of the formulas:

$$\text{moles}(n) = \frac{\text{mass}\,(m)}{\text{molar mass}\,(M)}$$
$$\text{mass}\,(m) = \text{moles}(n) \times \text{molar mass}\,(M)$$

Mass–mass stoichiometric problems can be solved in four steps:

1. Write a balanced chemical equation for the reaction, identifying the known (given) and unknown (required) quantities of substance.
2. Calculate the number of moles of the known quantity of substance present.
3. From the equation, find the molar ratio that states the proportion of known to unknown quantities in the reaction and use it to calculate the number of moles of the required substance.
4. Calculate the quantity (mass) of the required substance.

Mass–mass stoichiometry involves calculating the mass of one of the substances in a reaction by using the equation and the mass of one of the other substances. The steps are:

- Write a balanced equation, identifying the known and unknown quantities.
- Change known mass into moles.
- From the equation, use the molar ratio to predict the moles of the required substance.
- Change the moles into mass of the required substance.

Sample problem 15.3

Some sulfur is present in coal in the form of pyrites, FeS_2. This substance is also known as 'fool's gold'. When pyrites burns, it pollutes the air with one of its combustion products, sulfur dioxide, and produces solid iron(III) oxide. What mass of iron(III) oxide is formed from the complete combustion of 183.5 g of pyrites?

Solution:

STEP 1

Write the equation, identifying the known and unknown quantities of substance.

$$4FeS_2(s) + 11O_2(g) \longrightarrow 2Fe_2O_3(s) + 8SO_2(g)$$

Pyrites is the known quantity and iron(III) oxide is the unknown quantity.

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Unit 2	**Mass–mass stoichiometry** Summary screen and practice questions
AOS 2	
Topic 2	
Concept 6	

STEP 2

Calculate the number of moles of the known quantity of substance, FeS_2.

$$\begin{aligned} n(FeS_2) &= \frac{183.5}{55.8 + (2 \times 32.1)} \\ &= \frac{183.5}{120.0} \\ &= 1.529 \text{ mol} \end{aligned}$$

STEP 3

Find the molar ratio, FeS_2:Fe_2O_3, from the equation, and use it to calculate the number of moles of the required substance, Fe_2O_3.

The ratio is 4 : 2, which becomes 2 : 1.

$$\begin{aligned} n(Fe_2O_3) &= \frac{2}{4} \times n(FeS_2) = \frac{n(FeS_2)}{2} \\ &= \frac{1.529}{2} \\ &= 0.7643 \text{ mol} \end{aligned}$$

'Fool's gold' is chemically known as pyrites, FeS_2.

STEP 4

Calculate the required quantity (mass) of Fe_2O_3 formed by the reaction.

$$\begin{aligned} m(Fe_2O_3) &= n \times M \\ &= 0.7643 \times (2 \times 55.8 + 3 \times 16.0) \\ &= 122.0 \text{ g} \end{aligned}$$

Revision questions

8. **Calculate the mass of water that is produced when 2.8 g of methane is burned in air.**

9. **Propane burns in air according to the equation:**

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$$

- **(a)** **What mass of water is produced by 6.5 g of oxygen?**
- **(b)** **What mass of O_2 reacts with 1.7 g of C_3H_8?**
- **(c)** **What mass of CO_2 is produced by 0.50 moles of propane gas, C_3H_8?**
- **(d)** **How many grams of propane gas are needed to produce 5.92 g of $CO_2(g)$?**
- **(e)** **How many kilograms of $CO_2(g)$ are released into the atmosphere when the entire contents of a 5.0 kg cylinder of propane are used at a barbecue?**

10. **When 2.864 g of potassium iodide reacts completely with a solution of lead nitrate solution, a yellow precipitate of lead iodide forms according to the following equation:**

$$2KI(s) + Pb(NO_3)_2(aq) \longrightarrow PbI_2(s) + 2KNO_3(aq)$$

Calculate the mass of lead iodide that precipitates.

11. **A solution is made by dissolving an unknown amount of barium chloride in water. It is then added to a solution of potassium carbonate such that all the barium chloride reacts. The equation for this reaction is:**

$$K_2CO_3(aq) + BaCl_2(aq) \longrightarrow 2KCl(aq) + BaCO_3(s)$$

(a) **What information in the equation tells you that barium carbonate is a precipitate?**
(b) **If 4.582 g of barium carbonate is formed, calculate the mass of barium chloride that was originally dissolved.**

12. **Industrially, metallic iron may be obtained from an ore such as haematite, Fe_2O_3, after reaction in a blast furnace according to the equation:**

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(l) + 3CO_2(g)$$

Use a spreadsheet program or a graphics calculator to set up a spreadsheet to calculate the mass of iron that may be formed by this reaction, given an initial mass of haematite. The following template can be used as the basis for your spreadsheet.

	A	B	C	D
1	mass of haematite used (g)	moles of haematite formed	moles of iron formed	mass of iron formed (g)
2				

- **In cell B2, type = A2/159.7. Then click and drag from the bottom right corner of the cell down the column to fill the column.**
- **In cell C2, type = 2*B2. Then click and drag from the bottom right corner of the cell down the column to fill the column.**
- **In cell D2, type = C2*55.85. Then click and drag from the bottom right corner of the cell down the column to fill the column.**

Once your spreadsheet has been set up, you can enter values for the mass of haematite used in the reaction, and let your spreadsheet calculate the amount of iron you would expect to be produced.

(a) **What mass of iron would be produced by 23.7 kg of haematite?**
(b) **Explain the formulas set up in cells B2, C2 and D2.**
(c) **Add columns to your spreadsheet to calculate the mass of carbon dioxide that may be formed in the reaction.**
(d) **What mass of carbon dioxide would be produced by 805 kg of haematite?**

Limiting reactant calculations

When the amounts of reactants provided for a reaction are not in the mole ratio, amounts of one or more reactants are left over once the reaction has taken place.

The amounts of the products are limited by the amount of the reactant that is completely used up in the reaction. Let's return to our previous example:

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

Recall that a balanced equation indicates the mole ratios in which the reactants are used up and the products are formed. According to this equation, 1 mole of nitrogen gas reacts with 3 moles of hydrogen gas to produce 2 moles of ammonia gas.

If 1 mole of nitrogen and 4 moles of hydrogen are mixed and allowed to react, then all the nitrogen and 3 moles of hydrogen are used up. This means that 1 mole of hydrogen is left over. The reactant that is completely used up (in this case nitrogen) is called the **limiting reactant**. Any unreacted reactants (hydrogen in this example) are called **excess reactants**. The amount of product formed by the reaction is limited by the amount of the limiting reactant. In this case, the addition of more hydrogen to the mixture would have no effect on the amount of product formed, as there is no more nitrogen available to react with it.

Sample problem 15.4

Magnesium and oxygen react to form magnesium oxide according to the following equation:

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

If 24.31 g of magnesium and 20.00 g of oxygen are available for reaction, determine the limiting reactant, the reactant in excess and by how many moles, and the mass of magnesium oxide formed.

Solution: **STEP 1**

Convert the given information to moles.

$$n(\text{Mg}) = \frac{24.3}{24.3} = 1.00\text{ mol}$$

$$n(\text{O}_2) = \frac{20.00}{32.0} = 0.625\text{ mol}$$

In a limiting reagent calculation, first determine which substance is completely used up (the limiting reagent) and which is left over (the excess reagent). Calculations are then based on the moles of the limiting reagent.

STEP 2

Compare this mole ratio with the mole ratio in the equation. According to the equation, 2 mol of magnesium reacts with 1 mol of oxygen. Therefore, if 1.00 mol of Mg were to react we would need 0.500 mol of O_2. We have 0.625 mol O_2, which is more than enough. Therefore, Mg is the limiting reactant.

STEP 3

The amount of product formed is determined by the amount of limiting reactant. According to the equation, 2 mol of magnesium produces 2 mol of magnesium oxide. Therefore, 1.00 mol Mg would produce 1.00 mol MgO.

STEP 4

Calculate the mass of MgO produced.

$$\begin{aligned} m &= 1.00 \times M(\text{MgO}) \\ &= 1.00 \times 40.3 \\ &= 40.3\text{ g} \end{aligned}$$

Mass–concentration stoichiometry

Many chemical reactions involve interactions between solids and solutions. These reactions may be referred to as mass–concentration stoichiometry. Two solutions may be mixed to form a precipitate, or solids may dissolve in some solutions to form new products. Stoichiometric calculations that involve solids and solutions require the use of two formulas to calculate moles. When dealing with solids, we use the formula:

$$n = \frac{m}{M}$$

Mass–volume stoichiometry involves the same steps as mass–mass stoichiometry, except that the formula $n = cV$ is used.

When dealing with solutions, we use the formula:

$$n = cV$$

Mass–concentration stoichiometric problems may be solved in four steps:

1. Write a balanced chemical equation for the reaction, identifying the known (given) and unknown (required) quantities of substance.
2. Calculate the number of moles of the known quantity of substance present.
3. From the equation, find the molar ratio that states the proportion of known to unknown quantities in the reaction and use it to calculate the number of moles of the required substance.
4. Calculate the quantity of the required substance.

Note that these steps are essentially the same as for solving mass–mass stoichiometry problems.

Sample problem 15.5

Hydrofluoric acid etches glass and quartz according to the reaction:

$$SiO_2(s) + 4HF(aq) \longrightarrow SiF_4(g) + 2H_2O(l)$$

A quartz sculptor has a 500 mL container of 22.50 M hydrofluoric acid. What mass of quartz could be etched by the acid?

Solution:

STEP 1

Write the equation, identifying the known and unknown quantities of substance.

$$\underset{\text{unknown}}{SiO_2\,(s)} + \underset{\text{known}}{4HF(aq)} \longrightarrow SiF_4 + 2H_2O(1)$$

STEP 2

Calculate the number of moles of the known quantity of substance.

$$\begin{aligned} n(HF) &= cV \\ &= 22.50 \times 0.500 \\ &= 11.3 \text{ mol} \end{aligned}$$

STEP 3

Find the molar ratio from the equation and use it to calculate the number of moles of SiO_2 required.

$$n(SiO_2) : n(HF) \text{ is } 1 : 4$$

$$\begin{aligned} n(SiO_2) &= \frac{1}{4} n(HF) \\ &= 2.81 \text{ mol} \end{aligned}$$

STEP 4

Calculate the required quantity of SiO_2, using $m = n \times M$.

$$\begin{aligned} m &= 2.81 \times (28.1 + 2 \times 16.0) \\ &= 169 \text{ g} \end{aligned}$$

Hydrofluoric acid may be used to decorate crystal glasses and decorative items or etch markings into laboratory glassware. This acid can be kept in plastic reagent bottles.

Sample problem 15.6

A piece of aluminium is placed in a beaker containing 500 mL of H_2SO_4 solution and hydrogen gas is evolved. Given that the initial mass of Al was 15.14 g and its final mass was 9.74 g, calculate the concentration of the acid.

Solution:

STEP 1

Write the equation, identifying the known and unknown quantities of substance.

$$\underset{\text{known}}{2Al\,(s)} + \underset{\text{unknown}}{3H_2SO_4(aq)} \longrightarrow Al_2(SO_{4)3}(aq) + 6H_2(g)$$

STEP 2

Determine the mass of Al that was used in the reaction and calculate the number of moles this mass represents.

$$\begin{aligned} m(\mathrm{Al}) &= 15.14 - 9.74 \\ &= 5.40\ \mathrm{g} \\ n(\mathrm{Al}) &= \frac{5.40}{27.0} \\ &= 0.200\ \mathrm{mol} \end{aligned}$$

STEP 3

Find the molar ratio from the equation and use it to calculate the number of moles of H_2SO_4 required.

$$n(\mathrm{H_2SO_4}) : n(\mathrm{Al}) = 3 : 2$$

$$\begin{aligned} n(\mathrm{H_2SO_4}) &= \frac{3}{2} n(\mathrm{Al}) \\ &= \frac{3}{2} \times 0.200 \\ &= 0.300\ \mathrm{mol} \end{aligned}$$

STEP 4

Calculate the concentration of the H_2SO_4 required for the reaction.

$$\begin{aligned} c(\mathrm{H_2SO_4}) &= \frac{n}{V} \\ &= \frac{0.300}{0.500} \\ &= 0.600\ \mathrm{M} \end{aligned}$$

Revision questions

13. **Zinc metal is reacted with 400 mL of a 0.250 M solution of sulfuric acid, H_2SO_4. Calculate the mass of zinc sulfate formed.**

14. **Consider the following balanced equation:**

$$\mathrm{Cd(NO_3)_2(aq) + Na_2S(aq) \longrightarrow CdS(s) + 2NaNO_3(aq)}$$

- **(a)** **Calculate the mass of CdS produced from 235 mL of a 0.178 M solution of Na_2S with excess $Cd(NO_3)_2$ present.**
- **(b)** **What is the significance of $Cd(NO_3)_2$ being present in excess?**

15. **In a laboratory experiment, a strip of magnesium weighing 2.56 g was placed into 200 mL of a hydrochloric acid solution. Bubbles of gas were observed and identified as hydrogen. The next day, no bubbles were observed in the beaker so the magnesium strip was removed from the beaker, dried and reweighed. Its mass was recorded as 0.350 g.**

- **(a)** **Why were no bubbles observed in the beaker on the second day of the experiment?**
- **(b)** **Calculate the concentration of the acid.**
- **(c)** **Identify the sources of error in this experiment.**
- **(d)** **Outline the safety precautions that should be taken in this experiment.**

Gravimetric analysis

Gravimetric analysis is a form of quantitative analysis. It is analysis by mass.

Gravimetric analysis is a form of analysis by mass. It is a well-established technique that has been carried out in analytical laboratories for centuries.

At its simplest, it may involve drying a product to determine its moisture content. For example, we read about water being added to frozen chickens in order to increase the apparent weight at which they are sold. Gravimetric analysis can be used quickly to check whether this practice has occurred. In the area of agriculture, a soil scientist often needs to know the moisture content of a soil sample in order to determine the effectiveness of various soil conditioning procedures.

Gravimetric analysis can also be a much more sophisticated procedure in which knowledge of chemical reactions, solubilities and stoichiometry is used to determine the amount of a substance in a sample. Using chemical knowledge in this way makes gravimetric analysis a powerful tool that can be used by analytical chemists in a wide range of situations.

studyon

Unit 2
AOS 2
Topic 2
Concept 7

Gravimetric analysis
Summary screen and practice questions

A simple form of gravimetric analysis can be used to determine the water content of a sample. It involves gentle heating, usually in an oven, and subsequent weighing to determine the mass of water lost.

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eLesson
Gravimetric analysis
eles-2491

Analysing for water content

This simple procedure involves weighing an initial sample and then placing it in an oven at 110 °C. After heating, the sample is reweighed and then put back in the oven. Naturally it will weigh less due to the evaporation of some water. This cycle of heating and weighing is then repeated until no further decrease in mass is noted. This is called 'weighing to constant mass' and indicates that all the water in the sample has evaporated.

Sample problem 15.7

A 23.1 g sample of soil is heated to 105 °C for 60 minutes. It is then weighed and the process repeated a number of times. A mass of 21.0 g is eventually obtained, which does not change upon further heating.
Calculate the percentage of water in this sample of soil.

Solution: The mass of water lost from the sample is 23.1 – 21.0 = 2.1 g.

The percentage of water in the sample is therefore $\frac{2.1}{23.1} \times \frac{100}{1} = 9.1\%$.

studyon

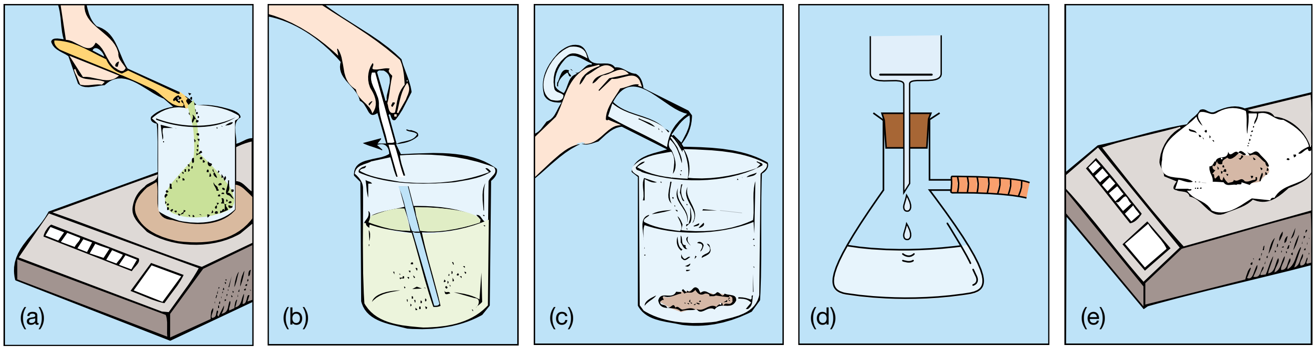
Unit 2
AOS 2
Topic 2
Concept 8

Calculating percentage mass from gravimetric analysis data
Summary screen and practice questions

Extending the gravimetric method

A more sophisticated version of gravimetric analysis can be used to quantitatively determine a component in a solid mixture.

In this procedure, the sample being analysed is usually dissolved in water and then treated with a chemical that forms a precipitate. (It is this precipitate that removes the ions required from solution so that the amount of ions may later be determined.) This precipitate is then carefully collected (usually by filtration), dried and weighed to constant mass as described earlier. From this mass and the original mass of material dissolved, the component may be determined quantitatively.



Gravimetric analysis often involves the following steps:
(a) weighing the sample to be analysed; (b) dissolving the sample in water; (c) adding a suitable chemical to form a precipitate; (d) filtering to collect the precipitate; and (e) repeated drying and weighing until a constant mass of precipitate is obtained.
Note: Sometimes, the substance being analysed is already dissolved, as with water samples. In such cases, the method starts at step (c).

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Interactivity
Gravimetric analysis
int-6357

This procedure is especially suitable for determining metal ions. It is also very useful for determining anions such as sulfate, which form some easily prepared insoluble salts (such as barium sulfate).

When designing a gravimetric procedure, a knowledge of the solubilities of the precipitates likely to be produced is important.

The steps in gravimetric analysis are:

- dissolve the test sample
- add a chemical to precipitate the component that is under analysis
- collect this precipitate
- dry and weigh the precipitate.

Simple stoichiometry can then be used to deduce the amount of the component in the original mixture.

TABLE 15.4 The solubility in water of compounds of common anions

Anion	Cations forming soluble compounds	Cations forming insoluble compounds
nitrates	all	—
chlorides bromides iodides	most	Ag^+, Pb^{2+} ($PbCl_2$ is moderately soluble in hot water.)
sulfates	most	Ba^{2+}, Pb^{2+} (Ag_2SO_4 and $CaSO_4$ are slightly soluble.)
carbonates	Na^+, K^+, NH_4^+	most
phosphates	Na^+, K^+, NH_4^+	most
sulfides	Na^+, K^+, NH_4^+	most (MgS, CaS, BaS, Al_2S_3 and Fe_2S_3 decompose in water.)
hydroxides and oxides	Na^+, K^+, Ba^{2+}	most ($Ca(OH)_2$ is slightly soluble.)

Sample problem 15.8

A pharmaceutical company wishes to test the purity of some commercial barium chloride.

A sample of the commercial barium chloride was weighed and found to have a mass of 10.0 g. After dissolving in water, excess sulfuric acid was then added to form a precipitate of barium sulfate. This precipitate was then filtered, dried and weighed. It was found to have a mass of 10.55 g.

Calculate the percentage purity of the barium chloride.

Solution: The equation for the formation of the precipitate is:

$$\mathrm{BaCl_2(aq) + H_2SO_4(aq) \longrightarrow BaSO_4(s) + 2HCl(aq)}$$

$$\begin{aligned}
m(\mathrm{BaSO_4}) &= 10.55\ \mathrm{g} \\
M(\mathrm{BaSO_4}) &= 233.4\ \mathrm{g\ mol^{-1}} \\
\therefore\ n(\mathrm{BaSO_4}) &= \frac{10.55}{233.4} \\
&= 0.045\,20\ \mathrm{mol} \\
\therefore\ n(\mathrm{BaCl_2}) &= 0.045\,20\ \mathrm{mol} \\
\therefore\ m(\mathrm{BaCl_2}) &= 0.045\,22 \times 208.3\ \mathrm{g} \\
&= 9.415\ \mathrm{g}
\end{aligned}$$

$$\begin{aligned}
\therefore\ \text{percentage barium chloride (in original sample)} &= \frac{9.415}{10.0} \times 100 \\
&= 94.2\%
\end{aligned}$$

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Interactivity
Common precipitation reactions
int-6358

Sample problem 15.9

A sample of groundwater from an old lead battery manufacturing site was analysed gravimetrically for its lead content.

A 2000 mL sample was treated with sodium sulfate solution until no further precipitate was observed to form. After filtering and drying to constant weight, 1.214 g of lead sulfate was obtained.

Calculate the level of lead ions in the groundwater in mg L^{-1}.

Solution: In cases such as this, an ionic equation is more practical than a full equation. This is because we are interested in the Pb^{2+} ions only. We can therefore write:

$$Pb^{2+}(aq) + SO_4(aq) \longrightarrow PbSO_4(s)$$
$$m(PbSO_4) = 1.21 \text{ g}$$
$$M(PbSO_4) = 303.3 \text{ g mol}^{-1}$$

Therefore, $n(PbSO_4) = \dfrac{1.21}{303.3} = 0.003\,99$ mol

$$n(Pb^{2+}) : n(PbSO_4) = 1 : 1$$

Therefore, $n(Pb^{2+}) = 0.003\,99$ mol

Therefore, $m(Pb^{2+}) = 0.003\,99 \times 207.2 = 0.827$ g

Since the amount analysed was 2000 mL (2 L), the concentration of Pb^{2+} equals:

$$\frac{0.827}{2} = 0.414 \text{ g L}^{-1}$$
$$= 414 \text{ mg L}^{-1}$$

Alternative solution:

This relies on the assumption that all the lead in the original sample precipitates. Of course, this should be true of *any* properly designed gravimetric process!

When the molar mass of $PbSO_4$ is evaluated, it becomes clear that, in the 303.3 g that represents 1 mole, 207.2 g is due to lead.

The fraction of lead in $PbSO_4$ is therefore $\dfrac{207.2}{303.3}$.

It follows that in 1.214 g of $PbSO_4$ there are $\dfrac{207.2}{303.3} \times 1.21 = 0.827$ g of lead.

The rest of the calculation is the same as above.

Sample problem 15.10

The zinc content in a water sample was analysed as follows.

A 1000 mL sample was treated with a solution of ammonium hydrogen phosphate, $(NH_4)_2HPO_4$, to precipitate all the zinc ions present. A precipitate of NH_4ZnPO_4 was obtained, which was then decomposed by heating to produce 2.918 g of $Zn_2P_2O_7$.

Calculate the concentration of zinc in the sample in g L^{-1}.

Solution: This appears to be much more complicated but is easily solved using the alternative method in sample problem 15.9. We assume that, no matter how many steps are involved, all the zinc in the original sample ends up in the final precipitate, $Zn_2P_2O_7$, that is weighed.

$$M(Zn_2P_2O_7) = 304.8 \text{ g mol}^{-1}$$

of which 130.8 g is zinc.

Therefore, fraction of zinc in $Zn_2P_2O_7 = \dfrac{130.8}{304.8}$.

Therefore, $m(Zn) = \dfrac{130.8}{304.8} \times 2.918$ g

$$= 1.252 \text{ g}$$

As this is already in 1 L, the concentration of zinc = 1.252 g L^{-1}.

Revision question

16. **The calcium content of a sample can be determined gravimetrically by first precipitating the calcium ions present as calcium oxalate. When this precipitate is heated, it decomposes completely to produce calcium oxide.**

In testing the purity of a calcium chloride sample, a chemist used gravimetric analysis and obtained 8.81 g of calcium oxide from an original 18.0 g sample of calcium chloride.

Calculate the percentage purity of the calcium chloride.

Examining sources of error

Although most gravimetric procedures have a clearly recognisable series of steps, the properties of the chemicals involved may necessitate some modifications to the method in particular circumstances.

A thorough appreciation of the method is also important in evaluating the limits of your results. If certain things happen during the analysis, it is important to predict how these may affect the final result. Table 15.5 shows some possible sources of error.

TABLE 15.5 Sources of error

Action	Effect on analysis result	Reason
insoluble materials not filtered out before forming the precipitate	overestimated	The apparent mass of the precipitate will increase.
not enough of the precipitate-forming chemical added	underestimated	Not enough of the precipitate will form as some of the ions that are being analysed will remain in solution.
forming a precipitate that is too soluble	underestimated	Not all of the ions being analysed will be in the precipitate.
forming extra precipitate due to the presence of other competing ions	overestimation	Too much of the precipitate will form.
weighing the precipitate before it is dry	overestimation	The water present will increase the apparent mass of the precipitate.
not rinsing the precipitate before drying it	overestimation	As the precipitate dries, other soluble chemicals will begin to crystallise out of the small amount of solution still trapped in the precipitate. These will add to the mass.*
adding too much of the precipitate-causing chemical	no effect	This is a necessary part of the method to make sure that all of the required ions are in the precipitate. The chemical must be in excess.
using too much water for the initial dissolving	no effect	This is a practical consideration — the more water you have, the longer the filtering step.

**Note:* To avoid this situation, the precipitate should always be washed with a small amount of pure solvent before it is dried and weighed. Too much water or solvent could cause some of the precipitate to redissolve. In this case, the final result will be underestimated.

Chapter review

studyon
Unit 2 **Analysing water for salts**
AOS 2 Summary screens
Topic 2 and practice questions

Summary

- Chemical analysis is the process of determining the substances present in a test sample. A range of techniques exist for doing this.
- The choice of a particular technique depends on many factors, the most important of which are the properties of the substance that is being tested for.
- Qualitative analysis is the process of determining which substances are present in a sample.
- Quantitative analysis is the process of determining how much of a substance is present.
- Strict protocols must be observed when collecting water samples for analysis so that the results obtained are accurate.
- The salts dissolved in water samples come from a variety of sources, including natural sources as well as the result of human activity. They include a range of salts as well as heavy metal compounds and organometallic substances.
- A contaminant is a substance that makes water unsuitable for an intended use.
- Contaminants enter waterways through a variety of methods including direct discharge, stormwater run-off, contaminated air contact and contact with contaminated groundwater.
- Stoichiometry is the study of the relative amounts — ratios — of substances involved in chemical reactions. This chapter introduced:
  - mass–mass stoichiometry
  - mass–concentration stoichiometry.
- The coefficients in a balanced chemical equation tell us the relative numbers of moles of reactants and products and can be used to find the mole ratio of one substance to another.
- A four-step approach may be used to solve stoichiometric problems.
  Step 1 Write a balanced chemical equation for the reaction, identifying the known (given) and unknown (required) quantities of substance.
  Step 2 Calculate the number of moles of the known quantity of substance.
  Step 3 From the equation, find the mole ratio that states the proportion of known to unknown quantities in the reaction and use it to calculate the number of moles of the required substance.
  Step 4 Calculate the required quantity of the substance.
- Whenever two or more reactants are given in a chemical reaction, the limiting reactant must be identified before we can calculate the maximum amount of product that may form. A limiting reactant is completely used up in a chemical reaction. The other reactants are known as excess reactants and are not used up.
- Instrumental colorimetry can be used for coloured solutions. It relies on measuring how much light of a complementary colour is absorbed by a solution.
- Atomic absorption spectroscopy (AAS) and UV–visible spectroscopy both rely on the accurate measurement of radiation of a particular frequency.
- AAS is suitable for determining many metals.
- All the instruments mentioned above may be calibrated for quantitative analysis by obtaining readings from a number of standards and then plotting a calibration curve.
- The choice of instrument for a particular analysis depends on the properties of the material that is under analysis.
- Gravimetric analysis is analysis by mass.
- A simple method of determining moisture content is to gently heat a sample and record the mass lost as the water is driven off.
- A more sophisticated version involves the formation of suitable precipitates, the amounts of which are stoichiometrically related to the amount of the component being analysed.
- This method involves weighing the sample to be analysed, dissolving the sample, forming the required precipitate, filtering the precipitate, and weighing the precipitate to constant mass.
- The exact details of a given method depend on the properties of the substances involved.
- If the properties of the substances are not carefully considered, or if mistakes are made, the final result will be affected. It is important to be able to predict how such situations will affect the final calculated result.

Multiple choice questions

1. Which of the following would be most likely to be a contaminant in a sample of surface water?
A Cl^-
B SO_4^{2-}
C Ca^{2+}
D Hg^+

2. Which of the following is an organometallic substance?
A methyl chloride
B methylzinc
C lead nitrate
D carbon dioxide

3. A field technician collected a creek water sample (CW) and added two preserving agents (A and B) to his sample. In his field kit was also a container of deionised water (DW).

The protocol also called for the preparation of a blank.

The substances that the blank should contain are:

A A, B and CW
B A, B and DW
C A, B
D DW only.

4. A calibration curve is a graph of instrumental reading against:

A mass
B time
C volume
D concentration.

5. A standard used in instrumental analysis is:

A a solution containing the sample being tested
B a solution of exactly 10.00 ppm
C a solution of exactly known concentration
D a list of prescribed tolerances for use of the instrument.

6. Magnesium reacts with oxygen according to the equation:

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

In this reaction:

A 2 g of magnesium react with 1 g of oxygen to produce 2 g of magnesium oxide
B 2 g of magnesium react with exactly 1 g of oxygen
C 2 moles of magnesium are needed for every mole of oxygen used
D 2 moles of oxygen are needed for every mole of magnesium used.

7. Before digital printing, sodium thiosulfate, $Na_2S_2O_3$, known as 'hypo' by photographers, was used to remove excess silver bromide, AgBr, in the liquid-based film-developing process according to the equation:

$$2Na_2S_2O_3(aq) + AgBr(s) \longrightarrow Na_3Ag(S_2O_3)_2(aq) + NaBr(aq)$$

In the *balanced* equation for the reaction between sodium thiosulfate and silver bromide:

A 1 mole of sodium thiosulfate solution produces 1 mole of sodium bromide solution
B 3 moles of sodium thiosulfate solution produce 2 moles of $Na_3Ag(S_2O_3)_2$
C 1 mole of sodium thiosulfate solution produces 2 moles of sodium bromide solution
D 1 mole of silver bromide produces 1 mole of $Na_3Ag(S_2O_3)_2$.

8. Phosphorus may be prepared from calcium phosphate according to the equation:

$$2Ca_3(PO_4)_2(s) + 6SiO_2(s) + 10C(s) \longrightarrow P_4(s) + 10CO(g) + 6CaSiO_3(s)$$

How much phosphorus can be produced if 1000 kg of calcium phosphate is used completely?

A 50 kg
B 100 kg
C 200 kg
D 400 kg

9. The fertiliser diammonium phosphate (DAP) is a popular source of phosphate fertiliser for farmers. It is produced by the reaction of ammonia with phosphoric acid according to:

$$2NH_3(g) + H_3PO_4(aq) \longrightarrow (NH_4)_2HPO_4(aq)$$

The mass of fertiliser that would be produced by the complete reaction of 100 kg ammonia in this process would be:

A 5.88 kg
B 388 kg
C 776 kg
D 1553 kg.

10. A gelatinous precipitate of iron(III) hydroxide may be prepared according to the equation:

$$FeCl_3(aq) + 3NaOH(aq) \longrightarrow 3NaCl(aq) + Fe(OH)_3(s)$$

A solution containing 16.23 g of $FeCl_3$ is mixed with 300 mL of a 1.00 M NaOH solution. Which of the following statements is correct?

A $FeCl_3$ is in excess.
B NaOH is in excess.
C $FeCl_3$ is the limiting reagent.
D Neither reagent is in excess.

11. We want to analyse a sample of salty water for its chloride ion content. This is to be done using gravimetric analysis. Which of the following chemicals could be added for the formation of a precipitate?

A $CaCO_3$
B $Ba(NO_3)_2$
C K_2SO_4
D $AgNO_3$

12. The heating and drying of a precipitate to constant weight ensures that:

A there is no water left in it
B impurities are removed from it
C trace amounts left in solution are precipitated
D absorption of gases from the atmosphere is minimised.

13. A sample of lawn food is analysed gravimetrically. As part of the procedure, a precipitate of barium sulfate is formed. However, when this is later weighed, it is still slightly damp. As a result of this, the calculated percentage of sulfate would:

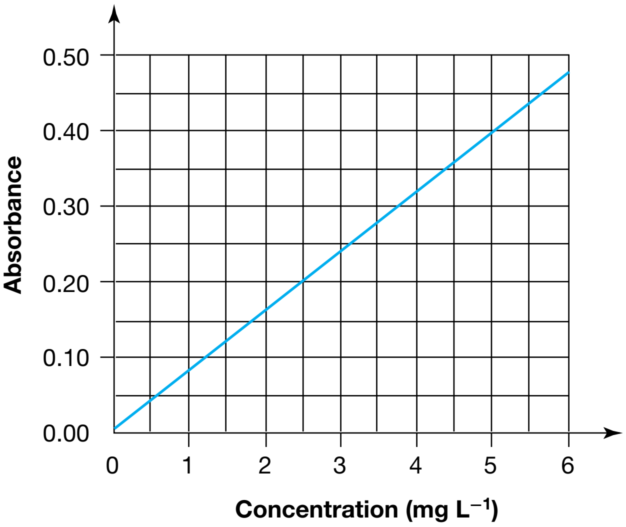
A be increased
B be decreased
C be unchanged
D vary in a random manner.

14. An atomic absorption spectrometer was used to determine the amount of iron lost by peas in the cooking process.

Using a suitable means of extraction, the iron was first extracted from a 5.0 g sample of uncooked peas to produce 100 mL of extract.

When tested in the instrument, the uncooked peas gave an absorbance reading of 0.20.

The AAS was then calibrated and its calibration curve is shown below.



The mass of iron (in mg) in the original sample of uncooked peas is:

A 0.21
B 0.25
C 2.1
D 2.5.

15. Using the technique of AAS, we want to accurately determine the concentration of some copper(II) sulfate solution. The concentration of this solution is known to be about 0.02 M.

A set of standards is available, the concentrations of which range from 2 to 16 mg (of copper) per litre.

By what factor should the copper(II) sulfate solution be diluted prior to being drawn into the instrument?

A 1
B 10
C 100
D 1000

Review questions

General analysis

1. Classify the following analyses as either qualitative or quantitative.
- (a) Analysis of the propellent from a spray can by an instrument, to check the claim that it does not contain chlorofluorocarbons, CFCs
- (b) Adding a piece of 'testape' to a diabetic's urine to estimate the glucose level present
- (c) Instrumental checking of the level of mercury in a sample of fish
- (d) Analysis of a hair-colouring preparation 'containing less than 1.5% aromatic nitro amines', to check the claim

2. A number of qualitative tests are used to identify various substances. Find out the tests that are used to identify the following.
- (a) Carbon dioxide gas
- (b) Starch
- (c) Oxygen gas
- (d) Protein in a food sample
- (e) Hydrogen gas
- (f) Water

Water solutes and sampling

3. (a) Define the term 'contaminant'.
- (b) Give an example of a situation in which water might be unsuitable for one particular use but suitable for another use.
- (c) Discuss your answer to part (a) in light of your answer to part (b)

4. Methylmercury can be a serious contaminant in water because it accumulates in fish and shellfish, which can subsequently be eaten by people.

Research and answer the following questions about methylmercury using the internet or other suitable references.
- (a) What do the terms 'inorganic mercury' and 'organic mercury' mean?
- (b) Describe how methylmercury may be produced in the environment.
- (c) What is Minamata disease? Where did this first occur and what was its cause?
- (d) Where have other outbreaks of this disease occurred and what were their causes?
- (e) Do any other metals form compounds similar to methylmercury?

5. **eBook*plus*** Procedures to be followed for the collection of water samples are listed in the document *Sampling and analysis of waters, wastewaters, soils and wastes*. See the **Sampling and analysis** weblink in the Resources section of your eBookPLUS to access this document. Use appendix A of the document to list some of the procedures that need to be followed to collect water samples for testing the following.
- (a) sulfate
- (b) magnesium ions
- (c) cyanide
- (d) Fe^{2+}
- (e) total mercury
- (f) pH

Stoichiometry

6. A solution containing 5.10 g of barium nitrate reacts completely with a solution of sodium sulfate. The unbalanced equation for this reaction is:

$$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \longrightarrow NaNO_3(aq) + BaSO_4(s)$$

(a) Balance the equation above by inserting the necessary coefficients and identify the precipitate formed.
(b) Calculate the mass of precipitate formed.

Mass–mass stoichiometry

7. Nitrogen in the cylinder of a car reacts with oxygen to produce the pollutant nitrogen monoxide.

$$N_2(g) + O_2(g) \longrightarrow 2NO(g)$$

(a) How many moles of nitrogen are needed to produce 1.52 mol nitrogen monoxide?
(b) How many grams of oxygen are needed for this reaction?

8. Barbecues burning charcoal briquettes are unsafe for indoor use because of the colourless, odourless, poisonous gas produced.
(a) Find the $n(O_2)$ gas that reacts with 3.5 g of charcoal briquettes (assume pure C) to produce carbon monoxide.
(b) If there is a plentiful supply of air, a safe colourless, odourless gas is produced. Find the mass of this gas produced if the same amount of charcoal is burned. You will need to write another equation.

9. When sodium and chlorine are reacted, table salt is produced.
(a) Write the balanced equation for the reaction.
(b) If 10.0 g of sodium is used, how much table salt would be produced?

10. In respiration, the equation for the reaction that produces energy in our bodies is:

$$C_6H_{12}O_6(aq) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l)$$

If 8.90 g of glucose is used, determine:
(a) the mass of oxygen needed
(b) the mass of carbon dioxide produced.

11. The silver used in jewellery and tableware becomes tarnished when exposed to air containing small amounts of hydrogen sulfide. The tarnish is a layer of silver sulfide.

$$4Ag(s) + 2H_2S(g) + O_2(g) \longrightarrow 2Ag_2S(s) + 2H_2O(g)$$

Calculate the mass of the tarnish when 0.025 g of silver is reacted.

12. In the final step of the Contact Process, where sulfuric acid is manufactured industrially, controlled addition of water to oleum, $H_2S_2O_7$, produces pure sulfuric acid. The equation for the reaction is:

$$H_2S_2O_7(l) + H_2O(l) \longrightarrow 2H_2SO_4(l)$$

Calculate the mass of sulfuric acid produced when 5.00 kg of oleum is dissolved in water.

13. The metal tungsten, used to make the filaments for incandescent light bulbs, can be obtained from its oxide by reduction with hydrogen.

$$WO_3(s) + 3H_2(g) \longrightarrow W(s) + 3H_2O(g)$$

(a) What mass of tungsten can be obtained from 200 g of its oxide?
(b) What mass of hydrogen is required?

14. In a class experiment, a strip of copper was placed in an aqueous solution of silver nitrate.
Initial mass of copper strip = 4.36 g
Final mass of copper strip = 2.21 g
(a) Write an equation for the reaction.
(b) What mass of silver formed in the reaction?

15. A 1.50 g sample of paint pigment was dissolved and the lead ions separated by precipitation as solid $PbSO_4$. The precipitate had a mass of 0.0806 g. Find the percentage of lead by mass in the pigment.

16. Aluminium burns in oxygen to produce an intense burst of light. Aluminium oxide is produced in the process. The equation for the reaction is:

$$4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$$

Calculate the mass of aluminium oxide produced from 5.0 g of aluminium.

17. As an emergency procedure, the *Apollo 13* astronauts used lithium hydroxide to remove carbon dioxide from the interior of their crippled spacecraft as it returned from the moon. Like all hydroxides, lithium hydroxide forms the appropriate metal carbonate when it reacts with carbon dioxide.

The equation for this reaction is:

$$2LiOH(s) + CO_2(g) \longrightarrow Li_2CO_3(aq) + H_2O(l)$$

(a) Calculate the mass of carbon dioxide that could be removed per kilogram of lithium hydroxide.
A possible alternative chemical for this process might have been the more readily available sodium hydroxide.
(b) Write the equation for the reaction between sodium hydroxide and carbon dioxide.
(c) From part (b), calculate the mass of carbon dioxide that can be removed per kilogram of sodium hydroxide.
(d) Use your answers to parts (a) and (c) to suggest a reason for the choice of lithium

hydroxide rather than sodium hydroxide in a spacecraft.

(e) Derive the ionic equations for both the reactions mentioned in this question.

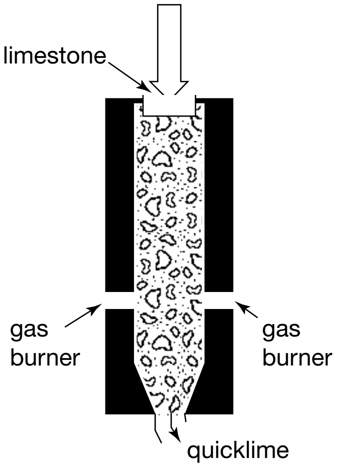
18. Limestone is an important raw material for industry. When heated in a kiln, it decomposes to form quicklime, CaO, and carbon dioxide. A typical kiln is shown in the figure below.

The equation for this reaction is:

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

(a) If limestone containing 83.5% calcium carbonate is used, calculate the mass of quicklime that would be produced from 100 tonnes of this limestone.

(b) This process is sometimes called 'lime burning'. Why is this not a correct term to use?



A typical limestone kiln. The production of lime from such kilns was one of Australia's earliest industries.

Mass–concentration stoichiometry

19. What volume of 1.3 M HCl just neutralises 2.5 g of $Ca(OH)_2$?

$$Ca(OH)_2(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + 2H_2O(l)$$

20. What mass of copper is required to react completely with 250 mL of 0.100 M $AgNO_3$?

$$Cu(s) + 2AgNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$

Analysing with instruments

21. Examine the graph on page 310.

(a) Explain why electrical conductivity can be used to measure the level of salt in the river.

(b) Describe the trends in salt content before the early 1980s and after the early 1980s.

(c) Why is it incorrect to interpret these results as evidence of sodium chloride content?

(d) Why is it more correct to interpret the results in terms of total salt content?

22. An atomic absorption spectrometer was used to check the iron content of a brand of commercially available soda water. To calibrate the instrument, four samples of known concentrations were used. The sample of soda water was then analysed. The following results were obtained (absorbance being measured in appropriate units):

Standard 1: Absorbance 0.053
Concentration 1.00 ppm
Standard 2: Absorbance 0.109
Concentration 2.00 ppm
Standard 3: Absorbance 0.157
Concentration 3.00 ppm
Standard 4: Absorbance 0.266
Concentration 5.00 ppm
Sample: Absorbance 0.121

(a) For a concentration of 0.00 ppm, what should be the value of the absorbance?

(b) Plot the calibration curve from the data above.

(c) Hence, estimate the concentration of iron in the soda water.

(d) Suggest a way in which a computer might replace part (b).

23. Flame emission spectroscopy is similar to atomic absorption spectroscopy, except that 'transmittance' is measured rather than 'absorbance'. Sodium levels in mineral waters were analysed by flame emission spectroscopy.

Sodium produces an orange light, with one of the wavelengths present being 589 nanometres (589×10^{-9} metres). The detector of the instrument was therefore set to measure the intensity of the emitted light at this wavelength.

Sample standards were run first, followed by two brands of mineral water. The results are shown in the table below.

Results from analysis of sodium levels in mineral water

Sodium (mg L^{-1})	% transmittance
0	0
5	10
10	22
20	44
30	66
40	85
50	100
Sample 1 (Australian)	86
Sample 2 (French)	12

(a) Plot a calibration curve using the sodium standards.
(b) From this curve, deduce the sodium levels in each of the samples analysed.

24. The level of mercury in various waterways can have serious implications for human health if foods such as oysters or flake produced in such waters are consumed. Consequently, various agencies regularly measure mercury levels in such waterways and in food samples produced from them.

In one particular analysis, a sample of water was analysed in an instrument and gave a reading of 0.375.

Without altering the setting of the instrument, standard samples containing known levels of mercury were then analysed, giving the results shown in the table below.

Results from analysis of water sample containing mercury

Concentration of mercury sample (mg L^{-1})	Reading
2	0.102
4	0.205
6	0.307
8	0.410
10	0.512

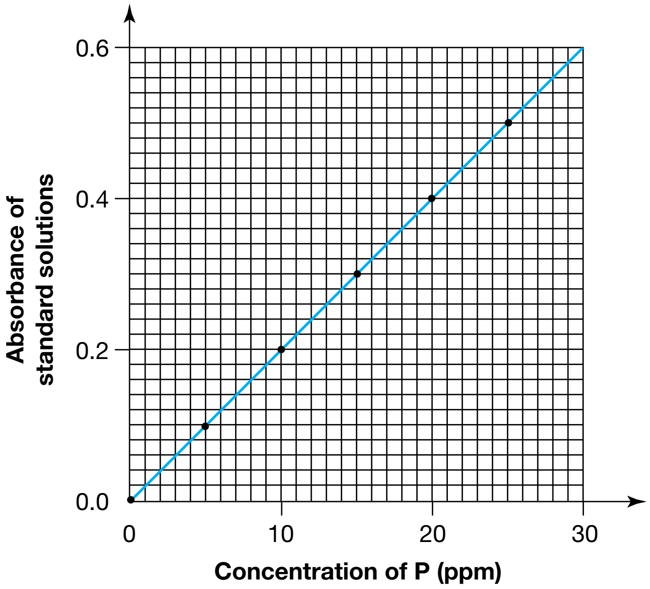
(a) Plot a calibration curve using the data in the table above.
(b) What is the level of mercury in the sample of water tested?
(c) What type of instrument do you think was used in this analysis?

25. The level of phosphorus (as phosphate) in a detergent may be determined colorimetrically by reacting it with a molybdate solution to develop an intense blue compound.

In the analysis of one particular brand of detergent, a 1.000 g sample was dissolved in 1000 mL. A carefully measured 20.0 mL sample of this was then treated with molybdate solution and heated to develop the blue colour.

When analysed in an instrumental colorimeter, a reading of 0.260 was obtained.

Four standard solutions containing phosphorus levels of 5, 10, 15 and 20 ppm were similarly treated and their absorbances measured to produce the graph shown above right.



(a) What is the concentration of phosphorus in the diluted solution?
(b) What is the percentage by mass of phosphorus in the detergent?
(c) What colour do you think the light used for this analysis would be?

Gravimetric analysis

26. In each of the following reactions, a precipitate is formed. Write the full balanced equation for each reaction and then write the corresponding ionic equation. (The precipitate formed in each case is shown in brackets.)
(a) Silver nitrate solution is added to sodium chloride solution. (silver chloride)
(b) Barium chloride solution is added to sodium sulfate solution. (barium sulfate)
(c) Lead nitrate solution is added to potassium chromate solution. (lead chromate)
(d) Sodium phosphate solution is added to silver nitrate solution. (silver phosphate)
(e) Dilute hydrochloric acid is added to lead nitrate solution. (lead chloride)

27. State whether the following substances are soluble or insoluble in water.
(a) sodium carbonate
(b) calcium carbonate
(c) barium sulfate
(d) silver chloride
(e) ammonium phosphate
(f) silver nitrate

28. The moisture content of an area of soil is one of its most important properties. A simple method for determining this is to take a soil sample and heat it in an oven at 105 °C until constant mass is achieved.

In a particular experiment, the following results were obtained.

Initial mass	124.829 g
Mass (after 60 min)	115.813 g
Mass (after 70 min)	109.614 g
Mass (after 80 min)	107.214 g
Mass (after 90 min)	107.212 g
Mass (after 100 min)	107.214 g

(a) Calculate the percentage of moisture in the soil sample.
(b) Why is the sample heated until constant mass is obtained?
(c) In a situation such as this, why would it be unwise to rely on the results from a single sample?

29. To determine the percentage of arsenic in a particular brand of pesticide, a 2.15 g sample is treated to precipitate all the arsenic as its sulfide, As_2S_3. If 0.353 g of precipitate is obtained, calculate the percentage of arsenic in the pesticide.

30. To test the percentage of sodium chloride in a sample of rock salt, a carefully weighed mass (0.997 g) was first dissolved in water. Silver nitrate solution was then added until no further precipitate was observed to be forming. This precipitate was filtered using a filtering crucible and then dried in an oven to constant weight. The mass of precipitate obtained was 2.359 g.
(a) What is the name of the precipitate formed in this process?
(b) Calculate the percentage of sodium chloride in the rock salt, assuming that it is the only source of chloride ions.

31. Sulfur in azalea fertiliser is present in the form of soluble sulfates. To check the percentage of sulfur in such a product, a 2.322 g sample was weighed and dissolved in water. Barium chloride was then added under carefully controlled conditions to precipitate out all the sulfate present as barium sulfate. After filtering, washing with water and drying, the mass of precipitate obtained was 0.564 g.
(a) Calculate the mass of sulfur in the precipitate and hence the percentage of sulfur in the sample as analysed.
(b) What is the effect on the final result if the precipitate is not washed with water before drying?

Experimental procedures and design

32. In each of the following situations, we want to analyse gravimetrically the component written in bold type. From the list in brackets, choose the ions that should be added to achieve this.
(a) **sulfate** (magnesium, aluminium, lead, ammonium)
(b) **chloride** (sodium, barium, calcium, silver)
(c) **carbonate** (potassium, lithium, barium, ammonium)
(d) **iron(III)** (nitrate, acetate, sulfate, hydroxide)

33. In gravimetric analysis, why are precipitates dried to constant weight?

34. 'Every compound is soluble to a greater or lesser extent.' Discuss the implications of this statement for gravimetric procedures. Would the above lead to a systematic or a random error? Explain.

35. Salinity is a serious problem in many of Victoria's irrigation areas. Design and describe an experiment whereby the level of salt (in $g\ L^{-1}$) in a water sample can be determined gravimetrically.

36. The following steps were considered by a student when designing a gravimetric procedure to estimate the level of magnesium carbonate in a sample of epsom salts. The steps have been written down in random order. Rearrange these steps into an order that would be acceptable, remembering that some steps might be repeated more than once and that some steps might not be required.
(a) Filter and collect the precipitate produced.
(b) Add a solution of potassium chloride.
(c) Accurately weigh out a sample of epsom salts.
(d) Accurately weigh the precipitate.
(e) Dissolve the epsom salts in a volume of water.
(f) Dry the precipitate.
(g) Measure out an accurate volume of water.
(h) Add a solution of barium chloride.

37. In a particular gravimetric procedure, a sample was weighed. It was then dissolved in water and the component of interest precipitated by adding a suitable chemical. This precipitate was then filtered and washed with a small quantity of water. Finally, it was dried in an oven before being weighed. The percentage (by mass) of the component was then calculated. Below are listed some possible faults that might occur in an analysis of this type. For each fault, predict the effect it would have on the final result.
(a) The precipitate is not dried completely before the final weighing.
(b) The precipitate is dried before it is washed with a quantity of water.
(c) The original sample is dissolved in an excessive volume of water.
(d) The experiment is performed at too high a temperature.

Exam practice questions

studyon
Unit 2
AOS 2
Topic 2
Analysing water for salts
Sit topic test

In a chemistry examination you will be required to write a number of short and extended response questions.

Extended response questions

1. A sample of limestone is analysed for its calcium carbonate content as follows. A 25.00 g sample is crushed and heated to a high temperature. It is then cooled, weighed and reheated until a constant mass is obtained. The mass remaining at the end of this process is 11.64 g.

(a) What gas is evolved during this process?

(b) Write the equation for the decomposition of calcium carbonate in the limestone at high temperature.

(c) Use the results obtained to calculate the mass of calcium carbonate in the limestone, and hence its percentage purity.

(d) State one important assumption in this method. **6 marks**

2. Phosphorus is an important plant nutrient and is therefore present in many fertilisers. It can be determined gravimetrically by carrying out a number of steps, eventually leading to its precipitation as insoluble $Mg_2P_2O_7$. In one such analysis, a 14.298 g sample of fertiliser yielded 4.107 g of $Mg_2P_2O_7$ precipitate.

(a) Calculate the mass of phosphorus in the precipitate.

(b) Hence, calculate the percentage of phosphorus in the fertiliser.

(c) State one important assumption that is made in this method. **4 marks**

3. As part of a research project into tidal effects in a river estuary, a student takes water samples at various locations and determines the concentration of salt (as sodium chloride) in each. Her method involves collecting 10.0 L of water from each location and boiling it to reduce the volume. Excess silver nitrate solution is then added to precipitate all the chloride ions present as silver chloride. This precipitate is then collected, dried and weighed to constant mass.

Following this procedure, one particular sample produced 4.463 g of precipitate.

(a) Write a balanced equation for the reaction that leads to the formation of the precipitate.

(b) Calculate the mass of sodium chloride in this particular sample.

(c) Hence, calculate the concentration of sodium chloride in this sample, in g L^{-1}.

(d) In what way does boiling the sample to reduce its volume affect your calculation?

(e) Some important assumptions are made by the student in this procedure. Name one. **7 marks**

CHAPTER

16 Analysing water for organic compounds

Many chemicals used in farming are organic in nature, and methods of analysis exist to monitor their levels in water and soil to make certain that such levels are within acceptable limits. If a crop being sprayed is for animal or human consumption, it is also important to test for any residues remaining in the food produced from such crops. Canola is a source of oil used for cooking, and in 'meal' form it is used in the livestock industry. Recently, it has also been used as a feedstock for the manufacture of biodiesel. In 2012, it was estimated that over 1.5 million hectares of canola was being grown in Australia, making it the third largest agricultural crop in the country.

The analysis of soil and water for pesticide residues is also important for 'organic' farms, which have to be free of pesticide residues and face strict certification guidelines. These often take years to comply with and involve demonstrating that their produce is pesticide free. Analysis of soil and water on such farms must demonstrate the absence of any residues from chemicals that might have been applied in the past.

YOU WILL EXAMINE:

- **how useful products such as insecticides and pesticides can enter water**
- **how unwanted substances such as dioxins are produced and how they can subsequently have an effect on water quality**
- **the principles of chromatography**
- **the use of high-performance liquid chromatography to determine the concentrations of water-soluble organic compounds**
- **oils spills and their causes, effects and treatment.**

> Apply yourself both now and in the next life. Without effort, you cannot be prosperous. Though the land be good, you cannot have an abundant crop without cultivation.
>
> *Plato*

Pesticides are applied to large areas as part of modern agricultural practice. They are needed to control pests that would otherwise reduce the yield and quality of the crops they protect. Here, a canola crop is being sprayed to protect it from insect infestation. It is important that such chemicals do not have unwanted effects on the environment and that they stay where they are applied.

Organic compounds in the environment

From simple molecules such as methane to incredibly complex molecules such as proteins, organic chemistry is all around us. Not surprisingly, organic substances have a wide range of properties. For example, some are very soluble in water while others are virtually insoluble. Many are harmless. Some are critical to life while others are among the most deadly substances known. Many are produced deliberately whereas others are produced as undesired by-products of chemical reactions.

Some organic chemistry revisited

Organic chemistry is the study of carbon compounds. It is generally understood that such compounds contain carbon–hydrogen bonds, thus making hydrogen the second major element that is present. Other elements may also be involved to a lesser degree, such as oxygen, nitrogen, the halogens, sulfur and phosphorus. These can have significant effects on the properties of organic molecules. As expected, the predominant bonding type is covalent molecular.

Naming

The ability of carbon to bond with itself to form single, double and triple bonds, as well as its ability to form cyclic structures, means that carbon can form a huge range of compounds. In chapter 8, we saw how these may be named according to a set of internationally agreed systematic rules.

More complicated organic chemicals have very complicated systematic names. For this reason, it is usual to refer to them by simpler, trivial names.

While every organic molecule can be assigned a formal name using such rules, these names can be extremely long for more complicated molecules. It is therefore common practice to refer to such compounds by either abbreviations or more common, shorter names. This is the case for many of the organic compounds used as part of our everyday lives. For example, the explosive TNT is systematically called trinitrotoluene, while the insecticide diazinon is systematically known as O,O-diethyl O-[4-methyl-6-(propan-2-yl)pyrimidin-2-yl] phosphorothioate!

Solubility in water

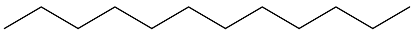
Organic compounds display the full spectrum of solubility in water. Many are extremely soluble while others are virtually insoluble. Many lie somewhere in between. It must be remembered however that, although some substances may be described as insoluble, they still dissolve to a miniscule extent.

Organic substances show a wide range of solubility in water. Some are very soluble while others are virtually insoluble. This is determined by the polarity of their molecules.

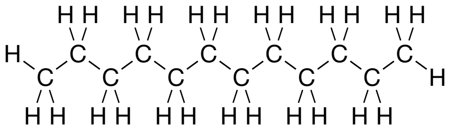
This difference in solubility can be explained by the presence of certain elements such as nitrogen and hydrogen and the shape of the molecule. These effects can combine to make certain molecules polar. Polar molecules dissolve readily in water because water is also polar. On the other hand, molecules containing large regions of carbon–hydrogen bonds are non-polar. These do not dissolve in water but are quite soluble in typical non-polar solvents such as oil and fat deposits. This illustrates the so-called ‘like dissolves like’ rule.

Representing molecular structures

Just as it can be complicated to name organic molecules, it can become complicated to draw them, showing every atom and bond present. A method has therefore been adopted to draw such molecules. The following examples illustrate how this is done.

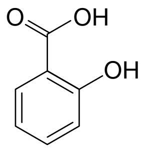


represents



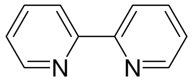
dodecane ($C_{12}H_{26}$)

Using this method, every vertex (and the ends) represents a carbon atom (unless otherwise shown). It is assumed that there are enough hydrogen atoms around these vertices to satisfy carbon's normal covalency of four.



salicylic acid ($C_7H_6O_3$)

Here there are 7 vertices and hence 7 carbon atoms. On three of these, four bonds are readily apparent. The other four appear to have only three bonds. At these locations, it is assumed that a hydrogen atom is present to make the number of bonds up to the four required for normal carbon valency.



2,2′-bipyridine ($C_{10}H_8N_2$)

Here the non-carbon vertices are clearly indicated as nitrogen. It can also be assumed that there are no hydrogen atoms at these locations as the correct number of bonds for nitrogen (3) are already shown.

What are organic compounds used for?

Modern society uses organic compounds for a large variety of purposes. The most obvious of these to the average person would be as fuels. Crude oil is extracted in enormous amounts and then refined to make our familiar transport fuels as well as a wide range of other products. Many of these other products are, in turn, used to make further products. This forms the large and diverse *petrochemical industry*. Other familiar products that rely heavily on organic chemicals in their production include, but are not limited to, plastics and polymers, pharmaceuticals, detergents, solvents, dyes and personal care products.

Another very large sector for organic chemicals is pesticides. From products that can be used at a personal level, such as insect repellent, to products used over millions of hectares in agriculture, this represents an enormous and dynamic sector in which organic chemistry plays a pivotal role. This group can be subdivided according to the specific pest being targeted. We therefore speak of insecticides, herbicides, fungicides, rodenticides and so on. Due to the large areas over which many of these chemicals are applied, there is always a danger that they might end up as contaminants in the soil and in water supplies if they are not applied with care and due diligence.

The American bald eagle was once on the endangered species list. Its reproduction was compromised by the production of thin-shelled eggs that were unable to withstand incubation. This was caused by the birds ingesting DDT.

Persistent organic pollutants

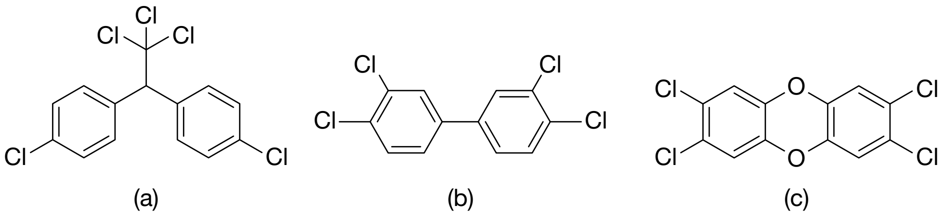
Unfortunately, there exists a range of organic compounds that can contaminate the environment, where they can persist for a long time. They have been referred to as 'hand-me-down poisons' due to their widespread use years ago before their negative properties were fully realised. As a group these are called persistent organic pollutants (POPs).

In 2004, a number of these POPs were recognised as posing a particularly unacceptable risk to the environment and to human health. The Stockholm Convention (of which Australia is a signatory) recognised an initial dozen such chemicals and classes of chemicals. Since then, another nine have been added. This treaty aims to reduce and eliminate, where possible, the use of such chemicals. One of the most well-known chemicals on this list is DDT. DDT was an insecticide used in large amounts throughout the 1940s, 1950s and 1960s to control a number of insects, including the anopheles mosquito, the carrier of malaria. In this role, DDT was hugely successful and the amount used quickly escalated. The publication of Rachael Carlson's famous book *Silent Spring* in 1962, however, drew attention to a number of emerging concerns, including an apparent decline in the reproductive rates of a number of bird species. Birds were producing eggs with very thin shells that broke as the parents sat on them. This in turn seemed to be associated with increased levels of DDT in the body fat of these birds. The use of DDT is now banned in many countries.

Persistent organic pollutants (POPs) are contaminants that are difficult to eradicate from the environment. Many have been used in large amounts in the past. The Stockholm Convention aims to eliminate or restrict the use of such chemicals. An original twelve, plus a further nine chemicals (or classes of chemicals) have been formally recognised by this convention.

Dioxins and PCBs are two further groups of these chemicals. PCBs have very specific uses in the electrical industry and are now being replaced with less dangerous alternatives. Dioxins on the other hand have no use at all but are often formed as unwanted by-products in other processes. Dioxins can even be produced by burning waste in domestic incinerators.

The structures of DDT, a typical PCB and a typical dioxin are shown below.



Structures of (a) DDT, (b) a typical PCB and (c) a typical dioxin

Substances that persist in the environment because they cannot be broken down are called *non-biodegradable*. On the other hand, those that can be broken down by natural means are called *biodegradable*.

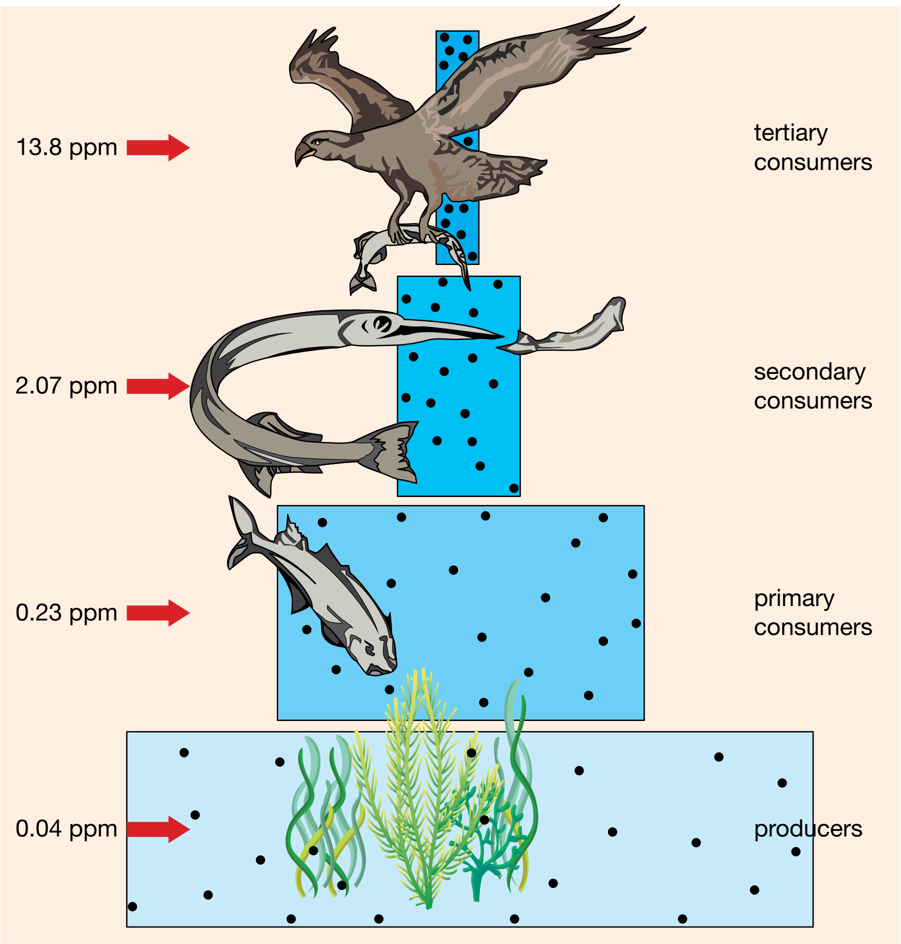
Biomagnification

Biomagnification is the process by which a substance becomes more concentrated in the tissues of organisms higher up the food chain. This results in the concentration of these substances in their bodies becoming higher than in the surrounding environment. Biomagnification can occur when the substance involved is:

- persistent in the environment, even at low levels. This occurs when a substance withstands environmental processes that would normally break it down.
- fat soluble, rather than water soluble. Most substances that become biomagnified are non-polar in nature and therefore are soluble in the typically non-polar fatty tissues of animals. This non-polarity also means that they are water insoluble, a property that contributes to their persistence.
- difficult to eliminate. An animal may not be able to metabolise or excrete something once it has entered its body. If the substance is in the fatty tissues, this is even harder.

Some organic substances can enter the fatty tissues of animals and become almost impossible to eliminate. The concentration of these substances increases as you go up the food chain in a process called biomagnification. At the top of the food chain, concentrations can be considerably higher than in the surrounding environment.

There are many well-documented cases of biomagnification. Mercury, especially in the form of methyl mercury, has already been mentioned in a previous chapter. DDT was another famous example. Many pesticides and their residues are also culprits in this regard.



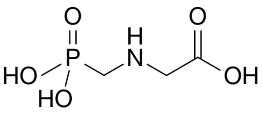
The concentration of DDT in the tissues of animals increases as we move up the food chain. Levels at the top of the food chain can be considerably higher than the background environmental levels. This is why persistent organic pollutants can be so dangerous.

How do organic contaminants enter water?

Organic contaminants may enter water supplies in a number of ways. Soluble organic compounds can dissolve directly in the water. Insoluble organic compounds may adhere to soil particles, which are then swept away by rain.

Organic substances can enter water supplies in a number of ways. If a substance is water soluble, it can dissolve and be transported in run-off water and storm water. A number of modern herbicides are deliberately designed to be water soluble to enable their uptake through the leaves or roots of plants. This is why there is a warning on many such products not to spray if rain is expected within 24 hours; rain would wash the herbicide away before it had penetrated the plant. Glyphosate is a good example of this.

For contaminants that are not water soluble, the main mechanism appears to be adsorption. This can occur on soil particles for example, which then might be washed into rivers and streams and eventually be deposited as silt in these watercourses. Even for insoluble materials, a very tiny amount may still dissolve and this process can go on for a very long time, leading to a small but persistent presence of these compounds in water samples.



Glyphosate — a common ingredient in many brands of herbicide

Fast-flowing rivers can transport soil particles long distances.

Adsorption can also occur in other scenarios as well. It may occur onto the surface of aerosol droplets. It may occur onto dust particles generated by activities such as ploughing fields and sowing seed. Prevailing weather conditions can then transport such adsorbed materials and contaminants long distances from their original sources. This is particularly true for POPs, which have been found in water and soil on every continent — even Antarctica!

Revision questions

1. **Use the structure of glyphosate shown on the previous page to write its molecular formula.**
2. **List the ways that organic chemicals can enter water bodies.**

Analysing for organic substances

As noted in the preceding sections, we often need to analyse water for the presence of organic chemicals. Many of these compounds have quite low solubilities, so any method chosen has to be very sensitive. Also many of these compounds are similar to one another, so any method chosen also has to have the ability to differentiate between them if required. One such method meeting these requirements is chromatography. At one end of the scale, this method can use materials as cheap as filter paper, while, at the other end, sophisticated instruments costing hundreds of thousands of dollars may be used.

studyon

Unit 2
AOS 2
Topic 3
Concept 1

Principles of chromatography
Summary screen and practice questions

Principles of chromatography

You may have noticed how certain food dyes, if mopped up with a damp piece of paper towel, spread out into different colours. You may have done simple experiments where a spot of ink is placed onto a piece of filter paper and water allowed to soak up and move the spot, eventually spreading it out into bands of different colours. These are both examples of chromatography.

Paper chromatography is a simple and inexpensive form of chromatography. The marker pens holding up the paper were used to draw a line across the bottom of the paper strip before immersion in the water. The green ink has been separated into blue and yellow components. The red ink has one single component, and the black ink has many.

Chromatography is the process where a mobile phase (which contains the mixture to be analysed) moves over a stationary phase. As this happens, a process of adsorption and desorption leads to some components of the mixture moving faster over the stationary phase than others.

In all forms of chromatography, a *mobile phase* is moved over a *stationary phase*. The mobile phase contains the substances that we are analysing. The stationary phase contains a material that has a high surface area. As the mobile phase moves over the stationary phase, substances in the mobile phase continually form weak temporary bonds with the surface in a process called *adsorption*. This is followed by the bonds breaking and allowing the substance to move on. This is called *desorption*. These bonds are similar to the weak bonds that form between molecules. Because there are differences in how strongly this occurs, substances move over or through the stationary phase at different speeds and are separated from one another as a result. In the case of the ink or food dyes mentioned on the previous page, the colours that move furthest are those that adsorb and desorb most weakly, while the colours with stronger adsorption and desorption do not move as far.

High-performance liquid chromatography (HPLC)

Adsorption is the process by which molecules adhere, or 'stick,' to the surface of other particles. It commonly involves liquid or gas molecules sticking to the surface of solid particles.

HPLC (originally called high-pressure liquid chromatography) is an adaption of the simple method described above. Invented in the early 1960s, advances in technology have resulted in this now becoming an extremely sensitive and widely used modern technique. Detection of concentrations in parts per million and parts per billion levels is routine. Advanced instruments can now detect even parts per trillion! Applications of HPLC include research, medicine, pharmaceutical science, forensic analysis, food analysis, drug detection in sport and environmental monitoring.

In HPLC, the stationary phase is a narrow diameter tube that is packed tightly with a finely divided powder. This provides the high surface area required for the process and is called a 'column.'

High-performance liquid chromatography (HPLC) is an excellent method for measuring the concentration of organic compounds in a sample.

studyon

Unit 2
AOS 2
Topic 3
Concept 2

High-performance liquid chromatography (HPLC) and calibration curves
Summary screen and practice questions

In HPLC, the separated components are captured after they have been eluted through the column, which is the steel tube on the left of this photo.

Columns containing a wide range of powders are available and can be chosen to give optimal results for a particular scenario. Columns packed with finely divided alumina or silica are common examples. The sample to be analysed

Retention time is the time taken for a substance to travel through the chromatography column.

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HPLC

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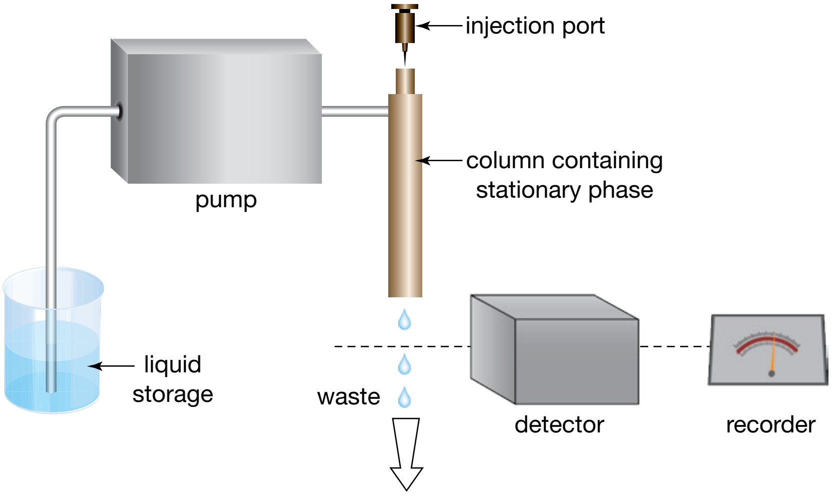
Calibration curves

med-0348

is injected at the top of the column as a liquid, and a further liquid (called the *eluent*) is then pumped through the column at high pressure. This liquid 'flushes' the mixture through the column, allowing adsorption and desorption of components in the mixture to take place. This is the mobile phase. In the same way as paper chromatography, this results in the components of the mixture moving through the column at different rates. Eventually each component reaches the end of the column and exits from it (this is called *elution*). The time taken for this to occur is called the *retention time* (R_t). A detector at this point records retention times in a variety of ways (for example, as a graph on a strip chart (called a *chromatogram*) or on a computer). The diagram below shows the essential components of a HPLC.



The essential components of a high-performance liquid chromatograph

Interpreting the results from HPLC

A common way of displaying the results of HPLC is as a graph showing the detector signal against time. Such a graph is called a chromatogram.

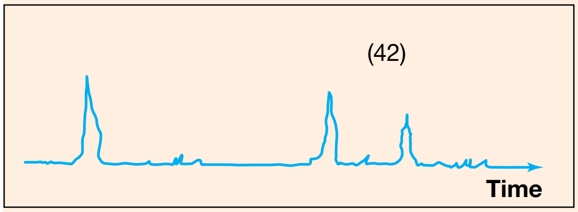
To understand how HPLC can be used to obtain quantitative measurements, let us consider a mixture containing three different substances: A, B and C. These produce three different signals from the detector as they elute from the column. If component C is the one that we are interested in, how can we distinguish its signal from those produced by A and B? The answer is simple. We inject a known sample of C by itself into the column and run it through under identical conditions. By noting retention times, we can then easily identify the correct signal from the mixture. (We assume that substance C takes the same time to move through the column by itself as it would in a mixture containing A and B.)

A calibration curve is a graph of concentration versus instrumental reading. For a HPLC, this reading is usually either peak height or peak area.

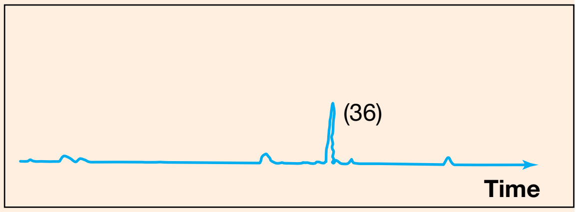
For quantitative analysis, the *size* of the signal is measured in some way. The higher the concentration of the substance causing the signal, the stronger the reading produced. Commonly, this reading is either *peak area* or *peak height*, terms that refer to the peaks produced on the chromatogram. This means that it is possible to produce a calibration curve in the same way as for atomic absorption spectroscopy (AAS). A number of standards (containing substance C) at accurately known concentrations are run through the instrument. The readings obtained are then graphed against concentration to produce a calibration curve. The concentration of substance C can then be easily determined from this graph.

The following diagram illustrates how this process might work.

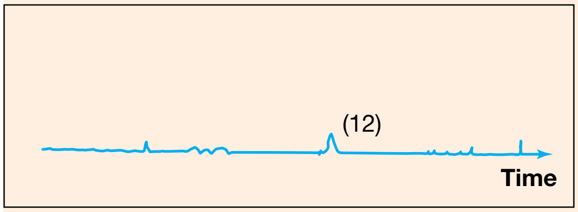
The HPLC process



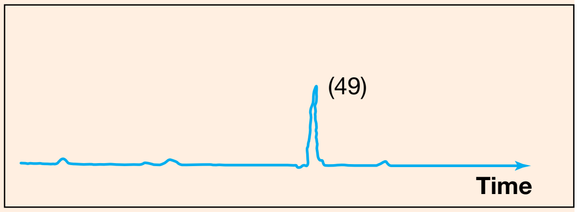
Chromatogram of mixture containing A, B and C (Which peak is which?)



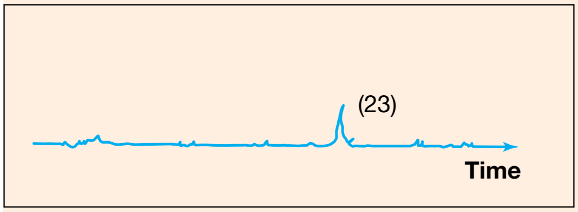
Chromatogram of C at 15 mg L^{-1}



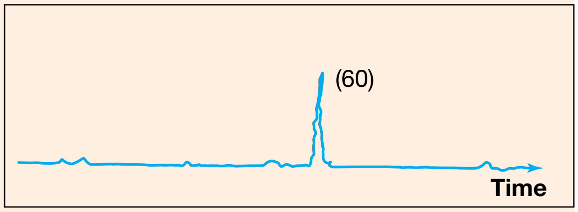
Chromatogram of C at 5 mg L^{-1}
(Peak C is now easily identified.)



Chromatogram of C at 20 mg L^{-1}

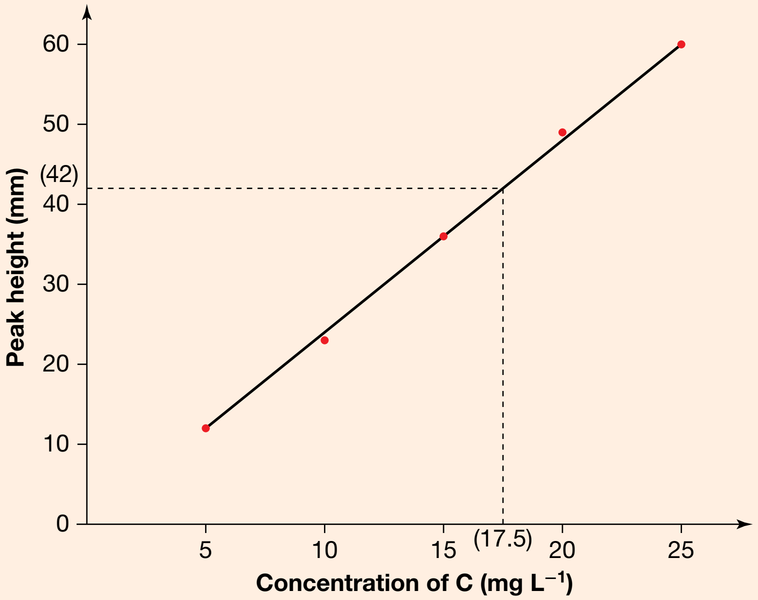


Chromatogram of C at 10 mg L^{-1}



Chromatogram of C at 25 mg L^{-1}

A calibration curve can now be drawn (peak height vs concentration).



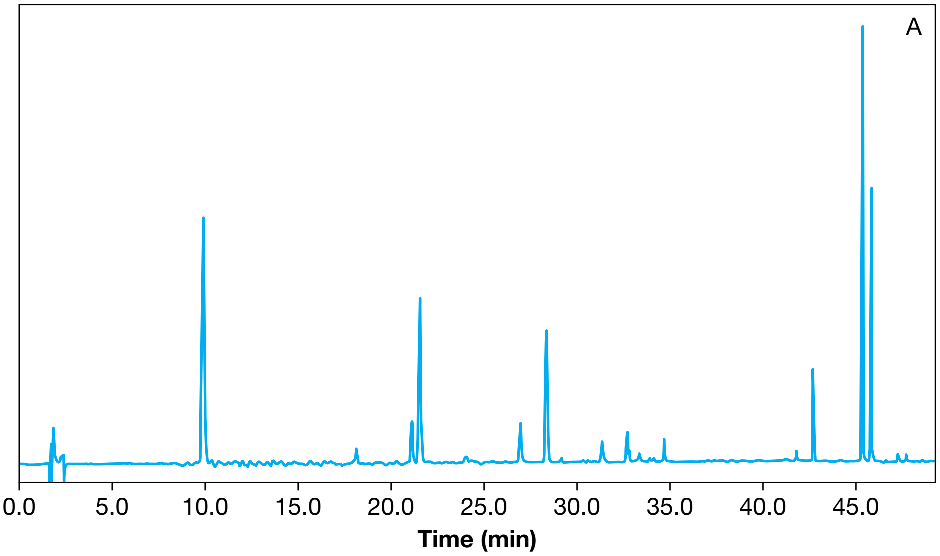
Finally, the peak height from the mixture can be read as a concentration.
Concentration of C = 17.5 mg L^{-1}

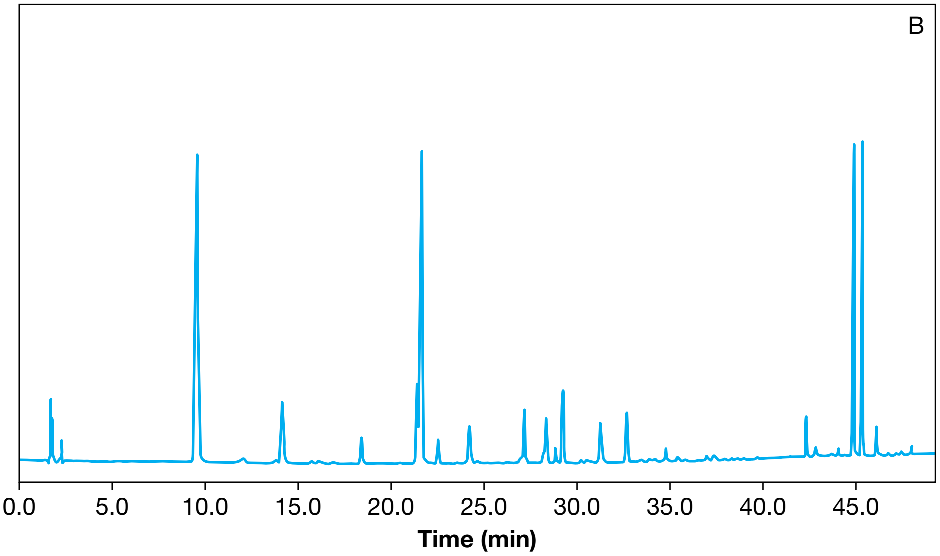
HPLC at work

HPLC is ideally suited to measuring the levels of dissolved organic substances in water. It has the sensitivity to detect very low levels of non-polar species. For more soluble organic substances, the sample is merely diluted first so as not to 'overload' the instrument, and the result is then adjusted to allow for the dilution. The variety of choice available for both the column and the eluent mean that the instrument can be optimised for a particular analysis.

Sample problem 16.1

A sample of lemon oil was suspected of being adulterated. To test this suspicion, HPLC was used to produce chromatograms from a sample of pure oil (A) and from the suspect sample (B). These are shown below.





Has this oil been adulterated?

Solution: Yes. Although the chromatograms are very similar, there are some obvious extra peaks in B. The most obvious is the peak at about 14 minutes.

Sample problem 16.2

Atrazine is a herbicide that is often applied to sugar cane crops to kill weeds. In order to determine the concentration of atrazine in run-off water after rain, representative samples were collected after rainfall over a period of time. These were then analysed using HPLC. The results of one such sample are tabulated below together with results from a set of standards.

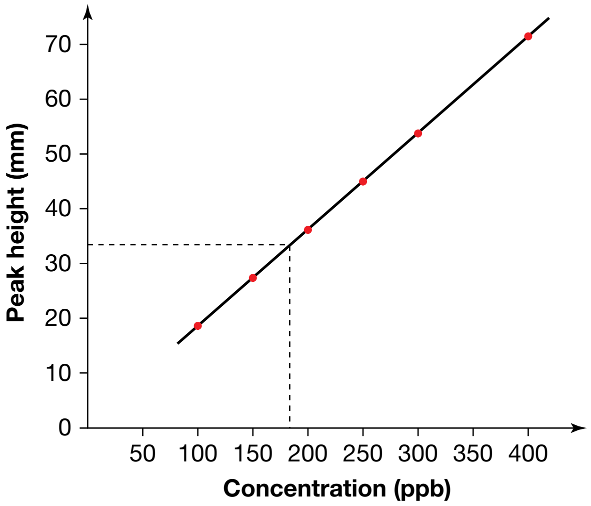
Atrazine standard concentration (ppb)	Peak height
100	21
150	28
200	36
250	45
300	54
400	72
Run-off sample	34

(a) By examining the table above, make a quick estimate of the atrazine level in the run-off sample.

(b) Draw the calibration curve from this data and use it to accurately determine the level of atrazine.

Solution: (a) The reading suggests that the level should be somewhere between 150 and 200 ppb.

(b)



From the calibration curve above, the level of atrazine is 180 ppb.

Revision question

3. **A sample of orange juice was analysed for its ascorbic acid (vitamin C) content. A 100 mL sample was diluted accurately to 1000 mL in a standard flask. After thorough mixing, a 100 mL portion was filtered and prepared for HPLC analysis.**

The results from this analysis, together with those from a set of standards are shown in the table on the next page.

Ascorbic acid concentration (mg L^{-}1)	Peak area
10	1690
15	2420
20	3105
25	3797
30	4504
Sample	3380

(a) Plot the calibration curve for this data.

(b) From part (a), determine the concentration of ascorbic acid in the diluted sample.

(c) Hence, calculate the concentration of ascorbic acid in the undiluted juice sample.

Oil spills

There are many causes of oil spills both on land and at sea. Sometimes, there are natural causes but often they are the result of human activity.

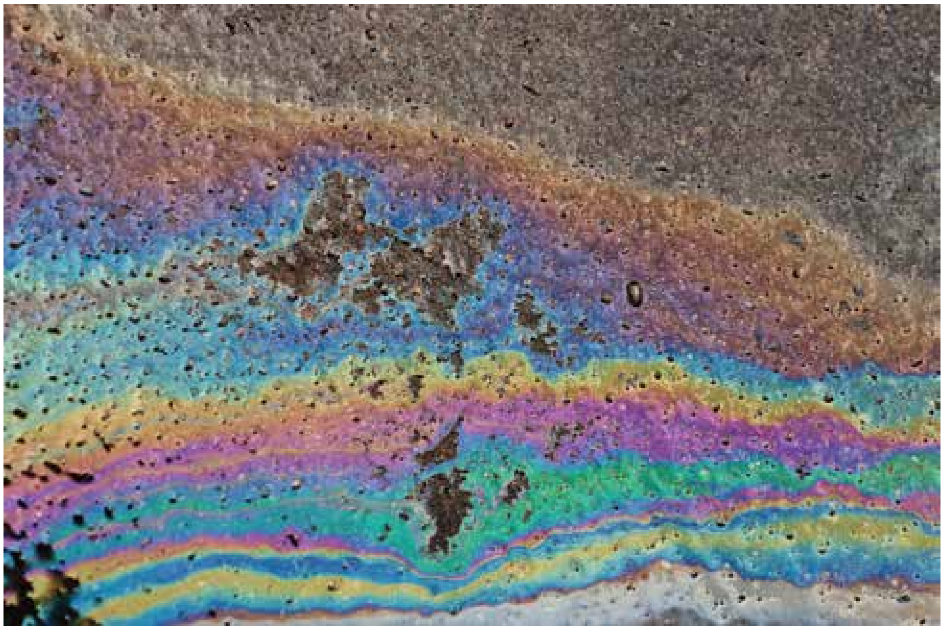
On 20 April 2010, the Deepwater Horizon oil rig exploded while drilling for oil in the Gulf of Mexico. Over the following weeks, it was estimated that as much as 780 000 m^3 of oil was spilled into the Gulf of Mexico, making it the largest spill in American history. In 1989, the tanker *Exxon Valdez* ran aground in Prince William Sound, Alaska and spilled over 40 000 m^3 of oil into a sensitive marine environment. Oil spills such as these make headlines because of their scale and the environmental damage that they cause. The causes of these incidents can be attributed to breakdowns and accidents. Most major oil spills occur at sea, a fact directly attributable to the huge amounts that are transported around the world and to the large amount of offshore drilling today. Additionally, marine spills are difficult to contain and usually affect large areas if left uncontained.

Most oil spills occur at sea due to the huge amounts that are transported around the globe.

However, oil spills occur in many other ways, on land as well as at sea. Often, these represent rather small amounts that happen over a period of time and hence they do not grab the headlines as the two examples above did. Examples of these include:

- Ground seepage. Underground oil deposits make their way to the surface, usually through faults in the underlying rock layers. Once on the surface, they can collect in pools, run off into water courses or evaporate into the atmosphere.

- Road run-off. Oil leaking from vehicles drops onto the road surface. This is then washed away when it rains and is one of the main reasons wet roads are so slippery after a spell of dry weather. This oil can often be seen as a rainbow sheen on the road as it becomes wet.



Road surfaces become coated with oil, which runs off into drains after rain.

- Leakage from pipes and storage.
- Dumping of used oil. It is now illegal to dump used oil (e.g. sump oil from motor vehicles). However, oil from a number of old dumping sites can still leak into the environment today.

Other serious causes of oil spills include natural disasters and deliberate acts. One of the two biggest spills in history occurred in 1991 when oil was deliberately released and set on fire as part of the First Gulf War.

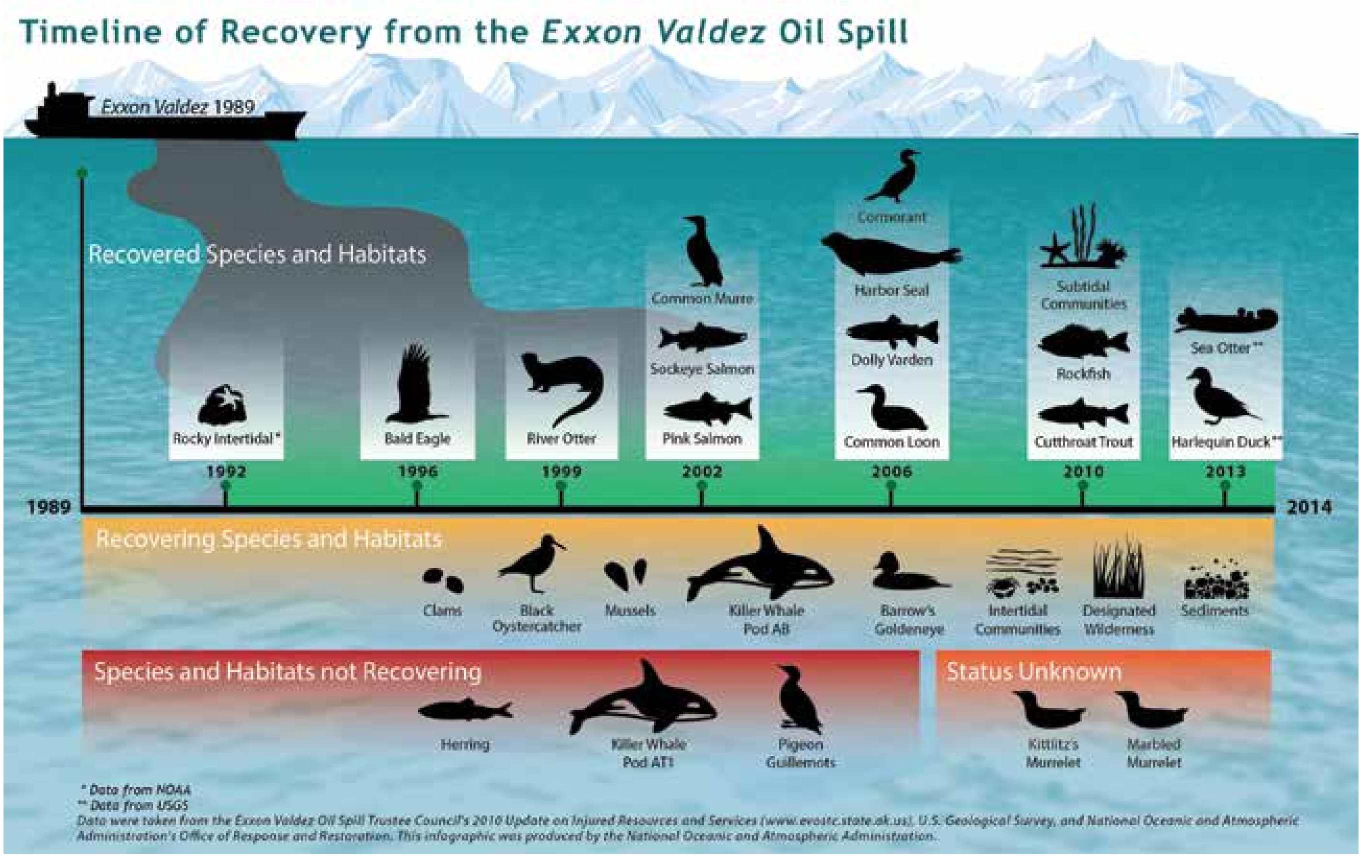
Effects of a marine oil spill

Oil spills have a large number of effects on the physical, biological, human, economic and aesthetic environments.

The effects of an oil spill depend on the type of oil that is spilled and where it occurs. Although nowhere near the biggest in history, the *Exxon Valdez* spill is remembered for the huge impact it had on wildlife in the sensitive area of Alaska that it occurred.

When oil is spilled at sea, the following sequence of events usually occurs. Because oil is a mixture of mostly non-polar hydrocarbons, these do not dissolve in the water but float on it instead. Over time, the oil spreads out, due to its low surface tension, to form a thin layer called a ‘slick’. Slicks are usually very thin but cover a very large area. Some of the more volatile hydrocarbons in the oil evaporate, and the oil left behind tends to form a more viscous sludge. The action of waves and other physical phenomena then break this up and some clumps may even sink, while other portions of the spill may eventually be washed up on nearby land. Over time, sunlight, bacteria and biological processes begin to break down the oil. On land, it may gradually become covered by natural processes such as silting and sand deposition.

These spills are directly harmful to many forms of marine life such as fish, shellfish, birds and mammals. Birds and mammals are affected when their feathers and fur become contaminated by oil and are no longer water repellent. They may drown or suffer hypothermia as a consequence. Ingestion of oil as they try to clean themselves can also be harmful. The effects can be more subtle, however. Affected plants may cause indirect effects further up the food chain. These effects are felt for many years.



* Data from NOAA
** Data from USGS
Data were taken from the Exxon Valdez Oil Spill Trustee Council's 2010 Update on Injured Resources and Services (www.evostc.state.ak.us), U.S. Geological Survey, and National Oceanic and Atmospheric Administration's Office of Response and Restoration. This infographic was produced by the National Oceanic and Atmospheric Administration.

The effects of an oil spill can be widespread in terms of area, species affected and time. On 24 March 1989, the tanker *Exxon Valdez* spilled almost 42 million litres of oil into Alaska's Prince William Sound. This timeline shows the recovery of the 28 types of animal, plants and marine habitats that were seriously affected by the spill.

How are oil spills treated?

Oil spills may be treated in a number of ways. The particular method (or methods) chosen depend on the circumstances of each spill.

No two oil spills are the same. Weather, type of oil, amount spilled, the natural environment and proximity to human habitation all affect how a spill is treated. A response usually involves one or more of the following procedures:

- Leave it alone. This may be considered if the spill is relatively small and a large distance from land.
- Containment. Floating surface barriers may be deployed to trap floating oil within their perimeter. These can be effective in calm conditions but not when waves are large or when they drift up on shore. Devices such as skimmers or special absorbing materials are often used in conjunction to remove the trapped oil.
- Application of dispersants. These surfactants help to disperse the oil into droplets and encourage it to mix with water. The increased surface area that results can also help the natural degradation processes of the oil.
- Accelerated biodegradation. Bacteria and other microorganisms that naturally break down the oil into fatty acids and carbon dioxide can be added and encouraged to proliferate by adding nutrients. This is sometimes used where oil washes up on land.
- Burning. Sometimes the oil is deliberately ignited, preferably close to the source of the spill and before more volatile components have evaporated.

Floating barriers such as these booms are often used to help contain an oil spill.

Chapter review

studyon

Unit 2 **Analysing water for organic compounds**
AOS 2
Topic 3 Summary screens and practice questions

Summary

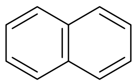
- Organic compounds are used in many situations and in large amounts as part of modern living.
- Due to their widespread use, organic compounds may find their way into water supplies and other bodies of water.
- Carbon compounds get into water supplies in a number of ways. Two common methods are by dissolving (sometimes in only small amounts) and by adhering to soil particles, which are subsequently washed into rivers and lakes.
- These compounds can be classified as biodegradable or non-biodegradable. Non-biodegradable substances may persist in the environment for an unacceptably long time.
- A group of 21 substances or classes of substances have been identified as persistent organic pollutants (POPs). The use of these substances is now banned or severely restricted in many countries.
- DDT is an example of a chemical that was applied extensively as an insecticide but is now on the list of POPs.
- Biomagnification is the process by which levels of POPs build up in the fatty tissues of higher order organisms.
- High-performance liquid chromatography (HPLC) is a technique ideally suited to analysing the concentration of organic substances in water. It is both versatile and sensitive enough to detect the low concentrations that are sometimes present.
- Like all forms of chromatography, HPLC involves a mobile phase moving over a stationary phase. The process of adsorption and desorption that takes place results in components of a mixture moving at different rates over the stationary phase.
- Retention time is the time taken for a substance to move through a chromatography column.
- An important step in using HPLC to determine unknown concentrations is the production of a calibration curve. This involves running a set of standards of accurately known concentration through the instrument. The reading produced for each standard is noted and the results used to draw a graph of concentration against reading. This is the calibration curve.
- Oils spills can occur on land as well as at sea. Spills at sea are more mobile and can affect large areas.
- Oil spills at sea can be treated by a variety of methods, either singly or in combination. These methods include leaving it alone, containment, skimming, absorbing, using dispersants, encouraging biodegradation and burning.

Multiple choice questions

1. Which of the following is an organic substance?
- **A** H_2O
- **B** $C_{12}H_{24}$
- **C** CO_2
- **D** $CuSO_4.5H_2O$

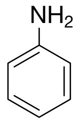
Questions 2 and 3 refer to the four compounds shown below. For ease of reference, these have been labelled A, B, C and D.

A



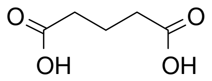
naphthalene

B



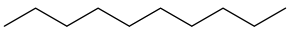
aniline

C



glutaric acid

D



decane

2. Which two would be expected to be water soluble?
- **A** A and B
- **B** B and C
- **C** C and D
- **D** D and A

3. Which two would be expected to be insoluble in water?
- **A** A and B
- **B** B and C
- **C** C and D
- **D** D and A

4. The use and phasing out of persistent organic pollutants is governed by the:
- **A** Geneva Convention
- **B** Stockholm Convention
- **C** Helsinki Convention
- **D** Dublin Convention.

5. Which of the following statements about dioxins is true?
- **A** Dioxins have been used in the past as insecticides.
- **B** Dioxins are inorganic chemicals containing heavy metals.
- **C** Dioxins are restricted to areas where there are large chemical industries.
- **D** Dioxins are unwanted by-products that have no useful applications.

6. Biomagnification is the process whereby the concentration of a pollutant is increased:
- **A** as it is exposed to sunlight and microbial action
- **B** in organisms in each step up a food chain
- **C** as remains undergo decomposition
- **D** due to a population explosion of a particular species.

7. Despite their low solubility in water, many organic contaminants can persist in water bodies for a long time. A reason for this is:
- **A** they adsorb strongly to the surface of soil particles in river beds
- **B** other liquid impurities in the water dissolve them
- **C** they are broken down by natural microbial action in the soil
- **D** they accumulate in the bodies of aquatic organisms, which then die.

8. All forms of chromatography involve a:
- **A** transitory phase moving over a mobile phase
- **B** transitory phase moving over a stationary phase
- **C** mobile phase moving over a transitory phase
- **D** mobile phase moving over a stationary phase.

9. In the quantitative use of a HPLC, which of the following would *not* be done?
- **A** Measuring the area of the peaks produced
- **B** Drawing a graph of concentration versus time
- **C** Drawing a graph of concentration versus peak area
- **D** Standardising the instrument's readings against known concentrations.

10. When preparing a calibration curve from HPLC data, which of the following would *not* be appropriate to use?
- **A** Concentration
- **B** Retention time
- **C** Peak height
- **D** Peak area

11. An oil spill at sea would *not* be treated by:
- **A** containment
- **B** dispersants
- **C** burning
- **D** neutralisation.

12. Which of the following would be the most likely route by which significant land-based oil pollution might enter oceans?
- **A** By adhering to dust particles that are subsequently deposited at sea
- **B** By run-off from urban areas into drains and rivers and subsequently to the sea
- **C** By soaking into the ground and slowly draining to the sea
- **D** From washing off tyres of cars and boat trailers at launching ramps

Review questions

Organic structure and properties

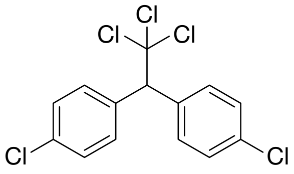
1. (a) Define the term 'organic chemical.'
(b) Which two elements are the most common in most organic chemicals?
(c) What other elements are often present in smaller amounts?

2. Many organic chemicals are soluble in water while many others are insoluble. What features in their molecular structures account for this difference?

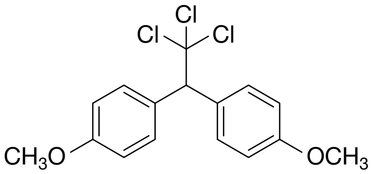
3. An examination of the formula of a complicated organic molecule reveals that it contains carbon, hydrogen, iodine and sulfur. Is this molecule likely to be water soluble? Explain.

4. In applications of organic chemistry, many compounds are known by abbreviations or by more common names, rather than by their systematic names, as studied in chapter 9. Examples of this include glyphosate and TNT. Give a simple reason for this.

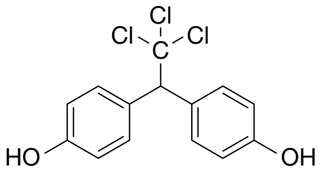
5. A problem with DDT is that it is not biodegradable. Because of this, it builds up in the fatty tissues of organisms and can reach quite high levels in organisms higher up food chains. Scientists have attempted to overcome this problem by producing variants that are still persistent but can be metabolised by animals into water-soluble products. One such insecticide is methoxychlor. The diagram below shows the structures of DDT, methoxychlor and a typical metabolite of methoxychlor.



(a) DDT



(b) methoxychlor



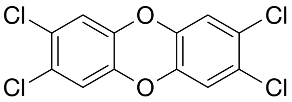
(c) a typical metabolite of methoxychlor

(a) Why is it desirable that an insecticide be persistent in the environment, at least to a reasonable degree?
(b) Explain why the product of methoxychlor metabolism shown above would be water soluble and therefore less likely to build up in fatty tissue.

Persistent organic pollutants

6. Use a suitable resource to research the history and use of DDT. Report your findings as a poster, PowerPoint presentation or some other visual means. Your report should include reasons why its use today is either banned or severely restricted.

7. Dioxins are a class of persistent organic pollutants. Shown below is the structure of a typical dioxin.



(a) Write the molecular formula for this dioxin.
(b) How are dioxins formed?
(c) Explain why the expected concentration of dioxins in water would be very low.
(d) Give two reasons why dioxins accumulate in human fatty tissue.

Chromatography

8. Define the following terms as they apply to chromatography.
(a) Stationary phase
(b) Mobile phase
(c) Chromatogram
(d) Retention time

9. A student used paper chromatography to produce a chromatogram from black ink, as shown below. The ink was added to the paper near the bottom, and water was used as the mobile phase.

(a) List the colours shown in order from most strongly adsorbed to least strongly adsorbed.
(b) Predict what would happen if a longer piece of paper was used.

10. A sample of supposedly pure compound A is suspected of being contaminated with a small amount of compound B. To check this claim, a small amount is dissolved and then analysed using HPLC.
(a) Draw a possible chromatogram produced by this sample, paying particular attention to the heights of the peaks.
(b) There is a small chance that only one peak might be obtained.
Suggest a possible reason for this.

Quantitative HPLC

11. Explain why a set of standards do not need to be evenly spaced to produce a calibration curve.

12. An analyst is designing a procedure to use HPLC to check the level of a water-soluble chemical in a common herbicide. The claim being tested is that the level is 120 g L^{-1}. The configuration of the HPLC requires that the concentration of this chemical be between 1 and 10 g L^{-1} for reliable analysis.

Using the product claim as a guide and assuming that a 100 mL standard flask is to be used, calculate the volume of herbicide that would need to be poured into the flask and then diluted so that a final concentration of 5 g L^{-1} is achieved.

13. The level of glucose, $C_6H_{12}O_6$, in a sample of human blood was analysed by HPLC. After suitable extraction, the sample was analysed, along with a set of known standards. The results are shown in the table below.

Glucose concentration (g L^{-1})	Peak area
0.4	1650
0.8	3280
1.0	4100
1.2	4920
1.6	6550
Sample	3840

(a) Draw the calibration curve using the data above.
(b) What is the concentration of glucose in the blood sample (in g L^{-1})?

The normal concentration of glucose in human blood should be 4.4 to 6.1 millimoles per litre or millimolar (mmol L^{-1} or mM).

(c) Convert your answer from part (b) into mM.
(d) Is the glucose within the normal range?

14. 'Snuffem' is a water-soluble insecticide sold for domestic use to kill ants and spiders. Its label claims that its active ingredient is 'taylachlor', at a level of 2.5 mg per 100 g.

To test this claim, a 10.045 g sample was dissolved in 100 mL of ethanol and filtered to remove the residue. The volume was then reduced to 13.5 mL by gentle heating over a water bath so as not to destroy the 'taylachlor'. Standards containing 10, 15, 20 and 30 mg L^{-1} were prepared. These standards and the condensed sample were then subjected to HPLC and gave the following results.

Concentration of 'taylachlor' (mg L^{-1})	Peak area
10	200
15	290
20	380
30	560
Concentrated sample	310

(a) Determine the concentration of 'taylachlor' in the sample as tested.
(b) Calculate the total mass of 'taylachlor' in the concentrated sample.
(c) Is the mass of 'taylachlor' in the concentrated sample the same as in the original 100 mL sample? Explain.
(d) Calculate the determined level of 'taylachlor' in 'Snuffem' (as mg per 100 g). Comment on how this compares to the claim on the label.

Oil spills

15. Give three reasons an oil spill may occur:
(a) at sea
(b) on land.

16. Crude oil can be classified in a number of ways. For example, *light crude oil* is crude oil that has a high proportion of volatile components. Consequently, it evaporates relatively quickly. On the other hand, *heavy crude oil* contains a high proportion of less volatile substances.

Suggest a possible response scenario to:
(a) a small spill of light crude oil close to land
(b) a small spill of heavy crude oil close to land
(c) a large spill of light crude oil away from land.

17. Use a suitable resource to research the history and effects of a large oil spill. Report your findings in a format of your choice. Your report should cover the history of the incident, its size, its effects, how it was treated and how the affected area has rehabilitated.

Exam practice questions

studyon

Unit 2
AOS 2
Topic 3

Analysing water for organic compounds

Sit topic test

In a chemistry examination you will be required to write a number of short and extended response questions.

1. Caffeine is widely used in a number of pharmaceutical products.

HPLC was used to determine the caffeine content in a particular brand of tablet. One tablet, weighing 0.103 g, was crushed and dissolved in exactly 10 mL of ethanol. This was subsequently diluted to 1000 mL in a standard flask. Five caffeine standards were also prepared as shown in the table below.

All six solutions were subjected to HPLC under identical conditions. The results are shown in the table.

Concentration of caffeine (mg L^{-1})	Peak area
5	370
20	1340
40	2620
60	3910
80	5200
Test sample	3400

(a) Plot the calibration curve for this data. **5 marks**

(b) Determine the concentration of caffeine in the diluted sample as tested. **1 mark**

(c) Calculate the concentration of caffeine in the ethanol used for extraction. **2 marks**

(d) What was the mass of caffeine in the tablet that was tested? Express your answer in milligrams. **1 mark**

2. Insecticides can move from where they are applied into other non-intended areas of the environment by a number of mechanisms. Water plays an important role in two of these. Firstly, water can dissolve an insecticide and carry it away in solution as run-off. Secondly, run-off water can carry away soil particles whose surfaces contain adhered insecticide molecules.

The potential of an insecticide to move 'off site' can be predicted by three parameters. These are its *solubility*, the strength of its *adherence* to particles in the soil and *how long* it persists before it is broken down into other products. Adherence is measured by K_{oc} values. The higher the value, the more strongly the insecticide is bound to soil particles. Field half-life (FHL) is an indication of how long the insecticide persists after being applied.

In general, an insecticide with a low FHL is unlikely to move. A chemical with a higher FHL may move in solution, by adsorption to soil particles or both.

The following table shows these parameters for three insecticides.

Insecticide	Solubility (mg L^{-1})	K_{oc}	FHL (days)
diazinon, $C_{12}H_{21}N_2O_3PS$	60	1520	40
permethrin, $C_{21}H_{20}Cl_2O_3$	0.006	10^6	42
oxamyl, $C_7H_{13}N_3O_3S$	2.8×10^5	25	4

(a) Explain why diazinon is considered to have high potential for both solution and adsorption run-off. **3 marks**

(b) Despite its high solubility, oxamyl is considered to have low potential to move. Explain why this is so. **1 mark**

(c) Explain what is likely to happen to a field sprayed with permethrin if heavy torrential rain is experienced within the first few weeks after application. **3 marks**

(d) If the field in part (c) experiences gentle misty rain instead, is there likely to be a problem? Explain. **2 marks**

CHAPTER

17 Analysing water for acids and bases

> Anger is an acid that can do more harm to the vessel in which it is stored than to anything on which it is poured.
>
> *Mark Twain*

YOU WILL EXAMINE:

- how some of the substances that dissolve in water can act as acids and bases
- the potential sources of these acids and bases, both natural and as the result of human activity
- alkalinity as a measure of a water body's capacity to withstand sudden pH changes caused by acid input
- how stoichiometry may be applied to reactions in solution
- the dilution of solutions and calculations associated with this process
- the use of acid–base titrations to analyse the concentration of acids and bases in water
- how standard solutions can be prepared
- the range of indicators available for acid–base titrations.

To determine the concentration of an acid or a base in solution, a technique called titration is often used. One of the pieces of equipment used is a burette, which delivers an accurate amount of solution.

Isn't water neutral?

The answer to this question is yes — but also no!

While pure water has a pH of 7, normal surface water can have a pH anywhere between 6.5 and 8. For ground water (underground water), this can be as low as 6 and still be considered normal, while sea water has a pH of 8.1. As we have already learned, water is an excellent solvent; therefore, virtually all the Earth's water contains dissolved substances. Some of these are dissolved in quantity while others only sparingly. It is this range of solutes that alters the pH of the water.

How is water made acidic?

Water can be made acidic by familiar inorganic acids such as hydrochloric, sulfuric and nitric acids, which can enter the water naturally and as contaminants. Natural sources such as volcanoes and lightning flashes produce small amounts of sulfuric and nitric acids and additional generation of these two acids as a result of human activity is the main cause of acid rain. Even so-called 'unpolluted' rain is slightly acidic due to dissolved carbon dioxide from the atmosphere, making the weak acid, carbonic acid, H_2CO_3.

$$CO_2(g) + H_2O(l) \longrightarrow H_2CO_3(aq)$$

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

In addition to the above sources, large numbers of salts can make water acidic. The expanded definition of acids and bases by Brønsted and Lowry explains how such salts can act as acids, thus affecting the pH of water. Often, such salts contain cations, which are the conjugate acids of weak bases. Ammonium salts containing NH_4^+ ions are an example of this. In other cases, the anion may be amphiprotic but may react in water more easily as an acid than a base. For example, hydrogen sulfate ions can react in two ways:

as an acid: $HSO_4^-(aq) + H_2O(l) \rightleftharpoons SO_4^{2-}(aq) + H_3O^+(aq)$
as a base: $HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_2SO_4(aq) + OH^-(aq)$

Because the acid reaction is 'easier' than the base reaction, salts containing hydrogen sulfate ions tend to make water acidic.

Water is made acidic when natural acids or acids that are generated by human activity dissolve in it.

Certain metal cations can also make water acidic due to the behaviour of water molecules surrounding them. Aluminium ions and ions of transition metals such as iron, nickel and copper are good examples of these.

Finally, there is a wide range of organic materials, both natural and contaminants, that can act to make water acidic.

How is water made basic?

Many salts contain anions that react as bases in water. Often, such anions are the conjugate bases of weak acids. The carbonate ion, CO_3^{2-}, which is the conjugate base of the weakly acidic hydrogen carbonate ion, HCO_3^-, is a good example of this. Dissolved salts such as sodium carbonate and calcium carbonate therefore tend to make water basic.

Amphiprotic anions may play a role here as well. Hydrogen carbonate ions illustrate this point well. They can react with water in two possible ways:

as an acid: $HCO_3^-(aq) + H_2O(l) \rightleftharpoons CO_3^{2-}(aq) + H_3O^+(aq)$
as a base: $HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$

In this case, the base reaction is 'easier' than the acid reaction, so the net effect is to make the solution basic.

Water is made basic when natural bases or bases that are generated by human activity dissolve in it.

Many organic acids are weak acids. If these have a chance to react to form their conjugate bases, dissolving the products of such reactions in water also contributes to raising the pH.

A knowledge of how to adjust pH levels is critical for effective swimming pool management.

Alkalinity

Alkalinity is a measure of how well a body of water is able to withstand acid pollution.

In the environment, such as in a stream, it is important that changes in pH are not too large and/or happen too quickly. The pH of stream water is the net effect of all the substances that are dissolved in it. However, sometimes the addition of an acidic contaminant does not affect the pH as much as might be expected. This is because some of solutes in the water can act to resist changes in pH. Such substances are called **buffers**. Calcium carbonate and calcium hydrogen carbonate are examples of such substances. If an acid enters water containing these ions, it reacts with them and is neutralised. It is only after all the ions have been removed that any subsequent acid is free to significantly affect the pH. The higher the level of buffers in the water, the more acid pollution the stream is able to withstand.

Sources of acids and bases found in water

Many chemicals that occur naturally and many made by human activity can make water acidic. Three natural substances are sulfuric, nitric and carbonic acids. Sulfuric and nitric acids are also made in extra amounts by human activity. Mining activities and industrial waste also contribute. Many rocks and minerals also contain acidic material that dissolves in water as it flows over them.

The acids and bases found in water may come from a number of sources, both natural and as a result of human activity. They may be regarded as normal or contaminants depending on the circumstances.

Some of these sources include:

- natural sources such as volcanoes and lightning. These can produce sulfuric and nitric acids. Carbonic acid is formed as water dissolves carbon dioxide from the atmosphere.

Volcanoes are a natural source of nitric and sulfuric acids.

- rocks and minerals. Many minerals can act as acids or bases when they dissolve in water. Depending on the geology of an area, surface water and ground water may be naturally acidic or naturally basic because of this.
- decomposition of plant and animal materials, which can produce acids that dissolve in water
- burning coal and many other fuels, as well as the internal combustion engine, which produce sulfuric and nitric acids. Many fuels contain sulfur as an impurity, and the sulfur dioxide produced can ultimately form sulfuric acid. In car engines, the high temperatures produced combine nitrogen and oxygen from the air to form various nitrogen oxides. On exposure to water, these form nitric acid.
- waste water from mines and tailing sites. When sulfide ores and coal are mined, the sulfur they contain comes into contact with oxygen and water. Oxidation of the sulfur, followed by dissolution in water produces sulfuric acid. An additional problem is that this acidified water may then leach unwanted heavy metals from ores and rocks to cause further contamination.
- industry. Examples include the iron and steel industries where processes such as galvanising and electroplating use large amounts of both hydrochloric and sulfuric acids. Acids are also used to make dyes and tan leather. Bases such as sodium hydroxide are used to refine petroleum and to make soap. Organic acids are used to make detergents, and organic bases are used to make pharmaceuticals and various herbicides. In every case, careful attention to detail must be followed to ensure that the release of waste water from such activities meets acceptable EPA guidelines.

Water can be made basic by chemicals present in many rocks and minerals that dissolve as the water flows over them. Carbonates are typical examples of this. Pollution from industrial sources can be another source of bases in water.

Reasons for analysing

Water needs to be at, or at least close to, an optimum pH for many of its uses. Knowing how much acid or base in is the water is important when its pH needs to be adjusted. Avoiding one problem could well cause another if the water is 'overadjusted.'

Problems such as metal corrosion in plant equipment and the ability of fish to survive and reproduce are typical problems where water has become too acidic. The deposition of scale in pipes and cloudiness in water are typical problems caused by basic conditions. Owners of backyard swimming pools know that the effectiveness of their chlorine in sanitising the water depends on pH. If it is too high or too low, the chlorine is not as effective. The same applies to chlorine added to domestic drinking supplies. Many plants have optimum pH ranges for the soil in which they grow, which can be altered by the water used to irrigate them. There are cases where the opposite applies — where undesirable growth occurs because an ideal pH environment has been inadvertently created. Mosquito larvae live within a pH range of 3.3 to 4.7, and a pH range of 7.5 to 8.4 is optimal for algal growth. Adjusting pH values away from these ranges may solve these issues. In all of these cases, an analysis of the water involved is a key first step.

Another effect of incorrect pH can be *synergy*. This is where pH combines with one or more other factors to have an effect that is greater than the individual effects combined. For example, the presence of metals such as iron, aluminium, lead and mercury in water at the wrong pH can make the metals much more toxic than normally expected. Fish that can tolerate pH values as low as 4.8, for example, die at the more moderate pH of 5.5 if the iron content of the water is a mere 0.9 ppm.

Scale inside steam pipes can be avoided by altering the pH of the water.

Analysing for acids and bases

Acids and bases are usually analysed using *volumetric analysis*. As the name suggests, volumetric analysis uses measured volumes involving solutions. If one of the solutions is of accurately known concentration, the concentration of a second solution may be derived using stoichiometry. The more accurately that volumes and known concentrations are determined, the more accurate is the derived concentration. Great care is therefore taken in how this is done.

Volumetric analysis, which involves solutions of accurately known concentration and careful measurement of volumes, is a convenient method for analysing acids and bases in water.

Although *volumetric analysis* is the technique of choice for many acid–base determinations, it can also be applied to other types of reactions that occur in solution. For example, many redox reactions are analysed using this method.

Before we look at the specifics of how volumetric analysis is performed, it is first necessary for us to examine how stoichiometry can be applied to reactions in solution.

Solution stoichiometry

When some solutions are mixed, chemical reactions may occur. Solution stoichiometry, sometimes also called 'concentration–concentration' stoichiometry, involves reactions in solution such as precipitation reactions and the neutralisation of an acid and a base to form water and a salt. Solution stoichiometry uses the formula:

$$n = cV$$

Solution stoichiometry also uses the same steps as for mass–mass stoichiometry. The difference is that the formula $n = cV$ is used for both the given and the required substances.

to calculate moles of solute in solutions that may react or that may be produced. Solution stoichiometry problems can be solved in four steps:

1. Write a balanced chemical equation for the reaction, identifying the known (given) and unknown (required) quantities of substance.
2. Calculate the number of moles of the known quantity of substance present, using the formula $n = cV$.
3. From the equation, find the molar ratio of known to unknown quantities in the reaction and use it to calculate the number of moles of the required substance.
4. Calculate the quantity (i.e. concentration) of the required substance.

Sample problem 17.1

A solution of cloudy ammonia is analysed for its ammonium hydroxide concentration. A 20.00 mL sample is neutralised with 30.00 mL of 1.10 M hydrochloric acid. Find the molarity or concentration of the ammonium hydroxide in the cloudy ammonia.

Solution:

STEP 1

Write the equation, identifying the known and unknown quantities of substance.

$$NH_4OH(aq) + HCl(aq) \longrightarrow NH_4Cl(aq) + H_2O(l)$$

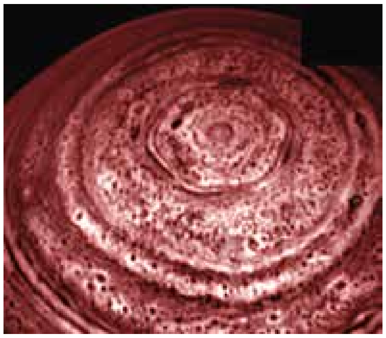
$$c_1 = ?\ M \qquad c_2 = 1.10\ M$$

$$V_1 = 0.020\,00\ L \qquad V_2 = 0.030\,00\ L$$

STEP 2

Calculate the number of moles of the known quantity of HCl.

$$\begin{aligned} n(HCl) &= c_2 \times V_2 \\ &= 1.10 \times 0.030\,00 \\ &= 0.0330\ mol \end{aligned}$$



Clouds of condensed ammonia hide the surface of Saturn. On Earth, ammonia exists in gaseous form. Ammonia gas dissolved in water is called ammonia; it is a common ingredient in many kitchen and bathroom cleaning agents.

STEP 3

Find the molar ratio from the equation and use it to calculate the number of moles of the required substance.

$$n(\mathrm{NH_4OH}) : n(\mathrm{HCl}) = 1 : 1$$

$$n(\mathrm{NH_4OH}) = n(\mathrm{HCl})$$

STEP 4

Calculate the quantity (concentration) of the required substance.

$$c(\mathrm{NH_4OH}) = \frac{n}{V} = \frac{0.0330}{0.02000}$$

$$= 1.65\ \mathrm{M}$$

Revision questions

1. **In an experiment to determine the concentration of a solution of hydrochloric acid, 15.5 mL of the HCl was neutralised by 25.0 mL of a 0.055 M solution of Na_2CO_3 in a conical flask.**
   - **(a) Write a balanced equation for the reaction.**
   - **(b) Determine the concentration of the hydrochloric acid.**
2. **What volume of a solution of 0.200 M potassium hydroxide, KOH, exactly neutralises 20.0 mL of a solution of 0.300 M hydrochloric acid?**
3. **What volume of a solution of 0.200 M potassium hydroxide, KOH, exactly neutralises 20.0 mL of a solution of 0.300 M sulfuric acid, H_2SO_4?**
4. **What volume of a solution of 0.200 M potassium hydroxide, KOH, exactly neutralises 20.0 mL of a solution of 0.300 M phosphoric acid, H_3PO_4?**
5. **The salinity of sea water can be measured by adding silver nitrate solution until all the chloride ions have been precipitated as silver chloride according to the equation:**

   $$\mathrm{Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)}$$

   A 20.00 mL sample of sea water required 22.40 mL of 0.500 M silver nitrate solution to precipitate all the chloride ions from solution. Calculate the molarity of the sodium chloride in the sample.

Dilution

When a solution is diluted by the addition of more solvent (e.g. water), the number of moles of solute remains the same. The addition of water to a concentrated solution does not alter the number of moles or the mass of the solute in that solution.

After a solution has been diluted, the number of moles of solute remains the same. So, if n_1 represents the number of moles of the initial or concentrated solution and n_2 represents the number of moles of the final or dilute solution, then we can say that:

Dilution of a solution changes its concentration. As the volume increases with dilution, the concentration decreases: $c_1V_1 = c_2V_2$.

$$n_1 = n_2$$

Using the equation $n = c \times V$, we can write equations for n_1 and n_2 as follows:

$$n_1 = c_1 \times V_1 \text{ and } n_2 = c_2 \times V_2$$

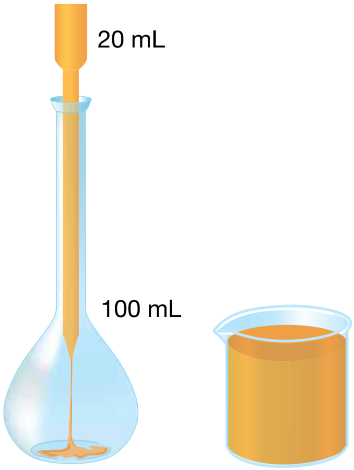
Since the values of n_1 and n_2 are equal, these equations can be combined to form the following formula:

$$c_1 \times V_1 = c_2 \times V_2$$

The following diagrams show how a 0.20 M $K_2Cr_2O_7$ solution is diluted to 0.04 M.

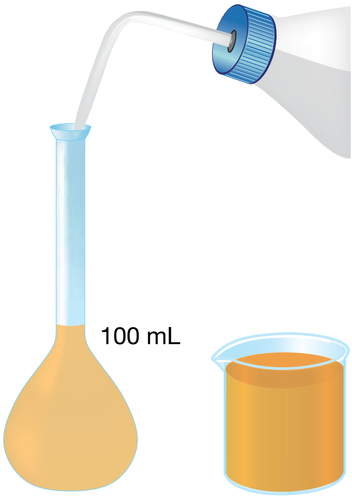
STEP 1

Measure 20 mL of 0.20 M $K_2Cr_2O_7$ into a pipette. Slowly add the $K_2Cr_2O_7$ to a calibrated 100 mL flask.



STEP 2

Rinse the pipette with water and also pour it into the calibrated 100 mL flask.



STEP 3

Stopper the flask and mix thoroughly, then add more H_2O to the 100 mL mark. The original 0.200 M $K_2Cr_2O_7$ has been diluted by a factor of 5, so the concentration decreases by a factor of 5. You now have 100 mL of 0.0400 M $K_2Cr_2O_7$.

Sample problem 17.2

What volume of 10 M stock hydrochloric acid is needed to prepare 250 mL of a 2.0 M hydrochloric acid solution?

Solution: **STEP 1**

Let c_1 and V_1 represent the quantities of concentrated solution and c_2 and V_2 represent the quantities of dilute solution.

$V_1 = ?\ \text{L}$ $\quad V_2 = 0.250\ \text{L}$

$c_1 = 10.0\ \text{M}$ $\quad c_2 = 2.0\ \text{M}$

STEP 2

Calculate the quantity of 10 M HCl needed to produce the required dilution.

$$\begin{aligned} c_1V_1 &= c_2V_2 \\ 10.0 \times V_1 &= 2.0 \times 0.250 \\ V_1 &= \frac{2.0 \times 0.250}{10.0} \\ &= 0.050\ \text{L} \\ &= 50\ \text{mL} \end{aligned}$$

Revision questions

6. **Calculate the volume of stock solution needed to prepare 250 mL of a 2.15 M solution of hydrochloric acid (hydrochloric acid stock solution = 12 M).**
7. **Calculate the concentration (molarity) of 2.0 L of 2.0 M HCl solution after dilution with 500 mL of water.**

Acid–base titrations

Acid–base titration is an experimental method that uses specialised equipment to determine the concentration of a solution. A requirement is a solution of known concentration with which to react the solution of unknown concentration. Solutions whose concentrations are known accurately are called standard solutions.

Titration is the experimental technique used for volumetric analysis.

Acid–base **titrations** are based on neutralisation reactions. They are a type of volumetric analysis where the unknown concentration of a solution is determined by reacting it with a solution of known concentration. The solution of known concentration is called a **standard solution** and is prepared by dissolving an accurate amount of solute in water using a volumetric flask that is calibrated to contain the specified volume. The standard solution is poured into a burette, which is used to deliver definite but variable volumes of liquid. In a titration, the volume of liquid measured by the burette is called a **titre**. The solution of unknown concentration is added to a conical flask using a pipette. A pipette is used to deliver a known volume of liquid, which is then called an **aliquot**.

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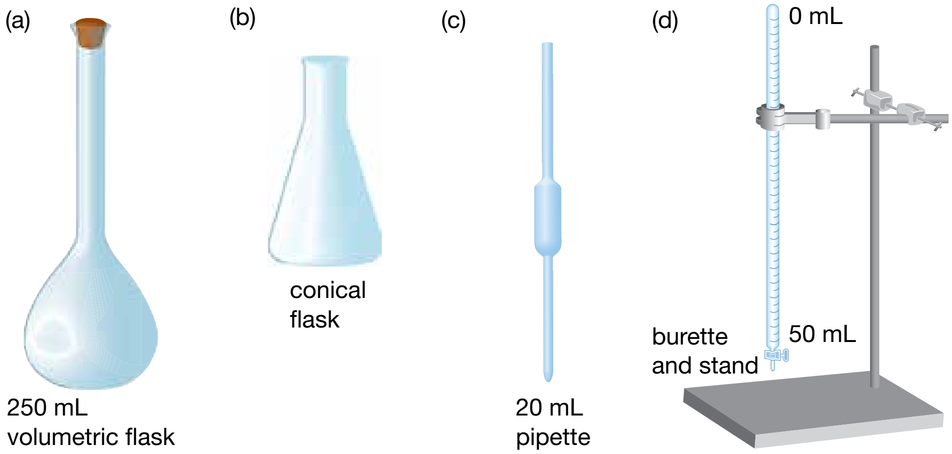
Unit 2 | **AOS 2** | **Topic 4** | **Concept 3**

Determining concentrations using titrations
Summary screen and practice questions

The correct use of a burette (left) is described on page 368. The photo on the right shows a common incorrect hand position for handling a burette.



Apparatus used in titration

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Acid–base simulation

A suitable indicator is added to the aliquot in the conical flask. The solution in the burette is slowly added to the aliquot until the indicator changes colour. This process is called titration. The point at which chemically equivalent amounts of acid and base (according to the equation) are present is called the

The equivalence point of a titration occurs when the acid and base are present in the correct stoichiometric proportions.

The end point of a titration occurs when the indicator changes colour. Indicators must be carefully chosen to ensure that this happens close to the equivalence point.

equivalence point. The point at which the indicator changes colour is called the **end point** and is usually about one drop after the equivalence point. Titres within 0.05 mL of each other are called **concordant** titres. Three concordant titres are needed to calculate the average titre.

A standard solution of hydrochloric acid is difficult to prepare because the concentrated form of hydrochloric acid produces fumes and there are dangers involved in trying to measure out a very accurate volume. For this reason hydrochloric acid is usually standardised with a known concentration of a suitable base. This involves a titration and allows accurate concentrations of hydrochloric acid to be determined. A standard solution of sodium carbonate is sometimes used as the base as its preparation is relatively easy.

Sample problem 17.3

1.33 g of anhydrous sodium carbonate is placed in a 250.0 mL (0.25 L) volumetric flask and water is added to the required mark.

This standard solution is poured into the burette and titrated against a solution of hydrochloric acid whose accurate concentration is to be determined. It was found that a 20.00 mL aliquot of HCl required a Na_2CO_3 titre of 23.50 mL for complete neutralisation.

Calculate the concentrations of:

(a) the standard solution of Na_2CO_3

(b) the HCl solution.

Solution: (a) Use $c = \frac{n}{V}$ to calculate the concentration of the standard solution.

$$n(Na_2CO_3) = \frac{1.33}{106.0} = 0.0155$$

$$c(Na_2CO_3) = \frac{n}{V} = \frac{0.0155}{0.2500} = 0.0620 \text{ M}$$

(b) Write the equation for the reaction, use it to find the molar ratio, and then calculate the concentration of the solution.

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

$$n(Na_2CO_3) = c \times V = 0.0620 \times 0.023\,50$$

$$n(HCl) = 2 \times 0.0620 \times 0.023\,50 = 2.91 \times 10^{-3} \text{ mol}$$

$$c(HCl) = \frac{n}{V} = \frac{2.91 \times 10^{-3}}{0.020\,00} = 0.146 \text{ M}$$

Therefore, the hydrochloric acid has a molarity of 0.146 M.

Sample problem 17.4

A solution of previously standardised sodium hydroxide was used to accurately determine the concentration of a sulfuric acid solution. An average titre of 18.56 mL was required to neutralise 20.00 mL of 0.0927 M sodium hydroxide solution.

Calculate the molarity of the sulfuric acid solution.

Solution: The equation for the reaction is:

$$H_2SO_4(aq) + 2NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

From this it can be seen that $n(H_2SO_4) = \frac{n(NaOH)}{2}$.

$$n(NaOH) = cV = 0.0927 \times 0.020\,00 = 0.001\,85$$

$$n(H_2SO_4) = \frac{0.001\,85}{2} = 0.000\,927$$

$$c(H_2SO_4) = \frac{n}{V} = \frac{0.000\,927}{0.018\,56} = 0.0499\text{ M}$$

The sulfuric acid therefore has a concentration of 0.0499 M (0.0499 mol L^{-1}).

Sample problem 17.5

A factory is suspected of polluting a stream with hydrochloric acid from its waste water. To gather evidence, a sample of stream water was collected downstream from the discharge point. Back in the laboratory, a 20.00 mL sample of this water was titrated against a solution of sodium carbonate that had been accurately diluted to 0.0010 M; 17.52 mL was required.

Assuming that the only acid in the sample was hydrochloric acid:

(a) calculate the molarity of hydrochloric acid in the sample

(b) calculate the concentration of the hydrochloric acid in g L^{-1}, mg L^{-1} and ppm.

Solution: The equation is:

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

(a) $n(Na_2CO_3) = cV = 0.0010 \times 0.017\,52 = 1.75 \times 10^{-5}$

$$n(HCl) = n(Na_2CO_3) \times 2 = 1.75 \times 10^{-5} \times 2 = 3.50 \times 10^{-5}$$

$$c(HCl) = \frac{n}{V} = \frac{3.50 \times 10^{-5}}{0.0200} = 0.001\,75\text{ M} = 0.0018\text{ M}$$

The molar concentration of the hydrochloric acid is 0.0018 M.

(b) $0.001\,75\text{ M} = 0.001\,75\text{ mol L}^{-1}$

$$M(HCl) = 36.5\text{ g mol}^{-1}$$

$$c(HCl)\text{ in g L}^{-1} = 0.001\,75 \times 36.5 = 0.0639\text{ g L}^{-1} = 0.064\text{ g L}^{-1}$$

$$c(HCl)\text{ in mg L}^{-1} = 0.0639 \times 1000 = 63.9\text{ mg L}^{-1} = 64\text{ mg L}^{-1}$$

mg L^{-1} is the same as ppm, so the concentration of HCl is also 64 ppm.

Revision questions

8. (a) Sodium oxalate, $Na_2C_2O_4$, can be used to make a standard solution. Describe how this is done.

(b) If 3.183 g of sodium oxalate is dissolved in water in a 250.0 mL volumetric flask, find the concentration of the solution.

9. (a) 1.461 g of dry sodium carbonate, Na_2CO_3, is dissolved in water in a 250 mL volumetric flask.

  (i) Why can this be called a standard solution?

  (ii) Calculate the concentration of the solution.

(b) 20.00 mL aliquots of this solution were titrated with nitric acid, HNO_3, three times. The average concordant titre was found to be 22.17 mL.

  (i) Write the equation for the reaction of sodium carbonate with nitric acid.

  (ii) Find the concentration of the nitric acid.

Standard solutions

A primary standard is a substance that can be used to produce solutions with precisely known concentrations. Primary standards must have certain important properties. A secondary standard can be prepared by standardising it (measuring its concentration) against a primary standard.

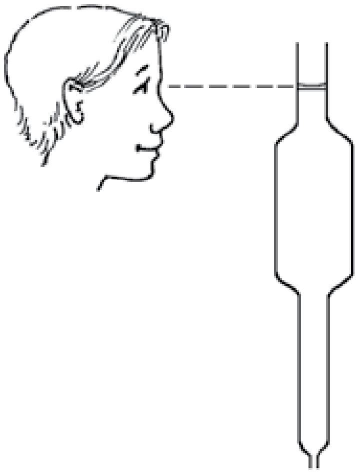
As already noted, a standard solution is one whose concentration is accurately known. There are usually two methods by which a solution may have its concentration determined accurately:

1. by reacting it with another solution whose concentration is known accurately
2. by taking a substance called a **primary standard** and dissolving it in a known volume of water. Primary standards are pure substances that satisfy a special list of criteria.

To qualify as a primary standard, a substance must have a number of the following properties:

- It must have a high state of purity.
- It must have an accurately known formula.
- It must be stable. In other words, its composition or formula must not change over time. This can happen, for example, as a result of storage or reaction with the atmosphere.
- It should be cheap and readily available.
- It should have a relatively high molar mass so that weighing errors are minimised.

Note that this means that not all substances are suitable for use as primary standards.



Reading the level of liquid in a pipette

Volumetric techniques

Successful volumetric analysis relies on accurate measurement of volumes and often, but not always, accurate knowledge of one of the initial concentrations. To this end, a number of very accurate volume-measuring pieces of equipment are used. However, their use is pointless unless the correct operating technique is practised. Correct use of some of the more common volumetric apparatus is outlined below.

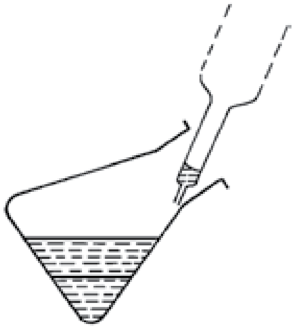
The pipette

A pipette is used to deliver a fixed volume of solution, called an aliquot. On each pipette is an etched mark above the bulb section. By drawing the liquid up above this mark, and then allowing it to drain down to it, the pipette delivers its specific calibrated volume.

Cleanliness is essential if consistent results are to be obtained. Before using the pipette, it should be rinsed with distilled water and then rinsed with a small portion of the solution it is about to contain. This ensures that the solution in use is neither contaminated nor diluted prior to being analysed.

To fill the pipette, a pipette filler should be used. This enables potentially dangerous solutions to be pipetted without the risk of the liquid being accidentally drawn into the mouth. The operation of pipette fillers varies from type to type.

Once the pipette is filled to the etched mark, the liquid is allowed to drain into the titration flask. This should be done by holding the jet of the pipette against the side of the flask and allowing the liquid to drain. After 15 seconds or so, a small amount of solution will still be present in the bottom of the jet. This should not be removed. Calibration of the pipette takes this small amount of remaining solution into account.



A small amount of solution should remain in the pipette if it is used correctly.

Common errors in using a pipette include:

- not keeping the tip of the pipette below the level of solution when filling. This causes the solution to rush up into the pipette filler and contamination may result.

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Unit 2	**Volumetric analysis equipment**
AOS 2	Summary screen
Topic 4	and practice
Concept 1	questions

- not reading from the bottom of the meniscus. In line with the usual procedure, pipettes are calibrated to the bottom of the curved liquid surface (meniscus).
- parallax error, caused by not lining up the etched mark at eye level.

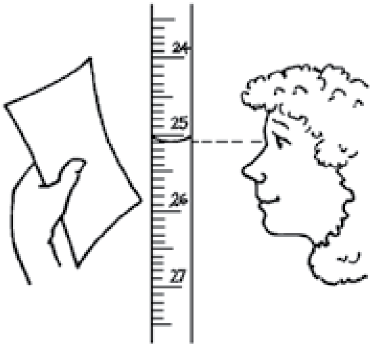
The burette

Burettes are used to deliver a variable volume, termed a titre, to the titration flask. They are marked with an accurate scale, which is usually calibrated to tenths of a millilitre. Common practice when reading these scales is to interpolate between these marks to obtain a reading to a hundredth of a millilitre (i.e. 2 decimal places). Simply by taking an initial reading, and then a final one, the difference between them corresponds to the volume delivered.

As with pipettes, clean equipment is essential for accurate results and a burette should be prepared in the same way as a pipette. That is, it should first be rinsed with distilled water, and then rinsed with a small portion of the solution it is to contain.

At this stage it is also a good idea to check that:

- the tap does not leak
- the tip is not blocked by a build-up of sediment.

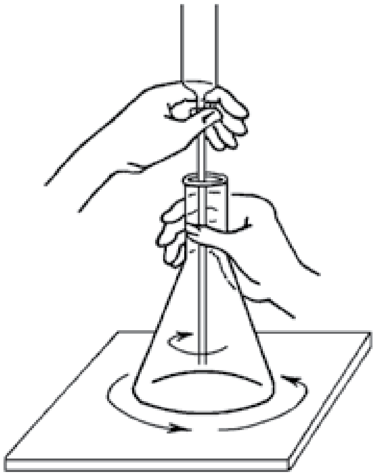


Reading a burette scale

If the burette appears to be satisfactory it should then be clamped vertically using a general purpose boss head and clamp, or a special burette clamp. The burette should then be filled with the desired solution. A small plastic funnel in the top should be used to prevent spillage. Besides being undesirable from a safety point of view, solution running down the outside may be added to the reaction flask without being measured, thus leading to error in the analysis. In most cases the solution is added until it is above the end of the scale and the funnel should then be removed. The tap at the bottom is then opened to drain the solution down until it is on the scale, and also to ensure that the jet of the burette is full. Note that the level of solution does not have to be exactly on the 0.00 mark — all that is required is that it be somewhere on the scale so that a reading can be taken.

All readings on a burette should be taken from the bottom of the meniscus and should be to 2 decimal places. When reading the scale, many people find that a piece of white paper held behind the scale is often of assistance. As with any scale reading, the eyes should be level with the level of the liquid to avoid parallax error.

In use, the 'left-hand turn — right-hand swirl' technique should be adopted. This involves wrapping the fingers and thumb of the left hand around the jet and tap and using them to open and close the tap. This leaves the right hand free to swirl the flask as the solutions are mixed. Although a little awkward at first, the technique is quickly mastered and is considered most suitable for this apparatus.



The 'left-hand turn — right-hand swirl' technique

To help determine colour changes during a titration a white tile should be placed under the titration flask. When you get near the anticipated end point, the solution should be added dropwise from the burette until the required colour change is observed. To further ensure accuracy, a wash bottle should be kept nearby so that any liquid that splashes onto the sides of the titration flask as it is being swirled can be washed back down into the liquid bulk.

Volumetric flasks

Volumetric flasks are also called standard flasks.

Volumetric flasks are used when either making up a standard solution or accurate dilution of an existing solution is required.

Preparation and use of volumetric flasks and pipettes is very similar. They are cleaned with distilled water and, once filled to an etched line, contain exactly their designated volume. However, an important difference is that, in filling, a volumetric flask is always filled to the mark from below. The volume of solution should never be allowed to go above the mark.

To make up a standard solution, the usual procedure is to:

- weigh out the required mass of solute in a special weighing bottle or small beaker
- transfer this to the volumetric flask using a dry funnel
- wash the weighing bottle and funnel a number of times, each time adding the rinsings to the flask
- add distilled water to approximately half the required volume, and then swirl to dissolve the solute
- add further distilled water until the bottom of the meniscus is level with the etched line on the neck of the flask
- fully invert at least 15 times to ensure that the concentration of the solution is uniform.

Indicators

The most important thing in acid–base titration is knowing when to stop. Acid–base indicators are critical in identifying this point. These are substances that have different colours depending on the pH of their surroundings. A range of indicators exist so that an appropriate choice may be made for a particular titration.

In an acid–base titration, the pH is not necessarily 7 at the equivalence point. If we remember that a conjugate acid and base are always formed in any acid–base reaction, it is possible that these new acids and bases have an effect on the pH. Knowledge of this pH for a particular titration means that an indicator can be chosen that changes colour at (or close to) the correct pH. Table 17.1 shows the details of some common indicators.

Acid–base indicators change colour over a narrow pH range. They allow us to stop a titration close to the equivalence point if they are chosen carefully.

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Unit 2
AOS 2
Topic 4
Concept 2

Acid–base indicators
Summary screen and practice questions

TABLE 17.1 Some common acid–base indicators

Indicator	Colour at lower pH	Colour at higher pH	pH range for colour change
methyl orange	red	yellow	3.1–4.4
methyl red	red	yellow	4.2–6.3
litmus	red	blue	5.0–9.0
bromothymol blue	yellow	blue	6.0–7.6
phenolphthalein	colourless	crimson	8.3–10.0

Looking at this table it can be seen that, if we know the pH at the equivalence point of a titration is 3.2, a suitable choice of indicator might be methyl orange. On the other hand, phenolphthalein would probably not be suitable.

It should be realised that other factors may also influence the choice of indicator. How *quickly* the pH changes at the equivalence point is one of these.

Alternatives to indicators

Sometimes, the end point of a titration may be gauged in other ways that do not involve indicators. Some of these are by:

- using a pH meter and stopping at the correct pH
- doing a potentiometric titration. This uses two electrodes and measures the changes in voltage produced between them as the titration proceeds.
- measuring conductivity. If the product of an acid-base reaction is a precipitate, the conductivity in the titration flask falls as ions are removed in the precipitate (remember that the other product of an acid-base reaction is water). As soon as all precipitation has finished, the conductivity increases again as excess ions are now added from the burette.

Titrations can be performed automatically in many modern laboratories.

Total alkalinity

Total alkalinity is an important environmental measurement for bodies of water such as rivers and lakes. As mentioned earlier in this chapter, it is a measure of how well such bodies are able to withstand pollution by acids.

Acid–base titration is used to measure total alkalinity. Water samples are titrated with a strong acid such as sulfuric acid. For technical reasons, a pH of 4.5 is chosen for the end point. The resultant calculation then predicts the amount of calcium carbonate that would have had to be present to consume the same amount of acid. The results are then quoted in terms of this.

This provides an easy way of comparing changes in a body of water over time or to compare different bodies of water.

Sample problem 17.6

Ten 1 L water samples were collected from different locations in a lake and then mixed together. One litre of this mixture was then titrated with 0.050 M sulfuric acid until the pH equalled 4.5. 15 mL of acid was required.

Calculate the total alkalinity of this water as mg L^{-1} calcium carbonate equivalent.

Solution: Calculate the mol of H_2SO_4 added: $n(H_2SO_4) = cV = 0.050 \times 0.015 = 0.000\,75$

Assume that $CaCO_3$ was the only base to react with this. The equation is:

$$H_2SO_4(aq) + CaCO_3(aq) \longrightarrow CaSO_4(s) + H_2O(l) + CO_2(g)$$

$M(CaCO_3) = 100.1 \text{ g mol}^{-1}$

$n(CaCO_3) = n(H_2SO_4) = 0.000\,75$

$m(CaCO_3) = 0.000\,75 \times 100.1 = 0.075 \text{ g} = 75 \text{ mg}$

As 1 L was the volume analysed, the total alkalinity was 75 mg L^{-1} (calcium carbonate equivalent).

Note: This result should not be interpreted to mean that calcium carbonate can actually dissolve to the level that is calculated. As has been mentioned already, calcium carbonate is virtually insoluble. It is a theoretical concentration that is representative of all the bases that are present in the sample.

Examining sources of error

In volumetric analysis, just as in gravimetric analysis, it is important to understand the steps in a procedure and predict the effect on the final calculated result if the steps are not performed correctly. Table 17.2 shows some general areas in which mistakes could be made and what effects these mistakes have.

Other situations can also give rise to mistakes when the properties of the chemicals involved are misunderstood. For example, a common method of standardising hydrochloric acid solution is to react it with a standard sodium carbonate solution. Solid sodium carbonate needs to be anhydrous to qualify as a primary standard. However, if it is left exposed to the atmosphere for too long, it absorbs moisture and begins to rehydrate. Should this then be used to make up a standard solution, not all the mass weighed out would be sodium carbonate, due to the presence of water molecules. The resulting solution of sodium carbonate is therefore lower in concentration, and the effect on the calculated concentration of the hydrochloric acid would be an overestimation.

TABLE 17.2 Effect on the calculated result of some possible mistakes during volumetric analysis

	Effect if substance under analysis is in the . . .		
Situation	**burette**	**titration flask**	**Comments**
rinsing water left in burette	underestimated	overestimated	The burette solution is diluted with water, so more is used.
rinsing water left in pipette	overestimated	underestimated	The solution aliquot in the titration flask is diluted.
indicator chosen changes colour too soon	overestimated	underestimated	The choice of indicator can be critical.
water in titration flask	no effect	no effect	All necessary measurements are made before the chemicals are mixed with this water.

Chapter review

studyon

Unit 2
AOS 2
Topic 4

Analysing water for acids and bases
Summary screens and practice questions

Summary

- Water can have a wide pH range and still be considered normal. This is due to substances that may dissolve in it.
- Water can be made acidic or basic by the substances it comes into contact with. These include rocks and minerals, the atmosphere, natural sources, dissolved salts and the results of human activity, including contamination.
- Alkalinity is a measure of the ability of water to withstand sudden pH change caused by acid input.
- The principles of stoichiometry can be applied to reactions in solution (concentration–concentration stoichiometry). The formula $n = cV$ is used here.
- When solutions are diluted, the formula $n_1V_1 = n_2V_2$ is used to calculate required quantities and results.
- Acid–base titrations are a form of volumetric analysis that use concentration–concentration stoichiometry. They are used to calculate the unknown concentration of an acid or a base.
- Titrations use specialised equipment such as burettes, pipettes and volumetric flasks. Accurate techniques must be used to produce reliable results.
- Acid–base titrations require the use of correctly chosen indicators. These are substances that have different colours at different pH values.
- The equivalence point of a titration is the point at which chemically equivalent amounts of acid and base, according to the equation, are present.
- The end point of a titration is the point at which the indicator changes colour. The indicator should be chosen so that the end point is close to the equivalence point.
- A standard solution is a solution of accurately known concentration.
- A primary standard is a substance that can be weighed out and dissolved to make a solution, the concentration of which is known accurately.

Multiple choice questions

1. Which of the following statements about water is *not* true?

A Sea water with a pH of 7 would be considered to have been acidified.
B Ground water with a pH of 6 would be considered to be normal.
C Surface water with a pH of 6 would be considered to have been acidified.
D Surface water with a pH of 5 would probably contain a lot of dissolved calcium carbonate.

2. Ground water often contains a higher concentration of solutes than surface water. Which of the following would *not* be a reason for this?

A The solubility of particular solutes is always lower in surface water than it is in ground water.
B Ground water is in contact with rocks for long periods of time.
C Temperatures are often much warmer underground.
D Ground water moves very slowly through the rocks that store it.

3. Which of the following acids is least likely to contribute to broad-scale acid rain?

A H_2CO_3
B HCl
C H_2SO_4
D HNO_3

4. Ethanoic acid, CH_3COOH, is a weak acid. It is therefore expected that a solution of its sodium salt, sodium ethanoate, CH_3COONa, has a pH:

A less than 7
B equal to 7
C greater than 7 but less than 14
D greater than 14.

5. The hydrogen sulfite ion, HSO_3^-, is amphiprotic. When it is dissolved in water, two reactions are possible:

Reaction 1:
$HSO_3^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_3^{2-}(aq)$

Reaction 2:
$HSO_3^-(aq) + H_2O(l) \rightleftharpoons H_2SO_3(aq) + OH^-(aq)$

A solution of sodium hydrogen sulfite is found to have a pH of 4. From this it can be deduced that:

A mostly reaction 1 has occurred
B mostly reaction 2 has occurred
C both reactions 1 and 2 have occurred equally
D no reaction has occurred.

6. A student dissolved 2.00 g of sodium hydroxide in 1.00 L to produce solution A. The student then diluted 100 mL of solution A with water to produce 1.00 L of solution B. A 10.0 mL sample of solution B was then diluted to 100 mL with water to produce solution C.

What was the concentration of solution C?

A 0.500 M
B 0.0500 M
C 0.005 00 M
D 0.000 500 M

7. What volume of 11 M hydrochloric acid is required to produce 2 L of 1 M acid when diluted with water?

A	91 mL	**C**	600 mL
B	182 mL	**D**	1838 mL

8. Which of the following is a *not* requirement for a substance to be a primary standard?
- **A** It must have a high state of purity.
- **B** It must not be a danger to health.
- **C** It must have an accurately known formula.
- **D** It must be stable.

9. The end point in an acid–base titration is the point when:
- **A** the solution is neutral
- **B** the indicator changes colour
- **C** equal volumes of reactants have been mixed
- **D** reactants have been mixed in the appropriate stoichiometric ratio.

10. The equivalence point in an acid–base titration is the point when:
- **A** the solution is neutral
- **B** the indicator changes colour
- **C** equal volumes of reactants have been mixed
- **D** reactants have been mixed in the appropriate stoichiometric ratio.

11. The mass of potassium hydroxide, KOH, needed to produce 200 mL of a 0.25 M solution is:

A	0.050 g	**C**	50 g
B	2.8 g	**D**	2.8 kg.

12. A solution containing 0.001 mol of HCl reacts exactly with:
- **A** 10 mL of 0.1 M NaOH
- **B** 10 mL of 0.01 M NaOH
- **C** 1 mL of 0.1 M NaOH
- **D** 1 mL of 0.01 M NaOH.

13. The volume of 0.100 M H_2SO_4 required to react completely with 25.0 mL of 0.150 M NaOH solution is:

A	9.38 mL	**C**	37.5 mL
B	18.8 mL	**D**	75.0 mL.

14. A student obtained the following titres during an analysis: 18.90 mL, 19.02 mL, 18.97 mL, 19.12 mL, 18.98 mL.

The average of the concordant titres is:

A	18.95 mL	**C**	19.01 mL
B	18.99 mL	**D**	19.07 mL.

Review questions

Acids and bases in the environment

1. Normal rainfall has a pH of about 5.6. However, the lowest pH at which surface water can be considered normal is 6.5.
- (a) Explain how rainwater becomes acidic.
- (b) What must happen to rainwater after it hits the ground to change its pH?
- (c) Rainfall in some parts of the United States has been measured at a pH of 4.3. What is the term given to such rainfall?
- (d) Explain how the situation in part (c) might come about.

2. Normal surface water may have a pH anywhere between 6.5 and 8.5. Explain why there is such a range.

3. Using the internet or other suitable reference, find three naturally occurring rocks or minerals that would make water:
(a) acidic (b) basic.

4. Two factories, in two different locations, discharge acidic waste into nearby rivers. In many ways, such as volume and flow rates, these rivers appear very similar. However, for one location, the pH of the river downstream of the factory is significantly reduced, while for the other location there is virtually no change. Explain how this might be so.

Dilution

5. Calculate the volume of stock solution needed to prepare the following.
- (a) 500 mL of a 0.750 M solution of sulfuric acid (sulfuric acid stock solution = 18 M)
- (b) 200 mL of a 2.5 M solution of ammonium hydroxide (ammonium hydroxide stock solution = 15 M)
- (c) 350 mL of a 0.150 M solution of ethanoic acid (ethanoic acid stock solution = 17 M)

6. Calculate the concentration (molarity) of:
- (a) 100 mL of 0.1 M solution after dilution with 10 mL water
- (b) 16 mL of 0.20 M solution after dilution with 4.0 mL water
- (c) 750 mL of 1.50 M solution after dilution with 150 mL water.

7. What volume of water must be added to:
- (a) 100 mL of 15 M H_2SO_4 to dilute it to 5.0 M
- (b) 130 mL of 3.50 M HNO_3 to dilute it to 1.00 M
- (c) 170 mL of 2.60 M HCl to dilute it to 0.250 M?

8. A student is required to make up 500 mL of 0.100 M $H_2SO_4(aq)$ by diluting a 0.500 M solution of the acid. What volume of 0.500 M $H_2SO_4(aq)$ is required?

9. What volume of water is needed to produce a 2.00 M NaCl solution if 2.35 g of NaCl is present?

10. A 3.50 g mass of $MgCl_2$ is dissolved in 200 mL of water.
- (a) What is the concentration of the solution?
- (b) If 150 mL of water is added to the solution, what is the concentration of the new, diluted solution?

Solution stoichiometry

11. Sodium chloride reacts with silver nitrate to give a white precipitate of silver chloride.
- (a) Write an equation for the reaction.
- (b) What volume of 2.0 M silver nitrate reacts with 120 mL of 1.5 M sodium chloride?

12. 50 mL of 1.2 M sulfuric acid was added to 30 mL of sodium hydroxide in order to neutralise it. Find the molarity of the sodium hydroxide.

13. An iron ore sample may be analysed by preparing an iron(II) ion solution of the ore, which is then reacted with acidic permanganate solution as in the given equation. If 15.0 mL of 0.0100 M MnO_4^- solution is required to completely react with 10.0 mL of the Fe^{2+} solution, determine the concentration of the iron(II) ion solution.

$$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \longrightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$$

14. What volume of a 2.30 M solution of Na_3PO_4 is required to react completely with 560 mL of a 3.25 M solution of $Ca(ClO_3)_2$, according to the reaction:

$$3Ca(ClO_3)_2(aq) + 2Na_3PO_4(aq) \longrightarrow Ca_3(PO_4)_2(aq) + 6NaClO_3(aq)$$

15. What volume of 0.250 M HNO_3 reacts completely with 280 mL of 0.200 M H_2S according to the reaction:

$$2HNO_3(aq) + 3H_2S(aq) \longrightarrow 3S(s) + 4H_2O(l) + 2NO(g)$$

Acid–base titrations

16. The concentration of vinegar was determined by titration. 20.00 mL of vinegar required 43.20 mL of 0.350 M sodium hydroxide according to the following equation. Find the concentration of the vinegar.

$$CH_3COOH(aq) + NaOH(aq) \longrightarrow CH_3COONa(aq) + H_2O(l)$$

17. What volume of 0.460 M H_2SO_4 is required to neutralise 24.00 mL of 0.620 M NaOH?

18. 1.365 g of anhydrous sodium carbonate, Na_2CO_3, was dissolved in 250.0 mL of water. 20.00 mL aliquots of this solution were titrated with nitric acid. An average titre of 21.95 mL was needed. Find the concentration of the nitric acid.

19. A sample of water known to contain sodium hydroxide as the only base was analysed prior to discharge into a river. It was found that a 20.00 mL sample required 8.19 mL of 0.0100 M hydrochloric acid solution for neutralisation.
    (a) Calculate the molar concentration of sodium hydroxide in the water sample.
    (b) Express the answer from part (a) in g L^{-1}.

Exam practice questions

In a chemistry examination you will be required to answer a number of short and extended response questions.

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Unit 2
AOS 2
Topic 4

Analysing water for acids and bases

Sit topic test

Extended response questions

1. Anhydrous sodium carbonate is often used to make standard solutions of sodium carbonate. This solution may then be used in acid–base titrations to calculate the exact concentration of acidic solutions.

(a) Calculate the mass of anhydrous sodium carbonate, Na_2CO_3, that is required to make 500.00 mL of 0.0500 M solution. **3 marks**

(b) Exactly 25.00 mL of the solution from part (a) is reacted with hydrochloric acid. The equation for this reaction is:

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

It is found that 23.50 mL of the acid is required for exact neutralisation. **4 marks**

Calculate the molarity of the hydrochloric acid solution.

2. A company operates a batch process that uses sulfuric acid. At the end of each batch, the leftover acid must be diluted before it is released. The amount of water to be added to enable this dilution is carefully calculated following analysis of the leftover sulfuric acid in the batch liquid.

The terms of the discharge licence state that there should be no more than 1.00 g L^{-1} of sulfuric acid in the discharge water.

A 25.00 mL sample of the batch liquid was accurately diluted to 250.0 mL in a standard flask. 25.00 mL of this diluted sample was then titrated with 0.100 M standardised sodium hydroxide solution using a suitable indicator. An average titre of 18.52 mL was obtained.

(a) Write the equation for the reaction between sodium hydroxide and sulfuric acid. **1 mark**

(b) Calculate the number of moles of sodium hydroxide that were added. **1 mark**

(c) From your answer to part (b), calculate the number of moles of sulfuric acid in the diluted sample. **1 mark**

(d) Hence, calculate the molarity of sulfuric acid in the diluted sample. **1 mark**

(e) Calculate the molarity of sulfuric acid in the undiluted batch liquid. **2 marks**

(f) Calculate the molarity of the sulfuric acid that is allowed in the discharge water. **1 mark**

(g) Hence, calculate the number of litres of water that must be added to each litre of batch liquid to meet the terms of the discharge licence. Give your answer to 1 decimal place. **3 marks**

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PRACTICAL INVESTIGATIONS

EXPERIMENT 1.1 Flame colours
EXPERIMENT 3.1 Investigating calcite crystals
EXPERIMENT 3.2 Ionic models
EXPERIMENT 4.1 Metallic trees
EXPERIMENT 4.2 The effect of heat on metals
EXPERIMENT 4.3 The underwater garden
EXPERIMENT 5.1 Empirical formula of magnesium oxide
EXPERIMENT 5.2 Stoichiometry of a reaction
EXPERIMENT 5.3 The empirical formula of a hydrated salt
EXPERIMENT 6.1 Building molecular models
EXPERIMENT 8.1 Formation of esters
EXPERIMENT 9.1 Cross-linking an addition polymer to make slime
EXPERIMENT 11.1 Classification of chemical reactions
EXPERIMENT 12.1 Testing for solubility
EXPERIMENT 12.2 Specific heat capacity
EXPERIMENT 12.3 Determine a solubility curve
EXPERIMENT 12.4 A water modelling exercise
EXPERIMENT 13.1 The reactions of acids
EXPERIMENT 13.2 Indicators and pH
EXPERIMENT 13.3 Simple redox equations
EXPERIMENT 13.4 Corrosion
EXPERIMENT 13.5 Minimising corrosion (student design)
EXPERIMENT 13.6 The reactivity of metals and their salts
EXPERIMENT 13.7 Complex redox equations
EXPERIMENT 14.1 Preparation of a solution of known concentration
EXPERIMENT 14.2 Finding the pH of common household substances
EXPERIMENT 14.3 Carbon dioxide content of fizzy drinks
EXPERIMENT 15.1 Zinc content of cornflakes using AAS

Answers

Chapter 1

Page 6

1. 20 times
2. 140 times

Page 9

3. (a) 1 (b) 10 (c) 47 (d) 79
4. (a) 13 (b) 26 (c) aluminium
5. (a) B (b) Mg (c) Ar (d) Ca

Page 11

6. (a) 22 (b) 18 (c) $^{40}_{18}\text{Ar}$
7. number of protons = 35, number of neutrons = 44
8. $^{2}_{1}\text{H}$
9.

Element	Number of protons	Number of electrons	Number of neutrons
$^{12}_{6}\text{C}$	6	6	6
$^{56}_{26}\text{Fe}$	26	26	30
$^{40}_{18}\text{Ar}$	18	18	22
$^{235}_{92}\text{U}$	92	92	143
$^{238}_{92}\text{U}$	92	92	146
$^{19}_{9}\text{F}$	9	9	10

(a) $^{235}_{92}\text{U}$ and $^{238}_{92}\text{U}$ are isotopes.

(b) $^{238}_{92}\text{U}$ has 3 more neutrons than $^{235}_{92}\text{U}$.

10. (a) $^{14}_{7}\text{N}$, $^{15}_{7}\text{N}$ (c) 7 and 8 respectively
    (b) 7 protons (d) isotopes
11. The atomic number is always the same for a particular element, whereas the mass number can differ due to the existence of isotopes.

Pages 14–15

12. from top to bottom — symbols: L, M, N; maximum number of electrons: 8, 18, 32
13. (a) The Bohr model explained the emission spectrum of hydrogen.
    (b) The spectra of atoms more complex than hydrogen could not be explained satisfactorily.
14. He 2; Li 2, 1; B 2, 3; C 2, 4; N 2, 5; O 2, 6; F 2, 7; Ne 2, 8; Mg 2, 8, 2; Al 2, 8, 3; Si 2, 8, 4; P 2, 8, 5; S 2, 8, 6; Ar 2, 8, 8; Ca 2, 8, 8, 2
15. Shell models have the electrons in shells according to the electron configurations in question 14.
16. (a) 2, 5 (b) calcium

Pages 18–19

17. $_{3}\text{Li}$: $1s^22s^1$
    $_{19}\text{K}$: $1s^22s^22p^63s^23p^64s^1$
    $_{6}\text{C}$: $1s^22s^22p^2$
    $_{17}\text{Cl}$: $1s^22s^22p^63s^23p^5$
    $_{18}\text{Ar}$: $1s^22s^22p^63s^23p^6$
    $_{7}\text{N}$: $1s^22s^22p^3$
18. (a) $1s^22s^22p^63s^23p^3$
    (b) ^{32}P has 17 neutrons while ^{31}P has only 16.
19. (a) group 16 (b) period 3 (c) sulfur, S
20. (a) lithium — group 1, period 2
    (b) neon — group 18, period 2
    (c) sodium — group 1, period 3
    (d) aluminium — group 13, period 3
    (e) argon — group 18, period 3
    (f) calcium — group 2, period 4
21. (a) $Z = 20$: $1s^22s^22p^63s^23p^64s^2$
    $Z = 12$: $1s^22s^22p^63s^2$
    $Z = 4$: $1s^22s^2$
    $Z = 9$: $1s^22s^22p^5$
    Same group: $Z = 20$, 12 and 4 (group 2)
    (b) $Z = 5$: $1s^22s^22p^1$
    $Z = 6$: $1s^22s^22p^2$
    $Z = 8$: $1s^22s^22p^4$
    $Z = 16$: $1s^22s^22p^63s^23p^4$
    Same group: $Z = 8$ and 16 (group 16)
22. (a) He $1s^2$; Ne $1s^22s^22p^6$; Ar $1s^22s^22p^63s^23p^6$
    (b) Al $1s^22s^22p^63s^23p^1$
    (c) N $1s^22s^22p^3$
    (d) Cl $1s^22s^22p^63s^23p^5$

Page 19

23. Shells are composed of subshells.
25. (a) ground, (b) excited, (c) ground
26. (c) $1s^22s^22p^63s^2$
27. (a) manganese
    (b) calcium

Multiple choice questions

1. C 2. D 3. B 4. A 5. C 6. D 7. D 8. D
9. C 10. D 11. C 12. A 13. A 14. C 15. B 16. C

Review questions

1. Protons and neutrons have almost identical masses whereas the mass of the electron is negligible. Protons are positively charged. Electrons have negative charge and neutrons have no charge. Protons and neutrons are found in the nucleus whereas electrons are found in the space around the nucleus.
2. red blood cell 7000 nm, virus 30 nm, hydrogen molecule 0.15 nm
3. (a), (b)

4. helium atom, DNA molecule, bacteria, width of human hair, flea
5. (a) (i) 11 (ii) 23 (iii) 11 (iv) 12 (v) 11 (vi) sodium
   (b) (i) 9 (ii) 19 (iii) 9 (iv) 10 (v) 9 (vi) fluorine
   (c) (i) 14 (ii) 28 (iii) 14 (iv) 14 (v) 14 (vi) silicon
   (d) (i) 26 (ii) 56 (iii) 26 (iv) 30 (v) 26 (vi) iron
   (e) (i) 79 (ii) 197 (iii) 79 (iv) 118 (v) 79 (vi) gold
   (f) (i) 92 (ii) 235 (iii) 92 (iv) 143 (v) 92 (vi) uranium
6.

Name of atom	**Atomic number**	**Mass number**	**Protons**	**Neutrons**	**Electrons**
argon	18	36	18	18	18
sulfur	16	34	16	18	16
argon	18	38	18	20	18
phosphorus	15	31	15	16	15
lead	82	208	82	126	82
potassium	19	39	19	20	19
sulfur	16	36	16	20	16

7. isotopes
8. (a) 6 (b) 12, 13, 14 (c) 6, 7, 8 (d) 6
9. (a) *A*, *C*, *D*
   (b) $A = {}^{40}_{20}\text{Ca}$, $B = {}^{37}_{17}\text{Cl}$, $C = {}^{42}_{20}\text{Ca}$, $D = {}^{46}_{20}\text{Ca}$
10. chlorine ${}^{37}_{17}A$, ${}^{35}_{17}D$

    magnesium ${}^{26}_{12}B$, ${}^{25}_{12}E$, ${}^{24}_{12}G$

    cobalt ${}^{59}_{27}C$, ${}^{60}_{27}F$
11. (a) absorption of energy
    (b) A photon equal to the difference in energy of the two levels is emitted.
12. one, as each compound contains the same metallic cation, Ca^{2+}
13. When sodium atoms are excited, the electrons jump to a higher energy level. When they return to their normal energy levels, they emit the characteristic yellow light.
14. The ground state refers to the most stable state where all electrons are occupying their normal energy levels or shells. When an atom is excited, the electrons can absorb energy and jump to higher energy levels (the excited states) depending on the amount of energy absorbed.
15. An atom becomes an ion when it loses or gains electrons.
16. $2 + 8 + 18 + 32 = 60$
17. (a) oxygen (c) chlorine (e) potassium
    (b) aluminium (d) argon
18. (a) 2, 7 (b) 2, 8, 8, 1 (c) 2, 6 (d) 2, 8, 2 (e) 2, 8, 8, 2
19. (a) A shell is the main energy level where an electron may be found. A shell is made up of subshells. Each shell has a specific number of subshells.
    (b) An orbit is a regularly repeated path or circuit.
    An orbital is a region of space in which electrons move.
20. (a) 6 (c) 10
    (b) 2 (d) 18
21. 5
22. (a) Half-filled and filled subshells are more stable than other partly filled subshells
    (b) chromium $1s^22s^22p^63s^23p^63d^54s^1$
    (c) copper $1s^22s^22p^63s^23p^63d^{10}4s^1$

Exam practice questions

1. B 2. A 3. D 4. C
6. (a) (i) Ar $1s^22s^22p^63s^23p^6$
       (ii) Ca^{2+} $1s^22s^22p^63s^23p^6$
   (b) Argon is in period 3 because the third shell is being filled. It is in group 18 because there are 8 electrons in its outer shell. For elements containing more than 2 electrons in the outer shell, the group number can be obtained by adding 10 (for the transition elements) to the number of outershell electrons.

Chapter 2

Page 28

2. Mendeleev based his predictions on properties of elements in the same group.

Pages 31–2

4. (a) period 2, group 16
   (b) period 2, group 1
   (c) period 4, group 8
   (d) period 4, group 18
   (e) period 4, group 3
   (f) period 4, group 12
5. (a) O, oxygen (c) Fe, iron (e) Sc, scandium
   (b) Li, lithium (d) Kr, krypton (f) Zn, zinc
6. (a) $1s^22s^22p^6$ period 2, group 16
   (b) $1s^22s^22p^63s^23p^5$, period 3, group 17
   (c) $1s^22s^22p^63s^23p^64s^2$, period 4, group 2
   (d) $1s^22s^22p^63s^23p^63d^{10}4s^24p^5$, period 4, group 17

Page 33

7. (a) Zn, $1s^22s^22p^63s^23p^63d^{10}4s^2$

Page 37

8. (a) Ca (b) Be (c) Cl

Multiple choice questions

1. C 2. B 3. D 4. B 5. C 6. A 7. C 8. D 9. D
10. C 11. B

Review questions

2. (a) Cl and I have similar properties because they are in the same group.
   (b) Na. It is easier for sodium to lose one electron than for magnesium to lose two to form a stable, full outer shell. Because transition metals are smaller, electrons are more strongly held and less reactive than groups 1 and 2 metals.
   (c) Nickel is a transition metal.
   (d) Cl is the most reactive non-metal as it has the highest electronegativity of the elements listed. Only one electron is required to complete its outer shell.
   (e) Nitrogen is the most electronegative as it has a relatively high nuclear charge and little shielding effect of inner electrons.
4. increasing atomic number
5. Arranging atoms in order of increasing atomic mass would result in elements not being in groups with similar properties.
6. (a) Na and K would be the most similar as they are in the same group. Na and Cl would be the most different because Na is a metal and Cl is a non-metal and so they would have different electronegativities.
   (b) Cl and I would be the most similar as they are in the same group. Cl and Mg would be the most different because Mg is a metal and Cl is a non-metal and so they would have different electronegativities.

7. Electrons are further away from the attracting power of the nucleus.
8. (a) $1s^22s^22p^63s^23p^3$
   (b) non-metal
   (c) group 15
   (d) period 3
   (e) *p* block
   (f) nitrogen or arsenic
   (g) lower
   (h) larger
9. These elements are placed in groups with similar properties and in order of increasing atomic number.
11. (a) Atomic size decreases going across a period because the effective nuclear charge increases.
    (b) Atomic size increases going down a group because the number of occupied electron shells increases.
13. Elements with the same outer electron configuration are in the same group and so have similar properties: e.g. oxygen $1s^22s^22p^4$, sulfur $1s^22s^22p^63s^23p^4$, selenium $1s^22s^22p^63s^23p^63d^{10}4s^24p^4$.

Exam practice questions

1. Metals have low electronegativities because they have low effective nuclear charge whereas non-metals have high electronegativities because of the high effective nuclear charge.
2. (a) the electron attracting power of an atom
   (b) Electronegativity is the electron-attracting power. Ionisation energy is the amount of energy required to remove an electron from an atom.
3. (a) *D* (or *X*) (c) *B* (e) *A* (g) *T*
   (b) *X* (d) *J* (f) *G* or *J*

Chapter 3

Page 46

1. Ca and Al each form an ion in order to attain a full outer shell and, hence, stability.
2. (a) calcium Ca → $Ca^{2+} + 2e^-$
   2, 8, 8, 2 2, 8, 8
   aluminium Al → $Al^{3+} + 3e^-$
   2, 8, 3 2, 8
   (b) nitrogen $N + 3e^-$ → N^{3-}
   2, 5 2, 8
   fluorine $F + e^-$ → F^-
   2, 7 2, 8
3. (a) barium ion, Ba^{2+}
   (b) potassium ion, K^+
   (c) phosphide ion, P^{3-}
   (d) chloride ion, Cl^-
   (e) sulfide ion, S^{2-}
4.

Period	Group 1	Group 2	Group 13	Group 15	Group 16	Group 17
2	+1	+2	+3	−3	−2	−1
3	+1	+2	+3	−3	−2	−1

Page 49

5. (a) $Ca^{2+} + O^{2-} \rightarrow CaO$
   (b) $Be^{2+} + 2Cl^- \rightarrow BeCl_2$
   (c) $Li^+ + F^- \rightarrow LiF$
   (d) $2Al^{3+} + 3S^{2-} \rightarrow Al_2S_3$
   (e) $3Na^+ + N^{3-} \rightarrow Na_3N$
   (f) $Mg^{2+} + S^{2-} \rightarrow MgS$
6. (a) K^+ 2, 8, 8; Ca^{2+} 2, 8, 8; Al^{3+} 2, 8
   (b) F^- 2, 8; O^{2-} 2, 8; N^{3-} 2, 8
7. neon

Page 50

8. (a) The ions Mg^{2+} and $2Cl^-$ make up $MgCl_2$.
   The ions K^+ and Cl^- make up KCl.
   (b) $MgCl_2$ has the ratio 1 : 2 because two Cl^- ions are needed to balance the charge on one Mg^{2+} ion. KCl has the ratio of 1 : 1 because one K^+ ion balances one Cl^- ion.
   (c) The lattice is maintained by the strong electrostatic attraction between the positive and negative ions. This is called the ionic bond.

Page 52

9. (a) Calcium chloride has a high melting point as it is an ionic compound with strong electrostatic attractive forces holding the ions together in a three-dimensional lattice.
   (b) It will shatter when pressure is applied as ions of like charge repel when aligned with one another.
   (c) It will conduct electricity only in the liquid state where the ions are free to move. In the solid state, the ions are held strongly in place by the strong ionic bond and so will not conduct electricity.
10. (a) (i) NaCl has one Na^+ ion for each Cl^- ion; therefore, the ratio is 1 : 1.
        (ii) MgO has one Mg^{2+} ion for each O^{2-} ion; therefore, the ratio is 1 : 1.
    (b) MgO has a higher melting point as the electrostatic attraction is greater between its ions due to their higher charges.
12. (a) X_2Y (b) ionic (c) hard, brittle, high melting point

Page 54

13. (a) $AlCl_3$ (b) BaO (c) Na_2S (d) Mg_3P_2
14.

Ions	**K^+**	**Ca^{2+}**	**Al^{3+}**
F^-	KF	CaF_2	AlF_3
O^{2-}	K_2O	CaO	Al_2O_3
N^{3-}	K_3N	Ca_3N_2	AlN

15. potassium fluoride, calcium fluoride, aluminium fluoride, potassium oxide, calcium oxide, aluminium oxide, potassium nitride, calcium nitride, aluminium nitride
16. (a) potassium chloride
    (b) silver sulfide
    (c) sodium nitride

Page 56

17. (a) iron(II) sulfide
    (b) iron(III) sulfide
    (c) copper(II) chloride
    (d) tin(II) oxide
    (e) copper(I) oxide
    (f) lead(II) bromide
18. (a) SnF_4 (b) PbS (c) HgO (d) FeN (e) Cu_2S (f) SnO

Page 58

19. HSO_4^-, hydrogen sulfate; HSO_3^-, hydrogen sulfite; SO_4^{2-}, sulfate; SO_3^{2-}, sulfite; $S_2O_3^{2-}$, thiosulfate
20.

Cation	Anion	Empirical formula
Ag^+	Cl^-	AgCl
K^+	S^{2-}	K_2S
Mg^{2+}	O^{2-}	MgO
Al^{3+}	Br^-	$AlBr_3$
Fe^{3+}	CO_3^{2-}	$Fe_2(CO_3)_3$
Ba^{2+}	PO_4^{3-}	$Ba_3(PO_4)_2$
NH_4^+	SO_4^{2-}	$(NH_4)_2SO_4$

21. (a) Na_2SO_3 (b) $Ca(NO_2)_2$ (c) $Cu(HCO_3)_2$

22. (a) aluminium carbonate, $Al_2(CO_3)_3$
 (b) sodium nitrate, $NaNO_3$
 (c) mercury(II) phosphate, $Hg_3(PO_4)_2$
 (d) lead(II) sulfate, $PbSO_4$

23. sodium hydrogen carbonate, $NaHCO_3$; ammonium nitrate, NH_4NO_3; nitrate, NO_3^-; nitrite, NO_2^-; sulfite, SO_3^-

Page 60

25. (a) $Na_2CO_3.10H_2O$ (b) magnesium sulfate heptahydrate
 (c) $MgCl_2.6H_2O$ (d) barium chloride dihydrate

Multiple choice questions

1. A 2. C 3. A 4. C 5. C 6. D 7. A 8. B 9. C 10. B
11. A 12. C 13. C

Review questions

1.

Name of atom	Symbol for atom	Electron configuration of atom
lithium	Li	2, 1
beryllium	Be	2, 2
nitrogen	N	2, 5
oxygen	O	2, 6
fluorine	F	2, 7
sodium	Na	2, 8, 1
magnesium	Mg	2, 8, 2
aluminium	Al	2, 8, 3
phosphorus	P	2, 8, 5
sulfur	S	2, 8, 6
chlorine	Cl	2, 8, 7
potassium	K	2, 8, 8, 1
calcium	Ca	2, 8, 8, 2

Name of ion	Symbol for ion	Electron configuration of ion
lithium	Li^+	2
beryllium	Be^{2+}	2
nitride	N^{3-}	2, 8
oxide	O^{2-}	2, 8
fluoride	F^-	2, 8
sodium	Na^+	2, 8
magnesium	Mg^{2+}	2, 8
aluminium	Al^{3+}	2, 8
phosphide	P^{3-}	2, 8, 8
sulfide	S^{2-}	2, 8, 8
chloride	Cl^-	2, 8, 8
potassium	K^+	2, 8, 8
calcium	Ca^{2+}	2, 8, 8

2. (a) 2, 8, 8, 2 (b) 2, 7
 (c) Two fluorine atoms each accept one electron from a calcium atom, resulting in two fluoride ions, $2F^-$, and one calcium ion, Ca^{2+}.
 (e) the ionic bond — the strong electrostatic attraction between the oppositely charged ions
 (f) 1 : 2

3. Metal atoms lose electrons to attain a full outer shell and become stable ions or cations. The resultant charge of the ion is positive as there are more protons than electrons.

4. The ratio of the different ions in potassium oxide is 2 : 1. This is because two K^+ ions are necessary to balance the O^{2-} charge.

5. (a) $K^+ + F^- \longrightarrow KF$ (c) $Be^{2+} + 2Cl^- \longrightarrow BeCl_2$
 (b) $2Al^{3+} + 3O^{2-} \longrightarrow Al_2O_3$

6. Ionic salts conduct electricity in the molten and aqueous states as the ions are mobile. In the solid state they are rigidly held by strong ionic bonds and therefore cannot conduct electricity.

8. Ionic substances are usually solids and have high melting points due to the strong attractive electrostatic forces holding the ions together in a lattice.

9. (a) XY (b) ionic (c) $X^{2+} + Y^{2-} \longrightarrow XY$

10. (a) $MgSO_4$ (h) PbI_2 (o) $Ba(NO_3)_2$ (v) Cr_2O_3
 (b) ZnO (i) $KHSO_4$ (p) $Fe(OH)_3$ (w) $Ca(NO_3)_2$
 (c) $Fe(OH)_2$ (j) $(NH_4)_2CO_3$ (q) Na_2S (x) $LiCl$
 (d) $AgCl$ (k) Ag_2SO_4 (r) Al_2O_3 (y) KCN
 (e) AlN (l) $SnCl_2$ (s) CaH_2 (z) Na_2HPO_4
 (f) $CaCO_3$ (m) K_2SO_4 (t) $CuSO_4$
 (g) $Ca(HCO_3)_2$ (n) NaF (u) NH_4OH

11. (a) iron(III) oxide (g) aluminium carbonate
 (b) aluminium sulfate (h) sodium sulfate
 (c) calcium chloride (i) silver nitrate
 (d) magnesium nitrate (j) sodium hydroxide
 (e) barium sulfate (k) potassium nitrate
 (f) zinc chloride (l) lead(II) chloride

12. (a) sodium carbonate, Na_2CO_3
 (b) sodium hydrogen carbonate, $NaHCO_3$
 (c) calcium carbonate, $CaCO_3$

13. (a) $MgSO_4.7H_2O$ (c) $ZnCl_2.6H_2O$
 (b) $Na_2CO_3.10H_2O$ (d) $BaCl_2.2H_2O$

14. (a) barium chloride trihydrate
 (b) lithium chloride tetrahydrate
 (c) cobalt(II) chloride pentahydrate

Exam practice questions

1. B 2. A
3. (a) Ca^{2+} (b) P^{3-} (c) Al^{3+}

Chapter 4

Page 67

1. (a) 2, 1 (b) Li^+, 2
 (c) Solid lithium contains lithium cations in a ‘sea’ of delocalised valence electrons. Electrostatic forces of attraction hold the lattice together.
2. (b), (d), (e)

Page 69

5. (a) The metallic lattice consists of magnesium cations, Mg^{2+}, in a ‘sea’ of delocalised electrons. The electrostatic attraction between the cations and the electrons bonds them together.
 (b) The mobile electrons within the lattice can reflect light.

Page 71

6. (a) Solder can be made by heating 30% tin and 70% lead together.
 (b) substitutional

(c) Solder has a lower melting point than both tin and lead and, therefore, can be used to join metal parts together.

7. (a) brass (b) aluminium alloy (c) steel (d) steel

Page 74

10. (a) +1 (b) +2
11. Actual values: density = 1.53 g cm^{-3}, melting point = 39.3 °C, thermal conductivity = 58.2 W m^{-1} K^{-1}, electrical conductivity = 0.83 ohm^{-1} $m^{-1} \times 10^7$ and hardness (soft) of rubidium

Page 76

13. Gold and silver are used for jewellery because they are soft enough to be easily worked, they are lustrous and unreactive so will not corrode easily.
14. (a) Gold is a coloured metal.
    (b) Mercury is a liquid at room temperature.
    (c) Zinc compounds are white.
15. (a) iron (b) aluminium (c) copper (d) lead
16. Aluminium is cheaper and lighter than copper.

Page 78

19. It does not explain the differences in magnetism, density and strength of metals.
20. A metal that has large grains will have fewer dislocations and will bend more easily than a metal with small grains that has many dislocations.
21. (a) Tempering disrupts the metallic lattice, reducing brittleness while retaining hardness.
    (b) quenching and annealing
22. The work-hardened copper becomes more brittle; annealing it restores ductility.
23. The metal clip has been work-hardened — the crystal grains are now smaller and the metal is more brittle, making it more difficult to bend it back to its original shape.

Page 84

26. Fe_3O_4, iron(II, III) oxide; Fe_2O_3, iron(III) oxide; FeS, iron(II) sulfide; $Fe_2O_3.3H_2O$, iron(III) oxide trihydrate; $FeCO_3$, iron(II) carbonate

Multiple choice questions

1. C 2. C 3. D 4. A 5. B 6. A 7. B 8. A 9. B
10. C 11. B 12. A 13. D 14. C 15. B 16. D 17. C

Review questions

1. *B*
2. (a), (e)
5. (a) Because gold is unreactive, it can be found in its natural state.
    (b) Gold is a soft metal.
6. Silver bonds to the mercury and helps the amalgam set.
   Tin helps the amalgam stay free from tarnish.
   Mercury can be poured into the tooth easily.
   Copper and zinc give the amalgam strength.
   Gold will not tarnish or react with food or drink.
7. (a) A metal is a pure substance (an element), whereas an alloy is a mixture of metals or of a metal and a non-metal (such as carbon) heated to a molten state and then cooled.
8. (a) element
    (b) alloy
    (c) element
    (d) alloy
    (e) alloy
    (f) element
    (g) alloy
    (h) compound
    (i) compound
    (j) element
    (k) alloy
    (l) alloy
    (m) compound
9. The properties of an alloy are different from the properties of the constituent metals.
10. Alloys often have properties that make them better suited for a particular use than the original metal(s).
11. (a) *d* (b) *p* (c) *s* (d) *d* (e) *d* (f) *d*
12. A work-hardened metal is difficult to bend as it becomes brittle due to smaller crystal grains and increased dislocations.
13. Annealing is the process of heating a metal and then cooling it slowly. This produces larger crystals and a softer metal. Tempering is the process of warming a quenched metal and allowing it to cool slowly. This reduces its brittleness but retains its hardness.
14. Quenching is the process of cooling a red hot metal quickly by plunging into cold water. This produces smaller crystals and hence a harder but more brittle metal.
15. (a) dish-draining rack on a sink (c) tin-coated cans
    (b) bike chain (d) galvanised steel roofing
16. Observe its reaction with water or acid and compare the results with other metals whose positions on the activity series are known.

17 (a) $4K(s) + O_2(g) \rightarrow 2K_2O(s)$
   (b) $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$
   (c) $Sn(s) + 2HNO_3(aq) \rightarrow Sn(NO_3)_2(aq) + H_2(g)$

18.

	Nanogold	**Bulk gold**	**Both**	**Neither**
(a) reactive	✓			
(b) magnetic	✓			
(c) spreads through water	✓			
(d) range of colours	✓			
(e) good conductor			✓	
(f) very large surface area	✓			
(g) insoluble		✓		
(h) unreactive		✓		
(i) does not conduct				✓
(j) transition metal			✓	

19. (b) Iron is malleable, ductile, has a high melting point and is hard.
    (c) Iron is abundant, hard, has high tensile strength and is relatively easily worked.
    (d) Iron is alloyed to improve properties, e.g. prevent corrosion
    (e) 1. Coke reacts with oxygen in the air:
    $$C(s) + O_2(g) \rightarrow CO_2(g)$$
    2. Limestone decomposes:
    $$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$
    3. Carbon dioxide reacts with more coke:
    $$C(s) + CO_2(g) \rightarrow 2CO(g)$$
    4. This CO reacts with iron oxide in the ore:
    $$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$$

Exam practice questions

1. (a) C (b) A (c) B
2. (a) Iron ores are abundant and iron is relatively easily extracted from its ore compared with some other metals.
    (b) Stainless steel. It is more resistant to corrosion.
    (c) construction; to make containers; to make vehicles, machinery and tools; as a catalyst

Chapter 5

Page 94

1. Carbon-12 can be measured accurately because it is stable and abundant.
2. (a) closer to 7 as this isotope is more abundant
   (b) 6.94
3. 63.6 (64 to 2 significant figures)
4. (b) 24.3 (24 to 2 significant figures)
5. 70.8 (71 to 2 significant figures)
6. 10.8 (11 to 2 significant figures)
7. (a) 44.0 (b) 58.5 (c) 34.0 (d) 98.1 (e) 180.0

Page 96

8. (a) 48.0 g mol^{-1} (c) 56.1 g mol^{-1}
   (b) 253.8 g mol^{-1} (d) 36.5 g mol^{-1}

Page 97

9. A mole is a number such that 1 mole of any substance *X* consists of Avogadro's number, 6.02×10^{23}, particles of *X*.
10. (a) 1.4×10^{24} atoms of chlorine
    (b) 1.90×10^{25} atoms of chlorine
    (c) 2.1×10^{24} chloride ions
    (d) 6.0×10^{23} chloride ions
11. (a) 0.270 mol (c) 0.540 mol
    (b) 1.63×10^{23} molecules (d) 3.25×10^{23} oxygen atoms
12. (a) 8.6 mol (b) 8.6 mol (c) 35 mol

Page 98

13. (a) 2.6 mol (c) 0.94 mol (e) 0.513 mol
    (b) 0.055 mol (d) 0.034 mol (f) 2.4×10^2 mol
14. (a) 11 g (c) 1.31 kg (e) 0.73 g
    (b) 769 g (d) 3.25×10^3 g (f) 665 g
15. (b)

Page 99

16. (a) 1.57 kg (c) 2.60×10^{-8} g (e) 5.53×10^2 g
    (b) 0.140 g (d) 5.23×10^6 g
17. (a)

Page 100

19. (a) C 81.8%, H 18%
    (b) Na 19.2%, H 0.83%, S 26.7%, O 53.3%
    (c) Ca 25.4%, C 30.4%, O 40.5%, H 3.8%
    (d) H 3.7%, C 44.4%, N 51.9%

Page 102

20. (a) 41.1% (b) 62.9% (c) 53.1%

Page 103

21. C_2OH_2
22. C_3OH_3
23. $CaSO_4.2H_2O$
24. $MgSO_4.7H_2O$

Page 105

25. $C_4H_{12}N_2$
26. ethyne 26.0, benzene 78.0, formaldehyde 30.0, ethanoic acid 60.0, glucose 180
    Molecular formulas are whole number multiples of the corresponding empirical formulas.
27. $C_{10}H_{14}N_2$
28. $C_8N_4O_2H_{10}$
29. $C_5H_{10}O_2$

Multiple choice questions

1. D 2. B 3. D 4. D 5. B 6. D 7. D 8. B 9. A
10. B 11. A 12. B

Review questions

1. (b) 14 (c) 28, 29, 30 (d) 28.1 (28 to 2 significant figures)
2. 55%
3. (a) 6.0 mol (d) 2.0 mol (f) 190
   (b) 24.0 mol (e) 3.61×10^{24} (g) 84.0 g
   (c) 8.0 mol
4.

Name	Formula	Molar mass (*M*)	Moles (*n*)	Mass (*m*)
sodium hydroxide	NaOH	40.0 g mol^{-1}	0.0853 mol	3.41 g
carbon tetrachloride	CCl_4	154.0 g mol^{-1}	1.40 mol	216 g
sodium carbonate	Na_2CO_3	106.0 g mol^{-1}	1.00 mol	106 g
potassium chloride	KCl	74.6 g mol^{-1}	0.25 mol	19 g
ammonium phosphate	$(NH_4)_3PO_4$	149.0 g mol^{-1}	0.0568 mol	8.46 g

5. (a) 0.562 mol; 1.35×10^{24} total atoms
   (b) 0.011 mol; 7.2×10^{22} total atoms
   (c) 10.5 mol; 5.71×10^{25} total atoms
   (d) 0.20 mol; 3.5×10^{23} total atoms
   (e) 0.26 mol; 3.2×10^{23} total atoms
6. (a) 8.93 kg (b) 4.5 g (c) 195 g
7. (a) 3.1×10^{-4} mol (b) 1.9×10^{20}
8. (a) 3.4×10^{-4} mol (b) 2.1×10^{20} (c) 50 kg
9. (a) 2.841×10^{-3}mol (b) 1.71×10^{21}
10. (a) 0.648 mol (c) 210 g
    (b) 17.5 mol (d) 10.4 g
11. (a) 1.00×10^{22} (b) 3.16×10^{22}
12. (a) C = 40.0%; O = 53.3%; H = 6.7%
    (b) Mg = 12.0%; Cl = 34.9%; O = 47.2%; H = 5.9%
    (c) Fe = 27.9%; S = 24.1%; O = 48.0%
13. (a) 77.79% (b) 58.73% (c) 46.83%
14. (a) N = 46.7%; O = 53.3%
    (b) N = 30.4 %; O = 69.6%
    (c) N = 30.4%; O = 69.6%
    (d) N = 63.6%; O = 36.4%
15. (a) 309.0 g mol^{-1}
    (b) C = 81.6%; H = 8.7%; N = 4.5%; O = 5.2%
16. (a) N in ammonium nitrate = 35.0%; N in urea = 46.7%
    (b) Not all compounds with a high nitrogen content would be suitable for use as fertilisers because they may be insoluble, toxic or unable to be taken up by plant roots.
17. $C_4H_5N_2O$
18. $C_3H_7NO_2S$
19. COH_4
20. $Na_2B_4O_7$
21. empirical formula: NaS_2O_3; molecular formula: $Na_2S_4O_6$
22. empirical formula: CH_2O; molecular formula: $C_3H_6O_3$
23. $C_2H_4Br_2$

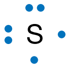
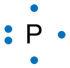
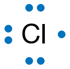
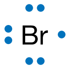
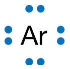
Exam practice questions

1. C
2. (a) 114.0 (c) 1.3×10^{22} (e) 15.8%
   (b) 365 g (d) 6.9×10^{23} (f) 126 g
3. empirical formula: C_2H_4O; molecular formula: $C_4H_8O_2$
4. Cu_2O
5. $ZnSO_4.7H_2O$
6. $C_{19}H_{38}O$

Chapter 6

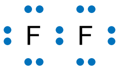
Page 114

1.

Group	Element	Valence electrons	Electron dot diagram	Lone pairs	Bonding electrons
16	S	6		2	2
14	Si	4		0	4
15	P	5		1	3
17	Cl	7		3	1
17	Br	7		3	1
16	Se	6		2	2
18	Ar	8		4	0

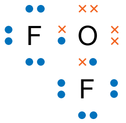
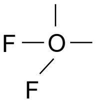
Page 116

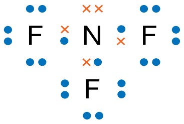
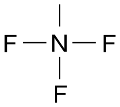
2. (a)  This is a triple covalent bond.
   (b) 2 (c) N≡N

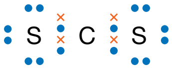
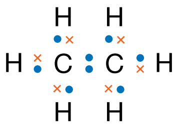
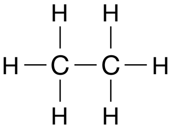
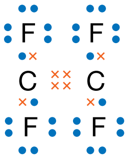
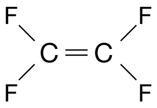
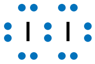
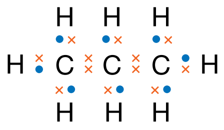
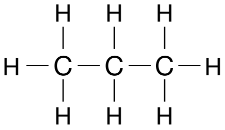
3.  F—F

4. iodine

Page 117

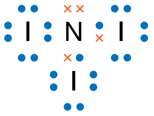
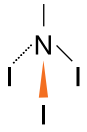
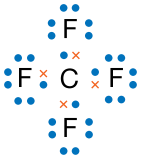
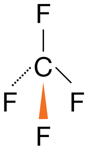
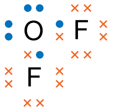
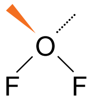
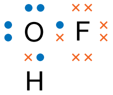
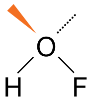
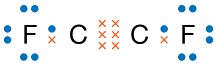
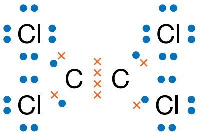
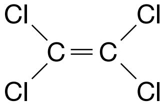
5. (a)  H—Br (b)  

6.  

7. (a)  S=C=S
   (b)  H—Cl
   (c)  
   (d)  
   (e)  I—I
   (f)  
   (g)  H—C≡N

Page 119

8.

Compound	Electron dot diagram of molecule	Number of lone pairs around central atom	Number of bonding pairs around central atom	Shape diagram and name of shape
NI_3		1	3	 pyramidal around N atom
CF_4		0	4	 tetrahedral around C atom
OF_2		2	2	 V-shaped around O atom
HOF		2	2	 V-shaped around O atom
C_2F_2		0 (for each atom)	2 (for each atom)	F—C≡C—F linear around each C atom
C_2Cl_4		0 (for each atom)	3 (for each atom)	 planar around each C atom

Page 121

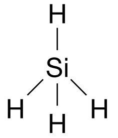
9. nitrogen trihydride, dihydrogen dioxide, trioxygen
10. (a) nitrogen trichloride
    (b) sulfur hexafluoride
    (c) oxygen dichloride
    (d) silicon tetrahydride
    (e) phosphorus trihydride
11. (a) CO (b) $SiCl_4$ (c) P_2O_5 (d) NO (e) SO_3
12. from top to bottom: N_2O; hydrogen chloride; tetraphosphorus decoxide; HF

Page 123

13. (a) H$\overset{\longrightarrow}{\text{—}}$F, polar (c) C$\overset{\longleftarrow}{\text{—}}$H polar
    (b) O$\overset{\longleftarrow}{\text{—}}$H, polar (d) N$\overset{\longleftarrow}{\text{—}}$H, polar (e) C—C, non-polar
14. (a) polar covalent (b) ionic (c) non-polar covalent
15. N—N, Cl—N, F—N, Li—F

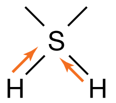
Page 124

16. (a) H$\overset{\longrightarrow}{\text{—}}$I polar (b)



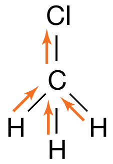
non-polar

(c) S$=$C$=$S non-polar (d)

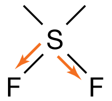


polar

(e)

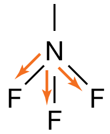


polar (f)

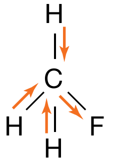


polar

(g)

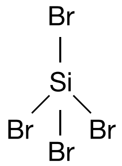


polar (h)



polar

(i)



non-polar (j) H $\overset{\longrightarrow}{=}$ Cl polar

(k) N≡N non-polar

Page 128

17. (a) dispersion forces (b) dispersion forces
    (c) dispersion and dipole–dipole forces
    (d) dispersion and dipole–dipole forces
    (e) dispersion forces and hydrogen bonding
18. Although dispersion and dipole–dipole intermolecular forces exist in all three substances, as the number of electrons in each molecule increases (from 18 in HCl, to 36 in HBr, and 54 in HI) so does the strength of the dispersion forces. This is reflected in the increasing boiling points.

Page 129

19. Cl_2 has the higher boiling point as dispersion forces increase with increasing number of electrons.
20. Helium atoms have only two electrons while argon atoms have 18. Thus the dispersion forces acting on argon atoms are stronger and its boiling point is higher.
21. HBr is a polar molecule and experiences dipole–dipole forces as well as dispersion forces, while Kr experiences only dispersion forces.
22. HF has hydrogen bonding, which is stronger than other dipole–dipole forces.
23. Glucose dissolves in water as both glucose and water are polar substances (like dissolves like). It does not conduct electricity in solid or aqueous form as it has no free ions or electrons.
24. They are all gases as only weak dispersion forces act between the molecules.
25. I_2 is a non-polar substance and therefore will dissolve in non-polar carbon tetrachloride more readily than in polar water.
26. (a) Candle wax is relatively soft and can be scratched as it is a non-polar covalent molecular compound with weak dispersion forces between its molecules.
    (b) It is not soluble in water as it is non-polar.
    (c) It is a non-conductor as it has no free-moving charged particles (electrons or ions).

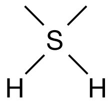
Multiple choice questions

1. D 2. A 3. A 4. A 5. D 6. B 7. C 8. C 9. D
10. A 11. B 12. B 13. A 14. B 15. A

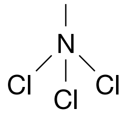
Review questions

1. (a) They already have complete outer shells of electrons.
2. (a) A fluorine molecule is composed of two fluorine atoms covalently bonded together. Each fluorine atom shares one electron pair so that both atoms attain a full outer shell of eight electrons, complying with the octet rule.
   (b) A hydrogen molecule is composed of two hydrogen atoms covalently bonded together and sharing two electrons between them. The octet rule is not used as only the first shell is involved, needing just two electrons to be full and therefore stable.
3. (a) group 14 (c) group 17
   (b) group 18 (d) group 15

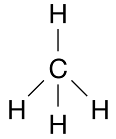
4. (a)



(b)

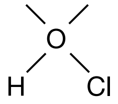


(c)

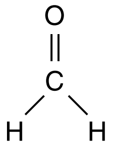


(d) Br—Br

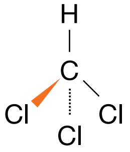
(e)



(f)

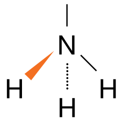


7. (a) O$=$O linear (b)



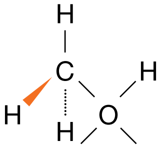
tetrahedral

(c) Br—Br linear (d)



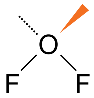
pyramidal

(e)

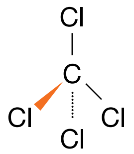


tetrahedral around C atom,
V-shaped around O atom

(f) H—C≡N linear (g)

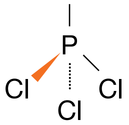


V-shaped

(h)


tetrahedral

(i) $H-C\equiv C-H$ linear (around each C atom)

(j)


pyramidal

(k) $N\equiv N$ linear

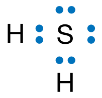
8. (a) tetrahedral
   (b) pyramidal
   (c) linear
   (d) V-shaped
   (e) pyramidal
   (f) planar around both carbon atoms
9. (a) non-polar
   (b) polar
   (c) polar
   (d) polar
   (e) ionic
10. (b) Each atom has the same attraction for the bonding pair of electrons.
    (c) Polarity depends on the shape of the molecule as well as bond dipoles. If the bond dipoles cancel each other out due to the shape of the molecule, it will be a non-polar molecule.
11. In question 7: b, d, e, f, g and j are polar.
    In question 8: b, c, d and e are polar.
12. (a) polar
    (b) non-polar
    (c) non-polar
    (d) polar
    (e) polar
    (f) non-polar
13. Ammonia, NH_3, has a molecular dipole due to its pyramidal shape. Carbon dioxide, CO_2, is non-polar because its shape is linear and the bond dipoles cancel each other out.
14. (a) Iodine has the higher number of electrons and hence stronger dispersion forces.
    (b) Fluorine is a gas at room temperature due to its weak dispersion forces. Iodine is a solid at room temperature as it has stronger dispersion forces.
15. CH_3F has dispersion and dipole–dipole forces acting between its molecules; however, dispersion forces and the stronger hydrogen bonding between molecules of CH_3OH cause it to have a higher boiling point.
16. (a) OCS has the stronger intermolecular forces (dispersion and dipole–dipole forces) while non-polar CO_2 has dispersion forces only.
    (b) HF has both dispersion forces and hydrogen bonding, which is stronger than other dipole–dipole forces, while HBr has dispersion and dipole–dipole forces.
17. $Ar < F_2 < HCl$
18. (a), (e), (f), (g), (h)
19. Although both methanol and ethanol have hydrogen bonding and dispersion forces, ethanol has the higher boiling point as it has more electrons and therefore stronger dispersion forces.
20. HCl is a covalent molecular compound. It has no free-moving charged particles (electrons or ions) to allow it to conduct electricity in the liquid state. However, when dissolved in water, it ionises to produce free mobile ions that can conduct electricity.
21.

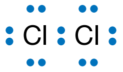
		Electrical conductivity		
Substance	**Chemical Formula**	**Solid**	**Molten**	**Dissolved in water**
magnesium nitrate	$Mg(NO_3)_2$	no	yes	yes
copper	Cu	yes	yes	insoluble
neon	Ne	no	no	no
methanol	CH_3OH	no	no	no
chlorine	Cl_2	no	no	no

22. from top to bottom: nitrogen dioxide; SO_2; CO; O_3; phosphorus; NH_3; methane (carbon tetrahydride); H_2S
23. (a) nitrogen tribromide
    (b) dinitrogen tetroxide
    (c) dinitrogen monoxide
    (d) dinitrogen pentoxide
    (e) dinitrogen trioxide
    (f) phosphorus trichloride
    (g) phosphorus pentachloride
    (h) diphosphorus pentoxide
    (i) sulfur hexafluoride
    (j) sulfur dichloride

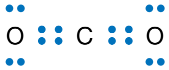
Exam practice questions

1. C 2. B 3. A

4. (a)


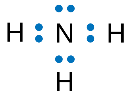
   (b)


   (c)

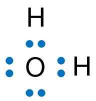

5. (a)


(b) linear (c) 4 (d) polar

   (e) non-polar (f) dispersion forces



6. (a) pyramidal
   (b) Nitrogen is more electronegative than hydrogen so the bonds are polar, leading to a polar molecule because of the overall uneven charge distribution.
   (c) While there are both dispersion forces and hydrogen bonds between ammonia molecules, ammonia exists as a gas at room temperature because these forces of attraction can be overcome by the kinetic energy (energy of motion) of the ammonia molecules.
7. (a) The lone pairs repel the bonding pairs of electrons.



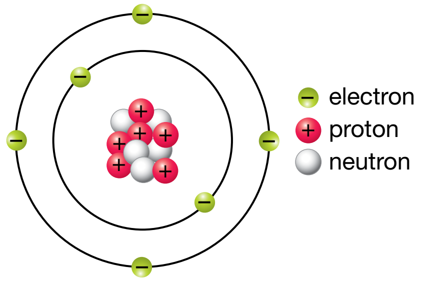
   (b) covalent intramolecular bonding, hydrogen bonding and dispersion forces
   (c) The density of ice is less than the density of water because the regular arrangement of the water molecules in ice takes up more space than the irregular arrangement of water molecules in liquid water.

Chapter 7

Page 136

1. (a) group 14, period 2

   (b)



   (c) $1s^2 2s^2 2p^2$

2. Carbon has four unpaired electrons, which can combine with other carbon atoms or other elements to form covalently bonded small or large molecules.

Page 139

3. Diamond is a network lattice with 4 carbon atoms covalently bonded in a tetrahedral shape around a central carbon atom. Graphite has carbon atoms in layers of hexagonal rings held together by dispersion forces, which means they are further apart than covalent bonds and so graphite is less dense.
4. (a) Graphite can be used for pencils as it has a covalent layer lattice structure. There are only weak dispersion forces between the layers, allowing them to slide over each other onto the paper.
   (b) Diamond has a covalent network structure with strong covalent bonding in three dimensions. This makes it very hard and hence it is useful for drilling.
5. There are no free electrons in the covalent network structure of diamond. Graphite, however, has delocalised electrons that can move across its layers under an electric field.

Page 141

7. Graphite is made of hundreds of thousands of layers of carbon rings. Graphene consists of single layers. Electrons can travel along a single plane more easily than through the many planes in a graphite lattice. Graphite is quite brittle whereas graphene is one of the strongest materials known. The properties of materials are very different when they are at the nanoscale level.
8. Carbon atoms are arranged in a sphere in buckyballs; carbon atoms are arranged in a cylinder made of sheets of hexagonal rings in nanotubes; graphene consists of flat sheets of hexagonal rings of carbon atoms.

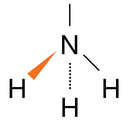
Page 143

10. (a) covalent molecular
    (b) metallic
    (c) covalent molecular
    (d) ionic
11. (a) *O* (b) *M* (c) *L* (d) *N*

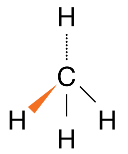
Page 145

12. (a) metallic
    (b) ionic
    (c) dispersion forces between atoms of argon
    (d) covalent
    (e) covalent
    (f) metallic
    (g) covalent
    (h) covalent
    (i) covalent
    (j) ionic
13. (a) covalent molecular
    (b) covalent network
    (c) ionic network lattice
    (d) metallic lattice
    (e) metallic lattice
    (f) covalent molecular
14. (a) covalent within the molecule; dispersion forces between molecules
    (b) covalent within the molecule; dispersion forces and hydrogen bonding between molecules
    (c) covalent within the molecule; dispersion forces between molecules
    (d) covalent within the molecule, dispersion forces and hydrogen bonding between molecules.
15. (a) O=C=O

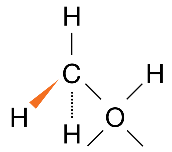
    (b)



    (c)



    (d)



16. H—F
17. (a) ionic bonding
    (b) Pressure forces like charges to align, causing repulsion and shattering of the crystal.
    (c) Solubility in water. Water molecules can move between the ions and free them by disrupting the rigid ionic lattice.
18. (a) Gold is used in jewellery as it is shiny, durable, highly ductile and malleable. These properties are due to its metallic lattice (cations in a 'sea' of delocalised electrons). The non-directional nature of the metallic bond allows distortion without disrupting the lattice, while the lustre is due to light being reflected from the delocalised electrons.
    (b) Graphite has a covalent layer lattice with weak dispersion forces between the layers. This makes it useful as a lubricant; it is slippery because the layers can slide over one another.

Multiple choice questions

1. D 2. D 3. D 4. D 5. C 6. A 7. D 8. B
9. B 10. C 11. A 12. C 13. C 14. C 15. B,C 16. D
17. B 18. D 19. D 20. C

Review questions

1. (a) Diamond has a covalent network lattice structure and graphite has a covalent layer lattice structure.
   (b) Both have strong covalent bonding between the carbon atoms in the lattice; however, graphite has strong covalent bonding in only two dimensions with weak dispersion forces between layers while diamond has strong covalent bonding in all three dimensions.
   (c) Both consist of countless atoms held together by covalent bonds.
   (d) Graphite is used as a lubricator, in pencil 'lead' and electrodes; diamond is used in jewellery, drilling and as an abrasive (cutting glass).
2. (a) lubricant — layers can slide over one another
   (b) There are only weak dispersion forces between the layers.
3. The purpose of the clay is to bind the graphite powder. The ratio of clay to graphite determines the hardness of the pencil. The more clay, the harder the pencil; the less clay, the softer the pencil.
4. The property of hardness due to the strong covalent bonding in three dimensions.

5. When enough pressure is applied to break the strong covalent bonds, the lattice is distorted and the diamond will shatter.
7. They all have a full outer shell.
8. Na, 2, 8, 1, and K, 2, 8, 8, 1, both have the same number of valence electrons (1), so their chemical properties are similar.
9. oxide O^{2-} , fluoride F^- , aluminium Al^{3+}
10. (a) ionic (b) *XY*$_2$
11. (a) C (b) E (c) A (d) D (e) B
12. (b) and (d)
13. (a) ionic bonding
    (b) covalent within molecule
    (c) ionic bonding
    (d) covalent within molecule, dispersion forces and hydrogen bonding between molecules in the solid state
    (e) covalent within molecules
    (f) covalent within molecules; dispersion forces between molecules in the solid state
    (g) covalent bonds in three-dimensional network lattice
    (h) metallic bonding
    (i) weak dispersion forces between atoms
    (j) covalent bonds within the molecule
    (k) covalent bonds within molecules, dispersion forces and hydrogen bonding between the molecules in the liquid phase
    (l) ionic bonding
    (m) covalent bonding within the molecule
    (n) ionic bonding
    (o) ionic bonding
    (p) metallic bonding
    (q) metallic bonding
14. (a) covalent molecular
    (b) metallic
    (c) ionic
    (d) covalent molecular
    (e) covalent molecular
    (f) ionic
    (g) covalent molecular
    (h) covalent molecular
    (i) ionic
    (j) covalent molecular
    (k) metallic
    (l) covalent molecular
    (m) covalent molecular
15. Discrete molecules are often volatile (evaporate) at room temperature due to weak intermolecular forces. Metallic, ionic and covalent network substances have strong bonds and remain solids at room temperature, hence do not have a smell.
16. (a) (i) covalent molecular — benzoic acid, paraffin wax and sugar; ionic — magnesium sulfate and sodium fluoride; metallic — aluminium
    (ii) aluminium
17. Ionic solids do not have free mobile electrons, whereas metals have free delocalised electrons that can move under an electric field.
18. (a) (i) ionic (ii) covalent molecular (iii) metallic
    (iv) covalent network
    (b) CH_3OH, NaF, Fe, SiO_2
19. KI, as ionic bonds are stronger and thus harder to break than the dispersion forces and hydrogen bonding between ammonia molecules.
20. (a) KF, $MgBr_2$, CaO (b) P_4O_{10}, CCl_4, C_2H_4, SiO_2
21. The metallic bonding is non-directional allowing distortion without disrupting the lattice of cations in the 'sea' of electrons. Calcium chloride is brittle because its rigid ionic lattice will shatter if pressure aligns the ions of like charge.
22. Diamond has a covalent network structure with strong covalent bonds in three dimensions. Tungsten has a metallic lattice structure consisting of close-packed cations in a 'sea' of electrons. The bonding is strong in both substances, and extreme heat is needed to break the bonds.
23. (a) covalent, covalent molecular lattice
    (b) ionic, ionic lattice
24. (a) *G* (b) *H* (c) *E*, *F*
25. *A* — covalent molecular
    B — metallic
    C — ionic
    D — covalent network

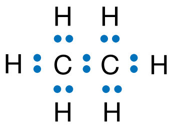
Exam practice questions

1. C 2. A 3. B 4. B
5. Silicon dioxide has a covalent network structure. Its strong three-dimensional bonding results in a rigid solid at room temperature. Carbon dioxide is a discrete covalent molecular compound. It has weak dispersion forces between its molecules and hence is a gas at room temperature.
6. (i) (a) dispersion forces (b) molecules
    (ii) (a) covalent (b) atoms
    (iii) (a) ionic (b) ions
    (iv) (a) covalent (b) atoms
    (v) (a) metallic (b) cations
    (vi) (a) metallic (b) cations
    (vii) (a) dispersion forces and hydrogen bonding (b) molecules
7. (a) HF (b) Na (c) KF (d) He (e) Si (f) H_2S
8. Graphene is a nanomaterial whereas graphite is a macromaterial.

Chapter 8

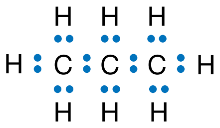
Pages 155–6

1. Ethane:



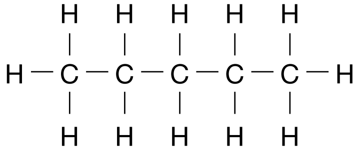
$$\begin{array}{ccccccc} & & H & & H & & \\ & & \bullet\bullet & & \bullet\bullet & & \\ H & : & C & : & C & : & H \\ & & \bullet\bullet & & \bullet\bullet & & \\ & & H & & H & & \end{array}$$

Propane:



$$\begin{array}{ccccccccc} & & H & & H & & H & & \\ & & \bullet\bullet & & \bullet\bullet & & \bullet\bullet & & \\ H & : & C & : & C & : & C & : & H \\ & & \bullet\bullet & & \bullet\bullet & & \bullet\bullet & & \\ & & H & & H & & H & & \end{array}$$

2. $C_{16}H_{34}$
3. $C_{25}H_{52}$
4.

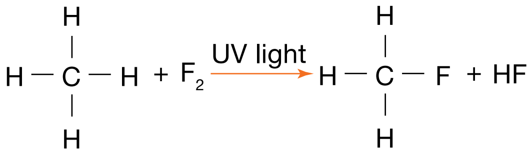


$$\begin{array}{ccccccccccccc} & & H & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & | & & | & & \\ & & H & & H & & H & & H & & H & & \end{array}$$

5. Butane is a larger molecule.

Page 156

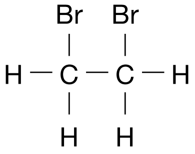
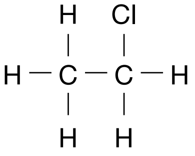
6. (a) $2C_2H_6(g) + 7O_2(g) \longrightarrow 4CO_2(g) + 6H_2O(g)$
    (b) $C_5H_{12}(g) + 8O_2(g) \longrightarrow 5CO_2(g) + 6H_2O(g)$
7.



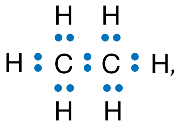
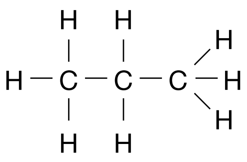
$$\begin{array}{ccccc} & & H & & \\ & & | & & \\ H & - & C & - & H \\ & & | & & \\ & & H & & \end{array} + F_2 \xrightarrow{\text{UV light}} \begin{array}{ccccc} & & H & & \\ & & | & & \\ H & - & C & - & F \\ & & | & & \\ & & H & & \end{array} + HF$$

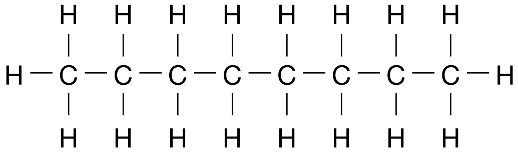
8. $2C_6H_{14}(g) + 19O_2(g) \longrightarrow 12CO_2(g) + 14H_2O(g)$

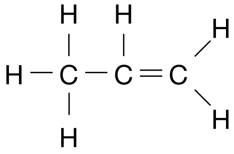
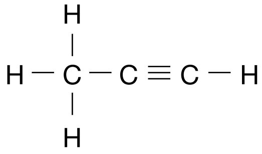
Page 157

10. three isomers
12. Methylpropane is branched and so molecules are further apart than in butane; therefore, the dispersion forces are weaker.
13. (a)  (b) 
14. A saturated compound has carbon atoms linked together by single bonds, e.g. ethane, propane. An unsaturated compound contains carbon–carbon double bonds or triple bonds, e.g. propene, 1-butene.

Page 159

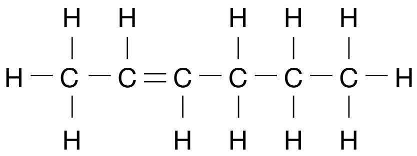
15. ethane, , tetrahedral around each carbon atom ethyne; H : C ::: C : H, linear
16. (a) 

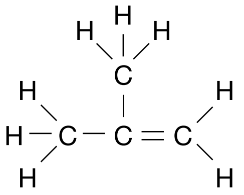
    (b) 

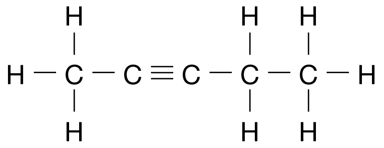
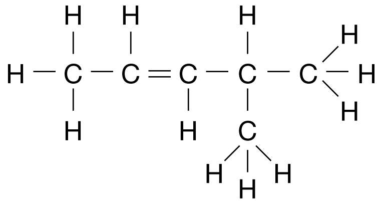
    (c)  (d) 
17. carbon dioxide and water
18. $2C_3H_6(g) + 9O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)$
19. $C_{21}H_{42}$
20. $C_{12}H_{24}$
21. the presence of the double bond
22. combustion, addition and self-addition reactions
23. (a) $C_4H_8(g) + 6O_2(g) \rightarrow 4CO_2(g) + 4H_2O(g)$

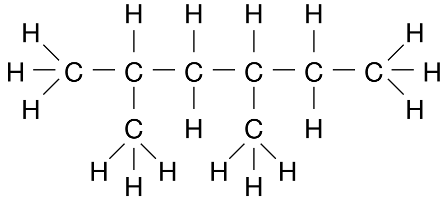
    (b) $C_6H_{12}(g) + 9O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)$
26. Dispersion forces increase with molecular size.

Page 161

27. (a) 

    (b) 

    (c) 
28. (a) 2,2-dimethylheptane (b) octane
29. (a) 

    (b) 

Page 162

30. (27) (a) $CH_3CHCHCH_2CH_2CH_3$

    (b) $CH_2CCH_3CH_3$

    (c) $CH_3CCCH_2CH_3$

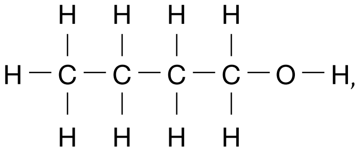
    (28) (a) $(CH_3)_3C(CH_2)_4CH_3$

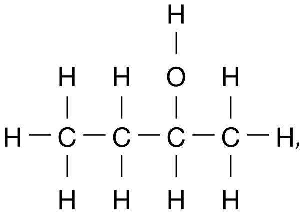
    (b) $CH_3(CH_2)_6CH_3$

    (29) (a) $CH_3CHCHCH(CH)_3CH_3$

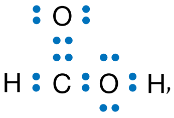
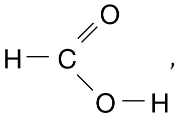
    (b) $CH_3CH(CH_3)CH_2CH(CH_3)CH_2CH_3$
31. $CH_3CH(CH_3)CH(CH_3)CH_2CH_2CH_3$

Page 163

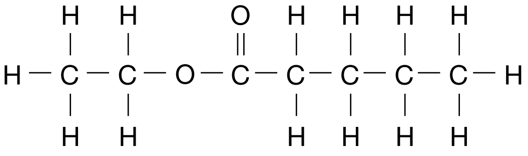
32. $C_4H_{10}O$, ,

    $CH_3CH_2CH_2CH_2OH$
33. $C_4H_{10}O$, , $CH_3CH_2CH(OH)CH_3$

Page 165

34. , , HCOOH
36. Each member of the series of carboxylic acids differs by $-CH_2$, e.g. HCOOH, CH_3COOH, CH_3CH_2COOH.

Page 166

37. ethyl ethanoate
38. 

39. The sulfuric acid acts as a catalyst. It reacts with the water formed.
40. COOH
42. Refer to table 8.2 on page 164.

Page 168

43. Lighter fractions are gases, which burn more readily than liquids.
44. The viscosity increases with increasing molecular size because of the stronger intermolecular dispersion forces and the increased possibility of entanglement of the long molecules.
45. The molecules with fewest carbon atoms in them are the smaller molecules. They are lighter and have fewer dispersion forces acting between them. They are gaseous.

Page 169

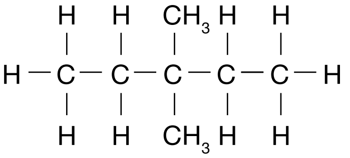
47. Catalysts increase the rate of reaction.
48. less heat required, quicker reaction time

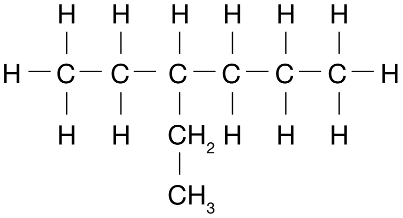
Multiple choice questions

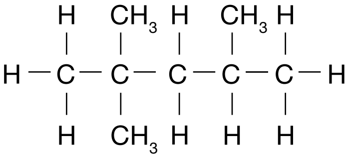
1. D 2. C 3. C 4. D 5. D 6. C 7. D 8. B 9. C
10. A, D 11. C 12. B 13. C 14. A 15. D 16. D 17. D 18. A

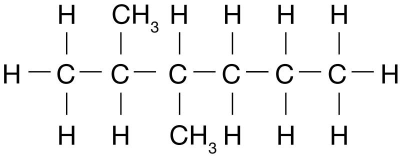
Review questions

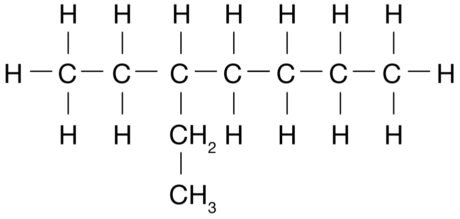
1. (a) chemical compounds that contain only carbon and hydrogen
   (b) propane, C_3H_8; propene, C_3H_6; propyne, C_3H_4
2. (b) $2C_3H_6(g) + 9O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)$
   $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
7. (a) $C_{22}H_{46}$ (b) $C_{17}H_{34}$ (c) $C_{13}H_{24}$
8. (a) 2-methylbutane
   (b) 3-methylpentane
   (c) 3-ethylhexane
   (d) 3,3-dimethylhexane
   (e) 2,4-dimethylhexane
   (f) 2,2,4,4-tetramethylpentane
9. *Note:* For structural formulas, *all* bonds need to be drawn.

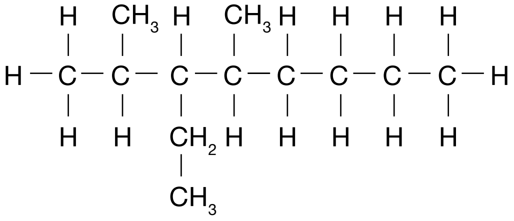
(a)

$$H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{CH_3}{\underset{CH_3}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$$

(b)

$$H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{\begin{array}{c}CH_2\\|\\CH_3\end{array}}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$$

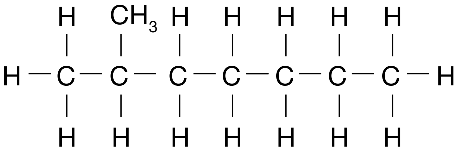
(c)

$$H-\overset{H}{\underset{H}{C}}-\overset{CH_3}{\underset{CH_3}{C}}-\overset{H}{\underset{H}{C}}-\overset{CH_3}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$$

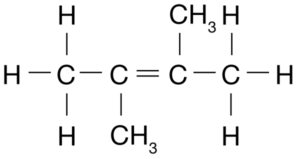
(d)

$$H-\overset{H}{\underset{H}{C}}-\overset{CH_3}{\underset{H}{C}}-\overset{H}{\underset{CH_3}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$$

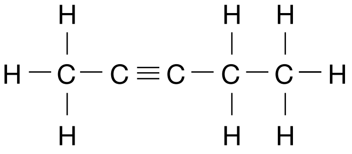
(e)

$$H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{\begin{array}{c}CH_2\\|\\CH_3\end{array}}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$$

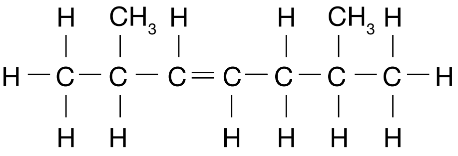
(f)

$$H-\overset{H}{\underset{H}{C}}-\overset{CH_3}{\underset{H}{C}}-\overset{H}{\underset{\begin{array}{c}CH_2\\|\\CH_3\end{array}}{C}}-\overset{CH_3}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$$

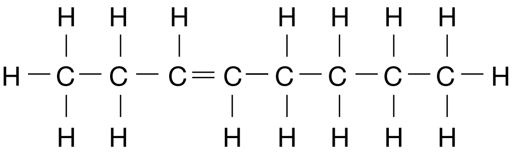
10. *Note:* For structural formulas, *all* bonds need to be drawn.

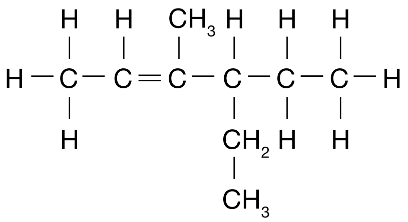
(a)

$$H-\overset{H}{\underset{H}{C}}-\overset{CH_3}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$$

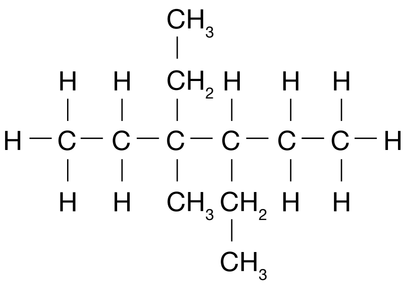
(b)

$$H-\overset{H}{\underset{H}{C}}-\underset{CH_3}{C}=\overset{CH_3}{C}-\overset{H}{\underset{H}{C}}-H$$

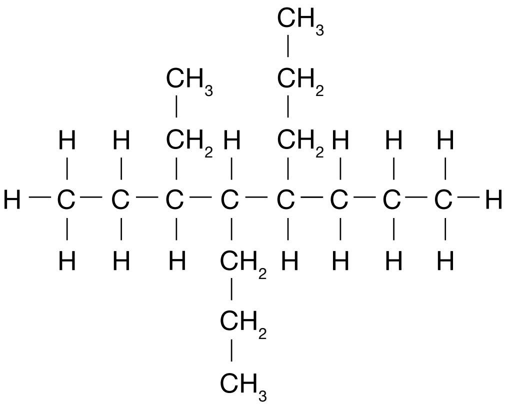
(c)

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

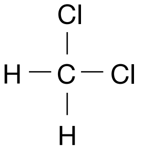
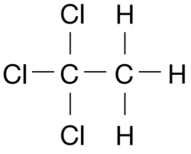
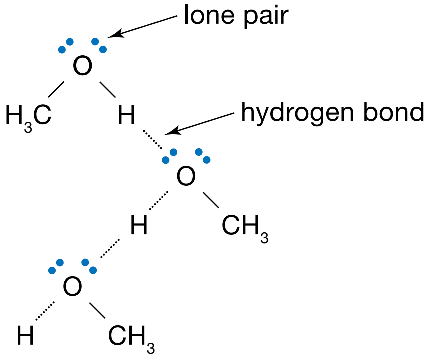
(d)

$$H-\overset{H}{\underset{H}{C}}-\overset{CH_3}{\underset{H}{C}}-\overset{H}{C}=\underset{H}{C}-\overset{H}{\underset{H}{C}}-\overset{CH_3}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$$

(e)

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

(f)

$$H-\overset{H}{\underset{H}{C}}-\overset{H}{C}=\overset{CH_3}{C}-\overset{H}{\underset{\begin{array}{c}CH_2\\|\\CH_3\end{array}}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$$

(g)


(h)


11. (a) 2,4-dimethyl-1-heptene
    (b) pent-1-yne
    (c) oct-3-ene
    (d) 3,4-dimethylpent-1-ene
    (e) 2,3-dimethylbutane
    (f) 4-ethyl-3-methylheptane
12. The longest chain has seven carbon atoms. The compound should have been named 3,3-dimethylheptane.
13. Successive members of the alcohol series differ by a CH_2 group.
15. (a)

(b)

16.


hydrogen bonding in methanol

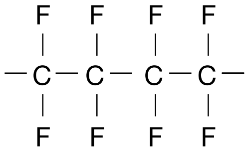
17. (b) Both ethane and ethanoic acid have 2 carbon atoms. The number of hydrogen atoms is different, and ethanoic acid contains 2 oxygen atoms. Ethane is a non-polar molecule and the atoms in the liquid state are held together by weak dispersion forces. Ethanoic acid is a polar molecule and is held together by hydrogen bonds. Ethane is insoluble and ethanoic acid is soluble in water.
18. (a) Esters have only weak dipole–dipole forces between the molecules and so are easily vaporised.
    (b) Smaller esters are soluble in water but, as the size of the non-polar section of the molecule increases, they become less soluble.
19. (b) condensation reaction
20. (a) methyl butanoate
    (b) propyl propanoate
    (c) ethyl methanoate
21. (a) $CH_3CH_2CH_2COOCH_2CH_3$
    (b) $CH_3COOCH_2CH_2CH_3$
    (c) $CH_3CH_2COOCH_2CH_2CH_2CH_3$
22. (a) $CH_3CHCHCH_3$ (c) $CH_3CH_2CH_2COOH$
    (b) $CH_3CH_2CH_2CH_2OH$ (d) $CH_2COOCH_2CH_3$
23. (a) 80.0% (b) 85.7% (c) 92.3%
24. C_3H_7
25. (a) C_3H_8 (c) alkane
26. C_2H_4O, $C_4H_8O_2$
27. propan-1-ol or propan-2-ol

Exam practice questions

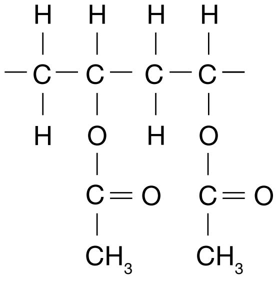
1. C
2. B
3. Boiling points reflect the strength of the intermolecular forces operating. The intermolecular forces operating are dispersion forces, which increase with molecule size.
4. (a) petrol (c) paraffin
   (b) petrol (d) further refine by heating

Chapter 9

Page 179

1. (a)


   (b) polytetrafluoroethylene, PTFE
   (c) Teflon, nonstick coating on frying pans and saucepans

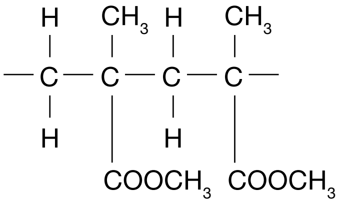
2. (a)


   (b) polyvinylacetate (c) glue

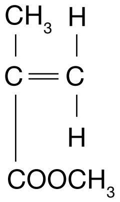
Multiple choice questions

1. A 2. C 3. B 4. C 5. C 6. B 7. C 8. A 9. C 10. B
11. A 12. B 13. C

Review questions

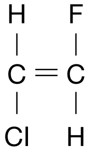
3. (a)


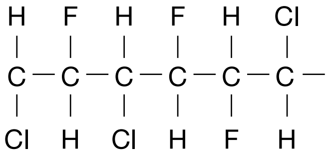
4. See figures (a) and (b) on page 184.

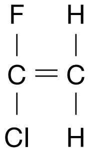
9. (a)


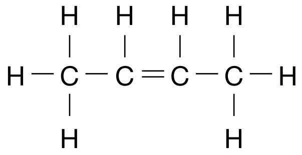
(b) dipole–dipole, dispersion

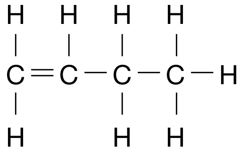
Exam practice questions

1. (a)


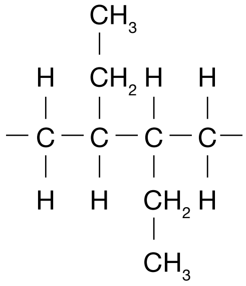
(b)


(c)


2. (a) (i)


(ii)


(b) but-2-ene

(c)


Chapter 10

Page 200

4. (a) 8×10^6
   (b) 1.3×10^{-3}
   (c) 3.844×10^8
5. (a) 6.942×10^3
   (b) 2.48×10^2
   (c) 4.9×10^{-4}
   (d) 3.17
   (e) 8.2×10^{-3}
   (f) 6.45×10^4
6. (a) 0.000 721
   (b) 573 900
   (c) 15.52
   (d) 7.112
   (e) 471
   (f) 0.000 002 577
   (g) 0.000 091 79

Pages 200–1

7. 11.43 mL
8. The reading is between 7 and 8 and is 7.33 mL.

Page 202

9. (a) 3 (b) 5 (c) 3 (d) 3
10. (a) 13 (b) 3.7 (c) 76.075 (d) 0.5 (e) 213 (f) 18.7
11. (a) 23.2 g mL^{-1} (b) 28.78 g (c) 3.2 mol (d) 29 g

Chapter 11

Page 209

1. (a) $Al(s) + O_2(g) \rightarrow Al_2O_3(s)$
   (b) $KClO_3(s) \xrightarrow{MnO_4} O_2(g) + KCl(s)$
2. (a) When solutions of lead(II) nitrate and sulfuric acid are mixed, solid lead(II) sulfate and aqueous nitric acid are produced.
   (b) When ammonia gas and oxygen gas react in the presence of a platinum catalyst, nitrogen monoxide gas and water vapour are produced.

Page 211

3. (a) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
   (b) $SiO_2(s) + 2C(s) \rightarrow Si(s) + 2CO(g)$
   (c) $4FeO(s) + O_2(g) \rightarrow 2Fe_2O_3(s)$
   (d) $16Cr(s) + 3S_8(s) \rightarrow 8Cr_2S_3(s)$
   (e) $2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$
4. $SiCl_4(l) + 2Mg(s) \rightarrow Si(s) + 2MgCl_2(s)$
5. (a) Hydrogen forms H_2 (rather than H_4) molecules.
   $CH_4 + O_2 \rightarrow CO_2 + 2H_2$
   (b) Iodochloride has the formula ICl.
   $I_2 + Cl_2 \rightarrow 2ICl$
   (c) Potassium oxide has the formula K_2O.
   $4K + O_2 \rightarrow 2K_2O$
   (d) Sodium chloride has the formula NaCl; bromine has the formula Br_2.
   $Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$

Page 213

6. (a) insoluble
   (b) soluble
   (c) insoluble
   (d) soluble
7. (a) all (c) nitrate and nitrite
   (b) none
8. (a) $2KOH(aq) + Ca(NO_3)_2(aq) \rightarrow Ca(OH)_2(s) + 2KNO_3(aq)$
   (b) $Na_2S(aq) + Pb(CH_3COO)_2(aq) \rightarrow 2NaCH_3COO(aq) + PbS(s)$
   (c) $2(NH_4)_3PO_4(aq) + 3CaCl_2(aq) \rightarrow 6NH_4Cl(aq) + Ca_3(PO_4)_2(s)$

Page 214

9. (a) $LiOH(aq) + HCl(aq) \rightarrow LiCl(aq) + H_2O(l)$
   (b) $NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$
   (c) $KOH(aq) + HCl(aq) \rightarrow KCl(aq) + H_2O(l)$

10. (a) $HNO_3(aq) + LiOH(aq) \rightarrow LiNO_3(aq) + H_2O(l)$
    (b) $H_2SO_4(aq) + Na_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g) + Na_2SO_4(aq)$
    (c) $H_2SO_4(aq) + CuO(s) \rightarrow CuSO_4(s) + H_2O(l)$
    (d) $2HCl(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)$

Page 215

11. (a) $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
    (b) $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$
    (c) $2C_2H_6(g) + 5O_2(g) \rightarrow 4CO(g) + 6H_2O(g)$
12. (a) very limited oxygen supply
    (b) $2C_4H_{10}(g) + 5O_2(g) \rightarrow 8C(s) + 10H_2O(g)$
    (c) $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$

Multiple choice questions

1\. A 2. D 3. C 4. C 5. B 6. D 7. B 8. D

Review questions

1. (a) Something appears to have changed visually; energy loss or gain is involved.
   (b) Reactants can be reobtained after a physical change; new substances are produced in chemical changes.
2. colour change (e.g. rusting); temperature change (e.g. rotting compost); light emitted (e.g. firefly); odour produced (e.g. unleaded petrol combustion); gas produced (e.g. fizzy tablet dissolved in water)
3. A chemical equation is a shorthand representation of a chemical reaction that must be balanced to show how the atoms have been rearranged to change the reactants into products. (Matter is not created or destroyed).
4. The law of conservation of mass applies to chemical equations since the total numbers and types of atoms in the reactants must equal the total numbers and types of atoms in the products.
5. (a) states of reactants and products; identification of reactants and products
   (b) how much of each reactant and product was involved in the reaction (i.e. reaction quantities or reaction conditions, such as temperature and catalyst); rate of reaction
6. (a) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
   (b) $6HCl(aq) + 2Al(s) \rightarrow 2AlCl_3(aq) + 3H_2(g)$
   (c) $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$
   (d) $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$
   (e) $12Na(s) + P_4(s) \rightarrow 4Na_3P(s)$
7. (a) $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$
   (b) $HCl(g) + NH_3(g) \rightarrow NH_4Cl(s)$
   (c) $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$
   (d) $2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$
8. (a) $2Cu + O_2 \rightarrow 2CuO$
   (b) $2HgO \rightarrow 2Hg + O_2$
   (c) $2AsCl_3 + 3H_2S \rightarrow As_2S_3 + 6HCl$
   (d) $Fe_2O_3 + 3H_2 \rightarrow 2Fe + 3H_2O$
   (e) $2NaCl \rightarrow 2Na + Cl_2$
   (f) $2Al + 3H_2SO_4 \rightarrow 3H_2 + Al_2(SO_4)_3$
   (g) $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$
9. (a) $2HNO_3(aq) + Zn(s) \rightarrow Zn(NO_3)_2(aq) + H_2(g)$
   (b) $2HCl(aq) + MgO(s) \rightarrow H_2O(l) + MgCl_2(aq)$
   (c) $2HCl(aq) + Ca(OH)_2(s) \rightarrow 2H_2O(l) + CaCl_2(s)$
   (d) $H_2SO_4(aq) + Na_2CO_3(s) \rightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$
   (e) $2Ca(s) + O_2(g) \rightarrow 2CaO(s)$
10. (a) $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(s) + H_2(g)$
    (b) $2Ba(s) + O_2(g) \rightarrow 2BaO(s)$
    (c) $2HCl(aq) + Mg(OH)_2(s) \rightarrow MgCl_2(aq) + 2H_2O(l)$
    (d) $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$
11. $Cl_2(g) + 2KI(aq) \rightarrow I_2(s \text{ or } aq) + 2KCl(aq)$
12. (a) Potassium metal reacts with water to produce a solution of potassium hydroxide, with the evolution of hydrogen gas.
    (b) Carbon can react with iron(II) oxide, when heated, to produce metallic iron and carbon monoxide gas.
13. (a) $3Ca(OH)_2(s) + 2H_3PO_4(aq) \rightarrow Ca_3(PO_4)_2(s) + 6H_2O(l)$
    (b) $3H_2SO_4(aq) + 2Al(OH)_3(s) \rightarrow Al_2(SO_4)_3(aq) + 6H_2O(l)$
    (c) $Fe(OH)_3(s) + 3HCl(aq) \rightarrow FeCl_3(aq) + 3H_2O(l)$
    (d) $BaCl_2(aq) + (NH_4)_2CO_3(aq) \rightarrow BaCO_3(s) + 2NH_4Cl(aq)$
    (e) $2AgNO_3(aq) + H_2S(g) \rightarrow Ag_2S(s) + 2HNO_3(aq)$
    (f) $K_2CrO_4(aq) + Pb(NO_3)_2(aq) \rightarrow 2KNO_3(aq) + PbCrO_4(s)$
14. (a) $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$
    $2C_3H_6(g) + 9O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)$
    $C_4H_8(g) + 6O_2(g) \rightarrow 4CO_2(g) + 4H_2O(g)$
    (b) $C_2H_4(g) + 2O_2(g) \rightarrow 2CO(g) + 2H_2O(g)$
    $C_3H_6(g) + 3O_2(g) \rightarrow 3CO(g) + 3H_2O(g)$
    $C_4H_8(g) + 4O_2(g) \rightarrow 4CO(g) + 4H_2O(g)$
15. (a) pentane: $C_5H_{12}(g) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(g)$
    2-pentene: $2C_5H_{10}(g) + 15O_2(g) \rightarrow 10CO_2(g) + 10H_2O(g)$
    2-pentyne: $C_5H_8(g) + 7O_2(g) \rightarrow 5CO_2(g) + 4H_2O(g)$
    (b) pentane: $2C_5H_{12}(g) + 11O_2(g) \rightarrow 10CO(g) + 12H_2O(g)$
    2-pentene: $C_5H_{10}(g) + 5O_2(g) \rightarrow 5CO(g) + 5H_2O(g)$
    2-pentyne: $2C_5H_8(g) + 9O_2(g) \rightarrow 10CO(g) + 8H_2O(g)$
16. (a) $2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$
    (b) $2C_8H_{18}(l) + 17O_2(g) \rightarrow 16CO(g) + 18H_2O(g)$
17. (a) $2C_3H_8O_3(l) + 7O_2(g) \rightarrow 6CO_2(g) + 8H_2O(g)$
    (b) $C_{12}H_{22}O_{11}(l) + 12O_2(g) \rightarrow 12CO_2(g) + 11H_2O(g)$
    (c) $CH_3COOH(l) + 2O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$

Exam practice questions

1. (a) $BaSO_4$
   (b) $Ba(OH)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaOH(aq)$
2. (a) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
   (b) $C_3H_8(g) + 5O_2\ (g) \rightarrow 3CO_2\ (g) + 4H_2O(g)$
   (c) $2C_3H_8(g) + 7O_2(g) \rightarrow 6CO(g) + 8H_2O(g)$
   (d) CO is poisonous.
   (e) C_3H_8. Three molecules of CO_2 are produced with C_3H_8 compared with one for CH_4.
   (f) C_4H_{10}

Chapter 12

Page 226

2. Heat is taken from the body for water to change from a liquid to a gaseous state.

Page 231

3. (a) soluble
   (b) soluble
   (c) insoluble
   (d) insoluble
   (e) soluble
   (f) insoluble
4. because of the polarity of the water molecules

Page 232

5. (a) $NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$
   (b) $Na_3PO_4(s) \rightarrow 3Na^+(aq) + PO_4^{3-}(aq)$
6. (a) $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$
   (b) $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$

Multiple choice questions

1\. C 2. D 3. A 4. B 5. C 6. A 7. B
8\. A 9. C 10. B 11. A or C

Review questions

4. (a) energy required for a change of state
   (b) energy required to increase temperature
   (c) Both values are high.
6. (a) Potassium chloride is ionic and water is polar.
   (b) Ammonia and water are both polar.
   (c) Methane is non polar but water is polar.
7. (a), (b), (f), (g), (h)
8. (a) yes (c) no (e) yes (g) no
   (b) yes (d) no (f) yes (h) no
9. (a) $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
   (b) $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$
   (e) $3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$
   (f) $Pb^{2+}(aq) + S^{2-}(aq) \rightarrow PbS(s)$
   (h) $Hg^{2+}(aq) + 2Br^-(aq) \rightarrow HgBr_2(s)$
10. (a) $NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$
    $OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
    (b) $K_2SO_4(aq) + Ca(NO_3)_2(aq) \rightarrow 2KNO_3(aq) + CaSO_4(s)$
    $SO_4^{2-}(aq) + Ca^{2+}(aq) \rightarrow CaSO_4(s)$
    (c) $Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow 2KNO_3(aq) + PbI_2(s)$
    $Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$
    (d) $HCl(aq) + NaHCO_3(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$
    $H^+(aq) + HCO_3^-(aq) \rightarrow CO_2(g) + H_2O(l)$
    (e) $3MgCl_2(aq) + 2Na_3PO_4(aq) \rightarrow 6NaCl(aq) + Mg_3(PO_4)_2(s)$
    $3Mg^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Mg_3(PO_4)_2(s)$
    (f) $K_2S(aq) + ZnCl_2(aq) \rightarrow ZnS(s) + 2KCl(aq)$
    $S^{2-}(aq) + Zn^{2+}(aq) \rightarrow ZnS(s)$
    (g) $(NH_4)_2CO_3(aq) + 2HNO_3(aq) \rightarrow$
    $2NH_4NO_3(aq) + H_2O(l) + CO_2(g)$
    $CO_3^{2-}(aq) + 2H^+(aq) \rightarrow H_2O(l) + CO_2(g)$
    (h) no reaction
11. (a) Ba^{2+} (c) Ba^{2+}
    (b) Pb^{2+} (d) Ba^{2+}
12. (a) Chlorides, bromides and iodides will precipitate silver but not zinc.
    (b) Sulfates precipitate barium but not magnesium.
    (c) Chlorides precipitate silver but not lead in hot water (see page 212).
    (d) Barium precipitates sulfates but not chlorides.

Exam practice questions

1. (a) the one with barium chloride
   (b) $BaCl_2(aq) + (NH_4)_2CO_3(aq) \rightarrow BaCO_3(s) + 2NH_4Cl(aq)$
   (c) $Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$
   (d) Spectator ions are present in the reaction but do not take part in the formation of the product. $Cl^-(aq)$ and $NH_4^+(aq)$
   (e) $NH_4Cl(s)$ is formed. As the water is evaporated, the solution forms solid crystals of ammonium chloride.
2. (a) Specific heat capacity is the quantity of heat energy that must be added to a specified mass of substance to raise it by a specified temperature.
   (b) sand
   (c) 11 200 J = 11 kJ
3. (a) Dissociation occurs when ions that are already present move apart.
   (b) Ionisation is the production of ions.
   (c) Sodium chloride can conduct electricity because its ions dissociate and become mobile.
   (d) Hydrochloric acid conducts electricity because it ionises in water to produce mobile ions.

Chapter 13

Page 240

1. $2HCl(aq) + Zn(s) \rightarrow ZnCl_2(aq) + H_2(g)$
   $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
2. (a) $H_2SO_4(aq) + Mg(s) \rightarrow MgSO_4(aq) + H_2(g)$
   $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$
   (b) $3H_2SO_4(aq) + 2Al(s) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$
   $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$
   (c) $2HNO_3(aq) + CuCO_3(s) \rightarrow Cu(NO_3)_2(aq) + CO_2(g) + H_2O(l)$
   $2H^+(aq) + CuCO_3(s) \rightarrow Cu^{2+}(aq) + CO_2(g) + H_2O(l)$
   (d) $H_2SO_4(aq) + 2KHCO_3(s) \rightarrow K_2SO_4(aq) + 2CO_2(g) + 2H_2O(l)$
   $2H^+(aq) + 2KHCO_3(s) \rightarrow 2K^+(aq) + 2CO_2(g) + 2H_2O(l)$
   (e) $2HNO_3(aq) + Na_2SO_3(aq) \rightarrow 2NaNO_3(aq) + SO_2(g) + H_2O(l)$
   $2H^+(aq) + SO_3^{2-}(aq) \rightarrow SO_2(g) + H_2O(l)$
   (f) $2H_3PO_4(aq) + 3PbS(s) \rightarrow Pb_3(PO_4)_2(aq) + 3H_2S(g)$
   $2H^+(aq) + PbS(s) \rightarrow Pb^{2+}(aq) + H_2S(g)$
   (g) $2HCl(aq) + CuO(s) \rightarrow CuCl_2(aq) + H_2O(l)$
   $2H^+(aq) + CuO(s) \rightarrow Cu^{2+}(aq) + H_2O(l)$
   (h) $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + H_2O(l)$
   $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

Pages 241–2

3. (a), (b), (d), (e), (g)

Page 243

4. example: $H_2SO_4(aq) + 2NH_4OH(aq) \rightarrow (NH_4)_2SO_4(aq) + 2H_2O(l)$

Page 245

5. (a) HSO_4^- (b) HS^- (c) S^{2-} (d) OH^- (e) NH_3
6. (a) H_2O (b) H_2CO_3 (c) HS^- (d) H_3O^+ (e) HCN
7. (c)
8. (a) $H_3PO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + H_2PO_4^-(aq)$
   $H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HPO_4^{2-}(aq)$
   $HPO_4^{2-}(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + PO_4^{3-}(aq)$
   (b) $H_3PO_4(aq)/H_2PO_4^-(aq)$; $H_3O^+(aq)/H_2O(l)$
   $H_2PO^{4-}(aq)/HPO_4^{2-}(aq)$; $H_3O^+(aq)/H_2O(l)$
   $HPO_4^{2-}(aq)/PO_4^{3-}(aq)$; $H_3O^+(aq)/H_2O(l)$
9. (a) $HS^-(aq) + H_2O(l) \rightarrow H_3O^+(aq) + S^{2-}(aq)$
   $HS^-(aq) + H_2O(l) \rightarrow OH^-(aq) + H_2S(aq)$
   (b) $HSO_4^-(aq) + H_2O(l) \rightarrow H_3O^+(aq) + SO_4^{2-}(aq)$
   $HSO_4^-(aq) + H_2O(l) \rightarrow OH^-(aq) + H_2SO_4(aq)$

Page 248

10. $HNO_3(aq) + H_2O(l) \rightarrow H_3O^+(aq) + NO_3^-(aq)$
    strong conductor of electricity
    $HS^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + S^{2-}(aq)$
    weak conductor of electricity
11. $(CH_3)_2NH + H_2O(l) \rightleftharpoons (CH_3)_2NH_2^+(aq) + OH^-(aq)$
12. (a) HNO_3/NO_3^-; HCl/Cl^-
    (b) Wear appropriate personal protective clothing.

Page 251

13. pH is a measure of how acidic or basic a solution is. It is calculated using $pH = -\log_{10}[H_3O^+]$.
14 A dilute solution has a small amount of acid or base dissolved per unit of volume. A concentrated solution has a large amount.
15. (a) pH = 5 (b) pH = 1
16. pH = 3
17. (a) pH =10 (b) pH = 13
18. pH = 3
19. (a) 3.0
    (b) 9.3
20. At 25 °C, $[H_3O^+] = 10^{-7}$ M; therefore, pH = 7.

Page 259

21. (a) Mg is being oxidised; S is being reduced.
    (b) Fe is being oxidised; Ag^+ is being reduced.
    (c) Br^- is being oxidised; Cl_2 is being reduced.
    (d) The reaction is not redox as there is no loss or gain of electrons.

Page 261

22. (a) oxidation: $Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$
    reduction: $Ag^+(aq) + e^- \rightarrow Ag(s)$
    (b) Ag^+ accepts electrons.
    (c) Pb is oxidised to Pb^{2+}.
    (d) Pb is the reductant.
    (e) Electrons are taken from Pb.
23. (a) H^+ accepts electrons.
    (b) H^+ is reduced to H_2.
    (c) Al is the reductant.
    (d) Electrons are taken from Al.
    (e) Al is oxidised to Al^{3+}.
    (f) H^+ is the oxidant.

Page 264

24. (a) $I_2(s) + H_2S(g) \rightarrow 2I^-(aq) + S(s) + 2H^+(aq)$
    (b) $2NO_3^-(aq) + 3H_2S(g) + 2H^+(aq) \rightarrow 2NO(g) + 3S(s) + 4H_2O(l)$
    (c) $Cr_2O_7^{2-}(aq) + 3H_2S(g) + 8H^+(aq) \rightarrow$
    $2Cr^{3+}(aq) + 3S(s) + 7H_2O(l)$
    (d) $5SO_3^{2-}(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow$
    $5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(l)$
    (e) $3Cu(s) + 2NO_3^-(aq) + 8H^+(aq) \rightarrow$
    $3Cu^{2+}(aq) + 2NO(g) + 4H_2O(l)$

Page 266

25. (a) $Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$ (reductant Mg, oxidant Cu^{2+})
    (b) no reaction
    (c) $2Al(s) + 3Pb^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Pb(s)$ (reductant Al, oxidant Pb^{2+})
    (d) no reaction
26. (a) $Fe(s) + Pb^{2+}(aq) \rightarrow Fe^{2+}(aq) + Pb(s)$
    (b) no reaction
    (c) $Zn(s) + Sn^{2+}(aq) \rightarrow Zn^{2+}(aq) + Sn(s)$
    (d) $2Al(s) + 3Zn^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Zn(s)$
27. (a) The aluminium metal may have had an oxide coating.
    (b) Use steel wool or emery paper to strip off the coating to expose the metal. Repeat the experiment.

Multiple choice questions

1. C	2. D	3. D	4. B	5. D	6. C	7. C	8. B	9. B
10. B	11. C	12. B	13. C	14. A	15. D	16. A	17. B	18. D
19. D	20. A	21. D	22. B	23. B	24. D	25. A	26. D	27. A
28. C	29. B	30. B	31. A	32. B	33. B	34. A		

Review questions

4. (a) acidic
    (b) acidic
    (c) both
    (d) acidic
    (e) acidic
    (f) basic
    (g) basic
    (h) both
    (i) basic
    (j) acidic
    (k) basic
5. (a) $2H_3PO_4(aq) + 3PbS(s) \rightarrow Pb_3(PO_4)_2(aq) + 3H_2S(g)$
    $2H^+(aq) + PbS(s) \rightarrow Pb^{2+}(aq) + H_2S(g)$
    (b) $2HCl(aq) + CuO(s) \rightarrow CuCl_2(aq) + H_2O(l)$
    $2H^+(aq) + CuO(s) \rightarrow Cu^{2+}(aq) + H_2O(l)$
    (c) $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$
    $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
    (d) $6HCl(aq) + 2Al(s) \rightarrow 2AlCl_3(aq) + 3H_2(g)$
    $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$
    (e) $Ca(s) + H_2CO_3(aq) \rightarrow CaCO_3(s) + H_2(g)$
    $Ca(s) + 2H^+(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s) + H_2(g)$
    (f) $2HNO_3(aq) + Li_2CO_3(aq) \rightarrow 2LiNO_3(aq) + CO_2(g) + H_2O(l)$
    $2H^+(aq) + CO_3^{2-}(aq) \rightarrow CO_2(g) + H_2O(l)$
7. Ionisation is the production of ions. Hydrolysis is the production of ions by reaction with water. Dissociation is the separation of ions already present.
8. HCl is a neutral molecule. The ions necessary for electrical conductivity are not produced until it reacts with water.
9. Uses include manufacture of glass, washing powder and detergents.
11. $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$
12. (b) (i) HCl (ii) HCO_3^- (iii) H_2SO_4 (iv) H_3O^+ (v) H_2O
    (c) (i) F^- (ii) CO_3^{2-} (iii) SO_4^{2-} (iv) OH^- (v) HS^-
13. (a) HS^-/S^{2-}; NH_4^+/NH_3
    (b) NH_4^+/NH_3; CH_3COOH/CH_3COO^-
    (c) HSO_4^-/SO_4^{2-}; H_3O^+/H_2O
    (d) $H_2PO_4^-/HPO_4^{2-}$; H_2O/OH^-
    (e) HNO_3/NO_3^-; H_2O/OH
14. $HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CO_3^{2-}(aq)$
    $HCO_3^-(aq) + H_2O(l) \rightleftharpoons OH^-(aq) + H_2CO_3(aq)$
16. (a) Ions necessary for conduction are produced by reaction with water.
    (b) electrolyte
    (c) Strong acids produce more ions.
17. Thalia is correct.
19. (a) (i) acidic (ii) neutral (iii) basic (iv) acidic
    (b) Providing concentrations are equal, (iv) is the strongest acid.
21. (a) a base (b) alkaline
23. (a) 3.0 (b) 3.3 (c) 3.4 (d) 8.2
24. Dilution is a linear scale. pH is a logarithmic scale.
25. (a) 11.0 (b) 10.7 (c) 10.6 (d) 5.8

27 Oxidation is a loss of electrons. Reduction is a gain of electrons.

28 a substance that oxidises something else and is itself reduced

29. (a) oxidant: Ce^{4+}; reductant: Sn^{2+}
    (b) oxidant: Pb^{2+}; reductant: Cd
    (c) oxidant: Cl_2; reductant: Al
    (d) oxidant: H^+; reductant: Mg
    (e) oxidant: H_2O; reductant: Na
30. (a) oxidation (b) reduction (c) reduction
31. (a) $Ag \rightarrow Ag^+ + e^-$; oxidation
    $O_2 + 4e^- \rightarrow 2O^{2-}$; reduction
    (b) $Fe \rightarrow Fe^{2+} + 2e^-$; oxidation
    $S + 2e^- \rightarrow S^{2-}$; reduction
    (c) $Pb \rightarrow Pb^{2+} + 2e^-$; oxidation
    $Cl_2 + 2e^- \rightarrow 2Cl^-$; reduction
    (d) $Hg \rightarrow Hg^{2+} + 2e^-$; oxidation
    $O_2 + 4e^- \rightarrow 2O^{2-}$; reduction
32. (a) $2Br^-(aq) + SO_4^{2-}(aq) + 4H^+(aq) \rightarrow Br_2(l) + SO_2(g) + 2H_2O(l)$
    (b) $2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$
    (c) $I_2(s) + H_2S(g) \rightarrow 2I^-(aq) + S(s) + 2H^+(aq)$
    (d) $3Cu(s) + 2NO_3^-(aq) + 8H^+(aq) \rightarrow$
    $3Cu^{2+}(aq) + 2NO(g) + 4H_2O(l)$
    (e) $Cu(s) + 2NO_3^-(aq) + 4H^+(aq) \rightarrow$
    $Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$
    (f) $3CuO(s) + 2NH_3(g) \rightarrow 3Cu(s) + 3H_2O(l) + N_2(g)$
    (g) $PbS(s) + 4H_2O_2(l) \rightarrow PbSO_4(s) + 4H_2O(l)$
    (h) $2Cr_2O_7^{2-}(aq) + 16H^+(aq) + 3CH_3CH_2OH(g) \rightarrow$
    $4Cr^{3+}(aq) + 11H_2O(l) + 3CH_3COOH(aq)$

33. Gold and silver are less reactive than zinc and iron.
34. Displacement occurs when metallic ions are converted to the metal by another metal. For example, zinc displaces copper ions in the reaction:

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

35. The better the reducing strength of a metal, the more reactive is the metal.
36. Groups 1 and 2 of the periodic table contain the most reactive metals.
37. $Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$
38. In each case, the reductant has been named first in the equation.
    (a) $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$
    (b) $Fe(s) + Sn^{2+}(aq) \rightarrow Fe^{2+}(aq) + Sn(s)$
    (c) $Pb(s) + 2Ag^{+}(aq) \rightarrow Pb^{2+}(aq) + 2Ag(s)$
    (d) $2Al(s) + 3Zn^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Zn(s)$
39. Tin will react with copper ions and will eventually wear away.
40. (a) $AgNO_3(aq)$, $CuSO_4(aq)$, $Pb(NO_3)_2(aq)$
    (b) $Fe(s) + 2Ag^{+}(aq) \rightarrow Fe^{2+}(aq) + 2Ag(s)$
    $Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$
    $Fe(s) + Pb^{2+}(aq) \rightarrow Fe^{2+}(aq) + Pb(s)$
41. $Na > X > Cu$
42. (a) $Sn(s) + Cu^{2+}(aq) \rightarrow Sn^{2+}(aq) + Cu(s)$
    $Cd(s) + Sn^{2+}(aq) \rightarrow Cd^{2+}(aq) + Sn(s)$
    (b) $Cd > Sn > Cu$
    (c) Only (iv) will occur.
43. (a) A^{+}, B^{2+}, C^{+}, D^{3+}
    (b) D is second in the series.
    (c) $B > D > A > C$
    (d) B^{2+}
44. (a) $E > A > F > C > D > B$
    (b) Mg, Zn, Fe, Sn, Pb, Cu
49. (a) any complete, undamaged coating
    (b) electrochemical protection, cathodic protection
50. (a) rust
    (b) rust
    (c) rust
    (d) no rust
55. (a) yes
    (b) yes, providing there is no damage to the coating
    (c) yes
    (d) no
    (e) yes

Exam practice questions

1. (a) can donate only one proton
    (b) $CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^{-}(s) + H_3O^{+}(aq)$
    (c) $CH_3COOH(aq) + NaOH(aq) \rightarrow NaCH_3COO(aq) + H_2O(l)$
    (d) The reaction in (b) does not go to completion.
2. (a) can donate or receive a proton
    (b) $HCO_3^{-}(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^{-}(aq)$
    $HCO_3^{-}(aq) + H_2O(l) \rightleftharpoons CO_3^{2-}(aq) + H_3O^{+}(aq)$
    (c) The first reaction occurs to a greater extent than the second.
3. (a) $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^{-}$
    (b) $Pb^{2+}(aq) + 2e^{-} \rightarrow Pb(s)$
    (c) Oxidant is Pb^{2+}.
    (d) Reductant is Sn metal.
    (e) No. Sn cannot reduce Mg^{2+} ions as it is lower on the activity series.

Chapter 14

Page 285

1. 12 g
2. 24 g
3. sodium sulfate

Page 288

5. (a) 17 g L^{-1}, 1.7×10^4 mg L^{-1} (b) 0.480 g L^{-1}, 480 mg L^{-1}
6. 0.0950 g
7. 4.00 mL

Page 290

8. 3.75 ppm
9. 0.352 g
10. 0.250 g
11. (a) 25.6%(m/m) (b) 0.50%(m/m) (c) 0.0720%(m/m)
12. 1.25 g
13. 258 g
14. 49%(m/m)

Page 291

15. (a) 0.400%(m/v) (b) 20%(m/v) (c) 0.0050%(m/v)
16. 5.63 g
17. 20 mL
18. 0.10 mL

Page 292

19. 2.20%(v/v)
20. (a) 37.5 mL (b) 500 mL
21. 86.3 mL

Page 293

22. (a) 0.8 M (b) 0.50 M (c) 4.29 M (d) 1.6 M

Page 295

23. (a) 2.0 M (b) $[Na^{+}] = 2.0$ M; $[OH^{-}] = 2.0$ M
24. 0.015 M
25. 0.417 M

Multiple choice questions

1. A 2. B 3. C 4. B 5. B 6. B 7. A
8. D 9. C 10. A 11. D 12. C

Review questions

2. (a) 4×10^{-4} g L^{-1} (b) 0.4 ppm
3. (a) (i) 48 °C (ii) 58 °C (iii) 24 °C
    (b) 65 °C
    (c) (i) 36 g/100 g (ii) 99 g/100 g (iii) 55 g/100 g
4. A saturated solution contains the maximum amount of solute that can be dissolved at a particular temperature. An unsaturated solution contains less. A supersaturated solution is an unstable situation formed on cooling where the solution may contain more than the apparent maximum.
5. (a) 11%(m/m) (b) 1%(m/m)
6. 40 g
7. 9.4 g
8. 2.9%(v/v)
9. (a) 400 mL (b) 140 mL
10. 50 g
11. (a) 18 g (b) 49 g (c) 2.8 kg (d) 1.3×10^2 g (125 g) (e) 30 g
12. (a) 0.66 M (d) 61.0 M
    (b) 3.22 M (e) 1.81 M
    (c) 6.7 M (f) 2.34 M
13. (a) 0.50 M (b) 0.340 M (c) 0.127 M
14. (a) 0.0030 mol (b) 0.45 mol
    (c) 0.092 mol (d) 0.5 mol

15. (a) 39 g (c) 128 g (e) 1.38 g
(b) 105 g (d) 13.9 g (f) 12.3 g
16. 16.4 g
17. (a) 0.298 M (b) 0.15 M (c) 0.266 M (d) 0.667 M
18. (a) 1.3 M
(b) $[NH_4^+] = 1.3$ M; $[NO_3^-] = 1.3$ M
19. (a) 2.14 M
(b) $[Al^{3+}] = 2.14$ M; $[Cl^-] = 6.42$ M
20. 0.025 M
21. 0.083 M

Exam practice questions

1. (a) V, W (b) T, X, Z (c) U, Y
2. (a) (i) 3×10^1 g L^{-1} (ii) 3×10^4 mg L^{-1}
(b) (i) 5×10^{-3} g L^{-1} (ii) 5 mg L^{-1}
(c) (i) 2.50 g L^{-1} (ii) 2.50×10^3 mg L^{-1}
3. (a) 0.1–2.0 ppb
(b) 1×10^9 L
(c) volume too large to process

Chapter 15

Page 308

1. (a) qualitative
(b) quantitative
(c) quantitative
(d) qualitative
2. (a) It could establish the source of pollution from an oil spill.
(b) The pattern of components in the oil, and their quantities, could be matched with samples from 'suspect' ships.

Page 310

3. Conductivity depends on the number of ions present, rather than the type of ions.

Page 312

4. to avoid contamination or interference with results

Page 317

5. 110 mg L^{-1}

Page 320

6. (a) 0.5 mol (b) 0.5 mol (c) 3.5 mol
7. (a) 0.5 mol (c) 0.1 mol (e) 0.05 mol (g) 8 mol
(b) 0.2 mol (d) 0.2 mol (f) 0.50 mol

Pages 321–2

8. 6.3 g
9. (a) 2.9 g (b) 6.2 g (c) 66 g (d) 1.97 g (e) 15 kg
10. 3.977 g
11. (a) the (s) symbol of state (b) 4.837 g

Page 325

13. 16.2 g
14. (a) 6.04 g
(b) All of the sodium sulfide solution is used up.
15. (a) All acid was used up. (b) 0.09 M
(c) incomplete drying of magnesium strip; inaccuracies in weighing magnesium and measuring acid volume
(d) Wear gloves, laboratory coat and safety goggles when handling acid; treat acid spills with copious quantities of water.

Page 329

16. 96.9%

Multiple choice questions

1. D 2. B 3. B 4. D 5. C 6. C 7. D 8. C
9. B 10. D 11. D 12. A 13. A 14. B 15. C

Review questions

1. (a) qualitative
(b) quantitative
(c) quantitative
(d) quantitative
2. (a) limewater test
(b) turns iodine dark blue
(c) glowing splint test
(d) turns a mixture of $CuSO_4$ and NaOH purple
(e) 'pops' when a small sample is ignited
(f) turns blue paper pink
6. (a) $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow 2NaNO_3(aq) + BaSO_4(s)$
Precipitate is barium sulfate.
(b) 4.56 g
7. (a) 0.760 mol (b) 24.3 g
8. (a) 0.15 mol
(b) $C(s) + O_2(g) \rightarrow CO_2(g)$; 13 g
9. (a) $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
(b) 25.4 g
10. (a) 9.49 g (b) 13.1 g
11. 0.029 g
12. 5.51 kg
13. (a) 159 g (b) 5.2 g
14. (a) $Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$
(b) 7.31 g
15. 3.67%
16. 9.4 g
17. (a) 0.921 kg
(b) $2NaOH(aq) + CO_2(g) \rightarrow Na_2CO_3(aq) + H_2O(l)$
(c) 0.550 kg
(d) It is more efficient per unit mass (an important consideration for a spacecraft).
(e) $2OH^-(aq) + CO_2(g) \rightarrow CO_3^{2-}(aq) + H_2O(l)$ (This applies to both reactions.)
18. (a) 46.8 tonnes
(b) It is decomposed; no oxygen is added.
19. 52 mL
20. 0.794 g
21. (a) Ions dissolved in water are required for it to conduct electricity. The more ions there are from dissolved salts, the greater is the conductivity.
(b) Before the early 1980s, there was an increasing trend. Since the early 1980s, there has been a decreasing trend.
(c) Ions from other salts are probably present also.
(d) Conductivity is related to the total amount of all ions present. It does not depend on which ions are present.
22. (a) 0
(c) 2.3 ppm
(d) Determine the equation for the line or curve of best fit.
23. (b) sample 1, 41 mg L^{-1}; sample 2, 6 mg L^{-1}
24. (b) 7.3 mg L^{-1}
(c) AAS
25. (a) 13 ppm
(b) 1.3%
(c) orange
26. (a) $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
(b) $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$
$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$
(c) $Pb(NO_3)_2(aq) + K_2CrO_4(aq) \rightarrow PbCrO_4(s) + 2KNO_3(aq)$
$Pb^{2+}(aq) + CrO_4^{2-}(aq) \rightarrow PbCrO_4(s)$

(d) $Na_3PO_4(aq) + 3AgNO_3(aq) \rightarrow Ag_3PO_4(s) + 3NaNO_3(aq)$
$3Ag^+(aq) + PO_4^{3-}(aq) \rightarrow Ag_3PO_4(s)$
(e) $2HCl(aq) + Pb(NO_3)_2(aq) \rightarrow PbCl_2(s) + 2HNO_3(aq)$
$Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

27. (a) soluble
    (b) insoluble
    (c) insoluble
    (d) insoluble
    (e) soluble
    (f) soluble
28. (a) 14.112%
    (b) to remove all water
    (c) It might not be a representative sample.
29. 9.99%
30. (a) silver chloride (b) 96.5%
31. (a) 3.34% (b) increased percentage of sulfur
32. (a) lead (b) silver
    (c) barium (d) hydroxide
33. to ensure that there is no water present
34. There will always be a loss due to some dissolution of the precipitate. The design of the experiment should choose a precipitate where this is minimised. This would be a systematic error.
35. Obtain a sample of the water, add excess silver nitrate solution to a measured volume of this water, filter the resulting silver chloride precipitate, dry and weigh it to constant mass, and do the calculation.
36. (c), (e), (h), (a), (f), (d)
37. (a) overestimated
    (b) overestimated due to precipitate containing traces of other chemicals
    (c) no effect
    (d) underestimated due to increased solubility of the precipitate at the higher temperature

Exam practice questions

1. (a) carbon dioxide
   (b) $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
   (c) 83.1%
   (d) Many answers are possible, e.g. calcium carbonate is the only carbonate present; the sample was dry to start with; all of the calcium carbonate decomposes to calcium oxide.
2. (a) 1.14 g
   (b) 8.00%
   (c) Many answers are possible, e.g. the precipitate is thoroughly dry; all of the phosphorus precipitates as $Mg_2P_2O_7$; the precipitating reagent doesn't form precipitates with other ions.
3. (a) $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
   (b) 1.821 g
   (c) 0.1821 g L^{-1}
   (d) no effect
   (e) Many answers are possible, e.g. sodium chloride is the only source of chloride ions in the samples collected; $AgNO_3$ doesn't form precipitates with other ions present; all of the chloride ions precipitate as AgCl.

Chapter 16

Page 343

1. $C_3H_8NP_2O_5$

Pages 348–9

3. (b) 22 mg L^{-1} (c) 220 mg L^{-1}

Multiple choice questions

1. B 2. B 3. D 4. B 5. D 6. B 7. A
8. D 9. B 10. B 11. D 12. B

Review questions

1. (a) carbon-based compounds
   (b) carbon, hydrogen
   (c) N, O and Cl, and possibly P and S
2. whether their molecules are polar
3. No. It will be non-polar given the elements it contains.
4. Some systematic names are too complicated.
5. (a) so that it has time to act
   (b) It contains oxygen–hydrogen bonds and is asymmetrical. Therefore, it has a degree of polarity. Water is polar, and fatty tissue is non-polar.
7. (a) $C_{12}H_4O_2Cl_2$
   (b) formed as unwanted by-products
   (c) Dioxins are non-polar.
   (d) Fatty tissue is non-polar and difficult to eliminate.
9. (a) various colours including blue and yellow
   (b) The colours would separate (resolve) further and you might be able to distinguish additional colours.
10. (b) Both substances might have the same retention time under the operating conditions of the HPLC.
11. You just need enough standards to produce a reasonable set of points on a graph from which a line of best fit can be drawn.
12. 4.2 mL
13. (b) 0.9 g L^{-1} (c) 5 mM (d) yes
14. (a) 16 mg L^{-1}
    (b) 0.22 mg
    (c) Yes. Only water has been removed by evaporation.
    (d) 2.15 g. The content is lower than claimed.

Exam practice questions

1. (b) 52 mg L^{-1} (c) 5200 mg L^{-1} (5.2 g L^{-1}) (d) 52 mg
2. (a) It has a high field half-life, high solubility and high adsorbance.
   (b) It has a low field half-life.
   (c) The chemical would probably be washed away adsorbed to soil particles, which would be moved by the large amount of run-off.
   (d) No. The soil particles would be unlikely to move, and it has a low solubility.

Chapter 17

Page 362

1. (a) $2HCl(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$
   (b) 0.18 M
2. 30.0 mL
3. 60.0 mL
4. 90.0 mL
5. 0.560 M

Page 363

6. 45 mL
7. 1.6 M

Page 366

8. (b) 0.09501 M
9. (a) (ii) 0.0551 M
   (b) (i) $Na_2CO_3(aq) + 2HNO_3(aq) \rightarrow 2NaNO_3(aq) + 2H_2O(l) + CO_2(g)$
   (ii) 0.09950 M

Multiple choice questions

1. D 2. A 3. B 4. C 5. A 6. D 7. B 8. B
9. B 10. D 11. B 12. A 13. B 14. B

Review questions

2. depends on what substances dissolve in it as it flows over the surface
4. One river has a higher alkalinity than the other.
5. (a) 21 mL (b) 33 mL (c) 3.1 mL
6. (a) 0.09 M (b) 0.16 M (c) 1.25 M
7. (a) 200 mL (b) 325 mL (c) 1.60 L
8. 100 mL
9. 20.1 mL
10. (a) 0.184 M (b) 0.105 M
11. (a) $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
    (b) 90 mL
12. 4.0 M
13. 0.0750 M
14. 528 mL
15. 149 mL
16. 0.756 M
17. 16.2 mL
18. 0.09387 M
19. (a) 0.00410 M (b) 0.164 g L^{-1}

Exam practice questions

1. (a) 2.65 g (b) 0.106 M
2. (a) $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
   (b) 0.001 85 mol
   (c) 0.000 926 mol
   (d) 0.0370 M
   (e) 0.370 M
   (f) 0.01 M
   (g) 35.3 L

Glossary

acid: a substance that can donate a proton to a base

acid rain: rain that has a pH lower than 5

addition polymerisation: the process in which monomers with at least one double bond react together to form a polymer by addition reactions

addition reaction: a reaction in which one molecule bonds covalently with another molecule without losing any other atoms

aliquot: the volume delivered by a pipette

alkali: basic substance

alkane: a simple, saturated binary compound of carbon and hydrogen atoms with single bonds between the carbon atoms. The general formula for all members of this homologous series is C_nH_{2n+2}.

alkenes: the family of hydrocarbons that contain one carbon–carbon double bond

alkynes: the family of hydrocarbons with one carbon–carbon triple bond

alloy: a substance with metallic properties that consists of two or more elements, at least one of which is a metal

alloying: mixing of a metal with one or more other elements (These other elements are usually metals.)

amphiprotic: a substance that can donate or accept protons

anhydrous: without water in the chemical structure

anion: a negatively charged atom or group of atoms

annealing: heating a metal in an inert gas to a moderate temperature, maintaining this temperature for several hours, then cooling it slowly

aqueous solutions: mixtures in which substances are dissolved in water

atactic: describes a polymer in which the methyl groups take random positions in front of and behind the carbon chain

atom: a neutral particle with a nucleus; the smallest sample of an element

atomic absorption spectroscopy: an analytical method for detecting metal species

atomic emission spectrum: a spectrum emitted as distinct bands of light of diagnostic frequencies by elements or compounds

atomic number: the number of protons in the nucleus of an atom of a particular element

Avogadro's number (N_A): the number of elementary particles (atoms) in exactly 12 g of carbon-12. This number is equal to 6.02×10^{23}.

ball bearing model: representation of a metal using ball bearings to show the arrangement of cations into grains

ball-and-stick model: representation of a molecule in which the atoms are shown as balls and the bonds as sticks

base: a substance with a pH greater than 7

binary: describes compounds made up of only two elements

biodegradable: able to break down via natural processes over the medium term

bond dipole: separation of charge in a polar covalent bond

bonding electrons: the pairs of electrons involved in forming a covalent bond

buffer: solution that resists a change in pH upon the addition of a small amount of acid or base

carbon nanotube: structure in the shape of a tube composed of carbon atoms

catalyst: a substance that alters the rate of a reaction without a change in its own concentration

catalytic cracking: the use of a catalyst to break bonds in long molecules to produce shorter molecules

cathodic protection: a method of protecting a metal from corrosion by connecting it to a power source and another metal that is less active. The electrons feed onto the metal being protected, making it a cathode so that it cannot be oxidised.

cation: a positively charged atom or group of atoms

central atom: the atom in a molecule with the most bonding electrons

chemical bond: arrangement of electrons between two atoms that generates a force, causing the atoms to be bound to each other

chemical changes: a reaction in which one or more substances are transformed into one or more new substances

chemical equation: notation used to describe a chemical reaction with the reactants on the left side of the arrow and the products on the right

coefficients: numbers placed in front of formulas to balance a chemical equation

combustion reaction: the addition of oxygen to a reactant, often producing heat

computer-generated model: representation of a molecule that is produced by a computer

concentration: the amount of solute that is dissolved in a known volume of solvent

concordant: individual titres within a certain limit

condensation polymerisation: polymerisation in which two monomers combine and a smaller molecule is eliminated

conjugate acid (of a base): the product that a base forms when it has accepted a proton from an acid

conjugate redox pair: two species that differ only by a certain number of electrons

copolymer: a polymer formed from the polymerisation of two monomers

corrosion: oxidation of a substance to form a new material

covalent bond: sharing of electrons between nuclei that bonds them together in a molecule

covalent layer lattice: a type of compound formed from layers bonded together

covalent molecular compounds: a molecular compound in which atoms of different elements share electrons with each other

covalent molecular element: element made up of identical atoms held together by covalent bonds

cross-linking: the bonding between two polymers

cyclic compounds: compounds containing a ring structure, such as benzene

delocalised: describes electrons that are not bound to any one atom but are free to move throughout a lattice

diatomic: describes a substance containing two atoms only

dilute: not concentrated

dipole–dipole interactions: weak bonding caused by the positive end of one dipole attracting the negative end of another dipole

diprotic: containing two protons

discrete: separate, distinct, not in an infinite array or lattice

dispersion force: the bond between adjacent molecules formed by instantaneous dipoles. This weak non-directional bonding is also known as van der Waals force.

displacement reaction: chemical reaction in which a more reactive element displaces a less reactive element from its compound

dissociation: the process by which ions separate when an ionic compound dissolves in a solvent

double bond: strong bond between two atoms formed by two pairs of electrons that are shared by the two nuclei

dry corrosion: corrosion of a metal in the air when water is not a reactant

elastomer: a polymer that can be stretched and return to its original shape

electrochemical protection: *see* **cathodic protection**; **sacrificial protection**

electrolyte: a liquid that can conduct electricity

electron configuration: the arrangement of electrons in the shells of an atom

electron dot diagrams: representation where the atom's nucleus and all innershell electrons are replaced by its element symbol and the outershell electrons are represented by dots around the symbol in a square arrangement

electronegativity: the electron-attracting power of an atom

electrons: the negatively charged entities that form part of an atom

electrostatic attraction: strong force of attraction between positive and negative ions that produces an ionic bond

element: chemical species consisting of atoms of a single type

empirical formula: formula that shows the simplest numerical ratio in which atoms are combined

end point: the experimentally determined equivalence point at which the indicator just changes colour or the pH curve becomes vertical

equivalence point: the point at which two reactants have reacted in their correct mole proportions in a titration

ester: an organic compound formed from a condensation reaction between an alcohol and a carboxylic acid

eutrophication: form of water pollution involving excess nutrients, such as nitrogen and phosphorus, leaching from soils, typically resulting in excessive growth of algae

excess reactant: reactant left over when the limiting reagent has been used up

excited: raised to a higher than ground-state energy level

fractional distillation: a method of separating the components of a liquid mixture that depends on the ease of vaporisation of the components

functional group: a group of atoms attached to or part of a hydrocarbon chain that influence the physical and chemical properties of the molecule

gravimetric analysis: analysis by mass of precipitate

ground state: the least excited energy level of an atom

half-equation: chemical equation showing either oxidation or reduction in a redox reaction

haloalkane: compound containing a halogen and an alkane

homologous series: a series of organic compounds with the same structure but in which the formula of one molecule differs from the next by a —CH_2 group

hydrated: containing water molecules within the structure

hydrates: adds water

hydrocarbons: compounds containing only carbon and hydrogen

hydrogen bond: the bond between a hydrogen atom covalently bonded to an atom of F, O or N and another molecule that also contains an atom of H, F, O or N

hydrolysis: a reaction involving the addition of water

hydronium ion: H_3O^+; the product of the reaction of a proton with water

indicators: substances that can indicate an acid or basic state and a range of pH values

intermolecular: between molecules

interstitial alloy: an alloy in which the smaller atoms fit into the spaces between the larger atoms

intramolecular: involving the internal structure of a molecule

ion: a charged atom

ionic bonding: bonding between cations and anions in an ionic compound

ionic compound: compound containing cations and anions in an ionic lattice structure or an ionic liquid, depending on ion size

ionic concentrations: concentrations of the anions and cations resulting from the dissolution of an ionic lattice; designated by square brackets

ionic equation of neutralisation: neutralisation equation without the spectator ions

ionisation: reaction in which a substance reacts with water to produce ions

ions: charged species

isoprene: *see* **rubber**

isotactic: describes a polymer in which the methyl groups are all on the same side of the carbon chain

isotopes: forms of an element with the same number of protons but different numbers of neutrons in the nucleus

isotopic symbol: representation of an element as ${}^{A}_{Z}E$, where E is the symbol for the element, A is the mass number and Z is the atomic number

latent heat of fusion: energy needed to change a fixed amount of water from the solid to the liquid state

latent heat of vaporisation: energy needed to change a fixed mass of water from the liquid to the gaseous state

law of conservation of mass: No detectable gain or loss in mass occurs in chemical reactions. Mass is conserved.

'like dissolves like' rule: Strongly polar and ionic solutes tend to dissolve in polar solvents, and non-polar solutes tend to dissolve in non-polar solvents.

limiting reactant: the reactant that is completely used up in a reaction

linear polymers: polymers that do not have side chains

lone pairs: *see* **non-bonding electrons**

mass number: the total number of protons and neutrons in the nucleus of a particular isotope of an element

mass spectrometer: an analytical instrument that determines the relative isotopic masses of the different isotopes of an element and abundance

metal: ductile, malleable, conducting element

metallic bonding: positively charged metal cations arranged in a lattice with delocalised valence electrons being able to flow around them

metallic lattice structure: crystal lattice made up of an array of cations

%(m/m): percentage mass per mass

molar concentration: *see* **molarity**

molarity: c; concentration of a solute in a solution in units of M, or mol L^{-1}

mole: 6×10^{23} particles of a substance

molecular formula: the actual number of atoms of each element in a molecule of a substance

molecule: group of atoms bonded together covalently

monomers: molecules that link together to form a polymer

monoprotic: containing one proton

multiple bond: bond formed when two atoms share two or more pairs of electrons

%(m/v): percentage mass per volume

neutralisation reaction: the reaction between an acid and a base

neutron: neutral subatomic particle in the nucleus of an atom

non-biodegradable: not able to break down in a natural environment

non-bonding electrons: electrons that are not involved in bonding

non-metals: non-ductile, non-malleable, non-conducting element

non-polar covalent bond: bond formed between atoms with the same electronegativity

non-polar gas: gas made up of molecules that are non-polar

non-polar molecules: molecules that do not have permanent dipoles or are symmetrical

nucleus: the central part of an atom, containing protons and neutrons

orbital: three-dimensional wave describing a bound electron

organic chemistry: the study of carbon-containing compounds and their properties

oxidant: an electron acceptor

oxidation: an increase in the oxidation number; a loss of one or more electrons

oxidation–reduction: *see* **redox**

parts per million (ppm): number of a particular component present within one million objects

percentage composition: list of percentages by mass of the elements in a compound

pH scale: $-\log[H_3O^+]$

photon: particle of light

physical changes: changes of state that do not form new substances

plastic: polymer that can be moulded when hot and retains its shape when cooled

polar covalent bond: bond formed when two atoms that have different electronegativities share electrons

polar gas: gas made up of molecules that are polar

polyatomic ion: charged ion composed of two or more atoms

polyethene: a polymer consisting of monomers of ethene

polymerisation: the formation of giant molecules by repeated monomers that are joined by covalent bonds

polyprotic: able to donate more than one proton to a base

precipitate: *see* **precipitation reaction**

precipitation reaction: a reaction where a solid with low solubility forms in a liquid and then settles to the bottom of a container

primary standard: a substance used in volumetric analysis that is of such high purity and stability that it can be used to prepare a solution of accurately known concentration

principal quantum number, *n*: quantum number that indexes the energy and size of an atomic orbital

product: chemical species obtained as the result of a chemical reaction

proton: a positively charged particle in the nucleus of an atom

qualitative analysis: use of experimental procedures to determine which elements are in a substance

quantitative analysis: use of experimental procedures to determine the percentage composition of a compound or the percentage of a component in a mixture

quenching: the process of plunging hot metal into water to rapidly cool it

reactant: chemical species that is transformed in a chemical reaction

reactivity series of metals: an ordered list of how readily the metals react with oxygen, water, steam, dilute acids and salt solutions

redox: describes a reaction that involves the transfer of one or more electrons between chemical species

reductant: electron donor

reduction: a decrease in the oxidation number; a gain of electrons

relative atomic mass: A_r; the mean mass of the naturally occurring mixture of the isotopes of an element on the relative atomic mass scale, on which the masses of particles are compared with the mass of the carbon-12 isotope

relative isotopic mass: RIM; the mass of an individual isotope of an element on the relative atomic mass scale, on which the masses of particles are compared with the mass of the carbon-12 isotope

rubber: a polymer formed from natural materials

rusting: wet corrosion of iron

sacrificial protection: protection of a metal from corrosion by attaching another more active metal to it. The more active metal reacts preferentially with corrosive elements and protects the less reactive metal.

salt: an ionic compound consisting of a metal ion and a non-metal ion, except oxides and hydroxides.

saturated hydrocarbon: a compound composed of carbon and hydrogen and containing only single covalent bonds

saturated solution: a solution in which the maximum amount of solvent has been dissolved

single bond: bond formed by one pair of electrons shared between two atoms

solubility: grams of a substance that will dissolve in 100 g of solvent at a particular temperature

solute: the dissolved substance in a solution

solution: a solute dissolved in a solvent

solvent: the liquid component of a solution

space-filling model: three-dimensional representation of a molecule that shows the relative sizes of atoms within the molecule and the distances between them

specific heat capacity: energy (measured in joules) needed to raise the temperature of 1 g of a pure substance by 1 °C

spectator ions: ions that take no part in a reaction but maintain electrical neutrality in a solution

standard solution: a solution that has a precisely known concentration

stoichiometry: the amount and ratio of reactants and products in a chemical reaction

structural formula: a diagrammatic representation of a molecule showing every bond

subatomic particles: electrons, protons and neutrons

substitution: exchange of one species for another

substitutional alloy: an alloy in which the atoms of the different elements are about the same size, so they can replace each other in the metal crystals

supersaturated solution: a solution in which a greater amount of solute is dissolved at a particular temperature than is predicted by a solubility curve temperature.

tempered: describes a quenched metal that is heated again but to a lower temperature than the initial heating

tensile strength: resistance to being stretched or drawn out and, therefore, to breaking

thermal cracking: the use of heat to break bonds in long molecules to produce shorter molecules

thermoplastic: describes polymers that soften on heating

thermosetting: describes polymers that do not soften on heating and char if heated strongly

thermosoftening: *see* **thermoplastic**

titration: process used to determine the concentration of a substance by using a pipette to deliver one substance and a burette to deliver another substance until they have reacted in exactly their mole ratio

titre: the volume of liquid measured by a burette during titration

triple bond: strong bond between two atoms formed by three pairs of electrons that are shared by the two nuclei

triprotic: able to donate three protons to a base

universal indicator: a mixture of separate indicators that show colour changes in solutions of a large range of pH values

universal solvent: water

unsaturated hydrocarbon: compound containing carbon and hydrogen but possessing one or more double or triple bonds

unsaturated solution: solution containing less than the maximum amount of solute that can dissolve

UV–visible spectroscopy: technique used to study compounds that absorb light in the UV–visible region

valence shell electron pair repulsion theory (VSEPR theory): a model whose main point is that the structure around a given atom in a molecule is determined principally by minimising electron repulsions

vulcanisation: the process in which sulfur is added to rubber and heated to cause cross-linking of polymer chains, increasing the strength of the rubber

%(v/v): percentage volume per volume

water of crystallisation: water molecules included in a crystal lattice structure

wet corrosion: corrosion of a metal by reaction with oxygen in the presence of water

work hardened: a property of metal brought about by beating the metal when hot

Index

Periodic table of the elements

	Period 1		
	1 Hydrogen **H** **1.0**	2 Helium **He** **4.0**	*Key* Atomic number Name Symbol Relative atomic mass

	Alkali metals	**Alkaline earth metals**	**Transition metals**						
	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9
Period 2	3 Lithium **Li** **6.9**	4 Beryllium **Be** **9.0**							
Period 3	11 Sodium **Na** **23.0**	12 Magnesium **Mg** **24.3**							
Period 4	19 Potassium **K** **39.1**	20 Calcium **Ca** **40.1**	21 Scandium **Sc** **45.0**	22 Titanium **Ti** **47.9**	23 Vanadium **V** **50.9**	24 Chromium **Cr** **52.0**	25 Manganese **Mn** **54.9**	26 Iron **Fe** **55.8**	27 Cobalt **Co** **58.9**
Period 5	37 Rubidium **Rb** **85.5**	38 Strontium **Sr** **87.6**	39 Yttrium **Y** **88.9**	40 Zirconium **Zr** **91.2**	41 Niobium **Nb** **92.9**	42 Molybdenum **Mo** **96.0**	43 Technetium **Tc** **(98)**	44 Ruthenium **Ru** **101.1**	45 Rhodium **Rh** **102.9**
Period 6	55 Caesium **Cs** **132.9**	56 Barium **Ba** **137.3**	57–71 Lanthanoids	72 Hafnium **Hf** **178.5**	73 Tantalum **Ta** **180.9**	74 Tungsten **W** **183.8**	75 Rhenium **Re** **186.2**	76 Osmium **Os** **190.2**	77 Iridium **Ir** **192.2**
Period 7	87 Francium **Fr** **(223)**	88 Radium **Ra** **(226)**	89–103 Actinoids	104 Rutherfordium **Rf** **(261)**	105 Dubnium **Db** **(262)**	106 Seaborgium **Sg** **(266)**	107 Bohrium **Bh** **(264)**	108 Hassium **Hs** **(267)**	109 Meitnerium **Mt** **(268)**

Lanthanoids

57 Lanthanum **La** **138.9**	58 Cerium **Ce** **140.1**	59 Praseodymium **Pr** **140.9**	60 Neodymium **Nd** **144.2**	61 Promethium **Pm** **(145)**	62 Samarium **Sm** **150.4**	63 Europium **Eu** **152.0**

Actinoids

89 Actinium **Ac** **(227)**	90 Thorium **Th** **232.0**	91 Protactinium **Pa** **231.0**	92 Uranium **U** **238.0**	93 Neptunium **Np** **(237)**	94 Plutonium **Pu** **(244)**	95 Americium **Am** **(243)**

The period number refers to the number of the outermost shell containing electrons.

Group 10	Group 11	Group 12	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
			5 Boron **B** **10.8**	6 Carbon **C** **12.0**	7 Nitrogen **N** **14.0**	8 Oxygen **O** **16.0**	9 Fluorine **F** **19.0**	10 Neon **Ne** **20.2**
			13 Aluminium **Al** **27.0**	14 Silicon **Si** **28.1**	15 Phosphorus **P** **31.0**	16 Sulfur **S** **32.1**	17 Chlorine **Cl** **35.5**	18 Argon **Ar** **39.9**
28 Nickel **Ni** **58.7**	29 Copper **Cu** **63.5**	30 Zinc **Zn** **65.4**	31 Gallium **Ga** **69.7**	32 Germanium **Ge** **72.6**	33 Arsenic **As** **74.9**	34 Selenium **Se** **79.0**	35 Bromine **Br** **79.9**	36 Krypton **Kr** **83.8**
46 Palladium **Pd** **106.4**	47 Silver **Ag** **107.9**	48 Cadmium **Cd** **112.4**	49 Indium **In** **114.8**	50 Tin **Sn** **118.7**	51 Antimony **Sb** **121.8**	52 Tellurium **Te** **127.6**	53 Iodine **I** **126.9**	54 Xenon **Xe** **131.3**
78 Platinum **Pt** **195.1**	79 Gold **Au** **197.0**	80 Mercury **Hg** **200.6**	81 Thallium **Tl** **204.4**	82 Lead **Pb** **207.2**	83 Bismuth **Bi** **209.0**	84 Polonium **Po** **(210)**	85 Astatine **At** **(210)**	86 Radon **Rn** **(222)**
110 Darmstadtium **Ds** **(271)**	111 Roentgenium **Rg** **(272)**	112 Copernicium **Cn** **(285)**		114 Flerovium **Fl** **(289)**		116 Livermorium **Lv** **(292)**		

The elements in the purple cells adjacent to the bold black border are neither metals nor non-metals. They are called metalloids.

64 Gadolinium **Gd** **157.3**	65 Terbium **Tb** **158.9**	66 Dysprosium **Dy** **162.5**	67 Holmium **Ho** **164.9**	68 Erbium **Er** **167.3**	69 Thulium **Tm** **168.9**	70 Ytterbium **Yb** **173.1**	71 Lutetium **Lu** **175.0**

96 Curium **Cm** **(247)**	97 Berkelium **Bk** **(247)**	98 Californium **Cf** **(251)**	99 Einsteinium **Es** **(252)**	100 Fermium **Fm** **(257)**	101 Mendelevium **Md** **(258)**	102 Nobelium **No** **(259)**	103 Lawrencium **Lr** **(262)**