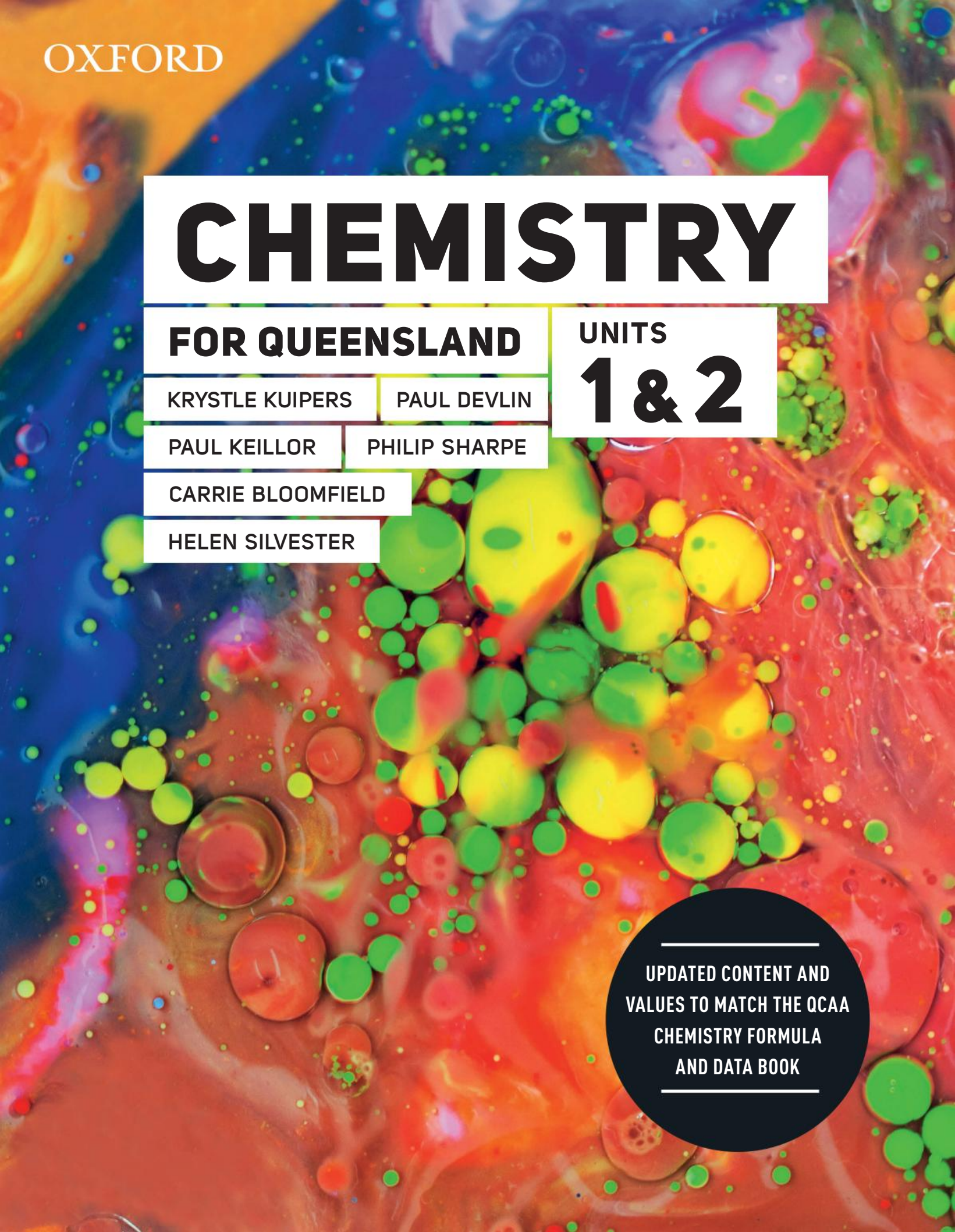
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FOR QUEENSLAND

UNITS 1 & 2

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Using *Chemistry for Queensland Units 1 & 2*

Chemistry for Queensland Units 1 & 2 has been written to meet the requirements of the QCAA Chemistry General Senior Syllabus. The first of the two-volume series, *Chemistry for Queensland Units 1 & 2* offers complete support for the Unit 1 & 2 teachers and their students, providing unparalleled depth and comprehensive syllabus coverage.

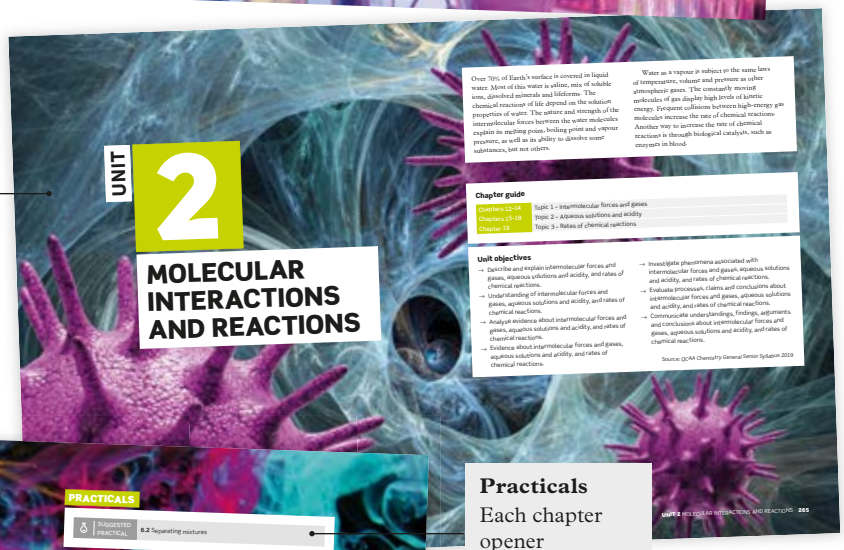
Key features of the Student Book:

Chemistry toolkit
The Student book begins with a stand-alone reference chapter that includes:

- assessment advice and structured examples
- a step-by-step guide to preparing for your exam
- methods for presenting and analysing chemical data.

Unit openers
Each unit begins with a unit opener that includes:

- an **overview of topics** in the unit
- **unit objectives** from the syllabus.



Chapter openers
Each chapter begins with a chapter opener that includes course objectives from the syllabus.

Practicals
Each chapter opener includes a list of mandatory and suggested practicals from the chapter to allow students to apply their knowledge to real-life activities.

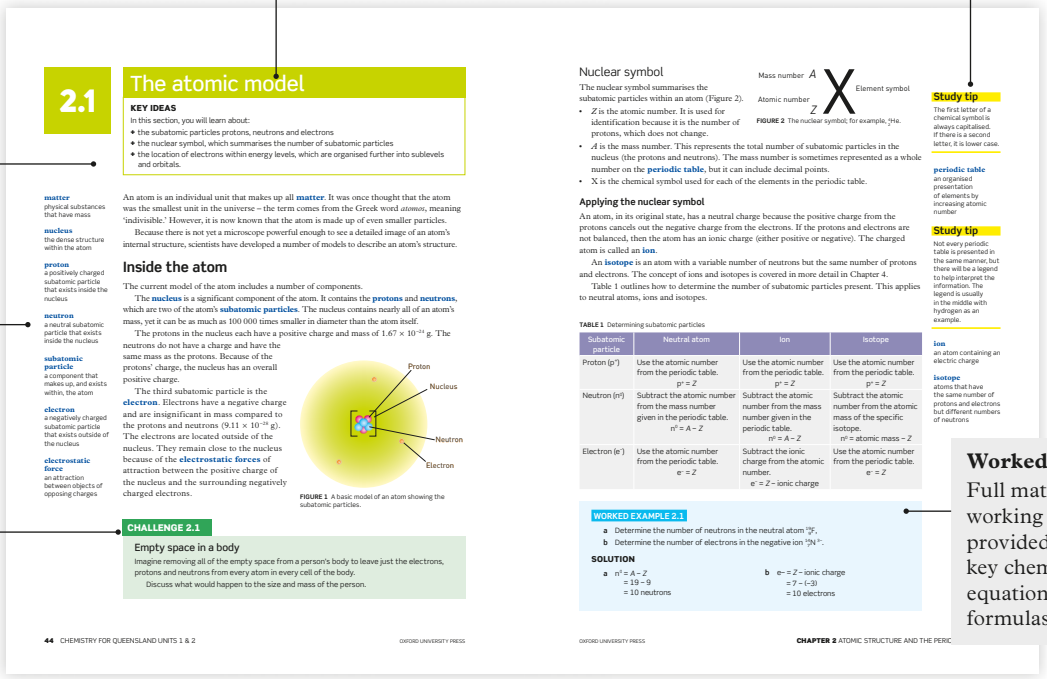
Section-based approach
Content is presented in clearly structured sections. Each **section** is clearly labeled and numbered to help navigation.

Study tip
Practical advice helps students improve their performance in assessment tasks.

Key ideas
A set of clear learning objectives from the syllabus prepares students for what they will study in every section.

Margin glossary
Key terms and definitions are highlighted to help students recall important information.

Challenge
Challenge boxes feature questions and tasks to foster critical thinking.



2.1 The atomic model

KEY IDEAS
In this section, you will learn about:
- the subatomic particles protons, neutrons and electrons
- the nuclear symbol, which summarises the number of subatomic particles
- the location of electrons within energy levels, which are organised further into sublevels and orbitals.

matter — physical substances that have mass
nucleus — the dense structure within the atom
proton — a positively charged subatomic particle that exists inside the nucleus
neutron — a neutral subatomic particle that exists inside the nucleus
subatomic particle — a component that makes up, and exists within, the atom
electron — a negatively charged subatomic particle that exists outside of the nucleus
electrostatic force — an attraction between objects of opposing charges

An atom is an individual unit that makes up all **matter**. It was once thought that the atom was the smallest unit in the universe – the term comes from the Greek word *atomos*, meaning 'indivisible.' However, it is now known that the atom is made up of even smaller particles.

Because there is not yet a microscope powerful enough to see a detailed image of an atom's internal structure, scientists have developed a number of models to describe an atom's structure.

Inside the atom

The current model of the atom includes a number of components.

The **nucleus** is a significant component of the atom. It contains the **protons** and **neutrons**, which are two of the atom's **subatomic particles**. The nucleus contains nearly all of an atom's mass, yet it can be as much as 100 000 times smaller in diameter than the atom itself.

The protons in the nucleus each have a positive charge and mass of 1.67×10^{-24} g. The neutrons do not have a charge and have the same mass as the protons. Because of the protons' charge, the nucleus has an overall positive charge.

The third subatomic particle is the **electron**. Electrons have a negative charge and are insignificant in mass compared to the protons and neutrons (9.11×10^{-28} g). The electrons are located outside of the nucleus. They remain close to the nucleus because of the **electrostatic forces** of attraction between the positive charge of the nucleus and the surrounding negatively charged electrons.

FIGURE 1 A basic model of an atom showing the subatomic particles.

CHALLENGE 2.1
Empty space in a body
Imagine removing all of the empty space from a person's body to leave just the electrons, protons and neutrons from every atom in every cell of the body.
Discuss what would happen to the size and mass of the person.

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Nuclear symbol

The nuclear symbol summarises the subatomic particles within an atom (Figure 2).
- Z is the atomic number. It is used for identification because it is the number of protons, which does not change.
- A is the mass number. This represents the total number of subatomic particles in the nucleus (the protons and neutrons). The mass number is sometimes represented as a whole number on the **periodic table**, but it can include decimal points.
- X is the chemical symbol used for each of the elements in the periodic table.

Mass number A, Atomic number Z, Element symbol X
FIGURE 2 The nuclear symbol; for example, $^{4}_{2}$He.

Applying the nuclear symbol

An atom, in its original state, has a neutral charge because the positive charge from the protons cancels out the negative charge from the electrons. If the protons and electrons are not balanced, then the atom has an ionic charge (either positive or negative). The charged atom is called an **ion**.

An **isotope** is an atom with a variable number of neutrons but the same number of protons and electrons. The concept of ions and isotopes is covered in more detail in Chapter 4.

Table 1 outlines how to determine the number of subatomic particles present. This applies to neutral atoms, ions and isotopes.

TABLE 1 Determining subatomic particles

Subatomic particle	Neutral atom	Ion	Isotope
Proton (p^+)	Use the atomic number from the periodic table. $p^+ = Z$	Use the atomic number from the periodic table. $p^+ = Z$	Use the atomic number from the periodic table. $p^+ = Z$
Neutron (n^0)	Subtract the atomic number from the mass number given in the periodic table. $n^0 = A - Z$	Subtract the atomic number from the mass number given in the periodic table. $n^0 = A - Z$	Subtract the atomic number from the atomic mass of the specific isotope. $n^0 =$ atomic mass $- Z$
Electron (e^-)	Use the atomic number from the periodic table. $e^- = Z$	Subtract the ionic charge from the atomic number. $e^- = Z -$ ionic charge	Use the atomic number from the periodic table. $e^- = Z$

WORKED EXAMPLE 2.1
a Determine the number of neutrons in the neutral atom $^{19}_{9}$F.
b Determine the number of electrons in the negative ion $^{14}N^{3-}$.

SOLUTION
a $n^0 = A - Z$
$= 19 - 9$
$= 10$ neutrons

b $e^- = Z -$ ionic charge
$= 7 - (-3)$
$= 10$ electrons

Study tip
The first letter of a chemical symbol is always capitalised. If there is a second letter, it is lower case.

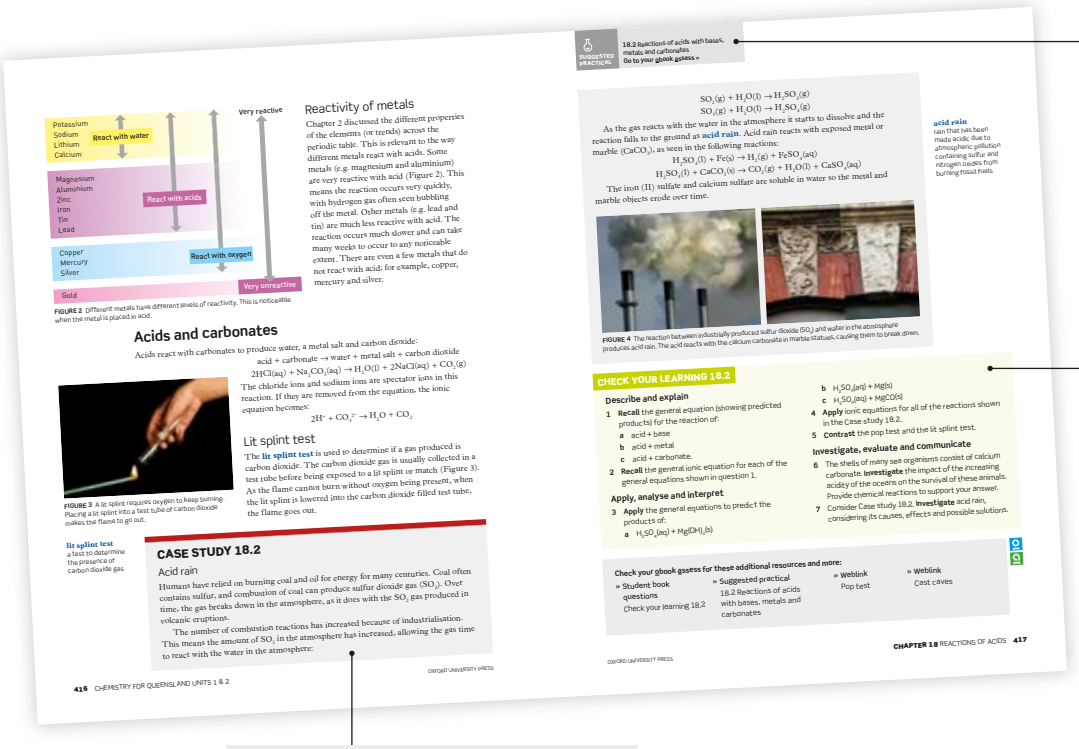
periodic table — an organised presentation of elements by increasing atomic number

Study tip
Not every periodic table is presented in the same manner, but there will be a legend to help interpret the information. The legend is usually in the middle with hydrogen as an example.

ion — an atom containing an electric charge
isotope — atoms that have the same number of protons and electrons but different numbers of neutrons

CHAPTER 2 ATOMIC STRUCTURE AND THE PERIC... OXFORD UNIVERSITY PRESS

Worked example
Full mathematical working is provided for key chemical equations and formulas.



Reactivity of metals

Chapter 2 discussed the different properties of the elements (or trends) across the periodic table. This is relevant to the way different metals react with acids. Some metals (e.g. magnesium and aluminium) are very reactive with acid (Figure 2). This means the reaction occurs very quickly, with hydrogen gas often seen bubbling off the metal. Other metals (e.g. lead and tin) are much less reactive with acid. The reaction occurs much slower and can take many weeks to occur to any noticeable extent. There are even a few metals that do not react with acid: for example, copper, mercury and silver.

Potassium, Sodium, Lithium, Calcium — React with water
Magnesium, Aluminium, Zinc, Iron, Tin, Lead — React with acids
Copper, Mercury, Silver — React with oxygen
Gold — Very unreactive
Very reactive
FIGURE 2 Different metals have different levels of reactivity. This is noticeable when the metal is placed in acid.

Acids and carbonates

Acids react with carbonates to produce water, a metal salt and carbon dioxide:

acid + carbonate → water + metal salt + carbon dioxide

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

The chloride ions and sodium ions are spectator ions in this reaction. If they are removed from the equation, the ionic equation becomes:

$2H^+ + CO_3^{2-} \rightarrow H_2O + CO_2$

Lit splint test

The **lit splint test** is used to determine if a gas produced is carbon dioxide. The carbon dioxide gas is usually collected in a test tube before being exposed to a lit splint or match (Figure 3). As the flame cannot burn without oxygen being present, when the lit splint is lowered into the carbon dioxide filled test tube, the flame goes out.

FIGURE 3 A lit splint requires oxygen to keep burning. Placing a lit splint into a test tube of carbon dioxide makes the flame to go out.

lit splint test — a test to determine the presence of carbon dioxide gas

CASE STUDY 18.2
Acid rain

Humans have relied on burning coal and oil for energy for many centuries. Coal often contains sulfur, and combustion of coal can produce sulfur dioxide gas (SO_2). Over time, the gas breaks down in the atmosphere, as it does with the SO_2 gas produced in volcanic eruptions.

The number of combustion reactions has increased because of industrialisation. This means the amount of SO_2 in the atmosphere has increased, allowing the gas time to react with the water in the atmosphere:

$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(g)$
$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(g)$

As the gas reacts with the water in the atmosphere it starts to dissolve and the reaction falls to the ground as **acid rain**. Acid rain reacts with exposed metal or marble ($CaCO_3$), as seen in the following reactions:

$H_2SO_4(l) + Fe(s) \rightarrow H_2(g) + FeSO_4(aq)$
$H_2SO_4(l) + CaCO_3(s) \rightarrow CO_2(g) + H_2O(l) + CaSO_4(aq)$

The iron (II) sulfate and calcium sulfate are soluble in water so the metal and marble objects erode over time.

FIGURE 4 The reaction between industrially produced sulfur dioxide (SO_2) and water in the atmosphere produces acid rain. The acid reacts with the calcium carbonate in marble statuary, causing them to break down.

acid rain — rain that has been made acidic due to atmospheric pollution containing sulfur and nitrogen oxides from burning fossil fuels

CHECK YOUR LEARNING 18.2

Describe and explain
1 **Recall** the general equation (showing predicted products) for the reaction of:
a acid + base
b acid + metal
c acid + carbonate.
2 **Recall** the general ionic equation for each of the general equations shown in question 1.

Apply, analyse and interpret
3 **Apply** the general equations to predict the products of:
a $H_2SO_4(aq) + Mg(OH)_2(s)$
b $H_2SO_4(aq) + Mg(s)$
c $H_2SO_4(aq) + MgCO(s)$
4 **Apply** ionic equations for all of the reactions shown in the Case study 18.2.
5 **Contrast** the pop test and the lit splint test.

Investigate, evaluate and communicate
6 The shells of many sea organisms consist of calcium carbonate. **Investigate** the impact of the increasing acidity of the oceans on the survival of these animals. Provide chemical reactions to support your answer.
7 Consider Case study 18.2. **Investigate** acid rain, considering its causes, effects and possible solutions.

Check your gbook assess for these additional resources and more:
» Student book questions — Check your learning 18.2
» Suggested practical — 18.2 Reactions of acids with bases, metals and carbonates
» Weblink — Pop test
» Weblink — Cast caves

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SUGGESTED PRACTICAL 18.2 Reactions of acids with bases, metals and carbonates
Go to your gbook assess »

Practical links
Mandatory and suggested practicals are linked in the relevant section of the Student book.

Check your learning
Activity boxes with questions and tasks are organised using cognitive verbs according to Marzano and Kendall's taxonomy.

Case studies
Real-life examples illustrate theoretical points being explained in the text.

Science as a human endeavour

Real-world contexts promotes curiosity and can be used as a starting point for research investigations.

SCIENCE AS A HUMAN ENDEAVOUR

5.4 Distribution of elements in the universe

KEY IDEAS

In this section, you will learn about:

- the analysis of the stars by using electromagnetic spectra, specifically their speed, temperature and composition.

Spectroscopy provides valuable information about the universe. Scientists have long been interested in the speed, composition and temperature of stars, galaxies and various other phenomena, that are well beyond our reach. A scientific understanding of the universe will lead to a better understanding of the life of our own star, the Sun.

Carbon-based organisms

Living organisms are often referred to as 'carbon-based' life forms because most of the molecules that form these organisms are made from elements bonded to carbon atoms. However, this does not mean that carbon is the most abundant element.

Within the human body, the most abundant element (or the most common element by mass) is oxygen. Approximately 65% of the human body is oxygen, whereas carbon only comprises 18.5%. The next two abundant elements are hydrogen (9.5%) and nitrogen (3.2%). Other elements occur in trace (very small) amounts; for example, calcium (1.5%), potassium (0.4%), phosphorus (1.0%) and sulfur (0.3%).

Our planet Earth

Different regions of Earth have different elements. The sea has a different elemental composition from Earth's crust, core and atmosphere.

The most abundant element in the atmosphere is nitrogen. The most abundant element in the oceans and Earth's crust is oxygen. The Earth's core consists mainly of an alloy of nickel and iron. Therefore, it can be misleading to state an overall composition of Earth because it does not reflect the differences between its many layers.

Overall, the most abundant elements on Earth are iron, oxygen, silicon, magnesium, sulfur, nickel, calcium and aluminium. These are all elements found in rock-based formations and it is unsurprising that oxygen has such a high abundance when it is located in rock-based compounds, water and the atmosphere.

Minerals in the Earth's crust

Based on what was known

Doppler effect

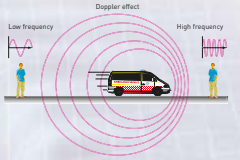

FIGURE 1 The Doppler effect explains why an ambulance siren sounds different to a pedestrian as the ambulance goes past.

The **Doppler effect**, named for Austrian physicist Christian Doppler, is essentially the change in frequency or wavelength observed by a bystander as the source of the wave moves past them. It is best observed when an ambulance passes a pedestrian on the side of the road. As the ambulance approaches, the pedestrian hears the sound of the ambulance siren as a high-pitched (and therefore high-frequency) 'yeee' sound. As the ambulance passes, the sound changes to a low-pitched (and therefore low-frequency) 'owww' sound. Therefore the pitch and frequency of sound shift relative to the motion of the ambulance and the position of the pedestrian.

doppler effect the change in a frequency of a wave as an object moves towards or away from an observer

The light from moving stars can also be observed to shift in frequency. An object moving towards Earth emits radiation that becomes squeezed. Therefore, the wavelength decreases and the frequency increases. As a result, its absorption spectrum shifts slightly towards the blue end. This is called **blue shift** (Figure 2). Conversely, an object moving away from Earth emits [illegible] that spreads [illegible]art. The [illegible] increases [illegible]quency. As a result, [illegible] spectrum [illegible]y toward [illegible]lting in **red** [illegible]e 2).

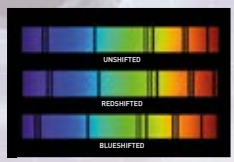

FIGURE 2 The blue shift of galaxies moving towards the Earth and the red shift of galaxies moving away from the Earth.

blue shift a shift in the spectrum towards the blue region, indicating that a galaxy is moving away from the observer

red shift a shift in the spectrum towards the red region, indicating that a galaxy is moving away from the observer

CHAPTER 5

Review

Chapter summary

5.1
- Electrons orbit an atom's nucleus in shells that have discrete energy levels. To transition to a higher energy level, electrons must absorb a photon of light with specific wavelength. Electrons are then in an excited state. An electron will emit a photon of light of specific wavelength when it returns to its ground state. The light both absorbed and emitted can be measured on the electromagnetic spectrum.

5.2
- The flame test is an analytical technique in which electrons are excited by heat from the blue flame of a Bunsen burner. When they return to ground level, the electrons emit a photon of light of a specific wavelength, which is measured on an emission spectrum.
- AAS is a technique that atomises a sample and shines light of a specific wavelength at the atoms. The light excites electrons to a higher electron shell and the electrons emit light when returning to the ground state. The emitted light is measured at a specific wavelength and the resultant absorbance is measured against concentration.
- To determine the concentration of a sample in AAS, a series of standards of known concentration must be analysed. The absorbance is recorded and used to create a calibration curve. The sample's absorbance is plotted on the calibration curve to determine the concentration.

5.3
- Mass spectrometry ionises a sample of an element and creates a spectrum according to the mass and abundance of the sample's isotopes. These are used to calculate the relative atomic mass of an element.
- If the element contains no more than two isotopes and A_r and RIM are known for both, then the percentage composition of the element can be calculated. The isotopes' percentage abundances are represented by x and 100 – x, as they must add up to 100. Once substituted into the equation, solve for x.

5.4
- The electromagnetic spectrum can be interpreted to determine the speed and direction of galaxy movement, the temperature of stars and the composition of stars.

Key terms

- absorption spectrum
- atomic absorption spectroscopy (AAS)
- blue shift
- calibration curve
- Doppler effect
- electromagnetic spectrum
- emission spectrum
- excited state
- flame test
- frequency
- ground state
- ionisation
- mass spectrometer
- mass spectrometry
- mass spectrum
- percentage abundance
- photon
- red shift
- relative atomic mass (A_r)
- relative isotopic mass (RIM)
- standard
- wavelength

Key formulas

relative atomic mass $A_r = \frac{(\text{RIM}_1 \times \text{abundance}_1(\%)) + (\text{RIM}_2 \times \text{abundance}_2(\%)) + (\text{RIM}_3 \times \text{abundance}_3(\%)) + \ldots}{100}$

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; *** = high.

Multiple choice

1 Which of the following is inconsistent with Bohr's model of electron configuration?
- A Electrons exist in orbitals.
- B Electrons cannot move into different shells.
- C Electrons orbit the nucleus of an atom.
- D For an electron to transition between orbitals, it must absorb or emit energy.

2 When an atom absorbs electromagnetic radiation of a specific frequency, electrons:
- A spin faster around the nucleus of an atom
- B move to an electron shell further from the nucleus
- C spin in the opposite direction
- D move into an electron shell closer to the nucleus.

3 An emission spectrum is produced when:
- A electrons orbiting the nucleus are promoted to a higher energy level
- B heat from a Bunsen burner is used to excite electrons
- C electrons leave an atom's electron shells and an ion is formed
- D electrons shed excess energy by releasing it as light.

4 On the electromagnetic spectrum, visible light is measure between the wavelengths of:
- A 400 and 750 nm
- B 300 and 700 nm
- C 500 and 800 nm
- D 200 and 600 nm.

5 What information can a mass spectrometer provide to a chemist?
- A the composition of elements in a sample
- B the concentration of atoms in a sample
- C the properties of molecules within a sample
- D the relative mass of atoms within a sample

Short answer

Describe and explain

* 6 **Explain** how absorption and emission spectra occur.

* 7 **Explain** the difference and the relationship between relative isotopic mass (RIM), relative atomic mass (A_r) and mass number.

Apply, analyse and interpret

* 8 **Differentiate** between the Lyman, Balmer and Paschen series.

*** 9 **Determine** the atomic mass number of the following elements from the isotope data given.

a Silicon has three isotopes.

TABLE 1

isotope	RIM	% abundance
^{28}Si	27.977	92.229
^{29}Si	28.976	4.683
^{30}Si	29.974	3.087

b Europium has two isotopes.

TABLE 2

isotope	RIM	% abundance
^{151}Eu	150.9198	47.81
^{153}Eu	152.9212	52.19

c Zinc has five isotopes.

TABLE 3

isotope	RIM	% abundance
^{64}Zn	63.929	48.63
^{66}Zn	65.926	27.90
^{67}Zn	66.927	4.10
^{68}Zn	67.925	18.75
^{70}Zn	69.925	0.62

*** 10 **Determine** the percentage abundance of the isotopes, given the following information.

a Bromine has a mass number of 79.9. Its isotopes have relative isotopic masses of 78.918 and 80.916.

b Indium has a mass number of 114.8. Its isotopes have relative isotopic masses of 112.904 and 114.904.

c Rubidium has a mass number of 85.5. Its isotopes have relative isotopic masses of 84.912 and 86.909.

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Chapter review

Each chapter review includes:

- a **summary of key learning** in each chapter
- revision questions written to target assessment through multiple-choice and short-answer questions
- key terms introduced throughout the chapter
- key formulas for the chapter.

Practicals

Each chapter contains a range of practical activities to reinforce theoretical aspects of the syllabus. All mandatory practicals are included in the practical manual and all suggested practicals are included in the obook assess.

CHAPTER 20

Practical manual

The QCAA Chemistry General Senior Syllabus outlines a number of mandatory and suggested practicals for completion in Units 1 & 2. Suggestions for methodology and materials have been supplied in this chapter and are outlined in the table on the next page. However, the following is not prescriptive; schools may complete the mandatory or suggested practicals in any other form suited to their resources.

The experiments in this chapter have been trialled, and obvious cautions of hazards given; however, teachers are legally obliged to carry out their own risk assessment prior to undertaking any practical activity.

FIGURE 1 A scientist loading the microscope stage while wearing the correct personal protective equipment

UNIT 1 PRACTICALS

MANDATORY	8.5 Finding the percentage yield of magnesium oxide
MANDATORY	9.3 Calorimetry: Measuring the enthalpy of a reaction
SUGGESTED	2.4 Investigating periodic trends using a database
SUGGESTED	2.5 Simulating Rutherford's gold foil experiment
SUGGESTED	5.2 Conducting a flame test
SUGGESTED	6.2 Separating mixtures
SUGGESTED	7.4 Testing for saturation
SUGGESTED	8.4 Investigating limiting agents
SUGGESTED	9.2A Applying Hess's law
SUGGESTED	9.2B Measuring the molar heat of a chemical reaction

UNIT 2 PRACTICALS

MANDATORY	12.2 Constructing 3D models of molecules
MANDATORY	14.1 Investigating the properties of gases
MANDATORY	17.1 Investigating the properties of strong and weak acids
MANDATORY	19.1 Investigating the rate of chemical reactions
SUGGESTED	13.2 Separating the compounds of a mixture by paper chromatography
SUGGESTED	14.2 Boyle's law
SUGGESTED	16.1 Precipitation reactions
SUGGESTED	16.2 Investigating the effect of temperature on solubility
SUGGESTED	18.2 Reactions of acids with bases, metals and carbonates

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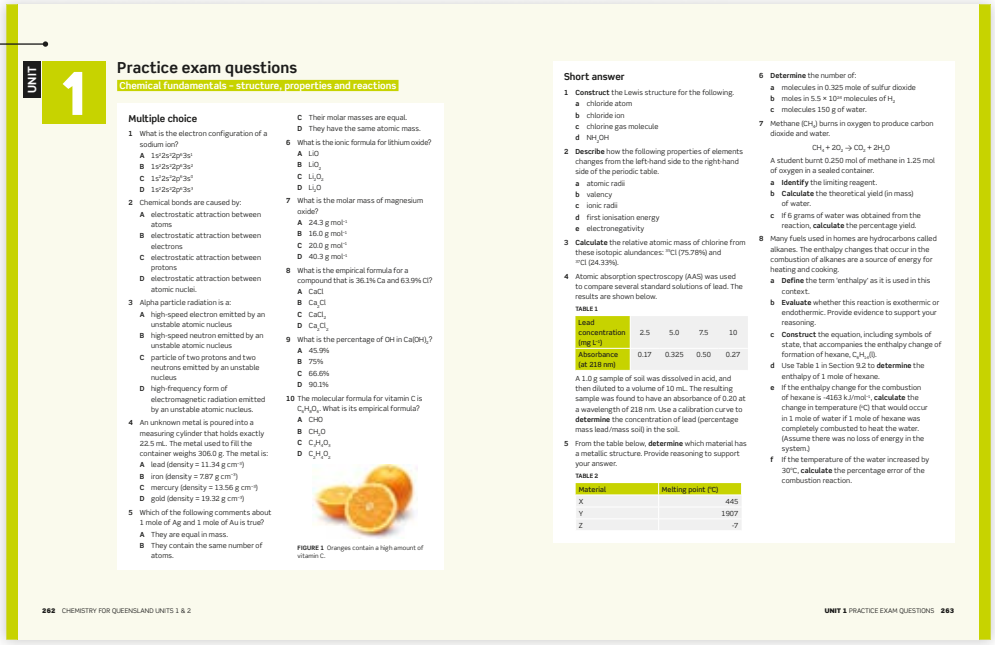
Practice exam questions
Questions can be found at the end of each unit to prepare students for their end-of-year external assessments.



UNIT 1

Practice exam questions

Chemical fundamentals – structure, properties and reactions

Multiple choice

1 What is the electron configuration of a sodium ion?
A $1s^2 2s^2 2p^6 3s^1$
B $1s^2 2s^2 2p^6 3s^2$
C $1s^2 2s^2 2p^6 3s^0$
D $1s^2 2s^2 2p^6 3s^2$

2 Chemical bonds are caused by:
A electrostatic attraction between atoms
B electrostatic attraction between electrons
C electrostatic attraction between protons
D electrostatic attraction between atomic nuclei.

3 Alpha particle radiation is a:
A high-speed electron emitted by an unstable atomic nucleus
B high-speed neutron emitted by an unstable atomic nucleus
C particle of two protons and two neutrons emitted by an unstable nucleus
D high-frequency form of electromagnetic radiation emitted by an unstable atomic nucleus.

4 An unknown metal is poured into a measuring cylinder that holds exactly 22.5 mL. The metal used to fill the container weighs 306.0 g. The metal is:
A lead (density = 11.34 g cm^{-3})
B iron (density = 7.87 g cm^{-3})
C mercury (density = 13.56 g cm^{-3})
D gold (density = 19.32 g cm^{-3})

5 Which of the following comments about 1 mole of Ag and 1 mole of Au is true?
A They are equal in mass.
B They contain the same number of atoms.
C Their molar masses are equal.
D They have the same atomic mass.

6 What is the ionic formula for lithium oxide?
A LiO
B LiO_2
C Li_2O_2
D Li_2O

7 What is the molar mass of magnesium oxide?
A 24.3 g mol^{-1}
B 16.0 g mol^{-1}
C 20.0 g mol^{-1}
D 40.3 g mol^{-1}

8 What is the empirical formula for a compound that is 36.1% Ca and 63.9% Cl?
A CaCl
B Ca_2Cl
C $CaCl_2$
D Ca_2Cl_3

9 What is the percentage of OH in $Ca(OH)_2$?
A 45.9%
B 75%
C 66.6%
D 90.1%

10 The molecular formula for vitamin C is $C_6H_8O_6$. What is its empirical formula?
A CHO
B CH_2O
C $C_3H_4O_3$
D $C_3H_2O_3$

FIGURE 1 Oranges contain a high amount of vitamin C.

Short answer

1 **Construct** the Lewis structure for the following.
a chloride atom
b chloride ion
c chlorine gas molecule
d NH_4OH

2 **Describe** how the following properties of elements changes from the left-hand side to the right-hand side of the periodic table.
a atomic radii
b valency
c ionic radii
d first ionisation energy
e electronegativity

3 **Calculate** the relative atomic mass of chlorine from these isotopic abundances: ^{35}Cl (75.78%) and ^{37}Cl (24.33%).

4 Atomic absorption spectroscopy (AAS) was used to compare several standard solutions of lead. The results are shown below.

TABLE 1

Lead concentration (mg L^{-1})	2.5	5.0	7.5	10
Absorbance (at 218 nm)	0.17	0.325	0.50	0.27

A 1.0 g sample of soil was dissolved in acid, and then diluted to a volume of 10 mL. The resulting sample was found to have an absorbance of 0.20 at a wavelength of 218 nm. Use a calibration curve to **determine** the concentration of lead (percentage mass lead/mass soil) in the soil.

5 From the table below, **determine** which material has a metallic structure. Provide reasoning to support your answer.

TABLE 2

Material	Melting point (°C)
X	445
Y	1907
Z	-7

6 **Determine** the number of:
a molecules in 0.325 mole of sulfur dioxide
b moles in 5.5×10^{24} molecules of H_2
c molecules 150 g of water.

7 Methane (CH_4) burns in oxygen to produce carbon dioxide and water.

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

A student burnt 0.250 mol of methane in 1.25 mol of oxygen in a sealed container.
a **Identify** the limiting reagent.
b **Calculate** the theoretical yield (in mass) of water.
c If 6 grams of water was obtained from the reaction, **calculate** the percentage yield.

8 Many fuels used in homes are hydrocarbons called alkanes. The enthalpy changes that occur in the combustion of alkanes are a source of energy for heating and cooking.
a **Define** the term 'enthalpy' as it is used in this context.
b **Evaluate** whether this reaction is exothermic or endothermic. Provide evidence to support your reasoning.
c **Construct** the equation, including symbols of state, that accompanies the enthalpy change of formation of hexane, $C_6H_{14}(l)$.
d Use Table 1 in Section 9.2 to **determine** the enthalpy of 1 mole of hexane.
e If the enthalpy change for the combustion of hexane is -4163 kJ/mol^{-1}, **calculate** the change in temperature (°C) that would occur in 1 mole of water if 1 mole of hexane was completely combusted to heat the water. (Assume there was no loss of energy in the system.)
f If the temperature of the water increased by 30°C, **calculate** the percentage error of the combustion reaction.

obook assess

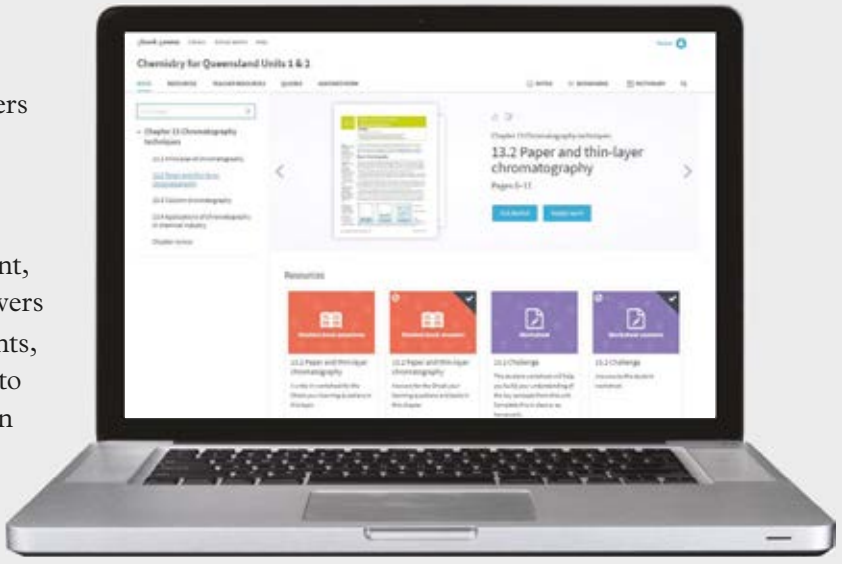
Chemistry for Queensland Units 1 & 2 is supported by a range of engaging and relevant digital resources via obook assess.

Students receive:

- a complete digital version of the Student book with notetaking and bookmarking functionality
- video tutorials demonstrating key skills
- write-in worksheets to accompany all mandatory and suggested practicals
- interactive auto-correcting multiple-choice quizzes
- a range of engaging weblinks to support understanding
- access to work assigned by their teacher: reading, homework, tests, assignments.

In addition to the student resources, teachers also receive:

- detailed planning resources
- Student book answers
- printable (and editable) sample assessment, including data tests and exams with answers
- the ability to set up classes, set assignments, monitor progress and graph results, and to view all available content and resources in one place.



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CHAPTER 1

Chemistry toolkit

Chemistry is an experimental science that studies matter, how substances can be combined or separated, and how substances interact with energy. 'Experimental' refers to how theories are developed to explain the nature and changes to matter. When a new theory is proposed, it can be tested using careful observations and measurements. It is useful to keep asking questions such as: 'What observations lead chemists to think this? What is the evidence?' If you keep asking these questions, you will be developing as a scientist.

While studying Chemistry Units 1 & 2, there are a number of helpful skills and ideas to learn. In particular, it will be important to know how to design an experiment using the scientific method; how to collect, present and analyse data; how to evaluate claims; and how to present results or findings in a proper scientific report.

This chapter is intended to be used as a 'toolkit' for those skills. It can be referred to at any point throughout the year as a resource that provides guidance for assessment, including the student experiment and the research investigation, as well as important tips for exam preparation.

FIGURE 1 Chemistry is an experimental science that involves a range of skills and equipment.

1.1 Why study chemistry?

KEY IDEAS

In this section, you will learn about:

- the ways in which chemistry is relevant to many areas of life and occupations
- how chemistry is essential for understanding the physical world.

The Australian Bureau of Statistics (ABS) is the government organisation responsible for carrying out the Australian Census every five years. As part of the census, the ABS collects information about occupations. According to the ABS, the work of chemists involves:

- conducting experiments and tests to identify the chemical composition and reactive properties of natural substances and processed materials
- analysing and conducting research to develop theories, techniques and processes, and testing the reliability of outcomes under different conditions
- developing practical applications of experimental and research findings

Around 9000 people in Australia are employed directly as chemists. However, many occupations require some understanding of chemistry, such as people who work in health or medicine, or people who need to know how to handle chemicals safely for their work. These occupations include doctors, nurses, nutritionists, dentists, pharmacists, painters, carpenters and builders, hairdressers and farmers.

chemistry
an experimental science, concerned with the study of matter, and how substances can be combined or separated and how substances interact with energy

model
a simplified description of a process or phenomenon, often in a mathematical form

Chemistry is often called the central science because an understanding of chemistry is required to make sense of phenomena in many other areas of science and technology, such as materials science (corrosion or the strength of materials), ecology (flow of nutrients through an ecosystem), environmental science (how pollutants may be dispersed in an ecosystem) and molecular biology (how genes are read and regulated on the molecular level).

Even if someone's work doesn't directly involve chemistry, there is no way to avoid chemistry in daily life. The modern world is the product of chemical technologies, from metals and plastics, to the food that keeps people alive.

Chemistry allows society to change the world by the formation of new substances and materials. The ethical question about how the discoveries in chemistry should be used is one that affects everyone because it directly influences the world we live in. To be able to meaningfully contribute to resolving those questions, an understanding of chemistry is required.

The best reason to study chemistry is because it can be a lot fun and very interesting to learn about the substances that make up our universe. Chemistry is not a complete and settled science. There are many further discoveries yet to be made and it opens up so many possibilities. By studying chemistry, you are giving yourself the opportunity to appreciate the wonders of the world around you in a new and different way.

Study tip

Think about why you are studying chemistry. Is it to satisfy your curiosity and learn more about the physical world? Do you have a particular career goal in mind? Keeping your ultimate goal in mind will help motivate you while studying.

TABLE 1 Main industry sectors employing chemists in Australia

Industry sector	Percentage of chemists in Australia (%)
Manufacturing	50.0
Professional, scientific and technical services	15.5
Retail trade	8.6
Health care and social assistance	6.7
Other industries	19.2

What chemists do

Chemists basically:

- observe and measure properties of substances
- explain and predict patterns in the behaviour of substances
- use and extend their knowledge to make new substances or purify and detect existing substances.

FIGURE 1 To carry out experiments safely, chemists use personal protective equipment (PPE) such as lab coats, safety glasses or goggles and gloves.

Chemists observe and measure properties of substances

Chemists attempt to understand the patterns in the way that different substances behave. They do this by making detailed observations and measurements. Students studying chemistry learn to take note of certain things that may indicate that a chemical change is taking place. These may be things like a colour change, bubbles of a gas or a solid forming, or heat being given off or absorbed. These are all changes that can be observed without specialised equipment.

In order to extend the range of chemical changes that can be investigated beyond those that can be detected by our physical senses, chemists also use instruments to make measurements. For example, a thermometer measures a change in temperature. A spectrometer can measure changes in the absorbance of light outside of the visible region of the spectrum. A scanning electron microscope enables images of individual atoms to be constructed. These instruments are tools that allow chemists to explore what is happening on the very small scale of atoms and molecules.

FIGURE 2 The observable colour change in this mixture over time indicates to chemists that a chemical change is occurring.

Chemists explain and predict patterns in the behaviour of substances

The field of chemistry is based on a very simple assumption: that there are patterns in the way different substances behave and these patterns can be discovered by observation. Advances in chemistry come through the use of **models** that explain these patterns. Models are simplified descriptions of processes or systems.

Models are always descriptions (often in mathematical form) of reality; they are predictions, not prescriptions of what should happen. If the model doesn't fit reality, then the model needs to change. Experiments are the way that chemists check that their models are 'good enough'. A 'good enough' model adequately explains and predicts the behaviour of substances under the defined conditions of the experiment.

For example, Russian chemist Dmitri Mendeleev attempted to classify different types of atoms (called elements) according to their

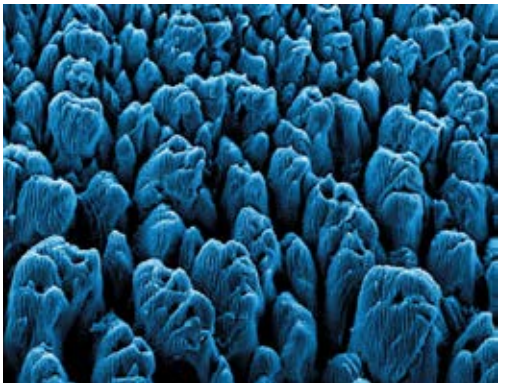

FIGURE 3 A scanning electron micrograph of the rippled microstructures on silicon.

properties. He noticed that there were repeating patterns, called 'periodicity', which led him to organise the elements in a periodic table. This enabled him to predict the existence of new elements that hadn't yet been discovered. Later scientists have used the periodic table to develop theories of how the arrangements of subatomic particles (called electrons) affect the properties of those atoms.

quantitative
relating to numerical data about a substance, an object or a phenomenon

The models that chemists use should be **quantitative** if possible. A quantitative model expresses relationships between different experimental variables using numbers and a formula. This allows the size of an effect to be predicted, not just whether something will happen or not. A model can be tested experimentally to determine whether it is correct or not and then modified or discarded if it does not give reliable predictions.

A useful model allows connections to be made between what is happening on the macroscopic level (the 'real' world level) and what is happening at the much smaller scale of submicroscopic particles. This is done using the specialised symbolic representations of chemistry and may also express relationships using mathematical formulas. Chemistry students need to be able to think about phenomena in all three ways (on the macroscopic and submicroscopic scales and symbolically) and to be able to move between them.

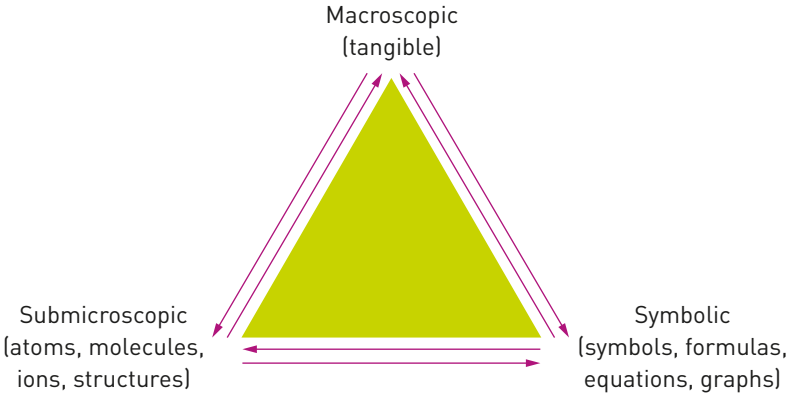

FIGURE 4 It is important to think about chemical phenomena in three ways: macroscopic, submicroscopic and symbolic.

CASE STUDY 1.1

Using models: the kinetic theory of gases

One example of a model used by chemists is the kinetic theory of gases. The theory is explained more in Chapter 14. By making some simple assumptions about the particles that make up a gas, it becomes possible to make a mathematical model that relates pressure, volume, temperature and the amount of gas in a sealed container.

The model was developed to explain the observations made by early chemists about the behaviour of gases. It allows chemists to describe their observations in terms of what is happening at the submicroscopic level of the particles making up gases. The model can be expressed mathematically in the form of the ideal gas equation:

$$pV = nRT$$

where p is pressure, V is volume, T is temperature (measured in special units called kelvin, K), n is number of moles (or the amount of the gas) and R is the ideal gas constant.

Chemists use and extend their knowledge to make new substances or purify and detect existing substances

As well as explaining how the world works, chemists change the world by developing new substances and materials that have never existed in nature before. These molecules may be developed to have a particular function, such as to treat a disease, to remove a pollutant from the environment, to make engines run smoother or more economically or to act as an electrode in a new type of battery. Sometimes, new substances are made so that chemists can test their theories experimentally.

The number and variety of substances found in nature is also a challenge to synthetic chemists (who make substances in the laboratory). Many useful substances were first found in nature, many with uses as drugs, or are modifications of those substances, but are now made by **synthetic chemistry** in the laboratory or in a pharmaceutical factory. Even if a useful substance is found in sufficient amounts from a natural source, it must be purified and separated from the other substances also present.

Analytical chemistry is the branch of chemistry concerned with whether a substance is present in a particular sample (detection), and if it is, how much of the substance is present (quantitation). Using advanced techniques, it is now possible to routinely detect substances in concentrations of parts per trillion (ppt), the equivalent of one drop of water in an Olympic-sized swimming pool.

FIGURE 5 Synthetic chemistry is conducted in laboratories or pharmaceutical factories.

synthetic chemistry
the branch of chemistry concerned with making substances in the laboratory and discovering new methods to transform substances into new substances

analytical chemistry
the branch of chemistry concerned with the detection and measurement of amount of substances

CHECK YOUR LEARNING 1.1

Describe and explain

1 **Describe** the three main aspects of what chemists do.
2 **Explain** the purpose of a 'model' in chemistry.

Apply, analyse and interpret

3 **Determine** how chemistry would be relevant to:
- **a** a wheat farmer from the Darling Downs
- **b** a marine biologist working on the Great Barrier Reef
- **c** an environmental toxicologist working in Mt Isa
- **d** an architect designing shade structures for a park in Townsville
- **e** a technician monitoring water quality for Queensland Urban Utilities at the South Pine dam
- **f** a perfume maker from Noosa
- **g** a council worker controlling mosquitoes in Cairns.

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- » Video
  Ask a scientist – Ms Jenny Powell
- » Video worksheet
  Ask a scientist – Ms Jenny Powell
- » Weblink
  QCAA syllabus

1.2 Data analysis

KEY IDEAS

In this section, you will learn about:

- categorising data as quantitative or qualitative
- presenting data in tables and graphs
- analysing data using features such as mean, mode, median and standard deviation.

The ability to present, analyse and interpret data is an important part of Chemistry. As part of this, it is important to understand the different types of data.

Data is typically categorised as quantitative or **qualitative**.

qualitative data about types or categories, which may be represented by names, symbols, codes etc.

TABLE 1 Qualitative versus quantitative data

	Quantitative data	Qualitative data
Description	Data about 'quantities' or values, which can be measured and expressed numerically. They are numerical variables.	Data about types or categories, which may be represented by names, symbols, codes etc. They are categorical variables.
Examples	Mass of a sample, number of molecules in a mole, volume of a gas.	Types of substances, colours of substances, malleability of metal.

Presenting data

After measurements are collected and data is obtained, these pieces must be put together in a useful way.

When presenting data, it is important to consider the best way to display that data. What will be the most informative way to present the data? It needs to be clear and easy to read.

Common types of data representation used in chemistry include tables, x–y scatter plots and bar graphs.

Tables

Tables are often used to summarise and organise data. A table should have a heading, as well as column and row headings. Where quantities are being reported, appropriate units should appear in the headings. Tables are best where there is a relatively small amount of data and the values are individually important, as shown in Table 2.

TABLE 2 Change in reaction temperature with time

Time (min)	Temperature (°C)
5	43
10	37
15	35
20	24

Scatterplots

Although tables are useful, graphs can be more useful where trends in data are being investigated for showing how one variable changes either over time or in response to changes in another variable.

x–*y* scatterplots are often used where there is a continuous trend in a measured value as one value is altered in a systematic way; for example, how the temperature of a reaction changes with time. To describe this graph, the *y*-axis is named first, the names of the two axes are separated by 'vs', which is the abbreviation for 'versus', and both axes names contain the units in brackets after the title: 'Temperature change (°C) vs Time (min) as shown in Figure 1.

Where there is a linear relationship, a **line of best fit** can be added to the graph, which can allow values between the extremes of the graph points to be estimated. However, a line of best fit should not be added without some thought. Not all relationships between variables are linear, although sometimes a linear relationship can be found by applying a mathematical operation to one or both variables.

line of best fit (trend line)
a line drawn on a graph joining as many points as possible and showing the general direction of the data; should be drawn with, approximately, an equal number of points above and below the line

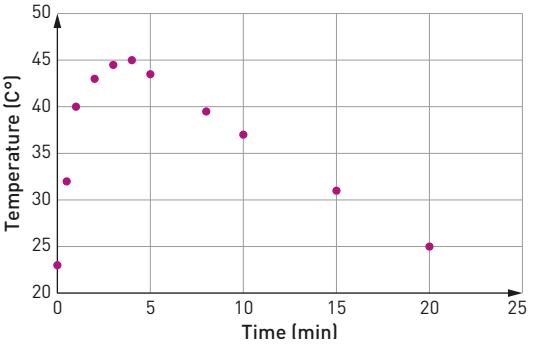

FIGURE 1 An *x*–*y* scatterplot showing how the temperature of a reaction (the dependent variable) changes with time (the independent variable).

Bar graphs and other charts

Bar graphs are useful for comparing different categories or samples; for example, the levels of dissolved phosphorus in different south-east Queensland water catchments (Figure 2).

There are many other types of charts, including pie charts and box-and-whisker plots. Pie charts (Figure 3) are good for comparing six or fewer categories, showing the proportion or percentage associated with each category.

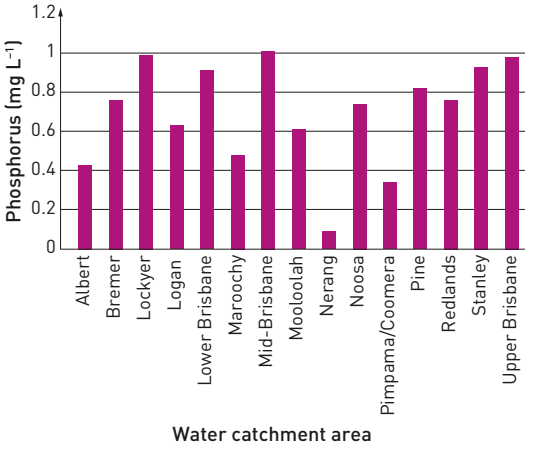

FIGURE 2 A bar chart showing phosphorus levels (mg L^{-1}) in south-east Queensland water catchments for 2017.

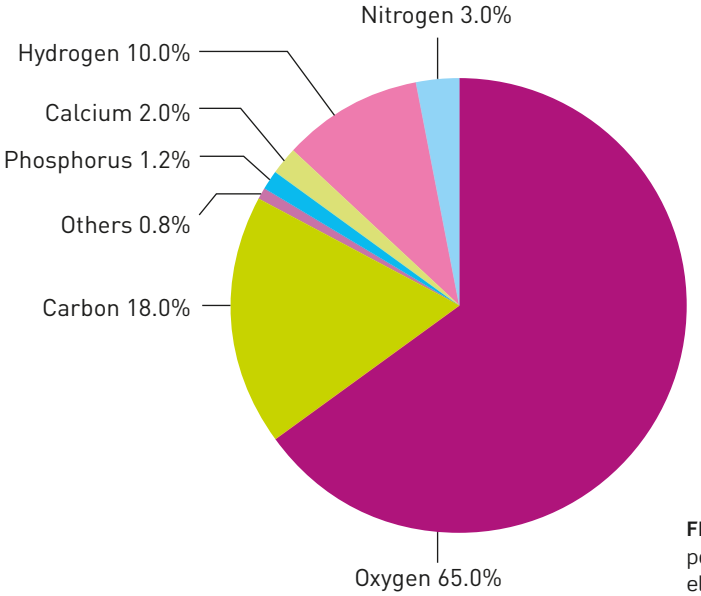

FIGURE 3 A pie chart showing the percentage distribution by mass of elements in the body.

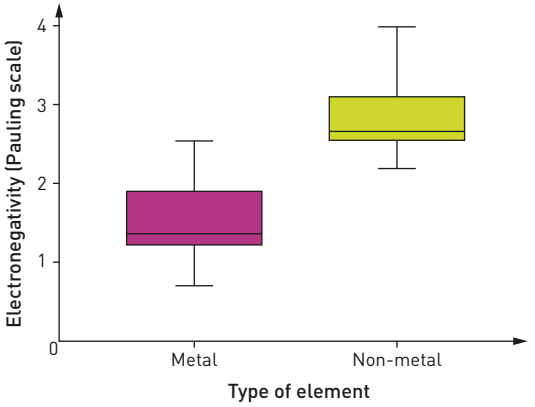

FIGURE 4 A box-and-whisker plot comparing the electronegativity of metals and non-metals. The boxes represent the middle quartiles (i.e. the values in the range 25–75%) and the whiskers represent the minimum and maximum values for each category. The median is the central line in the box, i.e. 50% of the values are above and 50% are below this value.

Box-and-whisker plots (Figure 4) are good for summarising a great deal of statistical information about groups of data, including the median and quartile values. Some spreadsheet programs allow you to graph data using several different types of charts. When choosing the type of chart to use, think about which type is most appropriate for the data and which will be most easily understood by a reader.

Analysing data

Scientific data needs to be analysed and interpreted. The goal of data analysis is to find patterns or trends, whereas data interpretation involves explaining those patterns.

Decisions need to be made about which data is most important or convincing and how that data should be treated in the process of making it into evidence to support a claim. The aim is to make any underlying trends or relationships clearer. For example, if someone were trying to decide what to wear today, they might collect data on the time of the year, whether the sun is shining outside, whether the weather forecast mentions rain, and what clothes are clean or dirty. That allows them to interpret what the most appropriate clothing would be.

Data is interpreted in order to construct a logical argument that explains the data. Scientists will talk in terms of which interpretation is the most plausible or most reasonable, as supported by the data. Chemists will draw on their experience and knowledge of what other experiments have reported in the past. Scientists will also acknowledge the limits of the conclusions that can be reached from existing experiments. Consider the graph in Figure 5. What is the *y*-value when the *x*-value is 45 mg L^{-1}?

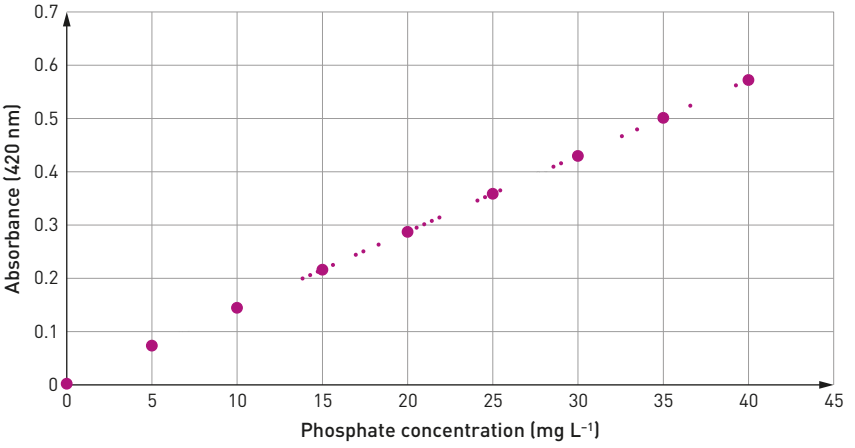

FIGURE 5 This graph shows the relationship between the amount of light absorbed by a solution at a wavelength of 420 nm and the concentration of phosphate ions in solution, after reaction with ammonium molybdate and ammonium vanadate in acidic solution.

A responsible scientist would say that it is impossible to know, because the given data does not allow you to say whether the linear trend continues. There is a limit on what conclusions can reasonably be made from the data available.

Scientists also use the principle of parsimony (sometimes called 'Occam's razor') in deciding between possible interpretations. This principle says that the simpler explanation is more likely (though not always) to be the correct one.

Measures of central tendency

The concepts of accuracy and precision are discussed on page 20. When there are a number of repeated measurements of the same value, it can be useful to have a way to discuss how close the values are to a central value – **central tendency**. For measurements that involve random errors, the values tend to cluster around a central value. Random errors are just as likely to result in underestimates and overestimates of the 'true' value, so by making repeated measurements, the values will tend to cluster around a central value.

central tendency the tendency for repeated measurements of the same value to be grouped around the mean, mode or median

Many natural variables tend to have a normal distribution. This is also sometimes referred to as a bell curve (Figure 6). This has the particular characteristics that the most likely value is the central value, which is the arithmetic mean or average of the results. The curve is symmetrical, with the probability of a particular value becoming lower as that value gets further away from the mean, in either direction. The curve can be more or less stretched out. This variation or spread in the results can be quantified by a value known as the **standard deviation**.

standard deviation a statistical value which expresses how spread out a group of values are or by how much they differ from the mean value for the group

For example, Table 3 shows the volume of aqueous hydrochloric acid needed to react with 20.0 mL of aqueous sodium hydroxide. Only six values are listed, but if the measurements were repeated several hundred times, then it is likely that a normal distribution of the values would be found around the mean.

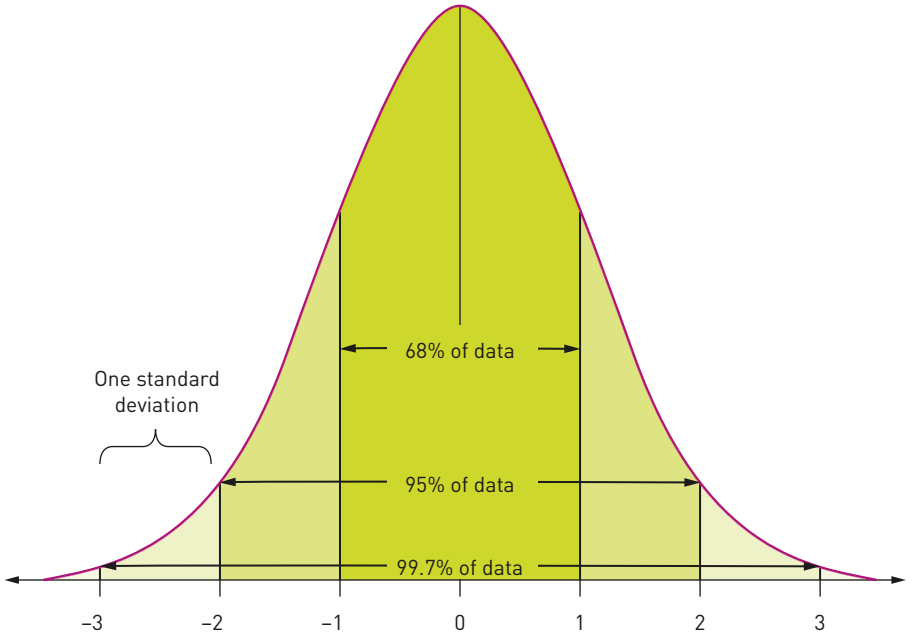

FIGURE 6 The normal distribution is also known as a bell curve.

TABLE 3 Volume of HCl(aq) required for reaction with 20.0 mL NaOH solution

Trial number	Volume of HCl(aq) (mL)
1	19.92
2	20.04
3	19.86
4	19.98
5	20.08
6	19.86

Mean

The arithmetic mean of a group of numbers is the sum of the numbers divided by the number of measurements. This can be expressed mathematically:

$$\bar{x} = \frac{\sum_{i=1}^{n} x_i}{n} = \frac{x_1 + x_2 + \cdots + x_n}{n}$$

where $\bar{x}$ is the mean, x_1, x_2, … x_n are the observed values, n is the number of measurements and Σ (capital sigma) means to add all the values (x) together.

This formula can be applied to the data in Table 3. The mean here should be reported to two decimal places, matching the precision in the original measurements.

$$\bar{x} = \frac{\sum_{i=1}^{n} x_i}{n} = \frac{(19.92 + 20.04 + 19.86 + 19.98 + 20.08 + 19.86)\ \text{mL}}{6}$$
$$= 19.95\dot{6}\ \text{mL}$$
$$\approx 19.96\ \text{mL}$$

One problem with the mean value is that it is sensitive to outlier values (those that are much larger or smaller than the other values). If there was an additional measurement of 40.00 mL, it would shift the average to 22.82 mL, which is quite far from the other values. (In practice, a value of 40.00 mL would probably be rejected as an outlier.)

Mode

The mode is another way of describing a group of values. The mode is the value that appears most often in a set of measurements. For the data in Table 3, this is 19.86 mL, because this value appears twice, but all the other values appear only once. It is possible for data to be bimodal if two values appear with the same frequency.

Median

The median value is the value that appears in the middle of a sorted list of numbers, if there is an odd number of items. Therefore, there are an equal number of values above and below the median value. If there is an even number of measurements, then the median is the average value of the middle pair of values. For the data in Table 3 (19.86, 19.86, 19.92, 19.98, 20.04, 20.08), this is 19.92 and 19.98 mL, and the median value is their average, 19.95 mL.

The median value can be very useful to report, particularly when there are extreme outliers in the data.

Standard deviation

The other parameter that describes a normal distribution is the standard deviation, which is a measure of how spread out the values are (Figure 7). Steeper bell curves have smaller standard deviations. Flatter bell curves have a larger standard deviation.

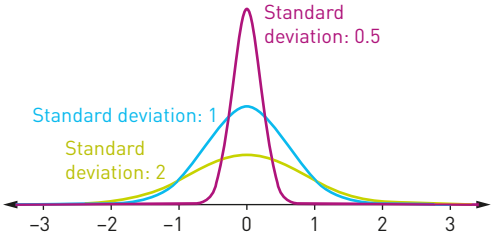

FIGURE 7 The effect of changing the standard deviation on the shape of normal curves.

The experimental standard deviation s is defined in mathematical terms by the formula:

$$s = \sqrt{\frac{\sum_{i=1}^{n} (x_i - \bar{x})^2}{n - 1}}$$

where $\bar{x}$ is the mean, x_1, x_2, … x_n are the observed values, n is the number of measurements and Σ (capital sigma) means to add all the values $(x_i - \bar{x})^2$ together.

Table 4 shows the results of these calculations on the experimental data given in Table 3.

TABLE 4 Steps in calculating the standard deviation of hydrochloric acid volumes

Trial number	Volume of aqueous hydrochloric acid solution (mL)	$x_i - \bar{x}$ mL	$(x_i - \bar{x})^2$ mL^2
1	19.92	19.92 – 19.957 = –0.037	0.001344
2	20.04	20.04 – 19.957 = 0.083	0.006944
3	19.86	19.86 – 19.957 = –0.0967	0.009344
4	19.98	19.98 – 19.957 = 0.0233	0.000544
5	20.08	20.08 – 19.957 = 0.12337	0.015211
6	19.86	19.86 – 19.957 = –0.0967	0.009344

$$s = \sqrt{\frac{(0.001344 + 0.006944 + 0.009344 + 0.000544 + 0.015211 + 0.009344)\ \text{mL}^2}{6 - 1}}$$
$$= \sqrt{\frac{0.042733\ \text{mL}^2}{5}}$$
$$= \sqrt{8.5466 \times 10^{-3}\ \text{mL}^2}$$
$$= 0.09245\ \text{mL}$$
$$\approx 0.09\ \text{mL}$$

This means that on average each data value deviates from the mean of 19.96 mL by 0.09 mL. The full working out has been shown here, but many calculators and spreadsheet programs also have statistical functions that will calculate these values more quickly and conveniently.

Correlation

Chemists often want to know how changing one variable in an experiment affects another variable. An example is shown in Figure 5, where the changes in absorbance are measured as the phosphate concentration in a solution is changed. Variables can be positively or negatively correlated. Positive **correlation** is when the value of one variable increases as the other variable increases. Negative correlation is when the value of one variable decreases, as the other variable increases (Figure 8).

correlation
a link between a change in the independent variable and a change in the dependent variable; this does not mean that the changing independent variable caused the change in the dependent variable

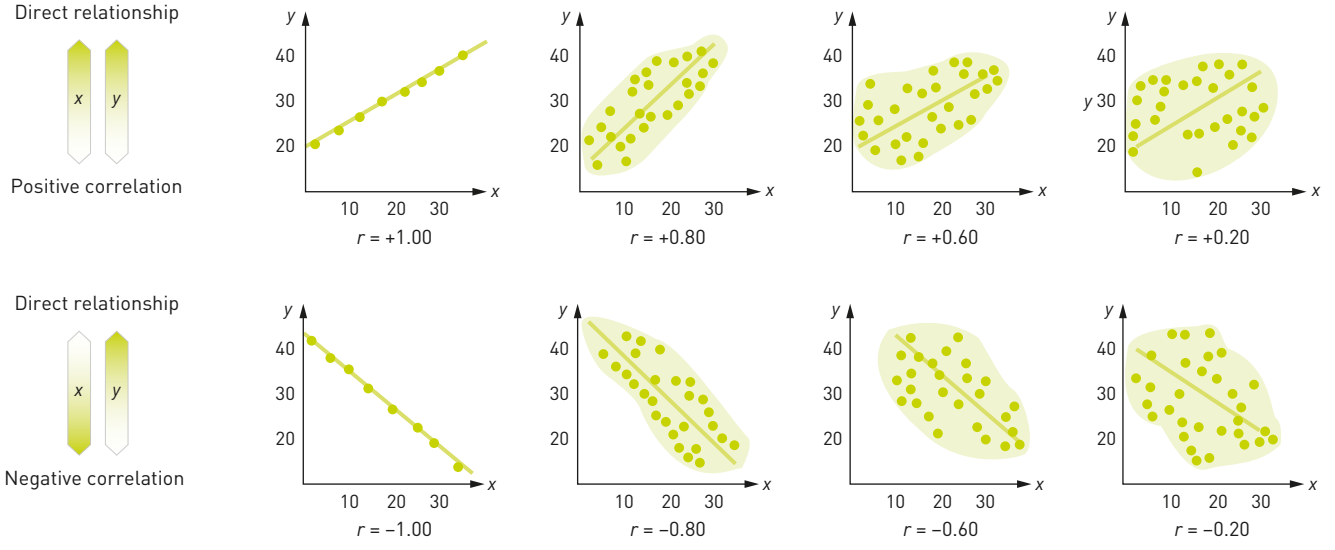

FIGURE 8 Examples of positive and negative correlation, with different values of the Pearson *r* correlation coefficient.

Study tip

If you have made an *x*–*y* scatter plot in Microsoft Excel, you can add a line of best fit to your graph by right-clicking on one of the data points. From the menu that appears choose 'Add Trendline', then choose 'linear' (but only if the relationship is actually a linear one!). Choose whether you want to force the line to go through the origin, by clicking 'Set Intercept'. Click 'Display Equation on Chart' and 'Display R-squared value on Chart'. Including these values in a report is a good way of showing how reliable measurements are.

It is common to talk about high or low correlation between variables. We can also calculate the size and direction of the relationship between two variables, in the form of a correlation coefficient. The correlation coefficient varies between –1 and 1. The sign of the correlation coefficient indicates whether the relationship is negative or positive, and higher absolute values indicate stronger correlation. A correlation coefficient with an absolute value of 1 shows a perfect fit between the variables, whereas a value close to zero indicates there is no relationship between the variables.

Chemists often investigate linear relationships between two variables and a formula can be applied to convert values into a linear form. Graphing the (x, y) pairs of values can visually reveal the relationship between the variables. For linear relationships, the Pearson r correlation coefficient can be calculated. This is related to the slope of a line of best fit, drawn through the points to minimise the distance of the points from the line.

The other useful feature of the Pearson correlation coefficient is that r^2 (the coefficient of determination) gives the proportion of the total variability of the y-values that is accounted for or explained by the x-values. For example, if the line of best fit has an r-value of 0.949, the r^2-value = $0.949^2 \approx 0.90$. In other words, 90% of the variation in the y-values is due to the x-values and 10% (100%–90%) is due to other factors. For many experiments in Chemistry it is possible to obtain r^2-values greater than 0.95, by careful experimental work and the use of highly reliable methods and equipment.

CHECK YOUR LEARNING 1.2

Describe and explain

1 **Identify** two examples of quantitative data.

2 **Identify** two examples of qualitative data.

3 There is a common saying in scientific circles that 'Correlation does not imply causation'. **Explain** what this means and give an example.

Apply, analyse and interpret

4 **Determine** the mean, mode and median of the following quantitative data set:
12.5 mg, 12.7 mg, 13.1 mg, 13.4 mg, 13.4 mg, 13.8 mg, 13.9 mg, 14.8 mg.

5 **Analyse** the following descriptions of data and choose the most appropriate type of plot or chart for displaying that data.

a concentrations of hydrogen peroxide measured at various times as a reaction occurs

b levels of caffeine found in 10 different energy drinks

c the percentage concentration of elements in the Earth's crust

d the energy of a single photon of light (E) compared to its frequency (ν), which can be expressed by the equation $E = h\nu$, where h is a constant

e a grouping of the elements by their decade of discovery

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Graphing data

» Increase your knowledge
Standard deviation

» Increase your knowledge
Measurements of central tendency

1.3 Student experiments

KEY IDEAS

In this section, you will learn about:

- the scientific method
- developing research questions
- drawing conclusions
- evaluating the validity of an experiment
- suggesting improvements for an experiment
- communicating findings scientifically.

In Chemistry Units 1–4, students are required to conduct their own experiment. This is an important part of the assessment, but it can also be a fun and interesting way to get more involved in chemistry. The goal of the student experiment is to design an experiment in order to answer a research question. The methodology and research question of the experiment may be based on an existing practical, but it must be extended, refined or redirected in some way. For example, it might be possible to improve the method of an experiment, or alter the method and materials slightly to test a different question.

Students need to go through the following steps.

1. Develop a research question.
2. Conduct background research for relevant scientific research that might help to modify the experiment.
3. Conduct a risk assessment for the modified experiment.
4. Conduct the experiment.
5. Collect qualitative and/or quantitative data.
6. Process and present the data.
7. Analyse the evidence.
8. Draw a conclusion.
9. Evaluate the validity of the experiment.
10. Suggest improvements to the experiment.
11. Communicate findings scientifically (e.g. in a scientific report, poster, presentation or article).

Study tip

If you are having trouble finding an experiment to extend, redirect or refine, use a practical you have already completed in class.

Some of these steps can be done in groups, such as developing a research question, conducting a risk assessment and conducting the experiment. However, it is important to draw conclusions, evaluate the process and communicate the findings individually.

To conduct the student experiment, it is crucial to have an understanding of the scientific method.

scientific method systematic exploration of a phenomenon or topic using observation, measurement and experiment to support, disprove or modify hypotheses

The scientific method

Scientists use the **scientific method** as an approach to problem-solving. It involves a series of steps that can be used to test hypotheses. The steps involved in the scientific method might not always be completed in order. For example, it might be necessary to repeat a test or to refine the experimental design to solve a problem or to correct a mistake.

variable
a condition or parameter that is changed or changes during an experiment

independent variable (IV)
the variable that is altered systematically in a controlled way by the experimenter to test the effect on a related dependent variable

dependent variable (DV)
the variable that is measured or observed during an experiment, whose change is in response to changes made in the independent variable

extraneous variable
a variable not being intentionally studied during an experiment that could have an unplanned or unintended effect on the results

controlled variable
a condition that the experimenter tries to keep constant during an experiment, in order to minimise its effect on measurements or experiment outcomes

confounding variable
an extraneous variable that changes systematically along with the variables being studied and offers an alternative explanation for the observations or measurements made

Study tip

Remember: the DV *depends* on the IV.

The scientific method involves a series of steps to test a hypothesis (Figure 1).

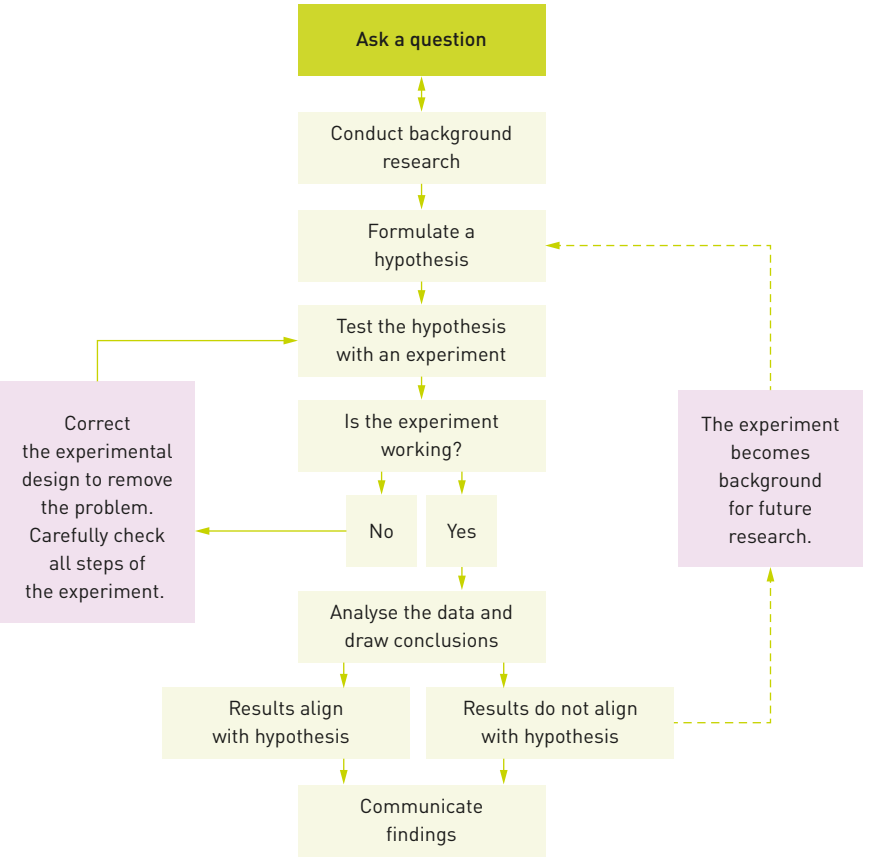

FIGURE 1 The scientific method

Variables

The term **variable** refers to quantities or qualities that can change in different conditions. They affect the results of the experiment. Experiments usually have an **independent variable (IV)** and a **dependent variable (DV)**.

- An IV is a variable, or factor, that is deliberately changed in the experiment.
- A DV is reliant on the experiment, or may change as a result of the experiment.

Other types of variables may need to be controlled so that they do not affect the outcome of an experiment. These include:

- **extraneous variables** – variables (other than the IV) that could alter the value of the DV. If the effects of the extraneous variable are removed from the experiment, or controlled, it becomes a **controlled variable**.
- **confounding variables** – variables that influence the IV and the DV in a way that affects the results of the experiment.

Making predictions

A crucial part of the student experiment is making predications. Before carrying out the practical work in the student experiment, it is important to make a prediction, or guess the answer to the research question, based on observations. Therefore, students must develop a hypothesis, this is the initial expectation about the outcome of the experiment. It can then be tested by conducting the experiment.

Hypotheses

Throughout Years 7–10, students will have developed **hypotheses** for experiments.

hypothesis
a proposed explanation used as a starting point for further investigation

A hypothesis should be written as a clear statement predicting how changes in the IV will affect the value of the DV.

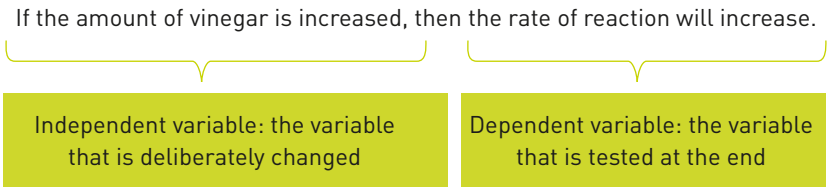

FIGURE 2 Predictions are often written as IF/THEN statements.

TABLE 1 The difference between a prediction and a hypothesis

Prediction	Hypothesis
The expected outcome of the experiment based on observation.	A proposed explanation for observations that can be tested, used as a starting point for investigation.
'If temperature is increased, then the rate of enzyme action will decrease.	'Increased temperature reduces the rate of enzyme action.'

Developing a research question

The research question is the basis of the student experiment and gives the research a focus, or clearer aim. Without a research question it would be much harder to know what to test, and what data to collect.

Typically, a research question should contain a variable that can be changed in the experiment (such as temperature, time and volume.) and a variable that can be measured.

For example, 'How does temperate impact the rate of the chemical reaction between Substance A and Substance B?'

In this example, the independent variable is the temperature and the dependent variable is the rate of chemical reaction.

When developing a research question, keep the following in mind.

- Is the question testable? The question must be able to be answered by using the scientific method to test the hypothesis.
- What is reasonably achievable? Time limits and resources can restrict what can be achieved. It is easy to underestimate how long experiments might take and to try looking at too many variables. Knowing what equipment is available in the school laboratory is important, because there is no point in planning experiments that rely on a piece of equipment that is not available. That doesn't mean to give up! There is usually more than one way to conduct an experiment.

When planning an experiment, think about what type of measurements will need to be made. What equipment is available to make those measurements (Table 2)?

Is the question sufficiently focused? If the research question can be answered with a simple 'yes' or 'no', then it is probably too narrow. If the question is too broad, there will not be enough time to answer it.

TABLE 2 Methods and equipment used to make measurements

Quantity to be measured	Some suggested methods or equipment
Mass	Balance – pan balance or analytical balance? How many decimal places are required?
Volume	Solid – displacement of water Liquid – measuring cylinder, burette, pipette, volumetric flask Gas – displacement of water from an upturned burette or measuring cylinder or eudiometric apparatus
Temperature	Thermometer or digital temperature probe
Enthalpy change	Coffee cup calorimeter, pan containing ice
Time	Stopwatch
Pressure	Monometer, digital gas pressure sensor
Solution concentration	Titration with a suitable indicator, e.g. acid–base titration. UV–visible spectrometer or colorimeter if light-absorbing Conductivity meter if solution of an ionic compound or electrolyte Gravimetry – selective formation of a precipitate and measurement of mass of filtered solid
pH	Indicator paper or pH test strips, or pH probe

Conducting background research

Before commencing the experiment, background research is needed. Is there a published method for carrying out the experiment, or can an existing method be altered? Are all the required chemicals and equipment available? Can the experiment be done safely? How will any chemical waste generated be disposed of?

One important reason to research at this stage is to consider if there are any better ways to conduct the experiment. For example, the amount of hypochlorite ion (OCl^-) in household bleach can be determined in several ways. One method is to react a hypochlorite solution with known amounts of hydrogen peroxide and potassium permanganate by a method called a titration. This requires analytical glassware, such as volumetric flasks, pipettes and burettes, and a balance that can weigh accurately to four decimal places. Another method is to use an instrument called a spectrophotometer to determine the concentration of hypochlorite ion by measuring how much ultraviolet light is absorbed by a solution. The second method is potentially faster and uses fewer chemicals, but it requires the right instrument to be available.

When modifying or developing an experiment, it can be helpful to have a more detailed understanding of the theories or scientific research. For example, if an experiment is testing the amount of sodium azide required to inflate the airbag in a car, then it helps to know the balanced reaction involved, as well as to know something about how the pressure, volume and temperature of a gas is related to the amount of gas present. That way, calculations can be included as part of the planning stage of the experiment, to predict what the maximum amount of gas produced would be.

More information on how to perform background research is covered in Section 1.4.

Conducting a risk assessment

It is important to assess the risk involved in the planned experiment to ensure that everyone is safe and that no harm will come to anyone involved.

Students must conduct a **risk assessment** before beginning the experiment. A risk assessment evaluates the materials, equipment and methods to be used in the experiment. The aim is to identify any possible hazards and control or remove them from the experiment design if possible, or to reduce the risk to acceptable levels. Chemistry involves learning how to safely handle potentially dangerous materials, which means being aware and planning for what can go wrong.

risk assessment an evaluation in a systematic way of the potential risks and their likelihood involved in doing an experiment or activity

All chemicals in Australia are supplied with a Material Safety Data sheet (MSDS). This contains a summary of information about the chemical, and guidance on how it should be used and stored safely, and any known hazards to people or the environment.

When planning, think about what will happen to any chemicals after the experiment is finished. Some chemicals can safely be disposed of down the sink, dissolved in plenty of water. Other chemicals must be collected in waste containers that are specially collected and disposed of.

A template to help conduct the risk assessment is available on the obook assess. Five key steps are involved in making a risk assessment.

1. Identify the potential hazards.
2. Determine who could be harmed and in what way.
3. Assess the risks and the action required. Decide how likely is it that a particular hazard could result in harm.
4. Record the risk assessment. This may be in your lab book or in a special folder or on a website.
5. Keep the risk assessment up to date if any changes are required. If the experiment is changed in some way or new information becomes available about a hazard, then the risk assessment needs to be reconsidered.

Collecting scientific evidence

Collecting evidence during the experimental process is a crucial part of the student experiment. Presenting and analysing evidence, making conclusions and communicating findings all depend on keeping a detailed record.

Keeping a logbook

It is a good idea to keep a logbook to record key information throughout the experimental process, such as observations and measurements.

A bound book makes a good logbook, so that papers do not go missing. Another option is to keep an electronic version of a logbook; however, it is important to make sure this is backed up and can be authenticated.

Poor record-keeping can mean that experiments have to be repeated, which is wasteful of times and materials. It could also result in distorted or incorrect findings and ultimately it may not be possible to answer the research question accurately. Professional scientists keep a record of their experiments in a laboratory notebook. For a chemist, a record of an experiment should contain enough detail that another chemist can understand what the experiment is about and repeat it if necessary. Entries in the laboratory notebook are entered as the experiment is being carried out, because it is too easy to forget things that may be important later.

FIGURE 3 Keeping a logbook to record data and experiment notes is crucial when it comes to writing the scientific report.

Things that should be recorded in a logbook include:

- the date
- equipment or instruments used
- the procedure that was followed
- amounts of any chemicals used, either in grams or millilitres. Chemists also use the mole for amount, which is explained later in this textbook
- observations or results
- any other key notes it may be important to remember while writing the report.

Numbers are always followed by a space and then their associated units (with the exception of percentages).

For example, what do the following numbers refer to?

'Iron oxide (12.0) was added to the aluminium powder (4.0) and the mixture was ignited behind a safety screen.'

Does 12.0 refer to 12 g, 12 kg or 12 tonnes or something else? Without the units it is impossible to be sure.

Quality of data

accuracy
a measure of how close the measured value is to the true or accepted value

precision
the consistency and reproducibility of a series of measurements

It is important to ensure that the data collected in an experiment is sufficiently reliable and accurate to answer the research question that is being asked.

There are two concepts that are important in relation to experimental measurements. The first is **accuracy** – how close the values are to the 'true' or accepted value. The second is **precision** – how precise the measurements are or how reproducible the results are or how close they are to each other. These two concepts are explained in Figure 4 in terms of a dart board. The very worst situation to be in is to have results that are inaccurate, but are precise, as this can be misleading as to the quality of the results.

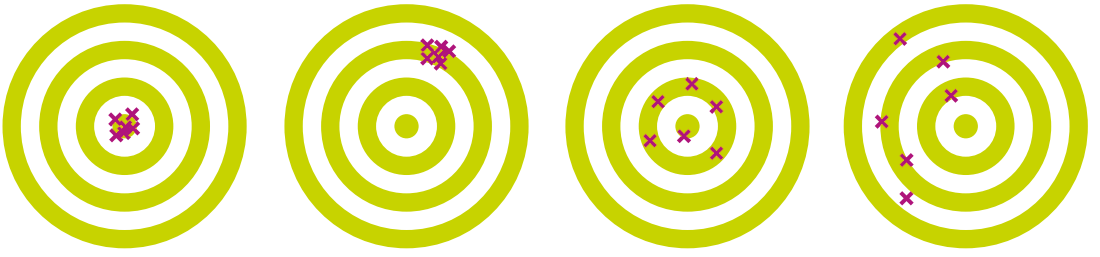

FIGURE 4 Accuracy is not the same as precision.

Ensuring that measurements are both accurate and precise enough involves thinking about the experiment before, during and after the experiment.

Quality assurance is about the processes involved in planning the experiment, before any collection of data. Questions that may need asking at this stage include the following.

- What level of precision can be expected from the equipment that is planned to be used? Is this sufficient? For example, if a solution with a volume of 100 mL is needed, should this be prepared in a 100 mL beaker (where the '100 mL' volume mark is ±10 mL and the volume could be between 90 and 110 mL) or in a 100 mL volumetric flask (where the '100 mL' volume mark is ±0.1 mL and the volume could be between 99.9 and 100.10 mL)?
- Is there a need to do **replicate experiments**, which means repeating measurements? If you only make one measurement, there is no indication as to the precision of your measurements. For titrations, it is usual to do at least one rough or approximate titration and at least three good titrations. Sometimes, limitations of time or reagents may mean it is not possible to do repeated experiments, but this has risks.
- How accurate are your instruments? When was the instrument last calibrated? A pH electrode and meter needs to be calibrated before starting an experiment, using buffer solutions with a set pH. What does the instrument manual say about the accuracy or precision of the instrument?
- If there are other variables besides the one being investigated that could alter the result, how are these other variables to be kept constant? For example, rates of reactions often increase with increases in temperature. If the aim of an experiment is to investigate how the rate changes with changes in the concentration of reagents, then the reaction must be kept at the same temperature throughout the experiment. This could be done by keeping the reaction in a water bath at a set temperature.

replicate experiment measurement that is repeated multiple times to give an indication of the precision of measurements; calibration to check that an instrument is giving accurate results and rectify if needed

negative control a test to make sure that an experiment is working as expected, which should give no result

positive control a test to make sure that an experiment is working as expected, which should give a positive result

blank solution a solution that does not contain the compound being measured, but contains all other compounds that would be present in the sample being tested; this allows the response due to those other compounds to be subtracted from other measurements

The quality of measurements should be monitored during the experiment. Plot your data as you go. If a numerical trend is expected, then is the data following that expected trend? Other ways of ensuring quality measurements are to use standards of known concentration, which can then be compared to solutions of unknown concentration.

It is also useful to have controls, both negative and positive. A **negative control** is one that should give no response. For example, if water samples were being tested for the presence of chloride ions (Cl^-), a negative control would be to do the test on distilled water, which does not contain Cl^-. A **positive control** is one that would be expected to give a response if the test or experiment was working properly. In the case of testing for Cl^-, tap water could be used as a positive control. It is also usual to use a **blank solution** (one that does not contain the substance being measured) to see what the measurement would be and to subtract it from the other measured samples. For example, if copper-containing solutions were made, it would be usual to make one solution with no added copper ions, to see what trace amounts of copper were present.

After the measurements have been made, it is essential to ask whether the results make sense. For example, if the level of sugar in cornflakes was being measured, a measurement of 200% sugar by mass would not make sense, because this would mean that the level of sugar in cornflakes was higher than the cornflake sample itself. Puzzling results like this can often be due to mistakes in calculations, such as a failure to convert millilitres to litres or forgetting to take dilutions into account.

Drawing conclusions

Once an experiment has been conducted and the data collected and analysed, the final critical step is to draw conclusions about the meaning of the results. This step is often neglected by students, perhaps because they have been thinking about the results so carefully they may regard the conclusions as obvious. Whether an experiment is a success or failure does not depend on whether a hypothesis was accepted or rejected, as long as the results are valid. An experiment is a failure if there are flaws in the design of the experiment or the reasoning that mean definitive conclusions are not possible.

It is useful to break down this process into three components.

1 What is my conclusion?
2 What evidence am I relying on for this conclusion?
3 What is my argument or reasoning process to reach this conclusion?

Evaluating the validity of the experiment

The basic question to be asked here is 'What confidence can be placed in the result?'. An experiment always involves some compromise, because there are limits on the time or materials available or the instruments or equipment available. There may have been flaws in the research method because the effect being investigated was too small to be reliably measured or detected with the existing experimental design. There could also be other variables influencing the experiment that have not been controlled. Was the experiment suitable to answer the questions proposed? Did the experiment test what it was meant to test? This means evaluating the equipment, the experimental method, the adequacy of the results and the analysis of the results, including any calculations or statistical tests used.

For example, if an experiment was conducted to measure the rate of a reaction, as indicated by a colour change, the validity of the experiment could be effected by the choice of either a clock, where time can be read to the nearest second, or a stopwatch, where time can be measured to the nearest hundredths of a second. If an acid–base titration was conducted using an indicator that changed colour two pH units away from the pH at the equivalence point, how would this change your confidence in the results?

Suggest improvements to the experiment

Sometimes an experiment does not go entirely to plan, or during the experiment a better way of testing might reveal itself. It is important to note these ideas for improvement so that they may be discussed when it is time to communicate findings and ideas. The conclusions that have been drawn may also form the basis for a new hypothesis or suggest a new research direction.

Communicating findings scientifically

Scientific writing has its own particular requirements and expectations, just as any other type of writing does. It is important to pay attention to those writing conventions.

The findings of experiments can be presented in a number of formats as a scientific report, scientific poster, conference presentation or a journal article.

When it comes time to communicate your findings, you need to plan what is to be written. Time spent in planning will save time overall and result in a better report, which is also easier to write. Think about the different sections in a report in order and plan each section (Table 3).

TABLE 3 Components of the scientific report

Section	Purpose
Title	Outlines the problem or research question under investigation
Introduction	Indicates what the report is about. It should include: • a rationale for the research question • findings from background research • identification and justification of modifications to the initial methodology • a hypothesis for the experiment that includes the IV and DV.

Section	Purpose
Context	A background paragraph that helps to give the reader the necessary knowledge required to understand the rest of the essay
Equipment list	Lists essential pieces of equipment or materials needed to carry out the experiment and information about concentration and amount of substances
Risk assessment	Outlines the possible hazards involved in the experiment, as well as the measures taken to control and/or remove them
Method	Lists clearly and concisely all steps involved in conducting the experiment. There should be enough detail to allow another person to conduct the same experiment and collect sufficient data. It may be necessary to include a diagram or photo of the experiment set-up
Results	Displays the collected qualitative and/or quantitative data from the experiment in a format that is able to be interpreted easily (e.g. table, graph, photo, descriptions, or any equations necessary)
Discussion	Includes: • an analysis of the data in relation to the research question and in identification of trends, patterns and/or relationships • an evaluation of the validity of the experiment through the identification of uncertainty and limitations of the evidence collected • suggestions for improvements to the experimental design. Do *not* restate the results in the discussion unless it is to analyse or interpret them in a more meaningful way.
Conclusion	States the findings in relation to the research question. Justify based on interpretation of the evidence. Discuss the reliability and validity of the experimental process undertaken. Suggest improvements and extensions to the experiment based on the analysis of the evidence.
Bibliography	Lists all the references used in the research. It may be presented in alphabetical order, using a consistent style for each entry.

It is important to proofread what you have written. Look for spelling and grammatical errors. Be alert for unnecessary repetition and sentences that are too long. Strategies include to start proofreading from the bottom of the page and read up one sentence at a time or to read the essay aloud. The first draft should never be the last draft.

CHECK YOUR LEARNING 1.3

Describe and explain

1 **Explain** the difference between a positive and a negative control.

2 **Describe** the main sections of a scientific essay and the purpose of each section.

Apply, analyse and interpret

3 **Identify** problems with the following research questions.

a How does vinegar affect a chemical reaction?

b What is the best way to collide two particles?

Investigate, evaluate and communicate

4 **Investigate** the meaning of the following abbreviations commonly used in chemistry.

a g
b aq
c UV
d IR
e K (more than one meaning is possible)
f EtOH
g kPa

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» Student book questions
Check your learning 1.3

» Video
Core experimental skills

» Video
Chemical safety

» Video
The student experiment

1.4 Research investigations

KEY IDEAS

In this section, you will learn about:

- performing background research
- evaluating claims
- writing scientifically.

In Chemistry Units 1–4, students are required to conduct their own research investigation. This is an important part of the assessment, but it can also be a great way to learn more about chemistry in action. In the research experiment, the goal is to modify or redesign an experiment in order to answer a hypothesis or question.

There are a number of steps involved in the research investigation.

1. Select a claim to be evaluated.
2. Identify the relevant scientific concepts associated with the claim.
3. Pose a research question addressing an aspect of the claim.
4. Conduct research to gather scientific evidence that may be used to address the research question and subsequently evaluate the claim.
5. Analyse the data to identify sufficient and relevant evidence.
6. Identify the trends, patterns or relationships in the evidence.
7. Analyse the evidence to identify limitations.
8. Interpret the evidence to construct justified scientific arguments.
9. Interpret the evidence to form a justified conclusion to the research question.
10. Discuss the quality of the evidence.
11. Evaluate the claim by extrapolating the findings of the research question to the claim.
12. Suggest improvements and extensions to the investigation.
13. Communicate findings in an appropriate scientific genre (e.g. report, journal article, essay, conference presentation).

As with the student experiment, it is important to have an understanding of the scientific method. Developing research questions, analysing evidence, drawing conclusions and evaluating the validity of an experiment are all crucial skills for the research investigation. These skills were explained in Section 1.3.

Beginning the task

A number of claims should be provided by your teacher. It is from these claims that a possible research question can be developed.

But where is the best place to start a research investigation? Often a quick overview of an area is required first, before more detailed questions and a research plan are drawn up.

Analyse the description of the task

What are the key words in the claim you are evaluating? These can be the starting point to search for information.

Also relevant is how much time is available to answer the question. A one-paragraph answer required within 10 minutes has very different requirements from a 10-page report required within 6 weeks.

For example, suppose that you have been hired by the Brussel Sprouts Marketing Board to investigate claims made about the ideal level of glucosinolates in brussel sprouts. The research question might then become: 'What is the ideal level of glucosinolates in brussel sprouts?'

From the sound of the word 'glucosinolate', you might reasonably suspect that it is the name of a type of chemical. 'Glucose' is the name of a familiar chemical, so it is likely that glucosinolates are related in some way to glucose.

The words 'ideal level' suggest that the idea of amount or concentration will also come up in answering this question. The question also implies that the amount of this chemical can vary.

FIGURE 1 Planning a research investigation first means that time taken in the laboratory doing experiments becomes more productive.

Now is the time to start brainstorming and to think about what background questions need to be answered. For example:

- What is a glucosinolate?
- What is the structure of these compounds (i.e. how are the atoms arranged and connected together)?
- Are these naturally occurring chemicals or synthetic (i.e. man-made)?
- How can these chemicals be detected and their amount measured?
- Is it possible to separate and purify them?
- Are they found in other sources besides brussels sprouts?
- What effect do glucosinolates have? Are they toxic? If they are toxic, at what level? How stable are they? Are they water-soluble? What compounds are they made from in brussel sprouts?
- What is known about what effects the levels of these chemicals?
- Has anyone made these compounds in a laboratory?
- Can I buy these chemicals?

FIGURE 2 Even brussel sprouts can be used for chemistry research!

These are initial questions only, to act as a guide for further research. Some of them may not be relevant if you are only being asked to put together a written report. Some will only be relevant if you need to carry out laboratory research to discover new information to answer the questions.

Conducting research

What sources of information can be used to answer this question? There is a great deal of chemical information on the internet, in libraries, in textbooks, encyclopaedias and specialised chemistry books and databases. The problem becomes knowing where to look to find the most reliable information.

Search engines

Typing 'glucosinolate' into a search engine will quickly bring up some information from Wikipedia about glucosinolates, including a chemical structure of the class of compounds. This is only a starting point for further research. Wikipedia and other general sites can be very useful for getting a quick overview of a topic, but must be used with caution.

FIGURE 3 Check that information obtained from a quick internet search is from a credible source.

Anyone can edit a Wikipedia page, whether they are an expert in the area or not. So any information from sites such as this should be independently verified from a reputable source.

Books

Some scientific information is published in books and encyclopaedias. These can be good for getting an overview or summary about what is known, or looking up a particular fact from a source known to be reputable. One potential drawback is that it is hard for books to be as up to date as online sources. Check that the information is still current and correct. How old is the source that you are using? Are there likely to have been recent developments?

Databases and collections

Chemical information is often collected into databases or collections of information, usually in the form of brief tables of data. There are several compilations of chemical data on different compounds and materials, which contain such information as melting and boiling points, densities and solubilities. The *Merck Index* and the C*RC Handbook of Chemistry and Physics* are examples of these types of collections of data. Increasingly, this information can also be found on the internet, in online versions of books or in scientific websites. For example, ChemSpider is a free resource put together by the Royal Society of Chemistry (UK) that covers around 64 million compounds. Information about household chemicals and products can be found at a website put together by the US government, called the Household Products Database. Manufacturers who sell chemical products in Australia are required to supply a Material Safety Data Sheet (MSDS) with their product, which gives information on what chemicals it contains and their properties, particularly the hazards associated with their use. Many suppliers of chemicals have online catalogues that contain basic information on the chemicals that they sell.

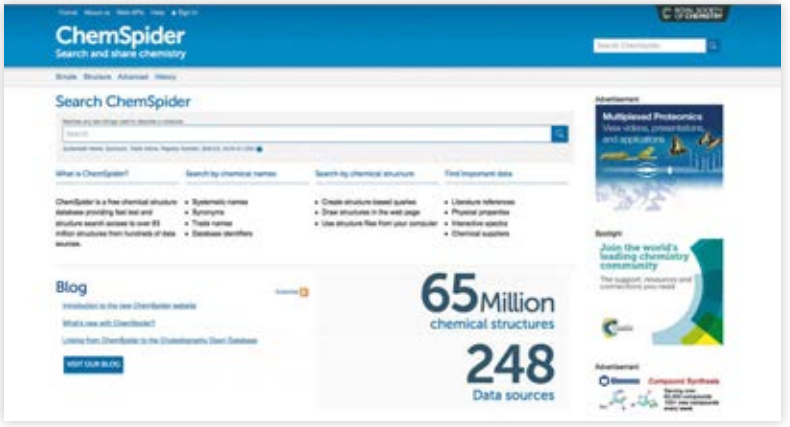

FIGURE 4 There is a link to the ChemSpider website on your obook assess.

Journals

Most scientific knowledge is published in scientific journals. These are collections of articles written by researchers on particular topics, collected and published at regular intervals. You will find the newest research reported in journals and the best place to start is Google Scholar. Often journals will specialise in a particular area, such as the *Australian Journal of Chemistry*. Other scientific journals publish articles from all areas of science, including the very prestigious journals *Nature* and *Science*. One disadvantage of scientific articles is that they are written for an expert technical audience who already usually knows a great deal about the area being discussed. They may be written in very dense, technical language with many abbreviations. It is a common mistake by students to use material from articles that they do not fully understand.

Open-source or open-access journals are accessible to the general public, but most journals require readers to pay a subscription or be a member of a library that subscribes. It can be frustrating to find a reference to an article that looks very relevant to your research question, only to find that payment is required to view the article.

The following points may help.

- Research articles often include a summary of the article at the start, called an **abstract**, which may contain all the information that you require. The abstract is usually free to view.
- Your school library may have an arrangement with a university library, which may give school students access to more resources. Ask your chemistry teacher or school librarian about this and if you live close to a university, you may be able to visit the library on campus.
- When searching in an online research articles database like Scopus or Science Direct, it is possible to limit search results to those found only in open-access journals.
- The State Library of Queensland also has online resources for high school students, as well as physical books and copies of journals.

abstract
a summary of the most important findings of a research article

When looking at a reference source, whether online or in print, be sure to ask whether it is a trustworthy source of information. Who has written or gathered the information? How current is the information? Is it consistent with other sources of information? Are there references to other sources that the author has relied on? For what purpose was the article written?

Evaluating claims

When evaluating claims in a piece of scientific writing the following may be useful.

1. Identify the claim.
2. List the evidence being used to support the claim. This evidence must be a fact or measurement. Where the evidence being relied on is work of other scientists, is the work attributed to an identified scientific source, such as a particular book, article in a scientific journal or a particular scientist or group? An opinion does not count as experimental evidence.
3. Evaluate the evidence. Is the evidence empirical; that is, is it based in observation and measurement? What is the source of the evidence? Are there any reasons that the evidence may be biased; for example, who funded the research? Is the evidence given internally consistent or does it contain contradictions?
4. Evaluate the claim. Should the claim be accepted or rejected or is there a need for more evidence, better evidence or a better explanation of how the conclusions were reached?

In particular, be careful not to confuse correlation with causation. Just because two variables are shown to be related does not necessarily mean that one causes the other. Suppose that the relationship between amount of pocket money received per week and whether students at a secondary school study chemistry was investigated and the results were graphed (Figure 5).

From the graph, it could be concluded that giving secondary school students more pocket money causes them to be more interested in chemistry. (Perhaps they are spending the extra money on chemistry books!) A more likely explanation is that because chemistry is studied in the later years of secondary school, the students who study it are older and older children often receive more pocket money.

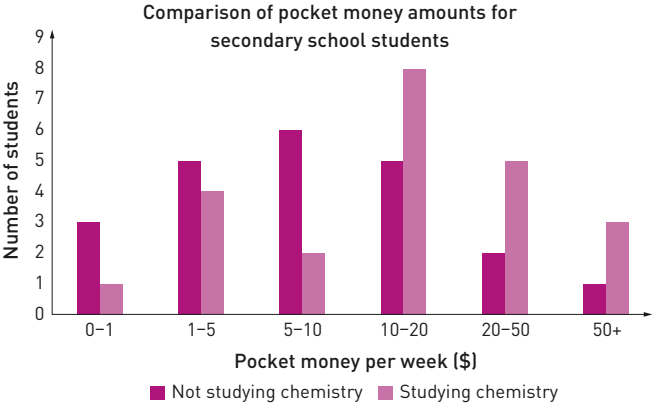

FIGURE 5 Comparison of the amount of pocket money paid to secondary school students, who do and do not study chemistry.

Writing the report

The findings of the investigation should be written up scientifically at the end of the process. A scientific report or essay should have the following components.

- An introduction paragraph tells the reader what the report will be about.
- A background paragraph gives the context to the essay and helps give the reader the necessary knowledge required to understand the rest of the essay.
- The body of the essay covers the main arguments in the essay, with each main idea appearing in a separate paragraph. One way to structure this is in terms of the claim being evaluated, the statement of evidence that supports or disproves the claim and then the argument as to how the evidence supports the statement.
- A conclusion summarises what has been said and stresses the most important points to the reader.
- A bibliography or reference list allows the reader to follow up on the sources that have been relied on in the essay.

Tips for writing scientifically

The following are tips to help you with writing scientifically.

- Use shorter sentences and words where possible. This will make your meaning clearer and argument easier to follow. When there are many subordinate clauses, each winding around a meaning lost somewhere in a muddle of obscurantism and verbosity, included to make the author seem more clever and important, this often instead results in the reader becoming lost, in a maze of impenetrable text, when the sentence really needs to be split up into shorter, more digestible sentences in order to facilitate comprehension of the written text. How hard was the preceding sentence to read?

  Instead, keep it short. Keep it simple.

  Poor: 'The aim of the present study by the author was to provide insight into …'

  Better: 'The aim of the experiment was to …'
- Use the first person 'I' or 'We' sparingly. This may be the opposite of the advice of many books on writing, but scientific writing, especially in chemistry, avoids the first person.
- If experiments are being reported, then use the past tense, as it is a report of work that has already been carried out. Often instructions to experiments are given in the present tense, but a lab report reports what was done. For example, 'The liquid was heated to 78°C, with the temperature measured using a thermometer'.
- When discussing analysis of data, then use the present tense. For example, 'The mean of the results in Table 2 is 0.85 mg L^{-1}, which overestimates the actual concentration of Zn^{2+} by 23%'.
- It is never possible to prove a fact definitively. For this reason, chemists will talk about evidence being consistent with a hypothesis, rather than proving the hypothesis. It is possible that results may show that a particular hypothesis should be rejected. If so, it is fine to say that.
- In referring to work done by others, there are a few verbs that are commonly used. These include 'claimed', 'suggested' and 'demonstrated' (in order of increasing confidence by the writer in what has been written).
- When using a key term or an unusual abbreviation for the first time, provide a definition. Some common abbreviations do not require an explanation, such as e.g., i.e., et al., ca. and etc. Use them correctly.

- Don't be vague. For example, don't use 'particle' when you mean 'atom'. If there is a specific chemical or scientific term, then use it.
- Always report numbers with their associated units. Consider using either scientific notation for very small or very large numbers or a unit with an appropriate prefix. For example, 0.000045 grams is better written as 4.5×10^{-5} grams or 45 micrograms.
- Pictures, graphs, equations, chemical structures and reactions can all play an important part in communicating meaning in a scientific essay. These should be referred to in the main part of the text by their particular number. For example, 'Absorbance values levelled off after 5.8 ppm, as shown in Figure 1'. 'The balanced reaction of calcium carbonate and sulfuric acid is shown in Reaction (8).'
- Use references to indicate in an essay where reliance is being placed on the work of others. Teachers will advise on the precise method of doing this. One method is to put the surname of the authors and the year of the publication in brackets (Smith and Jones, 2019). Another method is to include the reference as a superscript at the end of the sentence or phrase to which it relates. The full citation, including details of the authors and the details needed to find the reference, are included in a references section or bibliography at the end of the essay.

CHECK YOUR LEARNING 1.4

Describe and explain

1 **Identify** three sources for information that could be used during the research investigation.

2 **Explain** why some sources may not provide reliable information.

Investigate, evaluate and communicate

3 **Evaluate** the following introduction to a scientific report and rewrite it to make it clearer and easier to follow.

These AC molecules are very, very interesting because I eat them all the time. They can be a variety of colours from red through to blue and purple. The variety of colours from red through to blue and purple can also change depending on the pH of the extraction solution and we are going to use vinegar and baking soda solutions. Vinegar is acidic and baking soda is basic. Water is not. And they are not just in vegetables either any more. Anthocyanins are also used as naturally occurring food colours. In the experiment we are going to extract anthocyanins (ACs) from red cabbage, jacaranda flowers and blueberries. We can see what colours the different juice solutions will go. It was great, because I learnt to identify what type of anthocyanin was present and this was the aim of the experiment and we did that. Beetroot was used as a positive control, because it doesn't change colour, because it contains a different type of coloured compound.

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1.5 Exam preparation

KEY IDEAS

In this section you will learn about:

- planning time and making a timetable for study
- exam revision
- examination strategies.

Even though the external exam happens at the end of Units 3 & 4, it is important to develop good study and revision techniques now. Cognitive psychology is an area of science that investigates learning and learning strategies. Research in this area suggests that some revision strategies work better than others. For example, spacing out revision and testing understanding help learning and recall better than highlighting or rewriting notes. The goal is to adopt strategies that are proven to aid in both understanding and recall.

Be strategic about your learning. In a class or while reading a textbook, think about what details are essential to know and which are added for purposes of explanation or interest. Some things are worthwhile memorising because they make the rest of learning chemistry more efficient and easier. Psychologists have found that the amount of information that a person can keep in the front of their awareness is limited. This is called the 'working memory' hypothesis. When learning new information, the working memory can easily become overloaded and learning can become less efficient or even cease. For example, learning the symbol for sodium (Na) and its location in the periodic table is useful because this information will be used throughout high school chemistry and beyond. This means one less thing that has to be brought into active recall, allowing attention to be devoted to other parts of a question. Memorising the atomic mass of sodium to six decimal places (22.989 769 g mol^{-1}) is not useful because this information will always be provided if needed. The process of internal thinking about learning is called metacognition and is a key characteristic of the best students, who self-assess whether they really understand something.

FIGURE 1 Having a comfortable quiet place to study will make your revision time more effective.

Time management

During the busy senior years of high school, there are many competing demands on your time, not just from the other subjects being studied, but also from family, personal, social and employment. Time management has the goal of using the time available most efficiently for the best outcomes. It is recommended that average students spend three hours of personal study time each week for each subject being studied.

TABLE 1 A sample study timetable

Time	Monday	Tuesday	Wednesday	Thursday	Friday	Saturday	Sunday
6.00-7.00 am	Check all set homework completed	Instrument practice		Instrument practice		Run	
Breakfast Travel Classes Travel	Study class 50 min – Physics		Study class 50 min – Biology	Study class 50 min – Maths	Study class – 50 min Chemistry	8.00–8.20 am Instrument practice	
						9.00-10.30 am Lawn mowing	
						11.00 am – 1.00 pm Assignments, lab reports and essays	
						2.00–2.30 pm German	
						3.00–3.30 pm Chemistry	
4.00–4.30 pm	German	Chemistry	Maths	Concert rehearsal	Work		Cook dinner
4.30–5.00 pm	Chemistry	Biology	English				
5:00–5.30 pm	Maths	Physics	Chemistry	Physics			
5.30–6.30 pm	Dinner	Dinner	Dinner	Dinner			
7.00–8.00 pm	English	German	Physics	English			
8.15–9.15 pm	Biology	Maths	German	Biology			

A study timetable is useful for making sure that sufficient time is spent on your subjects across the week and that time is also available for other activities. There are some key factors to guide the planning of a study timetable.

- Shorter, more frequent study sessions result in better learning and recall. Cramming for tests may be initially effective, but results in very poor long-term recall. External examinations in Chemistry at the end of Year 12 will assume knowledge from throughout Senior Chemistry.
- Memory fades over time, so it is important to devote some time to refreshing earlier material. The more particular information and knowledge is brought into active recall, the easier it becomes to bring it into active recall at a later date.
- Time devoted to self-testing and answering practice questions is more beneficial than time spent highlighting or rewriting notes.

- Think about the time of day when you are the most alert, whether that is the morning, late afternoon or early evening. Make sure that you get enough sleep. Eight to 10 hours per day is the current recommendation for teenagers, based on a review of over 300 studies. Be aware that too much screen time from phones, devices or televisions in the evening can disrupt sleep patterns. Additionally, taking a 20-minute nap can help revive you if you are tired.
- The time spent on study should be devoted to just that, which means avoiding potential distractions like social media, phones or a distracting environment.
- Break up study sessions with short breaks to stretch and refocus. Set a timer if you know you get easily distracted and will delay getting back to study.
- Try to review your classes the day after they are held. The break will give you a better indication of what you need to review.

An example of a weekly timetable is shown in Table 1. Include any study periods at school.

Table 2 lists six learning strategies that should be helpful for effective learning in Chemistry.

There is some evidence that certain popular study methods are much less effective than their popularity with students would suggest. These are highlighting and underlining and rereading materials. This may be because they are also relatively low-effort strategies and are not effective in creating conditions that help long-term recall.

TABLE 2 Learning strategies

Strategy	Description	Application example	Why it is thought to help
Spaced practice	Spreading study schedule over time, rather than in one concentrated block.	Choose times to study and restudy key concepts (such as Lewis structures) on multiple occasions before an exam, rather than cramming the night before. Create palm card sets on topics from earlier in the term or semester and revise them on a regular basis.	Allows consolidation of long-term memories and aids recall.
Interleaving	Switching between topics while studying.	After doing problems on ionic bonding for a few minutes, switch to dipole–dipole interactions and hydrogen bonding.	Gives practice in identifying relevant information and selecting a solution strategy. Encourages greater use of long-term memory strategies.
Retrieval practice	Recalling previously learned information from long-term memory on repeated occasions.	Test yourself on what you can recall about solubility.	Testing by using long-term memory is thought to make accessing and recalling the information easier, by making the information more mentally organised. Also gives feedback on what areas need to be worked on.
Elaboration	Asking questions and explaining why or how things work that way.	Ask and explain why ethanol is soluble in all proportions in water, but pentanol is much less soluble.	Being prompted to create an explanation for a stated fact supports the integration of new information with what is already known.
Concrete examples	Illustrating abstract concepts with specific examples.	Compare the location of the nucleus and electrons in an atom to a pea in the middle of an oval.	Students can create memorable examples that are personally relevant.
Use of images (dual coding)	Combining words with pictures or diagrams.	Draw structures of polyatomic ions, given their formulas.	This can enhance mental organisation, improve understanding of information in a text form, help recall. Chemistry tests often involve diagrams and images.

For each week, set objectives for what you want or need to achieve in each subject. Revise class notes and the relevant section of the textbook. As you read, ask yourself whether you really understand what you are reading. Try explaining it to a family member or friend or a family pet if everyone else is too busy. If there are areas that are troublesome, then write down specific questions to ask a teacher or tutor. A general complaint that 'Solubility makes no sense' is much less useful than a specific 'I'm not understanding how to find the solubility of a salt, given the solubility constant'.

Decide on what is essential material to learn and create summaries to use for revision later. Palm cards help with learning the symbols and terminology of chemistry. Practice lots of problems, preferably of different types. Make sure that this practice is authentic, so that you are getting honest feedback about how you are going. Don't look back at your notes while doing a problem, because this can fool you into thinking that you are doing better than you really are. Do not just look at worked solutions. You need to do problems yourself, under close to exam conditions. Make sure that the calculator that will be used for exams is the one that you use in class and in study time and that you understand how to use the different functions.

Come back to topics at a later date and make sure that you still understand them and haven't forgotten important details. Being well practised is one of the best ways to beat test anxiety.

Study tip

Spending too much time studying can be as bad as too little. Maintain balance in the different aspects of your life to keep yourself motivated and not burn out. Remember that shorter, more frequent bursts of study are more valuable for learning and long-term recall.

During the exam

Examinations often have a perusal time at the start. Use this time to read questions and make a mental note of which questions to answer first. Rules may vary about what writing you can do during perusal time. If you are allowed to write on a rough piece of paper, make notes to yourself about the approach to use to solve problems.

Make sure that you have read and answered every part of a question. It is surprising how many times students answer a different question from the one actually asked on the examination paper.

Examination strategies

Developing a strategy for sitting exams is an important factor for success. Below is a list of tips to consider when honing your exam-taking skills:

1 **Pay attention to details**. That includes using the correct units and significant figures. Make sure that your working for any numerical problem is set out clearly and the logic of your working can be followed. Check your answers at the end. Does your answer make sense?

2 **Multiple-choice questions should only have one correct answer**. If you are not sure about an answer, then try and eliminate incorrect options. This increases your chances of choosing the correct answer. Unless there is a penalty for incorrect answers, you should answer every question.

   If you are struggling with multiple-choice questions try working backwards from the answers given. Read the question carefully. Do not answer too quickly and miss important qualifications or details in the main part of the question.

3 A number of studies have looked at whether you should change an answer that you are unsure about it, or whether the first answer was more likely to be correct. Surprisingly, it was found that when answers were changed, around 55% were from a wrong answer to a right answer, 25% were from a right answer to a wrong answer and 20% were from a wrong answer to another wrong answer. So, statistically speaking, you are better off changing your answer, if you are unsure.

4 For short-answer response questions, you will often be instructed to 'Briefly answer …'. This means to provide an outline or concisely discuss the topic. Do not simply restate the question. This wastes your time. The accuracy and number of the relevant concepts that you respond with is more important than the number of words you write, particularly if most of those words are not relevant. Do not be afraid to use diagrams, reactions and equations in Chemistry exams, particularly if it will make your answer more concise.

5 **Use chemical terminology and symbols correctly**. Do not use 'molecule' to refer to part of a molecule, when what is meant is a 'group', a 'functional group' or 'molecule fragment'. Do not refer to an atom as a molecule or an ion. When drawing chemical structures, avoid common mistakes such as pentavalent carbons (carbons with five bonds). Be aware of the differences between arrow types:

⇌ is an equilibrium arrow

↔ is a resonance arrow

These are two different, distinct concepts.

Draw structures neatly and legibly. The structure on the left in Figure 2 is much more clearly a benzene molecule than the right structure.

FIGURE 2 Drawing chemical structures neatly is an important part of answering chemistry questions. Messy structures can be ambiguous as to the intended meaning.

For longer questions, planning an outline of the answer is useful, to help organise your thoughts and to make sure that you have answered every part of the question.

Responding to cognitive verbs

In assessment tasks and examinations, students will encounter '**cognitive verbs**'. These verbs are '**task words**' that will provide information on what is expected to answer a question. A list of common cognitive verbs is included in Table 3. A full list is provided on your obook assess.

It is important to understand the difference between task words to answer a question correctly. For example a question that asks a student to '**compare**' is different to a command asking a student to '**contrast**'. One requires a student to show similarities and differences, while the other one asks to recognise differences by deliberately highlighting contrasting elements and provide a detailed account of the differences using examples.

Understanding exactly what a cognitive verb is asking for means that a student can provide exactly what the examiner is looking for. Examiners want to give students marks, but can only do so if the student provides the correct information. The cognitive verbs give students guidelines on how to answer the questions to receive full marks. For example, if a student **describes** data in their answer, but does not **analyse** the data, they will not receive full marks.

Below is a list of cognitive verbs that you might find on your assessments.

TABLE 3 List of cognitive verbs

Cognitive verb	Definition	Sample question
Analyse	dissect to ascertain and examine constituent parts and/or their relationships; break down or examine in order to identify the essential elements, features, components or structure; determine the logic and reasonableness of information; examine or consider something in order to explain and interpret it, for the purpose of finding meaning or relationships and identifying patterns, similarities and differences	**Analyse** the graph provided below to identify the relationship between bond length and bond energy.
Apply	use knowledge and understanding in response to a given situation or circumstance; carry out or use a procedure in a given or particular situation	**Apply** the Arrhenius definition of a base to show how $Ca(OH)_2$ qualifies as a base.
Calculate	determine or find (e.g. a number, answer) by using mathematical processes; obtain a numerical answer showing the relevant stages in the working; ascertain/determine from given facts, figures or information	**Calculate** how many molecules are in one mole of hydrogen.
Compare	display recognition of similarities and differences and recognise the significance of these similarities and differences	**Compare** a strong and a weak acid.
Consider	think deliberately or carefully about something, typically before making a decision; take something into account when making a judgment; view attentively or scrutinise; reflect on	**Consider** how a plant's ability to absorb nutrients would be affected if lime (a base) was added to the soil.
Contrast	display recognition of differences by deliberate juxtaposition of contrary elements; show how things are different or opposite; give an account of the differences between two or more items or situations, referring to both or all of them throughout	**Contrast** percentage uncertainty and percentage error in an experimental result.
Deduce	reach a conclusion that is necessarily true, provided a given set of assumptions is true; arrive at, reach or draw a logical conclusion from reasoning and the information given	**Deduce** the relative stability of the reactant and product from the enthalpy level diagram.
Describe	give an account (written or spoken) of a situation, event, pattern or process, or of the characteristics or features of something	**Describe** molar mass.
Determine	establish, conclude or ascertain after consideration, observation, investigation or calculation; decide or come to a resolution	**Determine** the molar mass of a pure substance if 3.5 moles of the substance has a mass of 59.584 g.
Discuss	examine by argument; sift the considerations for and against; debate; talk or write about a topic, including a range of arguments, factors or hypotheses; consider, taking into account different issues and ideas, points for and/or against, and supporting opinions or conclusions with evidence	'The simple rules for solubility of salts in water state that some ionic compounds are "slightly soluble".' **Discuss** this statement.

Cognitive verb	Definition	Sample question
Evaluate	make an appraisal by weighing up or assessing strengths, implications and limitations; make judgments about ideas, solutions or methods in relation to selected criteria; examine and determine the merit, value or significance of something, based on criteria	The optimum pH for water in a swimming pool is considered to be 7.4. A student claimed that this was a neutral pH. **Evaluate** this claim.
Explain	make an idea or situation plain or clear by describing it in more detail or revealing relevant facts; give an account; provide additional information	**Explain** why it is impossible for the percentage yield to be 109%.
Identify	distinguish; locate, recognise and name; establish or indicate who or what someone or something is; provide an answer from a number of possibilities; recognise and state a distinguishing factor or feature	**Identify** the range of values suggested by the measurement 10.3 ± 0.8 seconds.
Interpret	use knowledge and understanding to recognise trends and draw conclusions from given information; make clear or explicit; elucidate or understand in a particular way; identify or draw meaning from, or give meaning to, information presented in various forms, such as words, symbols, pictures or graphs	**Interpret** the data to conclude which reaction was producing oxygen at the fastest rate.
Justify	give reasons or evidence to support an answer, response or conclusion; show or prove how an argument, statement or conclusion is right or reasonable	**Justify** why an absorption or emission spectrum is referred to as a fingerprint of an element.
Predict	give an expected result of an upcoming action or event; suggest what may happen based on available information	An experiment found the density of nitrogen gas to be 1.2504 g L^{-1} at STP. **Predict** what would happen if the density of this gas was increased in a fixed volume container.
Propose	put forward (e.g. a point of view, idea, argument, suggestion) for consideration or action	**Propose** a way to minimise the error caused by the vibrations of people walking past an electronic balance.

Source: *QCAA Chemistry General Senior Syllabus 2019*

As shown in Table 3, some cognitive verbs have multiple definitions. It is therefore important to assess the whole question to determine specifically what the question requires.

Understanding the structure of questions

In addition to understanding cognitive verbs, it is also important to be familiar with the type of question that can be expected on the exam.

In Units 3 & 4, the external assessment will be divided into two papers that will test students on the following types of questions:

- multiple choice
- single-word responses
- sentence responses
- calculations
- paragraph responses.

A typical question will contain a number of components:

- **a cognitive verb** – a task word that will provide the instruction on how to respond to the question
- **quantifying words** – specific numbers of examples or definitions required. Pay attention to these words, as they can save you time on exams. For example if a question asks for one example, you do not need to provide more than that.
- **content words** – specific details about details, facts, processes, theories etc. that need to be considered in the answer.

For example:

1 **Describe** one advantage of paper chromatography.

2 Sugar cubes were used many years ago as a sweetener for tea and coffee, but now granulated sugar is more commonly used. **Justify** why this change was made.

3 **Explain** the two requirements of an 'ideal' gas.

CHECK YOUR LEARNING 1.5

Describe and explain

1 **Explain** which learning strategies from Table 2 you think will be most helpful in your study of chemistry and why.

2 **Describe** the benefits of spacing studying over several study periods, rather than studying in one long session.

Apply, analyse and interpret

3 **Distinguish** between a question that requires a student to 'identify' a process, and a question that requires a student to 'explain' a process.

4 **Distinguish** between a question that requires a student to 'compare' and a question that requires a student to 'contrast'.

5 **Distinguish** between a question that requires a student to 'describe' data, and question that requires a student to 'interpret' data.

Check your obook assess for these additional resources and more:

- » Student book questions
  Check your learning 1.5
- » Video
  How to prepare for your exam
- » Increase your knowledge
  Cognitive verbs
- » Increase your knowledge
  Study timetable template

CHAPTER 1

Review

Summary

1.1
- Chemistry is relevant to many areas of life and occupations. Chemists observe and measure properties of substances; explain and predict patterns in the behaviour of substances; and extend their knowledge to make new substances or purify and detect existing substances. Chemists also use models to explain observable phenomena.

1.2
- The scientific method is used as a basis for the student experiment. This requires a research question and background research to modify or extend an experiment. A hypothesis should be written as a clear statement predicting how changes in the independent variable will affect the value of the dependent variable.

1.3
- Data that is collected during the experiment should be presented appropriately (e.g. as tables, scatterplots or bar charts) and analysed. Methods of analysis include measures of central tendency, standard deviation and correlation. The analysis and interpretation of evidence allows conclusions to be drawn, to prove or disprove the hypothesis.

1.4
- Research investigations require students to develop a research question related to a particular claim. Students must then research the question using credible sources to evaluate claims, draw conclusions and communicate findings scientifically.

Key terms

- abstract
- accuracy
- analytical chemistry
- blank solution
- central tendency
- chemistry
- confounding variable
- controlled variable
- correlation
- dependent variable (DV)
- extraneous variable
- hypothesis
- independent variable (IV)
- line of best fit
- model
- negative control
- positive control
- precision
- quantitative
- replicate experiment
- risk assessment
- scientific method
- standard deviation
- synthetic chemistry
- variable

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; *** = high.

Multiple choice

1 The document that a seller of chemicals in Australia must supply with a chemical is called a:
   - **A** risk assessment
   - **B** certificate of analysis
   - **C** fitness for use warranty
   - **D** materials safety data sheet.

2 Which of the following study strategies is least effective, according to research?
   - **A** Proposing examples from everyday life that illustrate chemical principles
   - **B** Highlighting study notes
   - **C** Spacing multiple study sessions for an exam over several weeks
   - **D** Switching between different types of Chemistry problems while studying

3 Which of the box-and-whisker plots in Figure 1 corresponds to the data given below for volumes of hydrochloric acid used in a set of titrations?

42.0 42.5 42.8 44.0 44.0 43.0 45.0 49.0 52.0

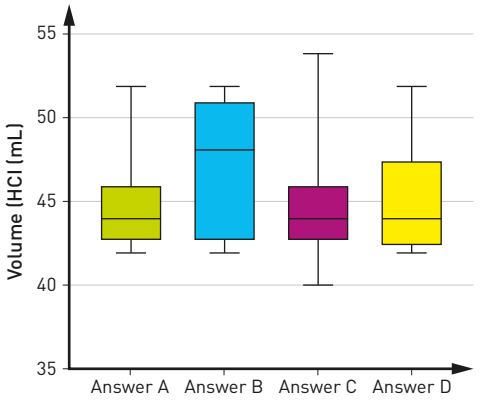


FIGURE 1 Multiple box-and-whisker plots

4 Which of the following is a definition of a dependent variable?
   - **A** Any variable that changes during an experiment
   - **B** A variable that the experimenter is seeking to keep constant, to minimise its effect on an experiment
   - **C** A variable measured or observed during an experiment, in response to an applied change
   - **D** A variable that is systematically and deliberately altered during an experiment in a controlled fashion

5 Of the steps below, which would come last in a student experiment?
   - **A** Process and present the data in a graph
   - **B** Write down observations in a logbook
   - **C** Conduct background research to help modify the experiment
   - **D** Evaluate the validity of the experiment

Short answer

Describe and explain

★6 **Describe** the differences between mean, median and mode in regards to experimental data.

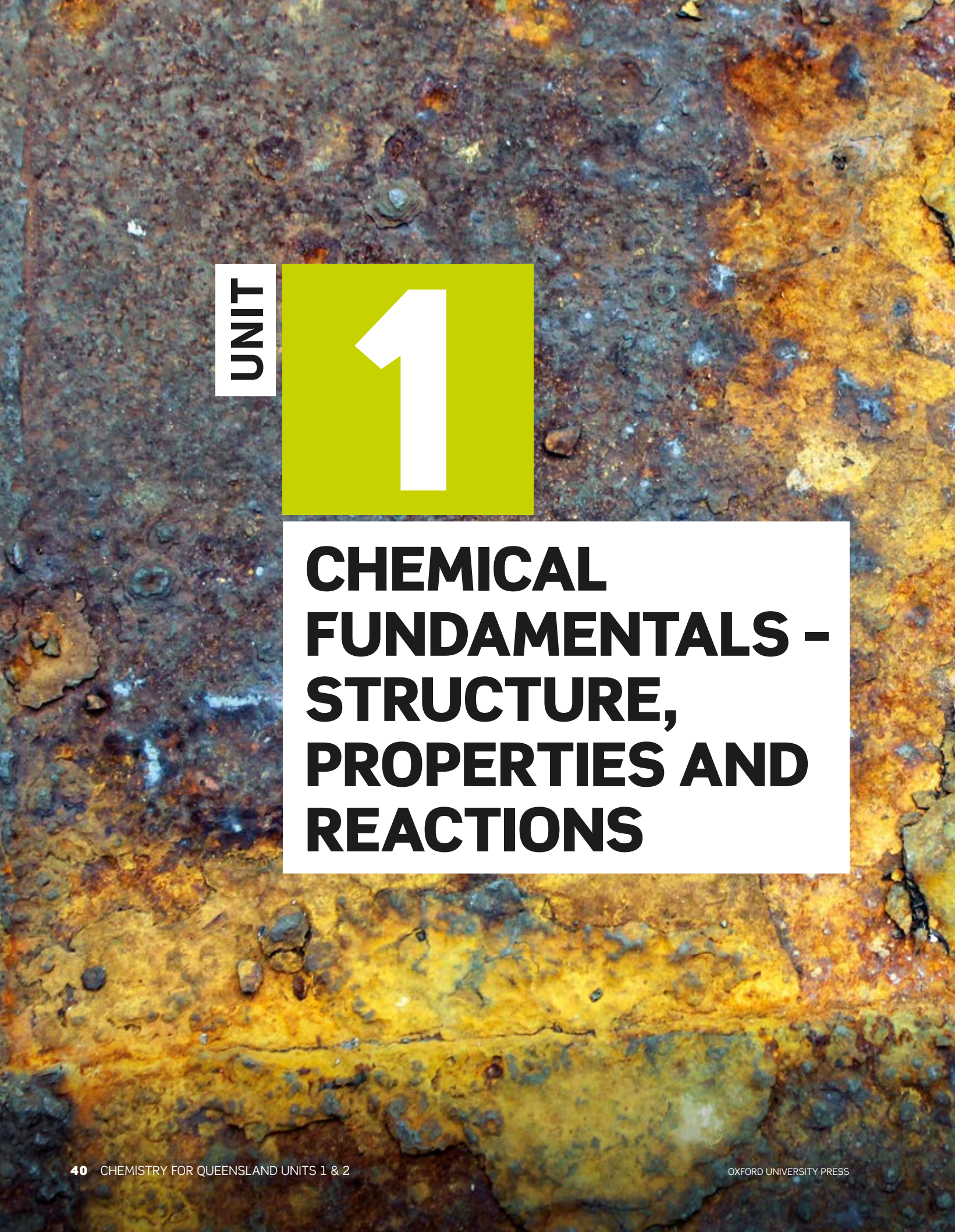
★7 **Explain** the reasons for making repeated measurements in an experiment.

Investigate, evaluate and communicate

★★★8 Use two different online chemistry sources to **investigate** basic information about caffeine (including its melting point, its proper chemical name and the type of elements it contains). **Consider** which database was easier to use and why? Which is the more reliable or trustworthy?

Check your obook assess for these additional resources and more:

- » Student book questions
  Chapter 1 Revision questions
- » Revision notes
  Chapter 1
- » assess quiz
  Auto-correcting multiple-choice quiz
- » Flashcard glossary
  Chapter 1



UNIT 1

CHEMICAL FUNDAMENTALS – STRUCTURE, PROPERTIES AND REACTIONS

All living and non-living objects on Earth are constructed from matter and energy that is neither created nor destroyed. All matter has unique physical and chemical properties. Molecules can break apart and re-form, building whole new structures with new sets of properties.

The ability to model the breaking and re-forming of bonds has allowed chemists to explain how plants from the past became combustible fuels. Burning these fuels releases excess energy in an exothermic reaction. This yields new molecules with new properties. During this process, molecules absorb energy from heat and light, and then release the energy in the spectrum of colours used to distinguish and identify unique matter.

Close examination of the properties of matter has revealed a series of patterns, where elements periodically group together in a manner predicted by their subatomic particles. These particles determine how atoms react and bond by sharing or donating electrons in an attempt to achieve stability with a full outer shell. Measuring such small particles or amounts of chemicals can be difficult, necessitating the use of mass, molar mass and moles of a product, combined with the mathematical calculation of uncertainty and errors.

Chapter guide

Topic 1	Properties and structure of atoms	Chapters 2–5
Topic 2	Properties and structure of materials	Chapters 6–7
Topic 3	Chemical reactions – reactants, products and energy change	Chapters 8–11

Unit objectives

→ Describe and explain the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

→ Apply understanding of the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

→ Analyse evidence about the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy.

→ Interpret evidence about the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

→ Investigate phenomena associated with properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

→ Processes, claims and conclusions about the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

→ Communicate understandings, findings, arguments and conclusions about the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

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FIGURE 1 Rust occurs when an iron or an alloy are exposed to moisture and air over a period of time.

CHAPTER 2

Atomic structure and the periodic table

The atom is the building block for all matter. The concept of the atom was first discovered in 450 BCE by the Greek philosopher Democritus. We can recreate the discovery by cutting a piece of paper in half with scissors, and then into quarters, and so on until it is no longer possible to use the scissors to cut the paper in half. Democritus called this the 'uncuttable'. Imagine using a more accurate and smaller-scale device to get down to a single atom. The remaining atom would measure 0.000 000 000 07 metres.

What exists within the atom determines the type of atom (or element) it will be. Each of the elements are organised into a periodic table, to aid scientists with predicting the properties of elements based on their position in the periodic table. The periodic table was first developed in 1869 by the chemist Dmitri Mendeleev and is now the most widely used reference for any scientist.

OBJECTIVES

- → Understand that atoms can be modelled as a nucleus surrounded by electrons in distinct energy levels held together by electrostatic forces of attraction between the nucleus and electrons; the location of electrons within atoms can be represented using electron configurations; and the structure of the periodic table is based on the electronic configuration of atoms.
- → Use and apply the nuclear symbol notation ${}^{A}_{Z}\mathrm{M}$ to determine the number of protons, neutrons and electrons in atoms, ions and isotopes.
- → Recall the relative energies of the s, p and d orbitals in energy levels to construct electron configurations for atoms and ions up to $Z = 36$ and recognise that the periodic table is arranged into four blocks associated with the four sublevels s, p, d and f.
- → Apply the Aufbau principle, Hund's rule and the Pauli exclusion principle to write electron configurations for atoms and ions up to $Z = 36$ and use orbital diagrams to represent the character and relative energy of orbitals.

→ Recognise the electron configuration of Cr and Cu as exceptions.
→ Recall that elements are represented by symbols and recognise that the structure of the periodic table is based on the atomic number and the properties of the elements.
→ Describe and explain that elements of the periodic table show trends across periods and down groups, including atomic radii, valencies, ionic radii, first ionisation energies and electronegativities as exemplified by groups 1, 2 and 13–18 and period 3.
→ Explain how successive ionisation energy data is related to the electron configuration of an atom.
→ Compare and explain the metallic and non-metallic behaviours of elements, including group trends and the reactivity for the alkali metals (Li–Cs) and the halogens (F–I).
→ Recognise that oxides change from basic through amphoteric to acidic across a period.

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PRACTICALS

SUGGESTED PRACTICAL	**2.4** Investigating periodic trends using a database
SUGGESTED PRACTICAL	**2.5** Simulating Rutherford's gold foil experiment

FIGURE 1 Models of atoms with a central nucleus surrounded by electrons.

2.1 The atomic model

KEY IDEAS

In this section, you will learn about:

- the subatomic particles protons, neutrons and electrons
- the nuclear symbol, which summarises the number of subatomic particles
- the location of electrons within energy levels, which are organised further into sublevels and orbitals.

matter
physical substances that have mass

nucleus
the dense structure within the atom

proton
a positively charged subatomic particle that exists inside the nucleus

neutron
a neutral subatomic particle that exists inside the nucleus

subatomic particle
a component that makes up, and exists within, the atom

electron
a negatively charged subatomic particle that exists outside of the nucleus

electrostatic force
an attraction between objects of opposing charges

An atom is an individual unit that makes up all **matter**. It was once thought that the atom was the smallest unit in the universe – the term comes from the Greek word *atomos*, meaning 'indivisible.' However, it is now known that the atom is made up of even smaller particles.

Because there is not yet a microscope powerful enough to see a detailed image of an atom's internal structure, scientists have developed a number of models to describe an atom's structure.

Inside the atom

The current model of the atom includes a number of components.

The **nucleus** is a significant component of the atom. It contains the **protons** and **neutrons**, which are two of the atom's **subatomic particles**. The nucleus contains nearly all of an atom's mass, yet it can be as much as 100 000 times smaller in diameter than the atom itself.

The protons in the nucleus each have a positive charge and mass of 1.67×10^{-24} g. The neutrons do not have a charge and have the same mass as the protons. Because of the protons' charge, the nucleus has an overall positive charge.

The third subatomic particle is the **electron**. Electrons have a negative charge and are insignificant in mass compared to the protons and neutrons (9.11×10^{-28} g). The electrons are located outside of the nucleus. They remain close to the nucleus because of the **electrostatic forces** of attraction between the positive charge of the nucleus and the surrounding negatively charged electrons.

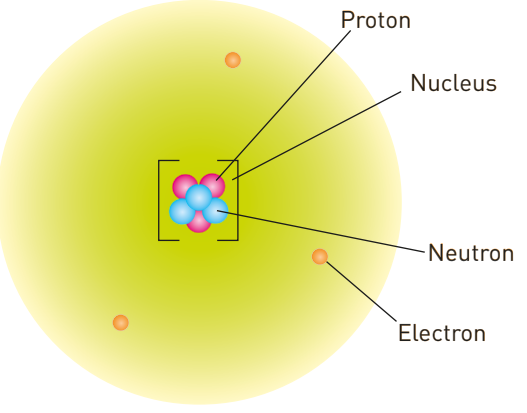

FIGURE 1 A basic model of an atom showing the subatomic particles.

CHALLENGE 2.1

Empty space in a body

Imagine removing all of the empty space from a person's body to leave just the electrons, protons and neutrons from every atom in every cell of the body.

Discuss what would happen to the size and mass of the person.

Nuclear symbol

The nuclear symbol summarises the subatomic particles within an atom (Figure 2).

- Z is the atomic number. It is used for identification because it is the number of protons, which does not change.
- A is the mass number. This represents the total number of subatomic particles in the nucleus (the protons and neutrons). The mass number is sometimes represented as a whole number on the **periodic table**, but it can include decimal points.
- X is the chemical symbol used for each of the elements in the periodic table.

FIGURE 2 The nuclear symbol; for example, $^{4}_{2}He$.

Applying the nuclear symbol

An atom, in its original state, has a neutral charge because the positive charge from the protons cancels out the negative charge from the electrons. If the protons and electrons are not balanced, then the atom has an ionic charge (either positive or negative). The charged atom is called an **ion**.

An **isotope** is an atom with a variable number of neutrons but the same number of protons and electrons. The concept of ions and isotopes is covered in more detail in Chapter 4.

Table 1 outlines how to determine the number of subatomic particles present. This applies to neutral atoms, ions and isotopes.

TABLE 1 Determining subatomic particles

Subatomic particle	Neutral atom	Ion	Isotope
Proton (p^+)	Use the atomic number from the periodic table. $p^+ = Z$	Use the atomic number from the periodic table. $p^+ = Z$	Use the atomic number from the periodic table. $p^+ = Z$
Neutron (n^0)	Subtract the atomic number from the mass number given in the periodic table. $n^0 = A - Z$	Subtract the atomic number from the mass number given in the periodic table. $n^0 = A - Z$	Subtract the atomic number from the atomic mass of the specific isotope. n^0 = atomic mass $- Z$
Electron (e^-)	Use the atomic number from the periodic table. $e^- = Z$	Subtract the ionic charge from the atomic number. $e^- = Z$ – ionic charge	Use the atomic number from the periodic table. $e^- = Z$

WORKED EXAMPLE 2.1

a Determine the number of neutrons in the neutral atom $^{19}_{9}F$,

b Determine the number of electrons in the negative ion $^{14}_{7}N^{3-}$.

SOLUTION

a $n^0 = A - Z$
$= 19 - 9$
$= 10$ neutrons

b $e- = Z$ – ionic charge
$= 7 - (-3)$
$= 10$ electrons

Study tip

The first letter of a chemical symbol is always capitalised. If there is a second letter, it is lower case.

periodic table
an organised presentation of elements by increasing atomic number

Study tip

Not every periodic table is presented in the same manner, but there will be a legend to help interpret the information. The legend is usually in the middle with hydrogen as an example.

ion
an atom containing an electric charge

isotope
atoms that have the same number of protons and electrons but different numbers of neutrons

CASE STUDY 2.1

Chemical symbols used in history

Alchemy was born in ancient Egypt, where mummification was an application of rudimentary chemical knowledge. By 332 BCE, the ancient Greeks had organised their understandings of chemistry into the four elements of nature – fire, earth, air and water. These four elements were given the symbols:

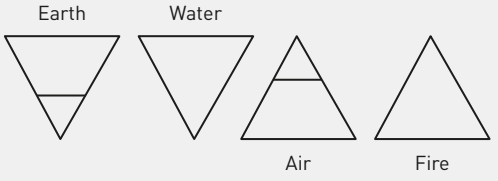

FIGURE 3 The ancient Greek symbols of the four elements of nature.

Classification of matter extended into common metals and other solids and liquids, because they were visible to the eye and readily available (Figure 4).

Up until the research performed by English chemist John Dalton (1766–1844), chemical symbols referred to both pure and mixed materials. Dalton used symbols for each of the elements that make up a material (Figure 5).

FIGURE 4 Alchemical symbols used to represent substances visible to the naked eye.

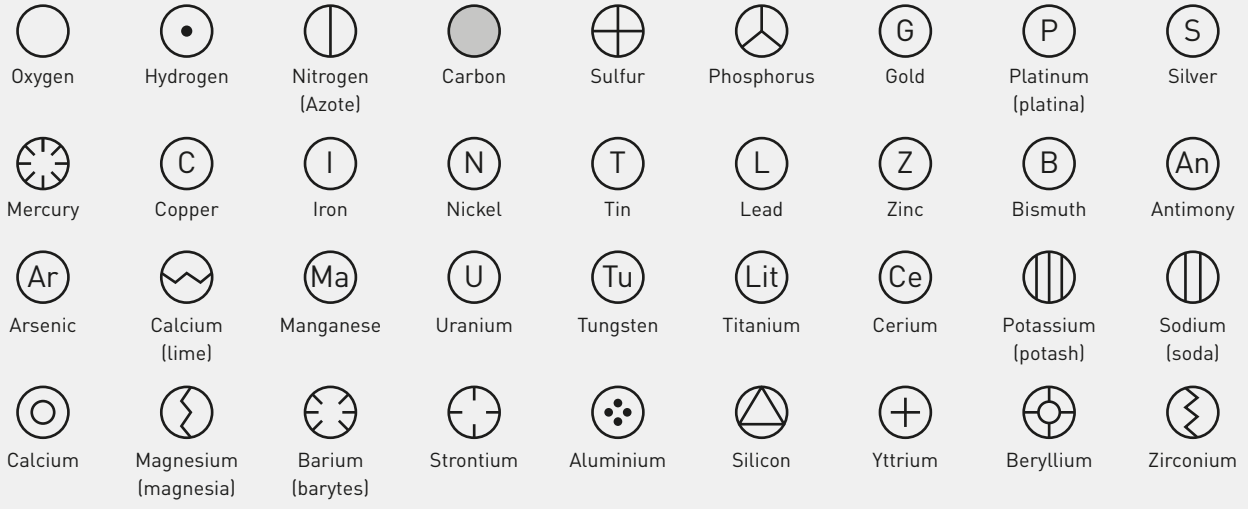

FIGURE 5 Dalton used symbols to represent pure elements.

Energy levels

Outside of the nucleus, the electrons are located within distinct energy levels at different distances from the nucleus. The electrons closest to the nucleus have the largest electrostatic attraction to the nucleus and the least amount of energy. The energy (n) levels are numbered from 1 to 7, and are further divided into sublevels.

As the energy levels increase, there is a tendency for electrons in the outer energy levels to spend more time on average further from the nucleus. This means that electrons in higher energy levels experience less of a pull from the positively charged nucleus.

There are four sublevels: s, p, d and f. Each sublevel has its own number of orbitals and distinct orbital shapes that denote the most probable location of its electrons. The s orbitals are spherical, the p orbitals are dumbbell-shaped, most of the d orbitals are cloverleaf shaped, and the f orbitals are varied in shape. As shown in Table 2, not every energy level contains every sublevel.

TABLE 2 Breakdown of sublevels present in each energy level

Energy level	Sublevel			
	s	p	d	f
1	✓			
2	✓	✓		
3	✓	✓	✓	
4	✓	✓	✓	✓

CHECK YOUR LEARNING 2.1

Describe and explain

1 **Describe** the properties of the three subatomic particles in an atom.

2 **Explain** why electrons in the 7s sublevel would have less electrostatic attraction to the nucleus than electrons in the 2s sublevel.

3 Use the periodic table to **demonstrate** the nuclear symbol for:

- **a** magnesium
- **b** titanium
- **c** bromine.

Apply, analyse and interpret

4 **Interpret** the nuclear symbol $^{23}_{11}Na$ to determine the number of protons, neutrons and electrons present in the neutral atom.

5 **Interpret** the nuclear symbol $^{27}_{13}Al^{3+}$ to determine the number of protons, neutrons and electrons present in the ion.

6 **Determine** how many electrons are equivalent to the mass of one proton.

Check your obook assess for these additional resources and more:

- » Student book questions
  Check your learning 2.1
- » Challenge
  2.1 Empty space in the body
- » Video
  Inside an atom
- » Interactive
  The atom

2.2 The periodic table

KEY IDEAS

In this section, you will learn about:

- the organisation of elements in the periodic table into vertical groups and horizontal periods, and into other sections such as metals and non-metals
- how the location of an element is linked to its properties.

element
a type of atom that is defined by its atomic number

Elements refer to each of the types of atoms that exist, defined by their atomic number; for example, carbon, hydrogen and iron. The periodic table organises elements in ascending order (by atomic number) to emphasise the regular repetition of chemical and physical properties of the elements. The structure of the periodic table provides information about an element simply by its location (Figure 1).

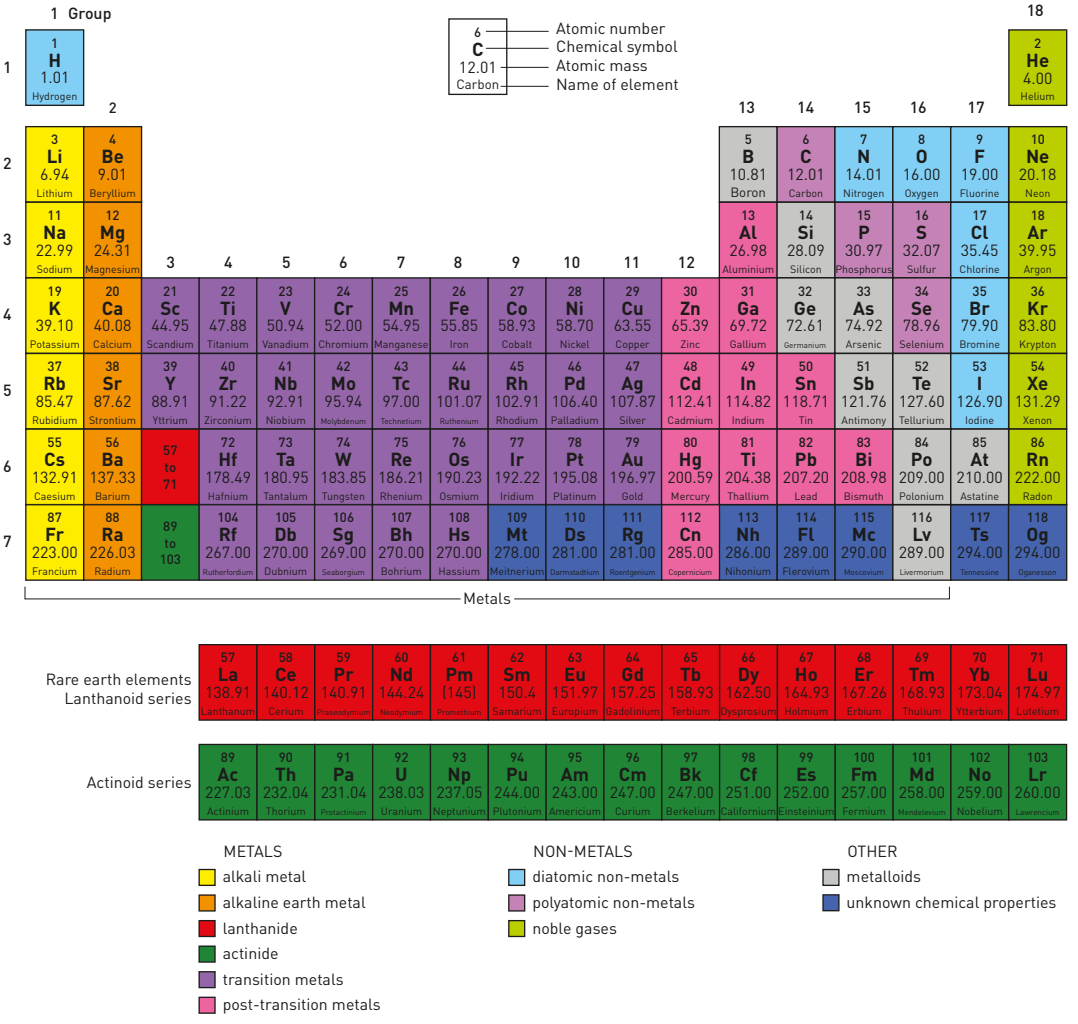

FIGURE 1 The periodic table organises elements by atomic number. See the Appendix for a larger periodic table.

Metals and non-metals

The periodic table consists mainly of metallic elements, with the non-metals located at the top right-hand side.

Metals are usually shiny solids with high **melting points**, heat and electrical conductivities and **densities**. Metals are also very **malleable**. The exception is mercury, which is a liquid at room temperature.

Non-metals typically have low melting and **boiling points**, do not conduct heat and electricity, and do not have a shiny surface. Examples of non-metals are oxygen, nitrogen, carbon, phosphorus, sulfur, chlorine, neon, helium and hydrogen (even though it is not usually located with the others).

Metalloids

Where the metals and non-metals overlap are the **metalloids**, which display both metallic and non-metallic properties. For example, silicon is a relatively good electrical conductor, but only when heated. Other metalloids are boron, germanium, arsenic, antimony, tellurium, polonium and astatine.

melting point
the temperature at which a solid becomes a liquid

density
how closely packed atoms are arranged; the mass of a substance per unit of volume

malleable
able to be pressed into other shapes

boiling point
the temperature at which a liquid becomes a gas

metalloid
an element that displays both metallic and non-metallic properties

Groups and periods

The periodic table also uses vertical columns (**groups**) and horizontal rows (**periods**) as indicators of an element's properties. There are 18 groups of the periodic table, and seven periods.

TABLE 1 Main groups of the periodic table

Group	Given name	Common properties
1	Alkali metals	Silver coloured, soft metals with a low density. They easily react with water and the halogens. Their reactivity increases down the group (Li–Cs), with caesium being the most reactive.
2	Alkaline earth metals	Silver coloured, hard metals with a slightly higher density than alkali metals. They react slowly with water, sometimes requiring hot water or even steam.
17	Halogens	Exist as non-metallic **diatomic molecules** in their natural state. They are highly reactive with alkali metals, and can be harmful or lethal to the environment. The halogens decrease in reactivity down the group (F–At) with fluorine being the most reactive.
18	Noble gases	Mostly unreactive.

Study tip

Not every periodic table indicates the metalloids. To help remember them, start at silicon and move down one, then one to the right, until reaching astatine. And don't forget boron up the top.

group
a vertical column of the periodic table

period
a horizontal row of the periodic table

diatomic molecule
a molecule containing two atoms of the same element or of different elements bonded together

CHALLENGE 2.2

The authority for naming elements

The International Union of Pure and Applied Chemistry (IUPAC) is an organisation that has represented chemists in individual countries since 1919. It is the recognised world authority for the naming of elements and compounds.

Discuss the significance of an organisation such as IUPAC, and why it is necessary for the advancement of science.

CASE STUDY 2.2A

Hydrogen and the *Hindenburg* disaster

A blimp is a long cigar-shaped gas-filled bag with a propeller and engine. Invented in 1852, the blimp was used for military and civilian purposes, including transatlantic travel in the early 1900s. Hydrogen was the gas used because it is Earth's lightest element, and can be obtained easily.

On 6 May 1937, a German passenger blimp, the *Hindenburg*, had just completed a round trip between Germany and Brazil, and was on route to the United States. This was intended to be the first of ten round trips between Germany and the United States. The *Hindenburg* began its descent in New Jersey. The landing involved dropping ropes at a high altitude and the blimp being winched down to the ground.

Bad weather made it difficult for the pilot to get into the right position. There were orders to drop weights from the blimp, and to valve the forward gas cells. Witnesses recall seeing fabric fluttering as if gas was leaking, and then mushroom-shaped flames destroyed the blimp. The *Hindenburg* was carrying 97 people; 35 of these people died as well as a worker on the ground.

Dozens of other hydrogen blimps were destroyed by fire, so the use of hydrogen for blimps was completely abandoned by the late 1930s. Blimps still exist but no longer use a flammable gas. Instead, they use helium, which although not as light as hydrogen, is still lighter than air, and is a noble gas – meaning it is unreactive. Today, blimps are used mainly for advertising, TV coverage, tourism and research purposes.

FIGURE 2 The *Hindenburg* was one of many blimps destroyed by fire because of the use of flammable hydrogen gas.

Other categories of the periodic table

There are a number of other categories that group elements within the periodic table, including transition metals, basic metals and rare earth elements.

Transition metals

Transition metals are located in groups 3–12. They often form colourful compounds; for example, manganese compounds can be pink, red, black, purple, blue, and green.

Basic metals

The basic metals are located underneath the non-metals and metalloids but to the right of the transition metals. Examples of basic metals are aluminium, gallium, indium, tin, thallium, lead, and bismuth. They can be recognised as metals because they are good heat and electrical conductors, and are lustrous, dense, malleable and ductile. However, the basic metals also have some similar non-metallic properties as the elements above them. Lead and gallium are soft, and some of the basic metals have lower melting points than is expected of a metal.

Rare earths

Known as the rare earth elements or inner transition metals, the two rows displayed underneath the periodic table consist of the lanthanoids (also known as lanthanides) (first row) and the actinoids (also known as actinides) (second row). They are considered special transition metals that are very dense and can be radioactive. Some are found naturally, and others are synthetic.

CASE STUDY 2.2B

Origins of element names

When comparing the name of an element with its chemical symbol, there is often an obvious link between the two. For example, carbon is C and nitrogen is N. However, sometimes the letters in the chemical symbol do not match the name. This is sometimes because the element once had a different name. Table 2 lists elements whose chemical symbols come from their original name.

TABLE 2 Origins of chemical symbols that do not match the names

Element name	Symbol	Original name	Origin
Antimony	Sb	Stibium	Latin
Copper	Cu	Cuprum	Latin
Gold	Au	Aurum	Latin
Iron	Fe	Ferrum	Latin
Lead	Pb	Plumbum	Latin
Mercury	Hg	Hydragyrum	Greek
Potassium	K	Kalium	Latin
Silver	Ag	Argentum	Latin
Sodium	Na	Natrium	Latin
Tin	Sn	Stannum	Latin
Tungsten	W	Wolfram	German

CHECK YOUR LEARNING 2.2

Note: Use the periodic table in the Appendix (page 459) to help answer these questions.

Describe and explain

1 **Describe** how the periodic table is organised.

2 **Identify** the ways mercury has properties that are exceptions to those of other metals.

Apply, analyse and interpret

3 **Compare** the terms 'atom' and 'element'.

4 **Categorise** an element in period 4 and group 17 as a metal, non-metal or metalloid. Use the periodic table to determine the answer.

Investigate, evaluate and communicate

5 **Investigate** alternative structures used to represent the periodic table.

6 Mendeleev used the organisation of his periodic table to predict the properties of undiscovered elements. **Evaluate** the significance of his predictions of scandium, gallium and germanium (Table 3).

TABLE 3 Mendeleev's predictions of scandium, gallium and germanium

Element	Property	Mendeleev's predictions	Observed properties
Scandium (discovered in 1877)	Atomic mass	44	44.95
	Oxide formula	M_2O_3	Sc_2O_3
	Density of oxide (g mL^{-1})	3.5	3.86
	Solubility of oxide	Dissolves in acids	Dissolves in acids
Gallium (discovered in 1875)	Atomic mass	68	69.72
	Oxide formula	M_2O_3	Ga_2O_3
	Density of metal (g mL^{-1})	6.0	5.96
	Solubility of oxide	Dissolves in ammonia solution	Dissolves in ammonia
	Melting temperature	Low	30°C
Germanium (discovered in 1886)	Atomic mass	72	72.61
	Oxide formula	MO_2	GeO_2
	Density of metal (g mL^{-1})	5.5	5.47
	Density of oxide (g mL^{-1})	4.7	4.70
	Melting temperature	High	900°C

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- **Challenge**
  2.2 The authority for naming elements
- **Weblink**
  Interactive periodic table
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  IUPAC periodic table

2.3 Electron configurations

KEY IDEAS

In this section, you will learn about:

- electron configurations and how they are used to represent the location of each of the electrons within the energy sublevels
- how electron configurations can be represented as full or condensed, and align with the element's location in the periodic table
- orbital diagrams, which nearly always comply with three rules: the Aufbau principle, the Pauli Exclusion principle and Hund's rule.

electron configuration
a representation of the electrons in an atom

An **electron configuration** shows where the electrons are located within an atom. Each energy level or shell surrounding an atom is given a letter (s, p, d and f). A breakdown of each energy level capacity is shown in Table 1.

TABLE 1 Breakdown of sublevel capacity

Sublevel	Number of orbitals	Number of electrons
s	1	2
p	3	6
d	5	10
f	7	14

Study tip

You need to know the order for filling electron configurations. It is not chronological, but there is a visual pattern shown in the diagonal diagram (Figure 1), to help remember the order.

In an electron configuration, the energy level is the number, the sublevel is the letter, and the number of electrons occupying the orbitals is the superscript. The order of energy levels and sublevels when creating an electron configuration is as follows:

$1s^2 \rightarrow 2s^2 \rightarrow 2p^6 \rightarrow 3s^2 \rightarrow 3p^6 \rightarrow 4s^2 \rightarrow 3d^{10} \rightarrow 4p^6$

CHALLENGE 2.3

A pattern for electron configurations

There are many ways to memorise a sequenced list of information. Analyse the diagram in Figure 1 to propose what mathematical pattern is present, starting at 2p.

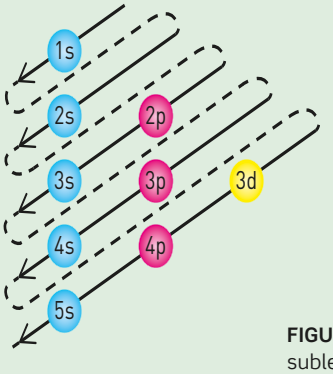

FIGURE 1 The order in which sublevels or subshells should be filled for electron configurations.

full electron configuration
an electron configuration that includes every electron in an atom

Types of electron configurations

A **full electron configuration** states the location of every electron in the atom, beginning at 1s and continuing until there are no more electrons.

WORKED EXAMPLE 2.3A

Determine the full electron configuration of silicon.

SOLUTION

Silicon has 14 electrons. Beginning at 1s, the full electron configuration is:

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

Remember the superscript dictates the number of electrons in each orbital.

So, 2 + 2 + 6 + 2 + 2 = 14 electrons.

condensed electron configuration
an electron configuration that only states electrons in the highest energy levels, based on the nearest noble gas

Noble gases are located in the far-right column of the periodic table: helium, neon, argon, krypton, xenon and radon. A **condensed electron configuration** is a way of abbreviating the configuration by stating the previous noble gas in square brackets and then stating the location of the remaining outer electrons in the atom.

The closest noble gas with a smaller atomic number is argon (18 electrons). Beginning at argon, the condensed electron configuration is: $[Ar]4s^2 3d^8$.

So, [Ar] + 2 + 8 = 10 electrons and [18] + 10 = 28 electrons

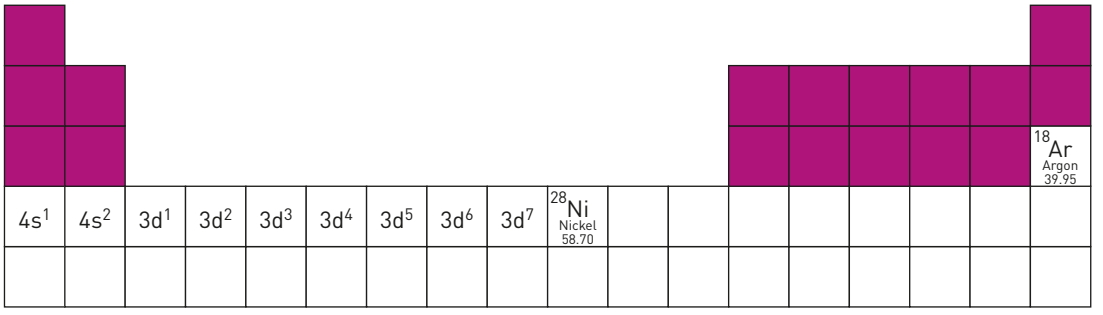

FIGURE 2 Representation of nickel's condensed electron configuration.

Layout of the periodic table

The periodic table can be divided into four blocks, labelled by the four types of sublevels: s, p, d and f (Figure 3).

Starting at the top-left corner of the periodic table (hydrogen), electron sublevels are filled from left to right across the periodic table as the atomic number increases by one each time. The energy sublevel filled (or partially filled) last will align with the location of that element within the periodic table.

For example, phosphorus is located in the p block of the periodic table, specifically within the second row of the p block (which refers to 3p). Therefore the electron configuration for phosphorus should end at 3p. Phosphorus has 15 electrons. Filling the sublevels in order, the full electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^3$.

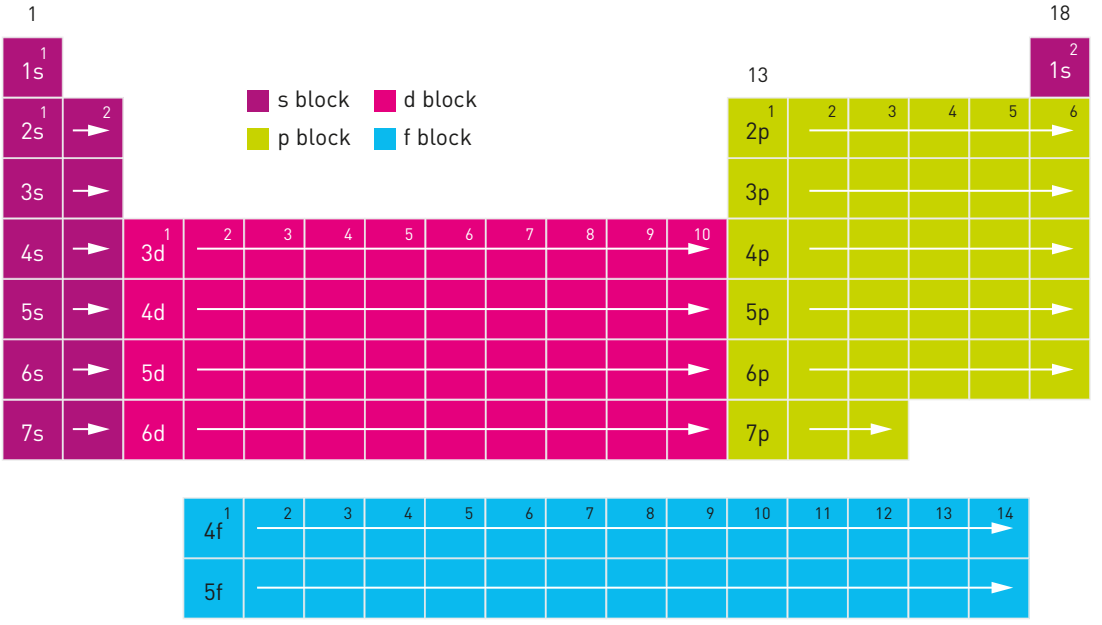

FIGURE 3 A periodic table organised into blocks to represent the s, p, d and f sublevels.

Study tip

Not every periodic table has the same layout, especially for the f block. It is important to note that 4f begins with ^{57}La and finishes with ^{70}Yb (14 elements in total). 5d begins with ^{71}Lu. 5f begins with ^{89}Ac and then 6d begins with ^{103}Lr.

Orbital diagrams

Orbital diagrams are a visual representation of electron configurations (Figure 4). They give more information specific to the individual electrons involved. Each square represents an orbital and sublevels that contain more than one orbital appear together. For example, d sublevels contain five orbitals so there will be five squares together. The orbitals appear in ascending order, based on the order in which they fill with electrons. There are three rules to follow when creating orbital diagrams as outlined here.

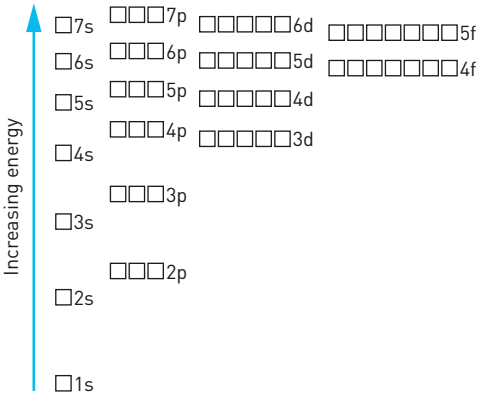

FIGURE 4 The energy levels of atomic orbitals. Each square represents an orbital.

Aufbau principle

The Aufbau principle states that electrons enter the orbital of the lowest energy first. Beginning at 1s, the electrons cannot be placed in any orbital outside of this order, although this strict order starts to break down further down the periodic table, as orbitals become closer in energy.

Pauli exclusion principle

The Pauli exclusion principle states that an atomic orbital can contain at most two electrons at any time. This means that an orbital can contain 0, 1 or 2 electrons. The first electron to enter an orbital is drawn as an upwards-pointing half arrow. A second electron must have an opposing direction to the first electron (Figure 5).

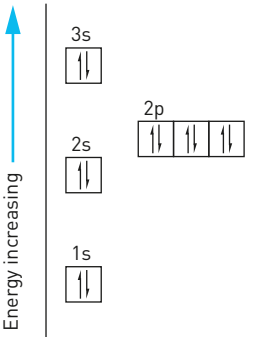

FIGURE 5 An orbital diagram for magnesium, which has 12 electrons. The first electron in each orbital is drawn as an upwards arrow, the second as a downwards arrow.

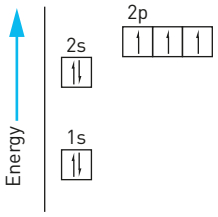

FIGURE 6 An orbital diagram for nitrogen, which has 7 electrons. This shows Hund's rule because p has three orbitals, with one upwards pointing electron in each.

Hund's rule

Hund's rule states that when electrons occupy orbitals of equal energy (p has three orbitals, d has five orbitals and f has seven orbitals), then one upwards pointing electron must be entered into each of the orbitals before a second electron can enter any of the orbitals (Figure 6).

WORKED EXAMPLE 2.3B

Apply the Aufbau principle, Pauli exclusion principle and Hund's rule to create an orbital diagram of Selenium, which has 34 electrons.

SOLUTION

- Selenium has 34 electrons, filling from the lowest energy sublevel (1s) according to the Aufbau principle.
- According to the Pauli exclusion principle, each orbital can contain up to two electrons (represented as arrows). The first electron in each orbital is an upward pointing arrow, and the second arrow in each orbital is downward pointing. The first 28 electrons are filled in this manner, completely occupying the orbitals up to the 3d energy sublevel.
- The final four electrons must occupy the 4p energy sublevel. Hund's rule states that all the upward pointing electrons must be represented first, which uses three of the four electrons. The final electron is the only downward pointing arrow in the 4p energy sublevel, and is located in the first orbital.

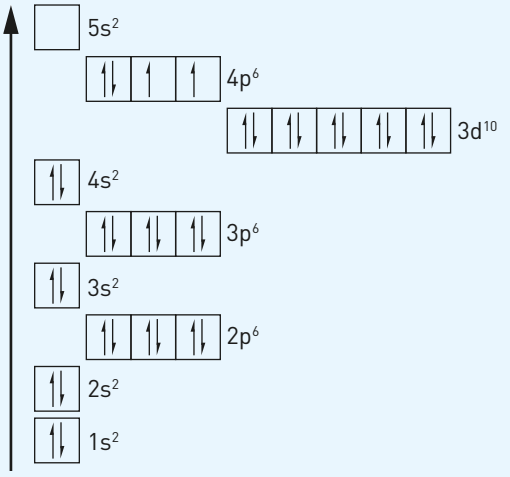

FIGURE 7 Orbital diagram for Se

- The total number of arrows indicates the number of electrons. There are 34.

Study tip

Students are required to construct electron configurations and use orbital diagrams for atoms and ions up to $Z = 36$.

Electron configuration exceptions

As with many other scientific relationships observed in chemistry, there are exceptions to the Aufbau principle. Chromium has 24 electrons and is located in the d block (more specifically it is the fourth element in the third row of the periodic table). The expected electron configuration of chromium is:

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

The actual electron configuration of chromium is:

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

TABLE 2 Effect of electron occupancy on stability

Level of stability	Degree sublevel is filled with electron	Example
Large	Completely full	$3d^{10}$
Medium	Half-full	$3d^5$
Small	Partially full	$3d^2$ or $3d^8$

The expected electron configuration has 4s full and 3d partially full, whereas the actual electron configuration has both 4s and 3d half-full and therefore is more stable overall. This is because some energy is required to pair up two electrons in the same orbital and the 4s and 3d orbital are very close in energy. The same situation applies to copper: 4s becomes half-full so that 3d can become completely full.

CHECK YOUR LEARNING 2.3

Describe and explain

1 **Identify** how many orbitals are present in each of the electron sublevels.
2 **Define** the condensed electron configuration for chromium.
3 **Construct** an orbital diagram for cobalt.

Apply, analyse and interpret

4 **Determine** the maximum number of electrons that can occupy each of the electron sublevels.
5 **Interpret** the full electron configuration to identify the element.
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^2$
6 **Interpret** the condensed electron configuration to identify the element.
[Ne] $3s^2\ 3p^6\ 4s^1$
7 **Determine** the number of electrons present in:
a 4p
b 5d
c the third energy level.
8 **Determine** the number of electrons present in the highest occupied sublevel for:
a tantalum
b oxygen.
9 **Identify** errors in the following orbital diagrams and suggest and how they should be fixed. Which rules for filling electron orbital have been broken in each instance.

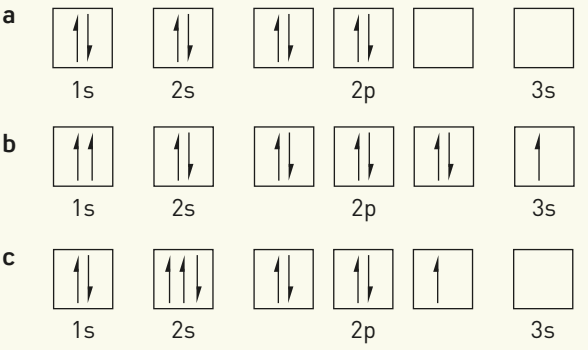

FIGURE 8

Investigate, evaluate and communicate

10 Using the order of energy levels and sublevels, **generate** the full electron configuration of boron, chlorine and lithium.
11 **Modify** the full electron configurations from question 9 into condensed electron configurations.
12 The complete electron configuration for bromine is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$. The accepted convention for writing electron configurations rearranges bromine's complete electron configuration into $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^5$ to give meaning to the readers. **Evaluate** the significance of this minor rearrangement.
13 **Investigate** the significance of using noble gases as reference points in condensed electron configurations.

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2.4 Periodic trends

KEY IDEAS

In this section, you will learn about:

- atomic radii and metallic character decreasing across a period and increasing down a group
- valency increasing across a period and staying consistent down a group
- ionic radii being smaller than atomic radii when the ion is positive, and larger than atomic radii when the ion is negative
- ionisation energy and electronegativity increasing across a period and decreasing down a group
- how oxides are acidic when oxygen is bonded to non-metals and basic when oxygen is bonded to metals.

Atomic radius

atomic radius used to describe the size of an atom, which is measured by halving the distance between the nuclei of adjacent atoms

picometre 10^{-12} or 0.000 000 000 001 metres

The **atomic radius** refers to the size of an atom. The outer edge of an atom is too difficult to measure because of the difficulty of locating the electrons precisely, so the distance between adjacent nuclei is measured (in **picometres**) and that quantity is halved.

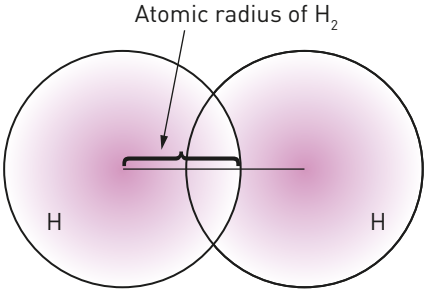

FIGURE 1 A representation of how the atomic radius is determined.

Atomic radii increase down a group because of the electron shielding effect. With each energy level that is added to an atom, there are many more electrons that need to occupy space, so the atom increases in size. For example, within the alkali metal group, lithium is smaller than sodium, which is smaller than potassium and then rubidium (Table 1).

Study tip

Think of adding another energy level as a person putting on another jumper or sweater. The extra clothing is going to take up some room even if it is thin. The more layers of clothing that are added, the bigger a person looks.

TABLE 1 Information about the alkali metals

Alkali metal	Atomic number	Atomic radius (pm)	Alkali metal	Atomic number	Atomic radius (pm)
Lithium	3	152	Rubidium	37	248
Sodium	11	227	Caesium	55	300
Potassium	19	280	Francium	87	348

Atomic radii decrease from left to right across a period because of electrostatic attractions. With each increase in atomic number, there is an increase in the number of positively charged protons (in the nucleus) and the attractive force that the negatively charged electrons experience. The electrostatic attraction between the additional subatomic particles pulls the electrons closer to the nucleus, and the atomic radius gets slightly smaller. The smallest atomic radius in any period is that of the noble gas, and the largest is that of the alkali metal.

Valency

valency the combining power of an element as determined by the number of electrons in its outer shell (valence electrons)

Valence electrons are the electrons occupying the outer most energy level of an atom. For example, silicon has an electron configuration of $1s^2\,2s^2\,2p^6\,3s^2\,3p^2$. The third energy level contains silicon's four valence electrons – two in the 3s orbital and two in the seperate 3p orbital.

Valency increases across a period, with the noble gas in any period having complete sublevels. This means noble gases are more stable than the other elements in the period, and largely unreactive. With an increase of one in the atomic number, there is an increase of one electron, and the need to begin filling the s sublevel in the next period.

Valency remains constant down a group. Every alkali metal has one valence electron in its s sublevel, and every halogen has seven valence electrons (two in its s sublevel and five in its p sublevel). The valency determines many of the elements' properties, including their reactivity.

Ionic charge and radius

Atoms tend to lose or gain electons to achieve complete s and p sublevels, particularly the first row of elements. This process is achieved by the easiest route possible. For example, aluminium has a valency of 3 ([Ne] $3s^1 3p^2$). To have completed s and p sublevels, aluminium can either gain five electrons or lose three electrons. A loss of three electrons requires less energy, and therefore is preferred.

Table 2 shows the general trend for ionic charge across the periodic table.

TABLE 2 General trend for ionic charge

Group	Ionic charge	Group	Ionic charge
1	+1	15	–3
2	+2	16	–2
13	+3	17	–1
14	+4/–4	18	0

Study tip

Angstrom (Å) is a unit of measurement equating to 10^{-10} m or 0.1 nanometres.

An atom with fewer than four valence electrons will lose electrons and form a positive ionic charge as a result. Positive ions are smaller than their neutral atom counterparts because they have lost space-occupying electrons (Figure 2).

An atom with more than four valence electrons will gain electrons and form a negative ionic charge as a result. Negative ions are larger than their neutral atom counterparts because they have gained space-occupying electrons.

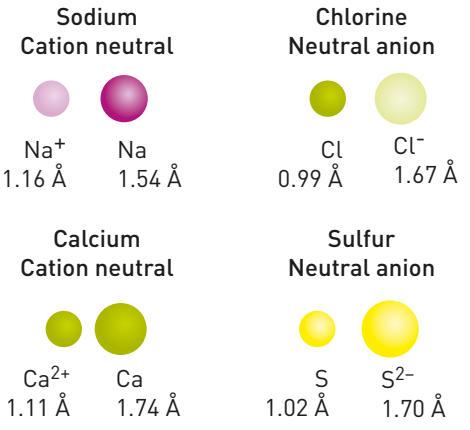

FIGURE 2 Cations are smaller than the neutral atoms. Anions are larger than the neutral atoms.

Ionisation energy

Ionisation energy (IE) is the energy required to remove an electron from a gaseous atom, forming a positively charged ion:

$$K(g) \rightarrow K^+(g) + e^-$$

ionisation energy the energy (in kJ mol^{-1}) required by a gaseous atom to remove an electron from that atom

IE increases across a period because the atom has significantly more electrostatic attraction between its nucleus and valence electrons. For example, sodium has an IE of 495 kJ mol^{-1} and chlorine has an IE of 1257 kJ mol^{-1}. Chlorine has significant electrostatic attractions to its seven valence electrons, and tends to become an ion by gaining an electron to have a complete p sublevel, rather than losing an electron to get further away from stability.

IE decreases going down a group because of the electron shielding effect (Figure 3). With each energy level that is added to an atom, there is additional distance between the negative valence electrons and the positive nucleus. The additional distance causes a weaker electrostatic attraction between the nucleus and the valence electrons.

Period	Group 1												13	14	15	16	17	18
1	H 1318	2																He 2370
2	Li 526	Be 900											B 800	C 1090	N 1400	O 1310	F 1680	Ne 2080
3	Na 495	Mg 735	3	4	5	6	7	8	9	10	11	12	Al 580	Si 780	P 1060	S 1005	Cl 1255	Ar 1527
4	K 420	Ca 590	Sc 630	Ti 660	V 650	Cr 660	Mn 710	Fe 760	Co 760	Ni 730	Cu 740	Zn 910	Ga 580	Ge 780	As 960	Se 950	Br 1140	Kr 1350
5	Rb 400	Sr 550	Y 620	Zr 660	Nb 670	Mo 680	Tc 700	Ru 710	Rh 720	Pd 800	Ag 730	Cd 870	In 560	Sn 700	Sb 830	Te 870	I 1010	Xe 1170
6	Cs 380	Ba 500	La 540	Hf 700	Ta 760	W 770	Re 760	Os 840	Ir 890	Pt 870	Au 890	Hg 1000	Tl 590	Pb 710	Bi 800	Po 810	At ...	Rn 1030
7	Fr ...	Ra 510																

FIGURE 3 First ionisation energies (kJ mol^{-1}) of some elements.

Successive ionisation energies

An atom's first IE is the energy required to remove one electron from its valence energy level. An atom's second IE is the energy required to remove a second electron from its valence energy level.

Therefore, the valency of an atom has a significant impact on the differences between an atom's successive ionisation energies. Alkali metals have one valence electron, so they have relatively small ionisation energies to remove the valence electron, which results in an ionic charge of +1 and achieves complete valence sublevels (with the electron configuration of the previous noble gas). A second IE is significantly larger because an electron is being removed from a sublevel that is full and stable, and there are greater forces of electrostatic attraction because the electrons are closer to the nucleus. With each successive IE, the value increases again, as the removal of more electrons means that there is an increasing imbalance between the number of protons and the number of electrons remaining, and each electron is experiencing a greater positive charge.

Alkaline earth metals have relatively small values for the first two ionisation energies, as they develop the same electron configuration of the previous noble gas. However, the third IE is greatly increased in value. Group 13 has three small ionisation energies and then a large fourth IE. This trend continues along the p block of the periodic table, aligning the number of small ionisation energies an atom has with the number of valence electrons it has (Table 2).

TABLE 3 Successive ionisation energies (IE) (kJ mol^{-1}) of some elements in the third row

Element	First IE	Second IE	Third IE	Fourth IE	Fifth IE	Sixth IE	Seventh IE
Na	502	4560					
Mg	744	1445	7730				
Al	584	1815	2740	11 600			
Si	793	1575	3220	4350	16 100		
P	1018	1890	2905	4950	6270	21 200	
S	1006	2260	3375	4565	6950	8490	27 000
Cl	1257	2295	3850	5160	6560	9360	11 000
Ar	1527	2665	3945	5770	7230	8780	12 000
			Increase →				Decrease ↑

WORKED EXAMPLE 2.4

Predict which element would have a greater IE. Justify your response.

a magnesium or phosphorus

b boron or indium

SOLUTION

a Magnesium and phosphorus are both located in period 3. Going across the period, the IE increases, so phosphorus would be expected to have a greater IE. This is confirmed by the fact that magnesium has two valence electrons so tends to lose electrons, and phosphorus has five valence electrons so tends to gain (not lose) electrons.

b Boron and indium are both located in group 3. Going down the group, the ionisation energy decreases, so boron would be expected to have a greater IE. This is because Boron only has two energy levels (so has a large electrostatic attraction to its valence electrons), whereas indium has five energy levels. The added distance weakens the electrostatic attraction to the valence electrons, so electrons are easier to remove from the atom.

Electronegativity

Electronegativity is the ability of a positive nucleus to electrostatically attract the valence electrons of the adjacent atom. With a large electronegativity, an atom can draw a neighbouring electron closer to itself in order to share that electron, or even gain the electron for itself. A small electronegativity hinders the atom's ability to share or gain neighbouring electrons.

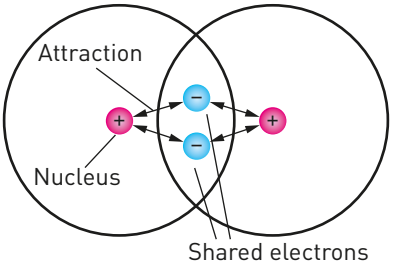

FIGURE 4 Electronegativity allows a positive nucleus to attract negative electrons so nearby atoms can share electrons and become stable.

Electronegativity increases across a period because of the electrostatic attractions caused by the increasing number of an atom's protons and valence electrons. The exception to this periodic trend is the noble gases because they have full valence sublevels and therefore no incentive to gain or share more electrons.

Electronegativity decreases down a group because of the electron shielding effect. As the distance increases between an atom's nucleus and its own valence electrons, the electrostatic attraction to the valence electrons and the electrons of the adjacent atom decreases.

electronegativity the attraction of a positively charged nucleus the negatively charged electrons of a neighbouring atom

Study tip

An atom has a greater electronegativity if that atom itself intends to gain an electron.

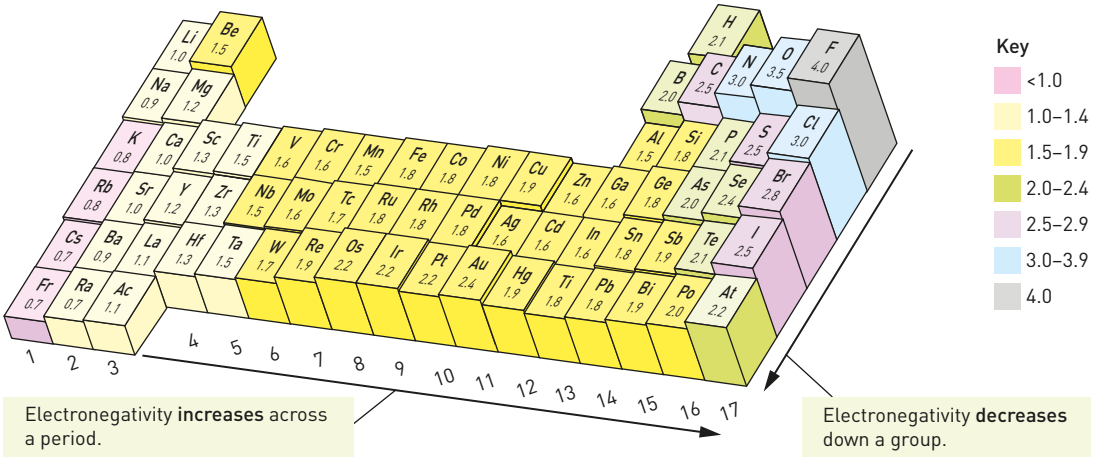

FIGURE 5 Electronegativity increases across a period and decreases down a group. Fluorine is the most electronegative element. Caesium and francium are the least electronegative. Group 18 (nobel gases) has been excluded from this periodic table because their unreactive nature means their electronegativities cannot be determined experimentally. The exception is krypton and xenon which react to form molecules in a small number of circumstances, and therefore have electronegativity values.

Other periodic trends

There are a number of other periodic trends, as shown in Figure 6.

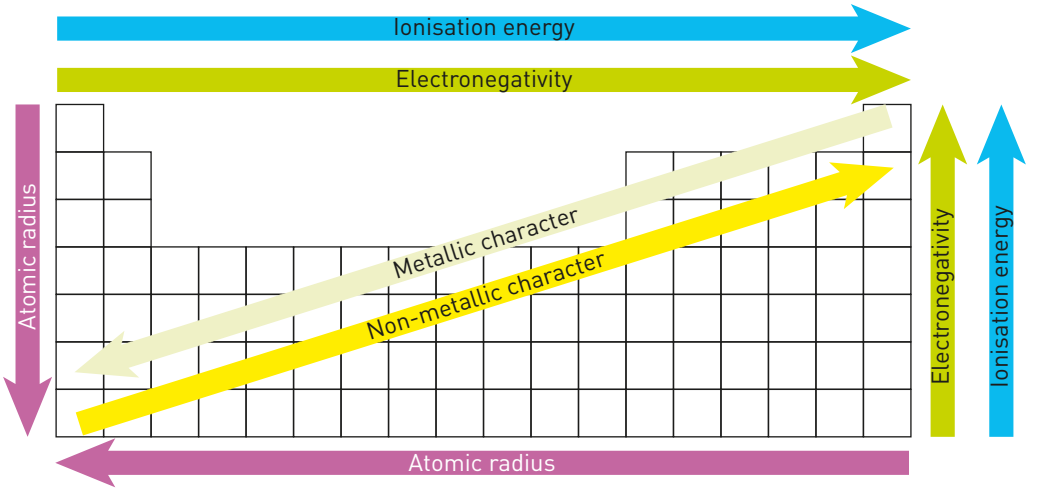

FIGURE 6 Summary of the periodic table trends.

Metallic character

Metallic character refers to the extent to which an element exhibits metallic properties, such as high melting and boiling points, conductivity, malleability and ductility.

Metallic character increases from the top right to the bottom left of the periodic table (Figure 6) and depends on the fact that most metals have large atomic radii and small electronegativities and ionisation energies. As a result, metals tend to lose electrons to other atoms when reacting, forming positive ions. The most reactive metal is francium; it is so reactive that many of its metallic properties have never been observed. Caesium, just above francium and the second highest in metallic character, is so reactive that it can spontaneously ignite in air.

basic
has a pH of higher than 7

amphoteric
can be basic or acidic

acidic
has a pH of lower than 7

Oxides

Oxides are compounds that contain oxygen and another element, which can be a metal or non-metal. The element's degree of metallic or non-metallic properties determines what type of oxide the compound will be. As metallic character decreases across a period, oxides change from **basic** through **amphoteric** to **acidic**.

Acidic oxides

Acidic oxides form between oxygen and non-metals. Examples are carbon dioxide (CO_2), diphosphorus pentoxide (P_2O_5) and nitrogen dioxide (NO_2).

When an acidic oxide reacts with water, it forms an acid:

CO_2	+	H_2O	→	H_2CO_3
acidic oxide		water		acid

When an acidic oxide reacts with a base it forms a salt and water:

CO_2	+	$2NaOH$	→	Na_2CO_3	+	H_2O
acidic oxide		base		salt		water

SUGGESTED PRACTICAL

2.4 Investigating periodic trends using a database
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Basic oxides

Basic oxides form between oxygen and metals. Examples are sodium oxide (Na_2O), magnesium oxide (MgO) and copper oxide (CuO).

When a basic oxide reacts with water, it forms a base:

Na_2O	+	H_2O	→	2NaOH
basic oxide		water		base

When a basic oxide reacts with an acid, it forms salt and water:

Na_2O	+	2HCl	→	2NaCl	+	H_2O
basic oxide		acid		salt		water

Amphoteric oxides

Amphoteric oxides have both acidic and basic properties. They are metals that can form more than one **oxidation state** (or **ionic charge**). Examples of metals capable of this are copper, zinc, tin, lead, aluminium and beryllium.

When an amphoteric oxide reacts with an acid, it forms a salt:

Al_2O_3	+	6HCl	→	$2AlCl_3$	+	$3H_2O$
amphoteric oxide		acid		salt		water

When an amphoteric oxide reacts with a base, it forms a salt:

Al_2O_3	+	2NaOH	+	$3H_2O$	→	$2Na(Al(OH)_4)$
amphoteric oxide		base		water		salt

oxidation state/ ionic charge
the difference between the positive charges from the protons and the negative charges from the electrons, represented as a superscript on the right- hand side of an elemental symbol

Study tip

When writing chemical equation they need to be balanced on both sides. A weblink for this can be found on your obook assess.

CHECK YOUR LEARNING 2.4

Describe and explain

1 **Explain** how an atomic radius differs from an ionic radius.
2 **Describe** the electron shielding effect.

Apply, analyse and interpret

3 **Determine** which element would have a greater ionisation energy.
a Al or Cl **b** N or Sb **c** Rb or As
4 **Determine** which element would have a greater atomic radius.
a K or Br **b** Be or Sr **c** Li or Te
5 **Determine** which element would have a greater electronegativity.
a Al or Cl **b** O or F **c** Na or N
6 **Analyse** the relationship between the location of elements on the periodic table and their electronegativities, using the data in Figure 5.

Investigate, evaluate and communicate

7 **Investigate** electron affinity and its applicability to the periodic trends. Write a paragraph detailing what it is, and how it changes throughout the periodic table. Source a graph from the internet showing atomic number against electron affinity to analyse in respect to each element's electron configuration.
8 **Evaluate** the significance that some metals with high oxidation states act as acidic oxides instead of basic oxides.
9 **Evaluate** why an atom containing four valence electrons (e.g. carbon) can either lose or gain electrons.

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2.4 Investigating periodic trends using a database

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2.4 Atom valence

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Balancing chemical equations

SCIENCE AS A HUMAN ENDEAVOUR

2.5 Historical development of the atomic model

KEY IDEAS

In this section, you will learn about:

- Democritus's ideas that everything is composed of atoms
- Dalton's theory that there are different kinds of atoms that can combine or mix
- Thomson's suggestion that there are negative electrons embedded within a positive atom
- Rutherford's discovery of the dense nucleus
- Bohr's suggestion that electrons exist within discrete energy levels
- Schrödinger's proposal of an electron cloud to account for the mathematical probability of an electron's location
- Chadwick's discovery of the neutron.

Throughout history, a number of key theories have influenced the way people think about atoms.

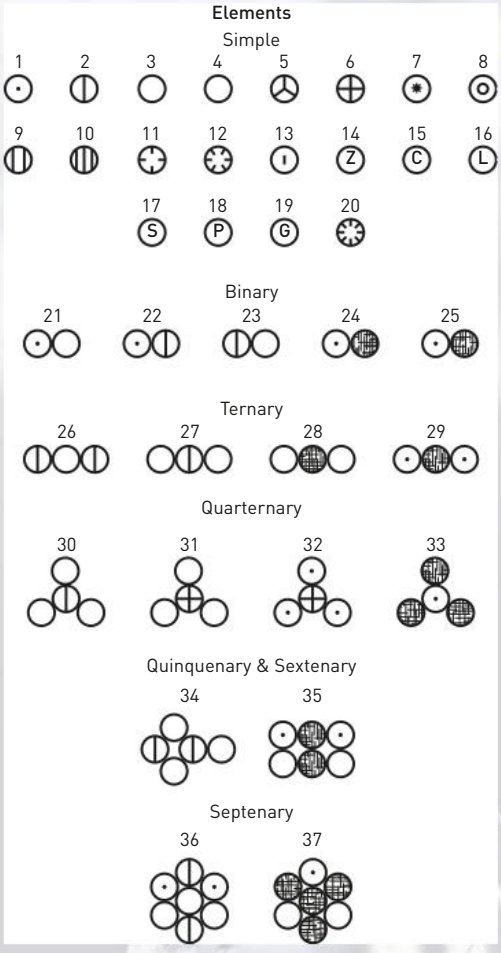


FIGURE 2 Dalton's proposal of how atoms combine to form molecules.

Democritus

Democritus (460–370 BCE) was an ancient Greek philosopher who focused on finding the purpose and reasoning behind matters in the world. Publishing more than 70 books in his lifetime, Democritus was the first to propose the idea that everything is composed of atoms (from the Greek word *atomos* meaning 'indivisible'). These atoms were considered solid, physically indivisible and indestructible. Democritus went further to say that atoms can differ in their size, shape, mass, position and arrangement. For example, he proposed that solid iron atoms are small and rough (to stick together), liquid water atoms are large and round (having a smooth and slippery surface), and gaseous air atoms are light and whirling (to be blown away).

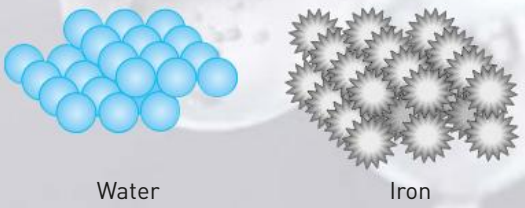


FIGURE 1 A representation of Democritus' proposal for the structure of water and iron.

Dalton

John Dalton (1766–1844), an English chemist and schoolteacher, conducted research in many areas of science, including gases and colour blindness, with which he was afflicted. Dalton's contributions to the development of the atomic model came from his use of scientific methodologies to study the ratios of elements in chemical reactions. In 1803, Dalton proposed four ideas.

1 All elements are composed of tiny, indivisible atoms.
2 Atoms of the same element are identical to each other. Atoms of different elements are different from each other.
3 Atoms can physically mix, or chemically combine, with one another in simple whole number ratios
4 Chemical reactions occur when atoms separate, join or rearrange with each other. They do not change into other atoms.

Thomson

Joseph John (J.J.) Thomson (1856–1940) was an English physicist and lecturer. He developed an experiment that passed electric current through gases under low pressure (within a sealed gas tube that contained a positively charged metal disc at one end of the container and a negatively charged metal disc at the other end). Thomson observed a glowing beam (cathode ray) that travelled from the negative disc to the positive disc. He theorised that the atom was not indivisible as previously thought, but contained particles within it that were negatively charged. He changed the gases and the metal discs, each time observing the glowing cathode ray. Thomson extended his experiment by placing a positive plate above the gas tube and a negative plate below the gas tube. This experiment showed the upwards deflection of the cathode ray towards the positive plate, proving that the particles he had found were negative.

Given that atoms themselves were neutral, Thomson theorised the 'plum pudding' model, where the atom itself was positive, and the negative electrons were located throughout, in a fixed position.

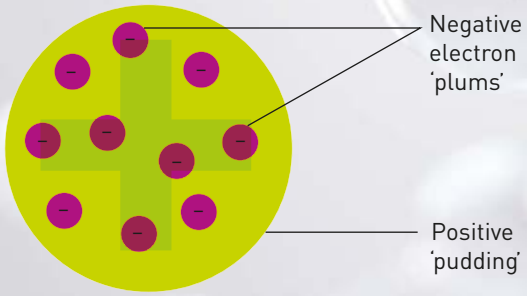


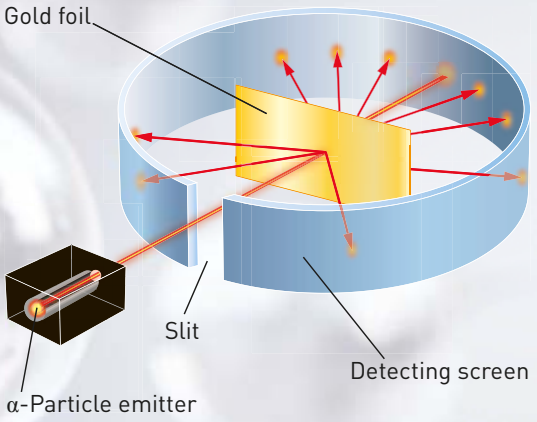
FIGURE 3 Thomson's plum pudding model of the atom.

Goldstein

Eugene Goldstein (1850–1930) was a German physicist who conducted similar work to Thomson. While not as well known, Goldstein is attributed with having discovered the proton. He found that while the electrons were travelling towards the positive metal disc there were also other particles travelling towards the negative metal disc. They consisted of the positively charged residual gas particles inside the tube, after the electrons were removed.

Rutherford

Ernest Rutherford (1871–1937) was a New Zealand-born physicist and student of J.J. Thomson. Rutherford discovered the idea of radioactive half-lives, and used his experience with **alpha (α) particles** in his most notable work, which led to the discovery of the nucleus. Rutherford set up an experiment in which he fired alpha particles at gold foil to study its scatter



alpha particle
a helium atom with removed electrons

FIGURE 4 Rutherford's experimental set-up in which he discovered that atoms are mostly empty space.

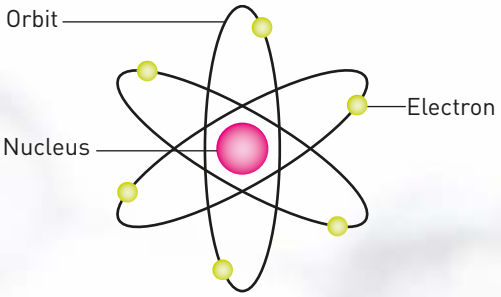


FIGURE 5 Rutherford proposed of a dense central nucleus surrounded by mostly empty space.

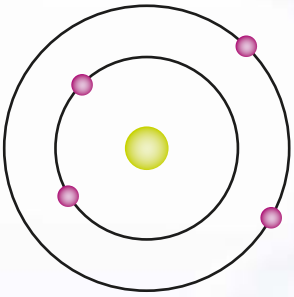


FIGURE 6 Bohr proposed that electrons exist in discrete levels.

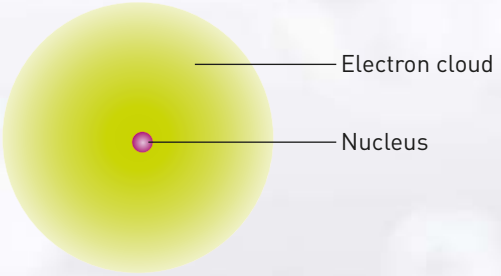


FIGURE 7 Schrödinger proposed that electrons were located in a cloud that surrounded the nucleus

quantum theory a theory that describes the behaviour of matter and energy at the atomic and subatomic level

pattern, expecting to see the alpha particles pass straight through with only a slight deflection due to repulsion of the positive protons.

Instead, Rutherford observed that most of the alpha particles travelled through the gold foil with no deflection. From this, he proposed that the atom is mostly empty space. A small number of the alpha particles had deflected, but the angle was significantly greater than he had anticipated. Rutherford suggested that this deflection was not the result of repulsion from the positive protons in the gold atoms, but that the alpha particles themselves had collided with a small and very dense structure within the atoms – the nucleus. Rutherford proposed that the nucleus had most of the atom's mass and carried a positive charge.

Bohr

Niels Bohr (1885–1962) was a Danish physicist and student of Rutherford. Beginning his early research in surface tension, Bohr extended his research into the development of an atomic model after he began working under Rutherford. Bohr proposed that the orbit of the electrons around the nucleus could be classified into different energy levels, and that the electrons could only occupy particular energy levels. Bohr's research eventually suggested that an electron could move between energy levels, but this theory was considered radical and was unacceptable to the greater scientific community at the time.

Schrödinger

Erwin Schrödinger (1887-1961) was an Austrian physicist who was a fundamental contributor to **quantum theory** and the next evolution in atomic models – the quantum mechanical model. He suggested that electrons, while contained within particular energy levels, do not follow predetermined orbits, but instead exist within an electron cloud. Schrödinger focused on using mathematical modelling to calculate the probability that an electron is in various locations around the nucleus. An area of higher probability of electrons being present is indicated by being most dense, and an area of lower probability is indicated by being less dense (Figure 7).

Schrödinger's research also introduced the concept of energy sublevels and the shape of atomic orbitals within the sublevels.

Chadwick

Study tip

Knowledge of atomic orbital shapes is not required.

James Chadwick (1891–1974) was another English physicist who contributed to the development of the atomic model – through his discovery of the neutron. In 1932, Chadwick fired alpha particles at beryllium atoms and observed the production of an unknown type of radiation particle that was neither positive nor negative. He proposed that the neutral particle was the approximate mass of a proton, contained within the nucleus (Figure 8).

SUGGESTED PRACTICAL

2.5 Simulating Rutherford's gold foil experiment
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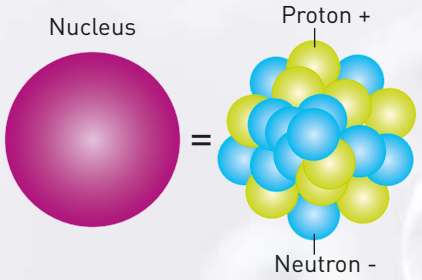

FIGURE 8 Chadwick's proposed that the dense nucleus contained both protons and neutrons.

CHALLENGE 2.5

Electron orbits

Explain how the ideas of electron orbits and orbitals differ between the Bohr and quantum mechanical models.

CHECK YOUR LEARNING 2.5

Describe and explain

1 **Describe** how the state of matter of iron at room temperature led to Democritus's interpretation of the iron atom.
2 **Summarise** the four ideas that Dalton proposed.
3 **Explain** the intended meaning for electrons being referred to as 'plums' in a 'pudding'.
4 Data from Rutherford's gold foil experiment is provided in Table 1. **Calculate** the percentage of alpha particle deflections with an angle of deflection of 75° or greater.

TABLE 1 Data from Rutherford's gold foil experiment

Angle of deflection (°)	Number of deflections
5	8 289 000
10	502 570
15	120 570
30	7800
45	1435
60	477
75	211
>105	198

Apply, analyse and interpret

5 Rutherford's experimental data showed a dense nucleus with the remainder of the atom as empty space. If all materials are made of mostly empty space, **reflect on** why it is not feasible to walk through walls.
6 Democritus did not employ a scientific method. **Consider** the significance of this in terms of his proposal on atoms.

Investigate, evaluate and communicate

7 Using a cathode ray experiment, Thomson would hold the positive plate above the cathode tube and the negative plate below the cathode tube.
   a **Sketch** the experiment.
   b **Predict** what would happen if the positive plate was placed on the left of the cathode tube, and the negative plate on the right of the cathode tube.
8 **Evaluate** the significance of the teacher–student links between many of the scientists that developed the atomic model.

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CHAPTER 2

Review

Chapter summary

2.1
- An atom can be modelled as a nucleus surrounded by negatively charged electrons in distinct energy levels (and sublevels) held together by electrostatic forces of attraction between the nucleus and electrons. Inside the nucleus are the positively charged protons and neutral neutrons.
- The nuclear symbol notation ${}^{A}_{Z}X$ can be used to determine the number of protons, neutrons and electrons in atoms, ions and isotopes.

2.2
- Elements are represented by chemical symbols and are located in the periodic table according to the atomic number and properties of the elements.
- The periodic table is divided into metals, metalloids and non-metals, as well as horizontal periods and vertical groups group 1 is the alkali metals, group 2 is the alkaline earth metals, groups 3–12 are the transition metals, group 17 is the halogens, and group 18 is the noble gases.

2.3
- Energy levels are divided into sublevels that have one or more orbitals. There are four types of orbitals, which have relative energies: s, p, d and f. Each orbital contains up to two electrons.
- Electron configurations can be represented in full or condensed formats, and are used to represent the location of the electrons in an atom or ion.
- The periodic table is based on the electron configurations of atoms, and is arranged into four blocks associated with the four sublevels s, p, d and f.
- The Aufbau principle, Hund's rule and the Pauli exclusion principle are applied when writing electron configurations and creating orbital diagrams.

2.4
- Periodic trends are relationships between groups of elements and their properties, when structured in increasing atomic number and organised into periods and groups.
- Atomic radii and metallic character decrease across the period and increase down the group.
- Valency increases across the period and stays consistent down the group. Ionic radii are smaller than atomic radii when the ion is positive, and larger than atomic radii when the ion is negative.
- Oxides are acidic when the oxygen is bonded to non-metals and basic when the oxygen is bonded to metals.
- Ionisation energy and electronegativity increase across the period and decrease down the group. Successive ionisation energies refer to the removal of the first valence electron, then second, third and so on.

2.5
- Experiments provided evidence that enabled scientists to develop models of the atom over time.
- Democritus proposed the idea that everything is composed of atoms, comparing iron to water.
- Dalton theorised there are different kinds of atoms that can combine or mix in differing whole number ratios.
- Thomson used cathode ray experiments to suggest that there are negative electrons embedded within a positive atom.
- Rutherford then targeted alpha particles at gold foil, and discovered the dense nucleus.
- Bohr suggested electrons exist within discrete energy levels.

- Schrödinger proposed an electron cloud to account for the mathematical probability of an electron's location.
- Chadwick discovered the neutron. Research on atomic models is constantly changing.

Key terms

- acidic
- alpha particle
- amphoteric
- atomic radius
- basic
- boiling point
- condensed electron configuration
- density
- diatomic molecule
- electron
- electron configuration
- electronegativity
- electrostatic force
- element
- full electron configuration
- group
- ion
- ionisation energy
- isotope
- malleable
- matter
- melting point
- metalloid
- neutron
- nucleus
- oxidation state/ionic charge
- period
- periodic table
- picometre
- proton
- quantum theory
- subatomic particle
- valency

Key formulas

Protons for atoms, ions and isotopes	$p^+ = Z$
Neutrons for atoms and ions	$n^0 = A - Z$
Neutrons for isotopes only	n^0 = atomic mass $- Z$
Neutrons for atoms and isotopes	$e^- = Z$
Electrons for ions only	$e^- = Z$ – ionic charge

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; *** = high.

Multiple choice

1 Identify the number of protons, electrons and neutrons respectively, based on the nuclear symbol $^{9}_{4}\mathrm{Be}$.

A 4, 5, 5 **B** 4, 4, 4
C 4, 4, 5 **D** 4, 5, 4

2 What is the number of protons, electrons and neutrons respectively, based on the nuclear symbol $^{24}_{12}\mathrm{Mg}^{2+}$?

A 12, 12, 12 **B** 12, 10, 12
C 12, 14, 12 **D** 12, 12, 24

3 Identify the correct nuclear symbol for cobalt.

A $^{58}_{27}\mathrm{Co}$ **B** $^{27}_{59}\mathrm{Co}$
C $^{112}_{6}\mathrm{C}$ **D** $^{59}_{27}\mathrm{Co}$

4 Identify the correct element based on the full electron configuration:
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^3$

A niobium **B** vanadium
C phosphorus **D** titanium

5 Identify the correct element based on the condensed electron configuration:
$[Ar]4s^2 3d^{10} 4p^3$

A arsenic **B** argon
C calcium **D** zinc

Short answer

Describe and explain

★ **6** **Describe** the steps required to create an orbital diagram.

★ **7** **Explain** why chromium and copper do not follow the standard electron configuration conventions.

★ **8** **Explain:**

a electronegativity

b ionisation energy.

Apply, analyse and interpret

★★ **9** There are two adjacent elements located in period 4 of the periodic table. The first element has a completed third energy level (including all of the sublevels) and the second element has at least one p electron in an unfilled shell. **Identify** the elements.

★★ **10** A sample of chromium with a mass of 51.996 g contains 6.02×10^{23} atoms. **Calculate** the mass of one atom of chromium.

★★ **11** **Interpret** the graph in Figure 1 of the first two energy sublevels: 1s and 2s. **Infer** which atomic model is represented. Give reasons for the response.

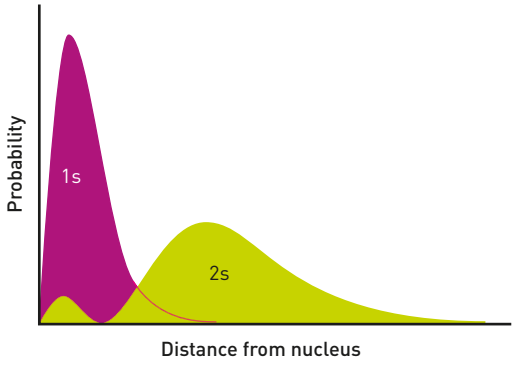

FIGURE 1 Radial probability distribution for 1s and 2s electrons

★★ **12** **Determine** the number of electrons present in:

a 2s

b 5f

c the second energy level.

★★★ **13** The following electron configurations are incorrect. **Identify** the error and **justify** why they ar errors.

a $1s^1\ 2s^2\ 2p^4$

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

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

Investigate, evaluate and communicate

★ **14** **Create** an orbital diagram of potassium.

★★ **15** Using the order of energy levels and sublevels, **create** the full electron configuration of silicon, rubidium and fluorine.

★★ **16** **Manipulate** the full electron configurations from question 15 into condensed electron configurations.

★★ **17** **Investigate** why corpuscles were renamed as electrons.

★★ **18** The full electron configuration given for nitrogen is $1s^2\ 2s^2\ 2p^3$. This represents nitrogen as a neutral atom. If nitrogen was ionised to N^{3-}, **determine** the full electron configuration that could best represent this.

★★★ **19 Evaluate** the effectiveness of the alternative periodic tables in Figure 2 for predicting properties and providing information.

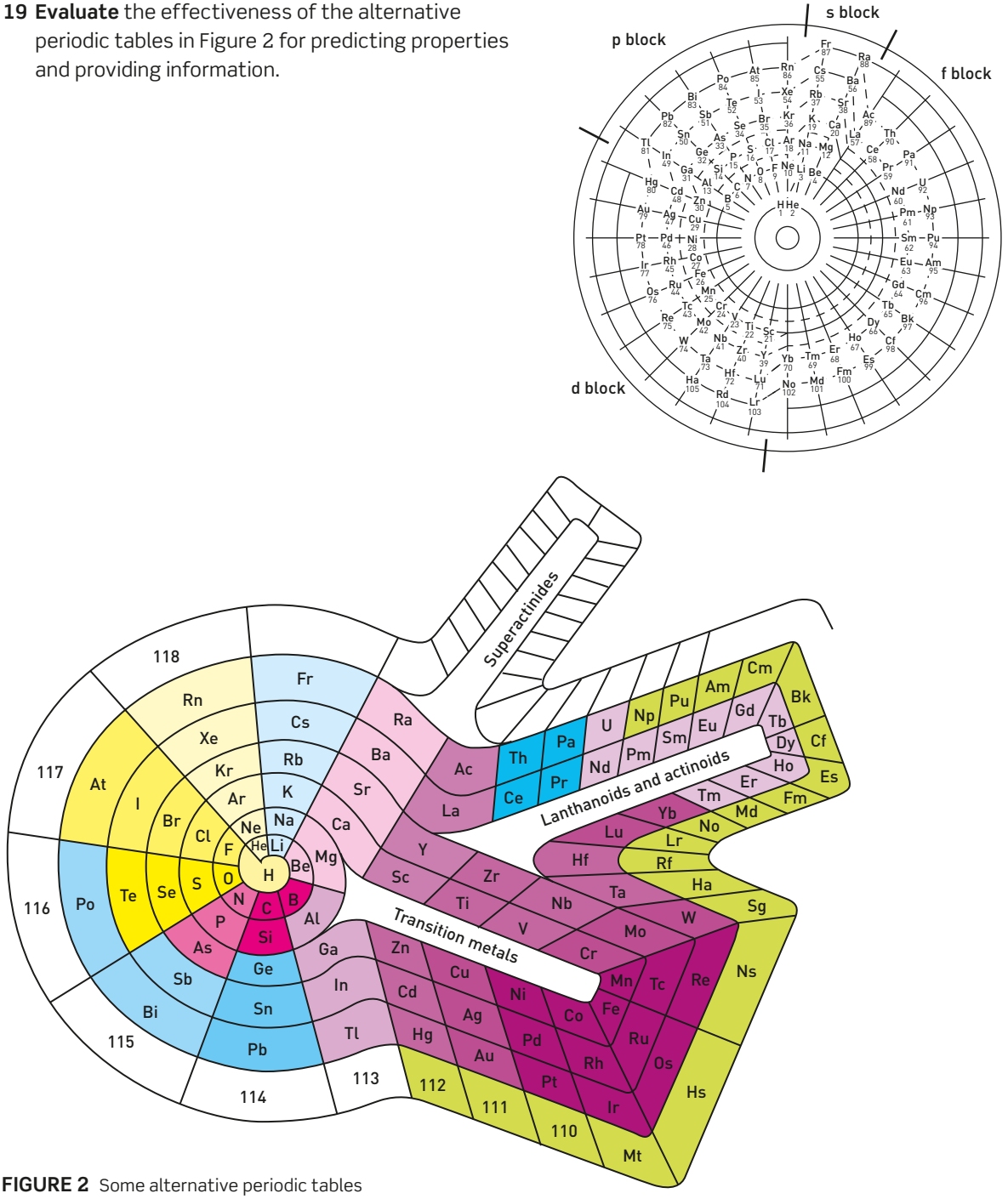

FIGURE 2 Some alternative periodic tables

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CHAPTER 3

Introduction to bonding

Atoms are the building blocks for all matter. In order to achieve stability, atoms form bonds with other atoms, and most substances are compounds rather than pure elements. Different types of bonds are involved in the formation of compounds, depending on the properties of the atoms that have been involved. Some elements, such as nitrogen and oxygen, can even bond with each other in a variety of ways to produce a variety of compounds.

OBJECTIVES

- → Recognise that the properties of atoms, including their ability to form chemical bonds, are explained by the arrangement of electrons in the atom and by the stability of the valence electron shell.
- → Understand that the number of electrons lost, gained or shared is determined by the electron configuration of the atom and recall that transition elements can form more than one ion.
- → Recognise that ions are atoms or groups of atoms that are electrically charged due to an imbalance in the number of electrons and protons and recognise that ions are represented by formulas which include the number of constituent atoms and the charge of the ion.
- → Understand that chemical bonds are caused by electrostatic attractions that arise because of the sharing or transfer of electrons between participating atoms and the valency is a measure of the number of bonds that an atom can form.
- → Determine the formula of an ionic compound from the charges on the relative ions and name the compound.
- → Deduce Lewis (electron dot) structures of molecules and ions, showing all valence electrons for up to four electron pairs for each atom.
- → Identify the numbers of bonding and lone pairs of electrons around each atom in a molecule.

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FIGURE 1 Atoms form bonds with other atoms.

3.1 Types of bonding

KEY IDEAS

In this section, you will learn about:

- electrostatic attractions causing atoms to gain, lose and share electrons with each other to form compounds and exist in a more stable state
- ionic bonding, which is the transfer of electrons from metallic cations to non-metallic anions, and the octet rule
- covalent bonding, which is the sharing of electrons between non-metals only
- polyatomic ions, which are covalently bonded molecules that are electrically charged.

valence
electrons located in an atom's outermost energy level only

ionic bonding
the process by which atoms transfer valence electrons to each other valence

cation
a metal atom that has lost one or more electrons and is positively charged

anion
a non-metal atom which has gained one or more electrons and has a negative charge

octet rule
atoms gain and lose electrons to reach eight electrons in their valence shells

The electron configuration of an atom can be separated into two basic components: the inner electrons and the **valence** (outer) electrons. Many of an element's properties result from the number of valence electrons in the element. For example, the reactivity of an element is determined by its electrostatic attraction to adjacent atoms. Its valence electrons are gained, lost or shared with these adjacent atoms in order to acquire a stable and low-energy noble gas electron configuration.

Ionic bonding

Ionic bonding is the process by which atoms transfer valence electrons to each other. The transfer takes place between an atom ready to lose one or more electrons, and an atom ready to gain one or more electrons.

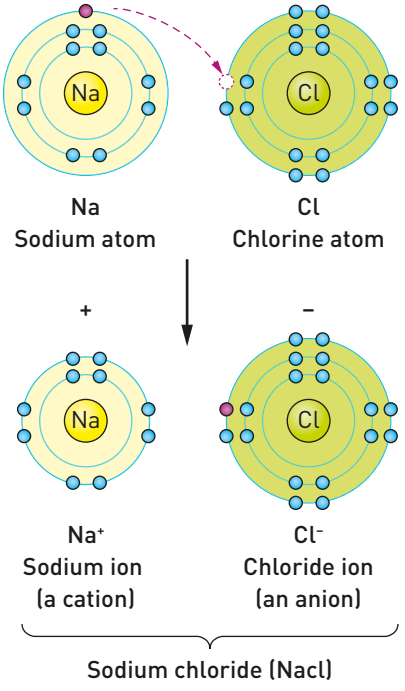

FIGURE 1 A sodium atom loses an electron to a chlorine atom to form the more stable sodium chloride.

If an atom loses one electron, it will develop an ionic charge because of the imbalance of protons and electrons. For example, sodium has 11 protons and 11 electrons. When sodium loses one electron, it develops a +1 ionic charge because it now only has 10 electrons, but 11 protons. The extra proton provides the excess positive charge. The loss of two electrons results in a +2 ionic charge, and the loss of three electrons results in a +3 ionic charge. Positive ions are called **cations**.

A gain of one electron causes a –1 ionic charge from the extra electron. For example, chlorine has 17 protons and 17 electrons; when it gains an electron to have 18 electrons, it develops a –1 ionic charge. Negative ions are called **anions**.

The octet rule

The **octet rule** refers to the tendency of C, N, O and F atoms to lose, gain or share electrons in order to achieve eight electrons in the valence energy level, which comprises the highest occupied s- and p-sublevels. In doing so, atoms acquire the same electron configuration as the nearest noble gas.

Table 1 shows how elements from the second period of the periodic table acquire the electron configuration of either helium or neon, according to the octet rule. Each element tends to either gain or lose electrons to acquire eight electrons, depending on the number of electrons it has in its valence shell and whichever process requires the least amount energy. This means lithium loses

an electron to achieve the electron configuration of helium, but oxygen gains two electrons to achieve the electron configuration of neon.

monatomic ion
an ion with only one atom

polyatomic ion
an ion with two or more atoms

TABLE 1 Determining ionic charges in the second row of the periodic table

Element	Number of valence electrons	Lose or gain electrons?	Ionic charge
Li	1 ($2s^1$)	✓ Lose 1 to be He × Gain 7 to be Ne	+1
Be	2 ($2s^2$)	✓ Lose 2 to be He × Gain 6 to be Ne	+2
B	3 ($2s^2$ and $2p^1$)	✓ Lose 3 to be He × Gain 5 to be Ne	+3
N	5 ($2s^2$ and $2p^3$)	× Lose 5 to be He ✓ Gain 3 to be Ne	−3
O	6 ($2s^2$ and $2p^4$)	× Lose 6 to be He ✓ Gain 2 to be Ne	−2
F	7 ($2s^2$ and $2p^5$)	× Lose 7 to be He ✓ Gain 1 to be Ne	−1

Study tip

Use the groups in the periodic table to remember common ionic charges. Group 1 is +1, group 2 is +2, group 13 is +3, group 15 is −3, group 16 is −2 and group 17 is −1.

Polyatomic ions

An ion is an atom or group of atoms that has an electric charge. An ion with only one atom is called a **monatomic ion**. A **polyatomic ion** contains two or more atoms (of the same or different elements). For example, OH^- is a polyatomic ion called hydroxide. Some common polyatomic ions are listed in Table 2. The naming of ions follows rules set by the International Union of Pure and Applied Chemistry (IUPAC).

TABLE 2 Common polyatomic ions

Polyatomic ion	Name of ion
NH_4^+	Ammonium
H_3O^+	Hydronium
CH_3COO^-	Ethanoate (acetate)
CO_3^{2-}	Carbonate
ClO_3^-	Chlorate
ClO_2^-	Chlorite
ClO^-	Hypochlorite
CrO_4^{2-}	Chromate
CrO_7^{2-}	Dichromate
HCO_3^-	Hydrogencarbonate
OH^-	Hydroxide
NO_3^-	Nitrate
NO_2^-	Nitrite
PO_4^{3-}	Phosphate
SO_4^{2-}	Sulfate
SO_3^{2-}	Sulfite

Ionic formulas

A monotomic ion is represented by the element symbol with the ionic charge as a superscript to the right of the element symbol.

If the ion is polyatomic, then the number of constituent atoms of each element is represented as a subscript to the right of each element symbol (but before the ionic charge superscript). However, elements with only one atom do not have a subscript of 1.

An ionic compound is often formed when electrons are transferred from one element to another element to form a cation and an anion. When representing an ionic compound as an ionic formula, write the cation first, followed by the anion (Figure 2). The ionic compound does not have a superscript ionic charge because it is overall neutral. Instead, subscripts for each of the elements in the ionic compound represent the lowest whole number ratio of the elements (Figure 3).

Mg^{2+} O^{2-}

FIGURE 2 The ionic formulas for the magnesium and oxide ions, which make up magnesium oxide.

CO_3^{2-}

FIGURE 3 The ionic formula for the carbonate ion.

WORKED EXAMPLE 3.1A

Determine the ionic formula when aluminium and oxygen bond together.

SOLUTION

Aluminium has 3 valence electrons so forms Al^{3+} when it loses the 3 electrons.

Oxygen has 6 valence electrons so forms O^{2-} when it gains 2 electrons.

Al loses 3 electrons, but O only gains 2 of them. The leftover electron goes to a second O. This O needs another electron, which is transferred by a second Al. This process keeps going until the same number of electrons is lost and gained, and both elements have lost or gained the number they require. In total, 2 Al and 3 O are required for the transfer of 6 electrons in total (Figure 4).

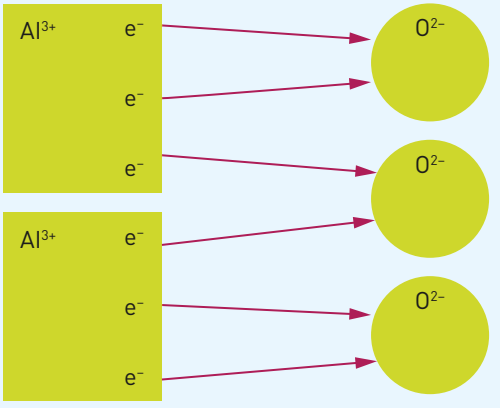

FIGURE 4 Two Al atoms and three O atoms are required for the transfer of 6 electrons in total to form aluminium oxide.

An easier way to work this out is by the swap and drop method.

The ionic charge from the cation is 'swapped' over to the anion and then 'dropped' to be a subscript. The number loses its positive or negative symbol since it is now representing the number of atoms of that element in the ionic compound, rather than an ionic charge. The same process happens for the ionic charge from the anion. It is 'swapped' to the cation and then 'dropped' to be a subscript for the cation without a positive or negative charge (Figure 5).

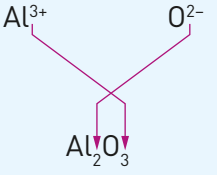

FIGURE 5 The swap and drop method for aluminium oxide.

WORKED EXAMPLE 3.1B

Determine the ionic formula when sodium and sulfate ions bond together.

SOLUTION

Sodium has 1 valence electron so becomes Na^+ when it loses the 1 electron.

Sulfate is a polyatomic ion with the formula SO_4^{2-}.

The ionic formula is Na_2SO_4 (Figure 6). (Note: We do not usually write a 1 to represent that there is one SO_4 in the ionic compound.)

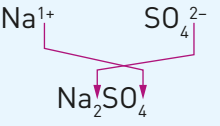

FIGURE 6 The swap and drop method for sodium sulfate.

WORKED EXAMPLE 3.1C

Determine the ionic formula when calcium and phosphate ions bond together.

SOLUTION

Calcium has 2 valence electrons so becomes Ca^{2+} when it loses the 2 electrons.

Phosphate is a polyatomic ion with the formula PO_4^{3-}. The ionic formula is $Ca_3(PO_4)_2$ (Figure 7).

(Note: Brackets have been put around the entire polyatomic ion to represent that there is 2 of everything inside the bracket – 2 phosphorus atoms and 4 × 2 = 8 oxygen atoms.)

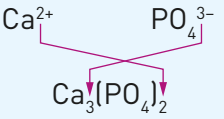

FIGURE 7 The swap and drop method for calcium phosphate

WORKED EXAMPLE 3.1D

Determine the ionic formula when calcium and sulfate bond together.

SOLUTION

Calcium has 2 valence electrons so becomes Ca^{2+} when it loses the 2 electrons.

Sulfate is a polyatomic ion with the formula SO_4^{2-}.

When a cation and anion have the same magnitude of charge (but opposite signs), the swap and drop method is not required. The ionic compound must be written as the simplest whole number ratio.

Ca^{2+} and SO_4^{2-} produce the ionic compound $CaSO_4$

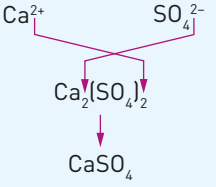

FIGURE 8 The swap and drop method for calcium sulfate

CHALLENGE 3.1A

Transition elements

Evaluate the statement 'Transition elements can form more than one ion.' to consider whether the term 'transition element' refers primarily to the transition metals, or whether it also refers to other elements in the periodic table. Use your research to list the ten elements that have the greatest number of ionic charges.

Naming ionic compounds

When naming ionic compounds, cations are listed first and anions are listed second. Metals maintain their elemental name as the cation. Polyatomic ions also maintain their name. But when the anion is monatomic, the last part of the name of the non-metals is removed and replaced with '-ide' (Table 3).

TABLE 3 Naming anions

Element name	Ionic formula	Base name	Anion name
Fluorine	F^-	Fluor	Fluoride
Chlorine	Cl^-	Chlor	Chloride
Oxygen	O^{2-}	Ox	Oxide
Sulfur	S^{2-}	Sulf	Sulfide
Nitrogen	N^{3-}	Nitr	Nitride
Phosphorus	P^{3-}	Phosph	Phosphide

WORKED EXAMPLE 3.1E

Name the following ionic compounds.

a CaO

b Al_2S_3

c Li_2CO_3

SOLUTION

a calcium oxide

b aluminium sulfide

c lithium carbonate

CASE STUDY 3.1A

Ions in our body

The human body requires a substantial amount of ions for crucial functions such as the contraction of the heart. These ions are obtained from our food. An ion deficiency due to a lack of particular elements in our food has a significant effect on our bodies.

Table 4 lists some of the more important ions we need every day. Trace amounts of other ions are also required by the body to function at capacity. These include iron (Fe^{2+}), copper (Cu^{2+}), zinc (Zn^{2+}) and iodide (I^-).

TABLE 4 Important ions in the human diet

Ion name	Ionic formula	Dietary sources	Associated functions	Effects of deficiency
Sodium	Na^+	Salt, additives, meat, fish, dairy, eggs, olives, pickled and smoked foods	Regulation and control of body fluid levels	Diarrhoea, anxiety, cardiac arrest (An excess causes hypertension, which can lead to stroke, heart attack, and kidney failure)
Potassium	K^+	Whole grains, meat, legumes, fruit and vegetables	Water balance in cells, muscle and nerve function, protein synthesis	Water imbalance, irregular heartbeat, cardiac arrest, tissue breakdown
Calcium	Ca^{2+}	Milk, leafy green vegetables, whole grains, egg yolk, legumes and nuts	Formation of bones and teeth, blood clotting, muscle function, transmission of nerve signals to cells	Rickets, osteoporosis, stunted formation of teeth
Magnesium	Mg^{2+}	Dairy, flour and cereal, dry beans, nuts, peas and leafy green vegetables	Nerve impulses, muscle contraction, enzyme activation	Tremors in muscles, heart spasms, convulsions
Chloride	Cl^-	Salt, seaweed, rye, tomatoes, lettuce, celery and olives	Balance of body fluid levels	Heavy sweating, chronic diarrhoea, vomiting
Phosphate derivatives	$H_2PO_4^-$ and HPO_4^{2-}	Dairy, cereal products and meat	Bone formation and energy production	Low energy, loss of appetite, numbing, burning or tingling extremities, bone pain, joint pain, swelling and stiffness

Covalent bonding

Covalent bonding occurs when atoms share electrons to form bonds. Covalent bonds typically form between two or more non-metals, and are influenced by the electronegativity values of each of the atoms involved.

covalent bonding
a sharing of electrons between atoms

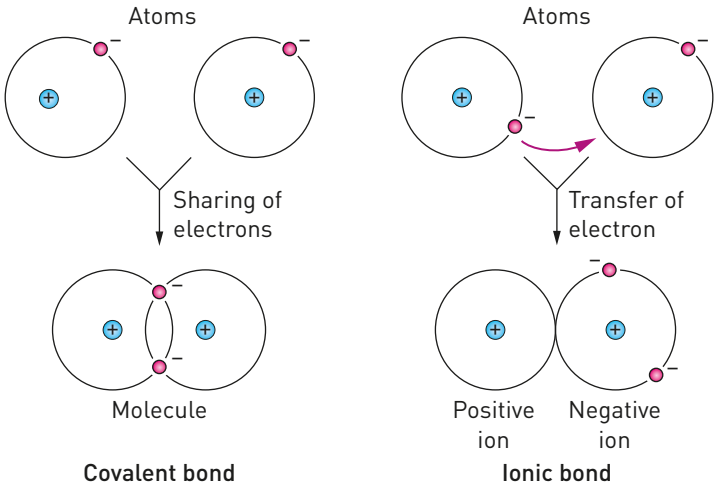

FIGURE 9 Covalent bonding and ionic bonding.

The electrons involved in covalent bonding are considered to belong to both of the atoms sharing the electrons. When investigating whether a molecule has followed the octet rule for each of its atoms, these shared electrons are counted twice for each atom.

WORKED EXAMPLE 3.1F

Determine the number of valence electrons each atom contributes to the covalent molecule CF_4 (Figure 10).

FIGURE 10 Tetrafluromethane (CF_4)

SOLUTION

Each carbon atom has 4 valence electrons. Each fluorine atom has 7 valence electrons.

Each bond contains 2 electrons, 1 from the carbon and 1 from the fluorine.

Carbon has 4 bonds, so it is sharing 8 electrons (and following the octet rule). 4 of these electrons are originally from the carbon atom, and the other 4 electrons are from the surrounding fluorine atoms.

Since each fluorine has 1 bond, it has shared one of its electrons and 1 of the carbon's electrons. Each fluorine still has its other 6 electrons: 6 + 2 = 8. Therefore, each of the 4 fluorine atoms follows the octet rule.

Polyatomic ions are electrically charged groups of atoms that are held together by covalent bonding. For this reason, a compound can contain both ionic and covalent bonding if there is a polyatomic ion present. For example, in calcium carbonate ($CaCO_3$), the three oxygen atoms each have covalent bonds to the central carbon atom, forming the carbonate polyatomic ion. The Ca^{2+} ion supplies the positive charges to make a compound that is overall electrically neutral.

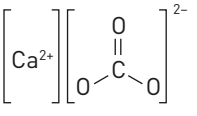

FIGURE 11 Calcium carbonate ($CaCO_3$).

Study tip

An ionic compound is made up of a metal and one or more non-metals. Covalent compounds have two or more non-metals. Make sure you check whether a metal is present before naming the compound.

CHALLENGE 3.1B

Naming covalent compounds

The naming of covalent compounds is slightly different from naming ionic compounds because the ratio of each element within the compound cannot be determined from the ionic charges. For this reason, the ratio forms part of the name. For example, N_2O is dinitrogen monoxide. Note that if the first element is singular, then the quantity is not referred to in the compound's name. For example, CO_2 is carbon dioxide (not monocarbon dioxide).

Use Table 5 to name the covalent compounds CCl_4, N_2Cl_4, CBr_4, SF_6 and P_2O_4.

TABLE 5 Naming covalent compounds

Number of atoms	Prefix used in name	Number of atoms	Prefix used in name
1	Mono	6	Hexa
2	Di	7	Hepta
3	Tri	8	Octa
4	Tetra	9	Nona
5	Penta	10	Deca

CASE STUDY 3.1B

Naming chemicals

Scientific names for chemicals are widely used by the scientific community and within educational institutions, but many chemicals that are used by the general public also have 'common' names. Most of the time, these common names reflect the physical properties of the chemical or its original discovery. Table 6 lists examples of chemicals that contain both scientific and common names.

Pharmaceutical chemicals are often referred to by their brand name. For example, ibuprofen is often called by its brand names Nurofen, Advil or Motrin, but its scientific name is 2-(4-(2-methylpropyl)phenyl)propanoic acid. Paracetamol (also commonly named acetaminophen) is often referred to by its brand names Panadol, Tylenol or Calpol but its scientific name is *N*-(4-hydroxyphenyl) ethanamide.

TABLE 6 Scientific and common names of some chemicals

Common name	Formula	Scientific name
Lime	CaO	Calcium oxide
Lye (caustic soda)	$NaOH$	Sodium hydroxide
Potash	K_2CO_3	Potassium carbonate
Table salt	$NaCl$	Sodium chloride
Laughing gas	N_2O	Dinitrogen monoxide
Baking soda	$NaHCO_3$	Sodium hydrogen carbonate
Muriatic acid	HCl	Hydrochloric acid
Hypo	$Na_2S_2O_3$	Sodium thiosulfate
Pool sunscreen	$C_3N_3O_3H_3$	Cyanuric acid

Common name	Formula	Scientific name
Vinegar	$C_2H_4O_2$	Ethanoic acid
Gypsum	$CaSO_4$	Calcium sulfate dihydrate
Quicksilver	Hg	Mercury
Limestone	$CaCO_3$	Calcium carbonate
Milk of magnesia	$Mg(OH)_2$	Magnesium hydroxide

CHECK YOUR LEARNING 3.1

Describe and explain

1 **Describe** the terms 'inner electrons' and 'valence electrons'.

2 **Describe** ionic bonding.

3 **Explain** how bonding is associated with the stability of the valence sublevels.

Apply, analyse and interpret

4 **Determine** the compound names for:

- **a** FrCl
- **b** NH_3
- **c** $Ca(NO_2)_2$.

5 Use the periodic table to **construct** the rest of Table 7.

6 **Analyse** the graph in Figure 12 to identify the relationship between bond length and bond energy.

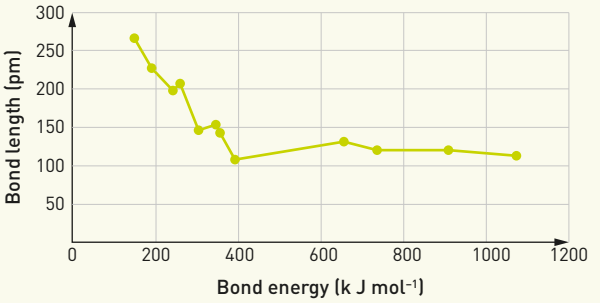

FIGURE 12 Bond length and bond energy of common covalent bonds.

7 Use the ionic formula $Cr_2O_7^{2-}$ as an example to **infer** what the subscripts and superscripts represent.

TABLE 7 Elements and their ions

Element	Group	Number of valence electrons	Number of electrons lost or gained	Ionic charge formed	Anion or cation
Rubidium					
		13			
Nitrogen					
		20			

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  Check your learning 3.1
- **» Challenge**
  3.1A Transition elements
- **» Challenge**
  3.1B Naming covalent compounds
- **» Weblink**
  IUPAC chemicals

3.2 Lewis structures

KEY IDEAS

In this section, you will learn about:

- how Lewis structures use element symbols and electron dots to represent an atom
- electron dots, which can be either singular bonding electrons or a lone pair of non-bonding electrons
- bonds comprising two electrons
- Lewis structures of ions, which take into account the ionic charge by either adding or removing electrons from the total amount used.

Lewis structure
visual diagrams to show the number of electrons in the valence shell

electron dot
represent the number of valence electrons in a Lewis structure

bonding electrons
valence electrons that are bonded to another atom

lone pair of electrons
valence electrons that are not bonded to another atom

double bond
a bond in which two atoms share four electrons

triple bond
a bond in which a bond in which two atoms share six electrons

Lewis structures (or **electron dot** structures) are visual diagrams that represent the valence electrons (from the highest s- and p-sublevels) present in atoms, ions and molecules. Electrons are drawn as dots surrounding the atom's elemental symbol, which represents the inner electrons that do not need to be individually shown because they do not take part in bonding.

FIGURE 1 The Lewis structure of a krypton atom.

Electron dots are added one at a time until the four sides of the element symbol are filled. The second dot for each side is then added, until all the atom's valence electrons are included. The exact location of the dots surrounding the element symbol is irrelevant, except for helium. Helium does not follow the octet rule because it has a full valence 1s sublevel of two electrons, so the two electrons are drawn together on one side.

TABLE 1 Lewis structures of helium and period 2 of the periodic table

Group 1	Group 2	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
							He:
Li•	•Be•	•B•	•Ċ•	•N•	•Ö:	•F:	:Ne:

Lewis structures of covalent molecules

The electrons in Lewis structures of covalent molecules are categorised as either singular **bonding electrons** or **lone pairs** of (non-bonding) electrons. To create a Lewis structure of a molecule, there are a several steps involved.

1. Count the total number of electrons in the molecule, from each of the atoms present.
2. Place the atoms' element symbols together, starting in the centre with the atom that has the most bonding electrons. Surround the central atom with the other atoms. The least electronegative atom is usually the central atom.
3. Draw a single line to connect the central atom with each of the surrounding atoms, to represent bonds. Draw two electron dots on each of these bonding lines.
4. Draw the lone pairs of electrons on the surrounding atoms to complete the octet.
5. Draw any remaining electrons on the central atom. If there is not enough electrons to finish the central atom's octet, then multiple bonds need to be formed: **double** or **triple bonds**.

Study tip

Place the bonding electrons of the surrounding atoms on the side of the element symbol nearest to the central atom. This will make the Lewis structure easier to draw.

Coordinate covalent bonds

A **coordinate covalent bond** is formed when the two electrons in the covalent bond have been sourced from only one of the atoms involved. For example, carbon monoxide (CO) has a triple bond between carbon and oxygen with six shared electrons – four from oxygen and two from carbon. Oxygen donates its lone pair (which would otherwise be non-bonding) to the triple bond. A coordinate covalent bond is represented by an arrow instead of a standard bond line (Figure 2).

FIGURE 2 In carbon monoxide, oxygen donates more electrons to the triple bond than carbon does.

coordinate covalent bond two electrons come from only one atom in a covalent bond

WORKED EXAMPLE 3.2A

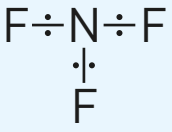
Determine the Lewis structure of NF_3.

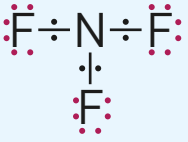
SOLUTION

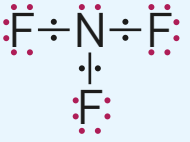
1 Each fluorine has seven valence electrons. Nitrogen has five valence electrons.
(3 × 7) + 5 = 26 electrons

2 Nitrogen has one lone pair of electrons and three bonding electrons. Fluorine has three lone pairs of electrons and one bonding electron. Therefore, nitrogen is the central atom.

F N F
F

3 

4 

5 

Study tip

Hydrogen's single electron is located in the 1s shell. It does not follow the octet rule because a full energy level will be 2 electrons.

WORKED EXAMPLE 3.2B

Determine the Lewis structure of HCN.

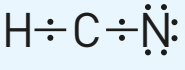
SOLUTION

1 Hydrogen has one valence electron, carbon has four and nitrogen has five.
1 + 4 + 5 = 10 electrons

2 Hydrogen has one bonding electron, nitrogen has three and carbon has four. Carbon is the central atom.

H C N

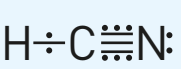
3

4 

Study tip

Electrons can be represented by symbols other than dots if their origin is confusing. For example, the Lewis structure of carbon monoxide could have dots for the carbon electrons and crosses for the oxygen electrons.

5 All 10 electrons have been allocated but carbon requires another 4 electrons to complete the octet. Each bond has 2 electrons, so a deficit of 4 electrons means either a triple bond or two double bonds need to be formed. Since hydrogen can only reach a doublet (rather than an octet), it cannot form multiple bonds. Therefore, there needs to be a triple bond between carbon and nitrogen.



This molecule is not considered to have coordinate covalent bonds because the carbon originally had 4 bonding electrons and nitrogen originally had 3 bonding electrons, so no donated sharing took place.

Lewis structures of ions

The difference between molecules and ions when drawing Lewis structures is the first step of the process. The ionic charge of monatomic and polyatomic ions requires the addition or subtraction of electrons from the total.

For example, an ionic charge of –3 means the total electrons is the atom's electron plus 3. An ionic charge of +1 means subtracting 1 electron from the atom's total number.

WORKED EXAMPLE 3.2C

Determine the Lewis structure of the monatomic ion oxide (O^{2-}).

SOLUTION

Oxygen has 6 valence electrons and a 2- charge, so add 2 electrons to give a total of 8 electrons.

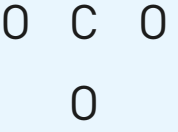
WORKED EXAMPLE 3.2D

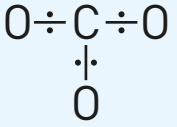
Determine the Lewis structure of carbonate (CO_3^{2-}).

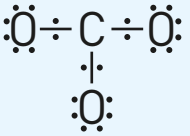
SOLUTION

1 Carbon has 4 valence electrons and each oxygen has 6. The ionic charge is 2–.
4 + (3 × 6) + 2 = 24 valence electrons

2 Carbon has 4 bonding electrons, and each oxygen has 2. Carbon is the central atom.

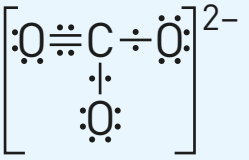


3 

4 

5 24 electrons have already been used, so a double bond is required to achieve an octet for carbon.

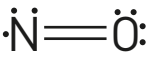
6 Square brackets are used to represent that the polyatomic ion has an overall ionic charge of 2–.



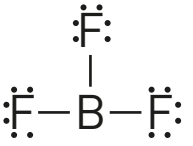
Exceptions to the octet rule

There are three main exceptions to the octet rule.

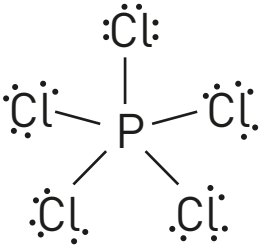
1 Molecules with an odd number of electrons that cannot be completely paired – these molecules are highly reactive, unstable radicals that react to fix the odd number of electrons. An example is NO.



2 Molecules with a central atom that has less than an octet that cannot form multiple bonds – for example, the fluorine atoms in BF_3 are so electronegative that they do not form a coordinate covalent bond.



3 Molecules with a central atom with more than an octet because the central atom (often from the third period and beyond) can accept additional electrons and it is bonded to highly electronegative surrounding atoms. An example is PCl_5.



CHECK YOUR LEARNING 3.2

Describe and explain

1 **Describe** what a Lewis structure represents by the element symbol and surrounding dots.

2 **Explain** why helium is an exception to the octet rule.

Apply, analyse and interpret

3 **Determine** the Lewis structure of:

- **a** phosphorus
- **b** strontium
- **c** boron.

Investigate, evaluate and communicate

4 Methane is a covalent molecule with the formula CH_4. Due to repulsion between the negatively charged electrons, the angles between the bonds are greater than 90°. **Evaluate** the effectiveness of the following two diagrams for their ability to accurately represent the three-dimensional model of methane.

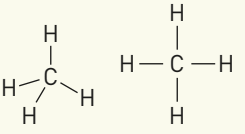


FIGURE 3 Two representations of methane.

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- » Video weblink
  Lewis structures
- » Weblink
  Lewis structures of molecules
- » Weblink
  Lewis structures of ions

CHAPTER 3

Review

Chapter summary

3.1
- When a metal atom is electrostatically attracted to a non-metal atom, a transfer of electrons takes place to form an ionic bond. The quantity of electrons lost by the metal usually equals the number of valence electrons it has. The quantity of electrons gained by the non-metal usually equals the difference between its valence electrons and 8, because it tends to react to achieve an octet for stability.
- The formulas of ionic compounds represent the ratio of atoms, and are named using the original metallic element name and a modified version of the non-metallic element name that ends with 'ide'.
- When non-metals are attracted to each other, they share electrons to form covalent bonds.

3.2
- Lewis structures use element symbols to represent the inner electrons of an atom and electron dots drawn around the element symbol to represent the valence electrons of an atom for the highest s- and p-sublevels that are occupied.
- Electrons can be drawn as singular bonding electrons to be shared by atoms in covalent bonds that require two electrons in total (one from each of the atoms). Electrons can also be drawn as lone pairs, which are considered non-bonding.
- Covalent molecules typically have single bonds and each of the atoms involved contributes half of the electrons required for bonding. There are exceptions when molecules require double or triple bonds, or when one of the atoms must donate both of the electrons of the covalent bond.
- Lewis structures of ions include the valence electrons of the atom(s) involved as well as the electrons accrued from the ionic charge (a negative charge means additional electrons and a positive charge means fewer electrons).

Key terms

- anion
- bonding electrons
- cation
- coordinate covalent bond
- covalent bonding
- electron dot
- double bond
- ionic bonding
- Lewis structure
- lone pair of electrons
- monatomic ion
- non-metal
- octet rule
- polyatomic ion
- triple bond
- valence

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; *** = high.

Multiple choice

1 Ionic bonding is characterised as:
- **A** sharing of electrons between a metal and non-metal
- **B** transfer of electrons between a metal and non-metal
- **C** transfer of electrons between two non-metals
- **D** transfer of electrons between two metals.

2 An ion is:
- **A** only an atom that is positively charged
- **B** an atom or group of atoms that are electrically charged
- **C** a group of atoms
- **D** an atom or group of atoms that are neutral.

3 How many times will an element with six valence electrons covalently bond?
- **A** once
- **B** twice
- **C** three times
- **D** four times

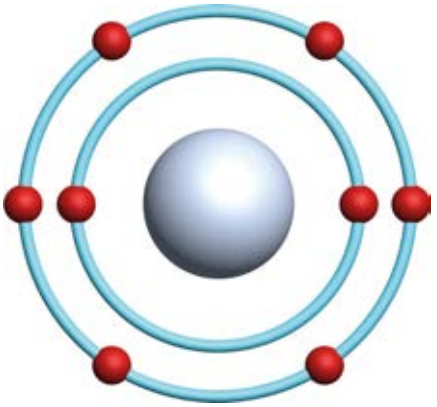


FIGURE 1 Oxygen atom with 6 valence electrons.

4 Nitrate is characterised as a polyatomic ion that contains:
- **A** one nitrogen and one oxygen
- **B** one nitrogen and two oxygen atoms
- **C** one nitrogen and three oxygen atoms
- **D** one nitrogen and four oxygen atoms.

5 BCl_3 is a compound called:
- **A** boron chloride
- **B** boron chlorine
- **C** boron trichloride
- **D** monoboron trichloride.

Short answer

Describe and explain

★ **6** **Describe** covalent bonding.

★ **7** **Describe** what sublevels a Lewis structure represents.

★ **8** **Explain** how the properties of atoms are determined by their electron arrangement.

★ **9** **Explain** the bonding involved for the electrons present in barium hydroxide.

★ **10** **Explain** the bonding involved for the electrons present in sodium thiosulfate ($Na_2 S_2 O_2$).

★ **11** **Explain** why an ion has an ionic charge.

Apply, analyse and interpret

★★ **12** **Determine** the ionic formulas for:
- **a** potassium sulfide
- **b** radium phosphide
- **c** aluminium nitride.
- **d** strontium chlorite
- **e** boron phosphate
- **f** ammonium sulfite.

★★ **13** Use the periodic table to **construct** the rest of Table 1. A larger version of the periodic table can be found in Appendix 1.

TABLE 1 Elements and their electrons

Group	Element	Number of valence electrons	Number of lone pairs of electrons	Number of bonding electrons
	Chlorine			
		16		
	Silicon			
		9		

★★ **14** **Determine** the Lewis structures of the following ions.
- **a** bromide
- **b** nitride
- **c** carbonate

★★ **15 Determine** the Lewis structure of:

a SCl_2

b CCl_2F_2

c $COBr_2$.

★★ **16 Determine** the Lewis structures of the following polyatomic ions.

a BF_4^-

b H_3O^+

c ClO_2

★★ **17 Determine** the ionic formula of:

a boron iodide

b magnesium oxide

c lithium phosphite.

★★ **18 Determine** the compound name for:

a N_2O

b BCl_3

c $(NH_4)_2CrO_4$.

★★ **19 Determine** the Lewis structure of:

a silicon

b caesium

c selenium.

★★ **20 Determine** the Lewis structure of the following monatomic ions.

a nitride

b sulfide

c iodide

★★ **21 Determine** the Lewis structure of the following molecules.

a N_2O_4

b SO_3

c PH_3

★★ **22 Determine** the Lewis structure of the following polyatomic ions.

a NH_4^+

b SO_4^{2-}

c ClO_3^-

★★★ **23 Consider** Table 2, which lists bond energies and bond lengths of some covalent bonds.

a **Determine** the relationship between types of bonds and their associated bond lengths and energies.

b **Determine** the relationship between types of elements and their associated bond lengths and energies.

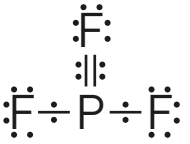
TABLE 2 Covalent bond energies and lengths

Covalent bond	Bond energy (kJ mol^{-1})	Bond length (pm)
C-O	356	143
C=O	736	121
C≡O	1074	113
C-C	347	154
C=C	657	133
C≡C	908	121
Cl–Cl	243	199
Br–Br	192	228
I–I	151	267

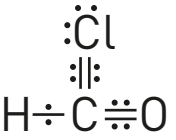
Investigate, evaluate and communicate

★★ **24 Evaluate** the following Lewis structures to identify the errors. Provide a correct representation of each molecule.

a

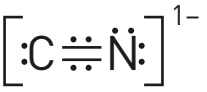


b

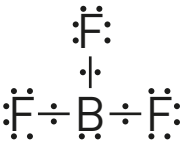


★★ **25 Evaluate** the Lewis structures provided to identify the errors. Provide a correct representation of each molecule.

a



b



★★★ **26** A polyatomic ion contains two central carbon atoms. The first is double-bonded to an oxygen that has two lone pairs of electrons. The same carbon is also bonded to a hydrogen. There is a single bond between the central carbon atoms. The second carbon is bonded to two hydrogen atoms as well as an oxygen (which has three lone pairs of electrons). **Decide** the polyatomic ion's name and represent it as a

Lewis structure and an ionic formula, complete with the charge.

★★★ **27** Electronegativity values determine the location of electrons in bonds. Electrons are closer to the more electronegative atom and further from the more electropositive atom. The difference between the atoms' electronegativities determines the type of bond or electrostatic attraction between the two electrons.

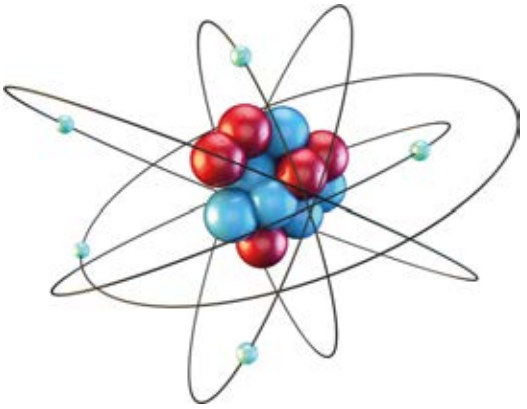


FIGURE 2 Boron has relatively close electrons as it has a high electronegativity.

A difference in electronegativity of 0–0.5 is considered a non-polar covalent bond because the electrons are equally shared by the atoms – the electrons are placed in the middle of the bond. A difference in electronegativity of 0.5–1.5 is considered a polar covalent bond because the electrons are not shared equally by the atoms – the electrons are closer to one atom. A difference in electronegativity of >2 is considered an ionic bond because the electrons have actually transferred completely to one of the atoms.

a Use the electronegativity values in Table 3 to **determine** the location of electrons in Lewis structures of Cl_2, H_2O and SiH_4. Pay particular attention to the placement of electron dots on the bonds.

TABLE 3 Electronegativity of elements

Element	Cl	O	H	Si
Electronegativity	3.2	3.4	2.2	1.9

b Ionic bonding takes place between a metal and non-metal, and covalent bonding takes place between two or more non-metals.

FIGURE 3 Copper is a metal and sulfur is a non-metal.

Evaluate the significance of not assigning a difference in electronegativity of 1.6–1.9 as either ionic or covalent. Use HF and NaBr as examples.

c **Investigate** the effect of polar covalent versus ionic bonds on a compound's properties.

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CHAPTER 4

Isotopes

The word 'isotope' comes from the Greek words *isos* (meaning 'equal') and *topos* (meaning 'place'). Therefore, the word literally translates to 'equal place', referring to how isotopes of an element have the same position in the periodic table. Although the concept of isotopes was first realised by radiochemist Frederick Soddy, the term to describe the concept was suggested by his physicist friend Margaret Todd.

An atom is considered to be carbon when it has six protons. The same atom is considered to be neutral when it has six electrons. Atoms can have different number of neutrons and these are called isotopes. Most carbon atoms (about 99%) have six neutrons, but about 1% have seven neutrons.

In this chapter, you will learn more about isotopes as atoms of the same element that have different numbers of neutrons.

OBJECTIVES

→ Recall isotopes are atoms of the same element that have different numbers of neutrons and can be represented in the form ^{A}X (IUPAC) or X-*A*.

→ Recognise that isotopes of an element have the same electron configuration and possess similar chemical properties but have different physical properties.

→ Understand that the relative atomic mass of an element is the ratio of the weighted average mass per atom of the naturally occurring form of the element to $^{1}/_{12}$ the mass of an atom of carbon-12.

→ Use appropriate mathematical representations to make inferences and to solve problems, including calculating the relative atomic mass of an element and percentage abundances of the isotopes of an element from data.

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FIGURE 1 Some isotopes are thought to have been created in the Big Bang, while others were synthesised in stars or supernovae; this image shows the Crab Nebula, a remnant of an exploded supernova.

4.1 Isotopes

KEY IDEAS

In this section, you will learn about:

- isotopes that are atoms of the same element but different numbers of neutrons
- how isotopes have similar chemical properties but different physical properties
- the relative atomic mass (used in the periodic table), which is the ratio of the weighted average of the atomic masses of all the naturally occurring isotopes of an element (taking into account their abundances on Earth) to one-twelfth the atomic mass of the ^{12}C isotope.

isotope
atoms of an element that have the same number of protons but different numbers of neutrons

Isotopes are atoms of an element that have the same number of protons but different numbers of neutrons. As a result, all isotopes of an element have the same atomic number but different mass numbers.

For example, hydrogen has three naturally occurring isotopes: protium, deuterium and tritium. Every hydrogen atom contains one proton, which is why the atomic number is 1. The mass number varies between 1 and 3 because the number of neutrons contained in a hydrogen atom varies between 0 and 2 (Figure 1).

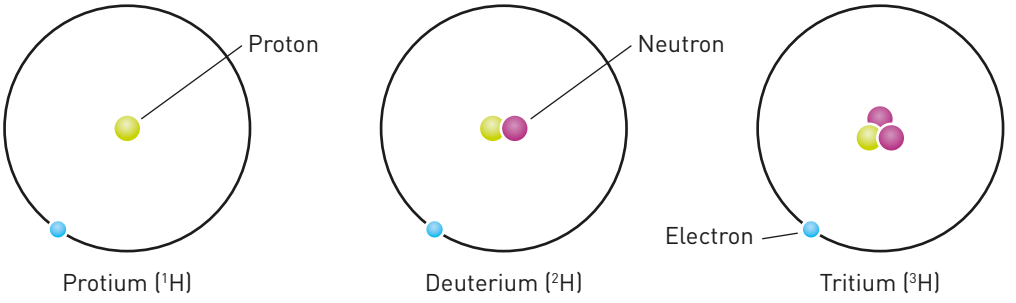

FIGURE 1 The three isotopes of hydrogen.

Isotopes can be identified by using two forms of representation (Table 1): ^{A}X or X-*A*. In both methods:

- *A* represents the mass number
- X represents the elemental symbol (or elemental name) used to identify the number of protons contained in the atom.

TABLE 1 Representations of selected isotopes

^{A}X	X-A
^{1}H	H-1 (hydrogen-1)
^{12}C	C-12 (carbon-12)
^{6}Li	Li-6 (lithium-6)

Study tip

The unit used for atomic mass is amu, which stands for atomic mass unit and refers to one-twelfth of the atomic mass of the ^{12}C isotope. Relative atomic mass does not have a unit because it is a ratio.

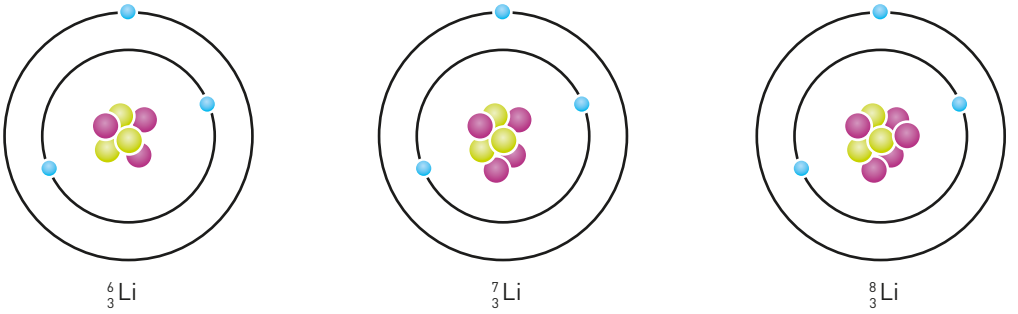

FIGURE 2 The number of neutrons differs between lithium isotopes.

WORKED EXAMPLE 4.1A

Represent the two naturally occurring isotopes of boron. Their atomic masses are 10.012938 amu and 11.009305 amu.

SOLUTION

The atomic masses rounded to the nearest whole numbers give you boron's isotopes.

The first boron isotope is ^{10}B or boron-10.

The second boron isotope is ^{11}B or boron-11.

Properties of isotopes

Isotopes of the same element have the same electron configuration when neutrally charged. For this reason, the chemical properties of isotopes of an element are similar.

However, physical properties are determined by the mass number. The varying number of neutrons causes isotopes to have different physical properties. For example, the physical properties of water change depending on which isotope of hydrogen is contained in the water (Table 2). H_2O contains two atoms of ^1H, and D_2O contains two atoms of ^2H (where D represents deuterium).

FIGURE 2 The physical properties of water change depending on the type of isotope it contains.

TABLE 2 Physical properties of different types of water

Physical property	D_2O (heavy water)	H_2O (light water)
Melting point (K)	276.97	273.15
Boiling point (K)	374.55	373.15
Density (g mL^{-1})	1.1056	0.9982
Dynamic viscosity (mPa s)	1.2467	1.0016
Surface tension (N m^{-1})	0.07187	0.07198
pH	7.44	7.0
Refractive index (at 589 µm)	1.32844	1.33335

Relative atomic mass

Most elements have two or more naturally occurring isotopes, yet only one number is included in the periodic table; this number is based on an average of the atomic masses of all the naturally occurring isotopes for an element. This number is called **relative atomic mass (RAM)** because it has been quantified relative to one-twelfth the atomic mass of the carbon-12 isotope.

relative atomic mass (RAM) the mass of an atom once the mass and percentage abundance of all isotopes have been considered, relative to the carbon-12 isotope

percentage abundance
the percentage of each isotope contained within a sample of element; it must always add up to 100%

Calculating the relative atomic mass of an element uses the atomic masses of all of the naturally occurring isotopes of the element and the natural abundance (usually represented as a **percentage abundance**) of each of the isotopes:

$$\text{Relative atomic mass} = \frac{(\%_1 \times \text{AM}_1)}{100} + \frac{(\%_2 \times \text{AM}_2)}{100} + \frac{(\%_3 \times \text{AM}_3)}{100} \ldots$$

where $\%_1$, $\%_2$ and $\%_3$, are the percentage abundances of isotopes 1, 2 and 3, and AM_1, AM_2 and AM_3, are the atomic masses of isotopes 1, 2 and 3. For elements with more than three isotopes, extra terms are needed. Synthetically prepared isotopes are not used in calculating an element's relative atomic mass.

Study tip

The term 'mass number' describes the sum of the number of protons and neutrons in an isotope. The term 'relative atomic mass' also describes the number of protons and neutrons but as an average number that considers the isotopes (and their natural abundances) of an element, and relative atomic mass appears in the periodic table.

WORKED EXAMPLE 4.1B

Calculate the RAM of chlorine using its two naturally occurring isotopes.

^{35}Cl Atomic mass = 34.968853 amu Abundance = 75.77%

^{37}Cl Atomic mass = 36.965903 amu Abundance = 24.23%

SOLUTION

$$\text{Relative atomic mass} = \frac{(\%_1 \times \text{AM}_1)}{100} + \frac{(\%_2 \times \text{AM}_2)}{100}$$

$$= \frac{(75.77 \times 34.968853)}{100} + \frac{(24.23 \times 36.965903)}{100}$$

$$= 26.4958999181 + 8.9568382969$$

$$= 35.45$$

Note: RAM does not have units because it is expressed as a ratio to one-twelfth of the atomic mass of ^{12}C.

WORKED EXAMPLE 4.1C

Calculate the percentage abundances of the isotopes of lithium using a RAM of 6.941. ^{6}Li has an atomic mass of 6.015 amu, and ^{7}Li has an atomic mass of 7.016 amu.

SOLUTION

$$\text{Relative atomic mass} = \frac{(\%_1 \times \text{AM}_1)}{100} + \frac{(\%_2 \times \text{AM}_2)}{100}$$

Let the abundance of ^{6}Li be X, and the abundance of ^{7}Li be $100 - X$.

$$6.941 = \frac{(X \times 6.015)}{100} + \frac{(100 - X) \times 7.016)}{100}$$

$$6.941 = 0.06015X + 7.016 - 0.07016X$$

$$0.01001X = 0.075$$

X = 7.49% and $100 - X$

= 92.51%

CHALLENGE 4.1

Carbon isotopes and rock art

Research how carbon isotopes can be used to decipher the age of ancient Aboriginal and Torres Strait Islander Peoples' rock art.

FIGURE 3 Rock art depicting fish, Nourlangie, Northern Territory, Australia

CHECK YOUR LEARNING 4.1

Describe and explain

1 **Describe** the term 'isotope' in relation to the number of protons, electrons and neutrons.

2 **Explain** how isotopes have the same chemical properties but different physical properties.

3 **Calculate** the relative atomic mass of lead from the information of the naturally occurring isotopes of lead listed in Table 3.

TABLE 3 Atomic masses and percentage abundances of the naturally occurring isotopes of lead

Isotopes of lead	Atomic mass (amu)	Percentage abundance (%)
^{204}Pb	203.973037	1.4
^{206}Pb	205.974455	24.1
^{207}Pb	206.975885	22.1
^{208}Pb	207.976641	52.4

4 **Calculate** the percentage abundances of the isotopes of thallium (RAM 204.4). The atomic masses of the isotopes are ^{203}Tl 202.972 amu and ^{205}Tl 204.974 amu.

Apply, analyse and interpret

5 **Analyse** the physical properties of H_2O compared with those of D_2O (Table 2).

6 **Apply** the two isotopes of copper as ^{A}X (elemental symbol) and X-*A* (element name) with atomic masses of 62.929599 amu and 64.927792 amu.

Investigate, evaluate and communicate

7 **Infer** the approximate percentage abundances of the following isotopes, by comparing the atomic masses of the isotopes to the relative atomic masses (Table 4). **Justify** your answer for each element.

TABLE 4 Relative atomic masses of selected elements and atomic masses of their respective isotopes

Element	RAM	Atomic mass of isotope 1	Atomic mass of isotope 2
Helium	4.0	3.016	4.002
Carbon	12.0	12.000	13.003
Nitrogen	14.0	14.003	15.000
Bromine	79.9	78.918	80.916

8 **Predict** the chemical properties of heavy water (D_2O).

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  4.1 Carbon isotopes and rock art
- » Weblink
  Isotopes in the real world
- » Weblink
  Carbon dating

SCIENCE AS A HUMAN ENDEAVOUR

4.2 Radioisotopes

KEY IDEAS

In this section, you will learn about:

- how radioisotopes undergo nuclear reactions to become more stable isotopes
- nuclear reactions that release radiation energy as a by-product, which can be dangerous to substances in the surrounding area.

radioisotope
an unstable isotope that undergoes radioactive decay

nuclear reaction
a process used by radioisotopes to become more stable

radiation energy
energy that is a by-product of nuclear reactions

Radioisotopes are isotopes that have unstable nuclei. Often synthesised under laboratory conditions, radioisotopes take part in **nuclear reactions** to increase their isotopic stability by changing the ratio of protons to neutrons in their nuclei and releasing large amounts of **radiation energy** in the process. A ratio of protons to neutrons of 1:1 is the most stable (Figure 1), as demonstrated by the commonly found stable isotopes $^{12}_{6}C$ (Figure 2), $^{14}_{7}N$ and $^{16}_{8}O$.

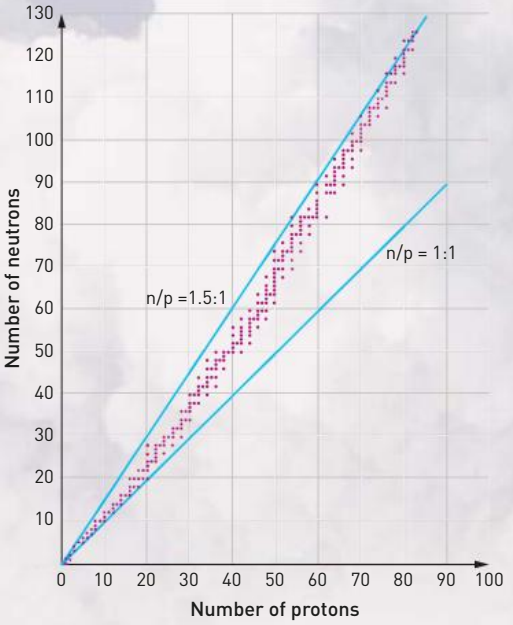

FIGURE 1 The band of stability for naturally-occuring isotopic nuclei.

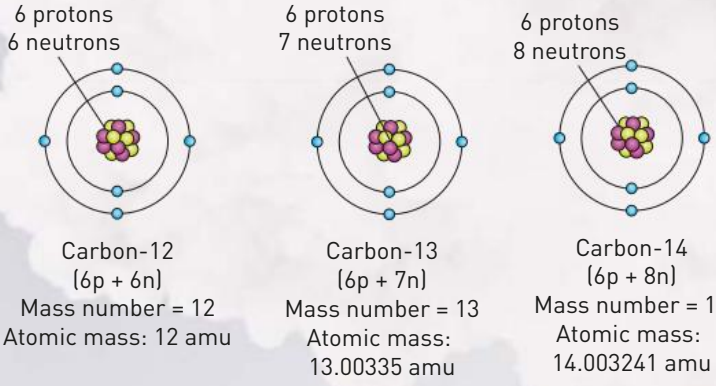

FIGURE 2 Isotopes of carbon. Carbon-12 and carbon-13 are both stable isotopes, whereas carbon-14 is an unstable isotope.

Types of radiation energy

The radiation energy released during the process of radioactive decay can be categorised as alpha, beta or gamma radiation.

- Alpha radiation (α) is positively charged because it consists of $^{4}_{2}He$ particles. Alpha radiation penetrates soft tissue but is easily stopped by a sheet of paper or dead cells on the surface of skin. Alpha radiation is dangerous when ingested.

  For example: $^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He$

  (initial isotope) → (new isotope) + (alpha radiation)

- Beta-negative radiation (β) is negatively charged because it consists of high-speed electrons. The small size of the electrons in beta radiation makes it significantly more dangerous. Beta-negative radiation can be stopped by aluminium foil or thin pieces of wood.

  For example: $^{14}_{6}C \rightarrow {}^{14}_{7}N + e^{-}$

  (initial isotope) → (new isotope) + (beta-negative radiation)

- Gamma radiation (γ) is a form of electromagnetic radiation similar to X-rays, but it has higher energy and shorter wavelengths than X-rays. Gamma radiation has no mass or charge so it can penetrate more easily than alpha or beta radiation. Concrete and lead can stop gamma radiation, although they are not completely effective.

  For example: $^{226}_{88}Ra \rightarrow {}^{222}_{86}Rn + {}^{4}_{2}He + \gamma$

  (initial isotope) → (new isotope) + (alpha radiation) + (gamma radiation)

Radioisotopes emit one or more of these radiations when they undergo nuclear reactions. A **half-life** refers to the time taken for a radioisotope to decay to half of its original quantity of atoms. As a result, radioisotopes that initially have large atomic masses have longer half-lives as a complex series of nuclear reactions takes place.

half-life
the time it takes for half of the atoms of a radioisotope sample to decay

For example, uranium-238 decays to 15 **radioactive intermediates** before becoming the stable isotope lead-206 (Figure 3). As a result, radioisotopes that initially have large atomic masses often undergo a series of nuclear reactions to form stable isotopes as products. The total time taken for all such nuclear reactions can be quite long.

radioactive intermediate
a radioisotope that is very unstable and has an extremely short half-life

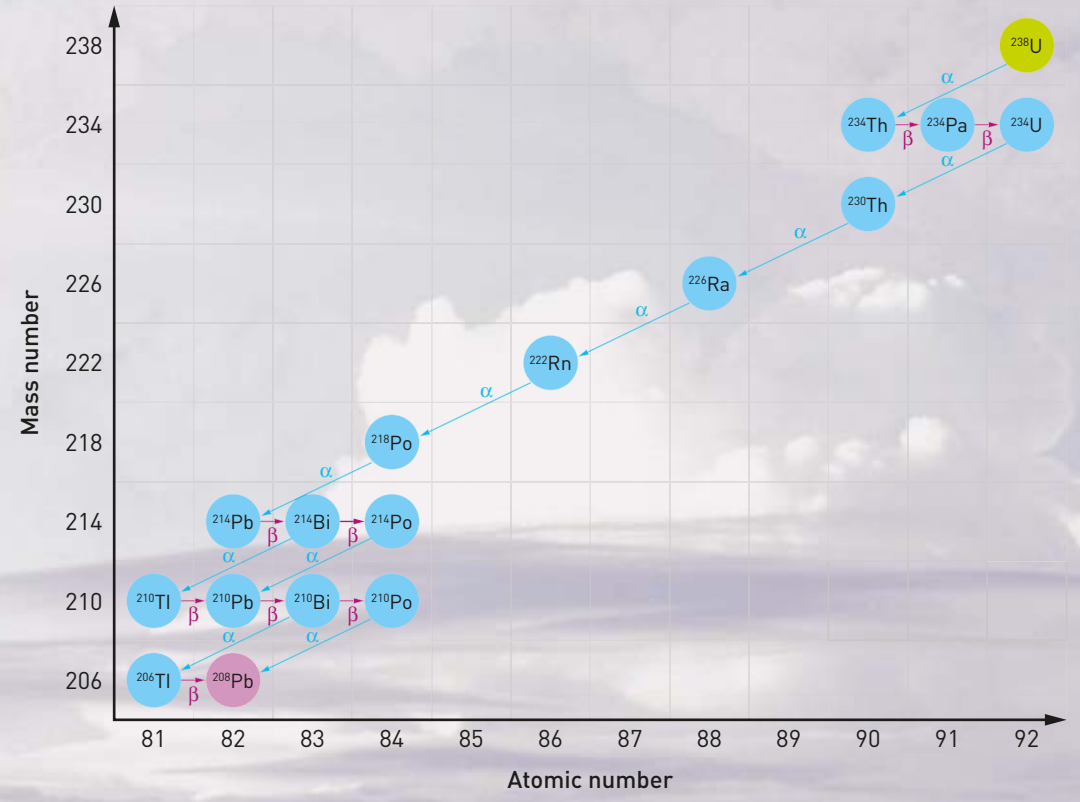

FIGURE 3 Radioactive decay series of ^{238}U.

Effects of radiation exposure

The radioactive energy produced by nuclear reactions affects the electrons of substances that it encounters. Electrons are involved in chemical bonding, so radiation exposure by living organisms breaks down crucial bonds within the body, including within the DNA molecules of each cell.

The follow-on effects of these broken bonds can be:

- interference with cell division because the DNA is irreparable so new cells cannot be formed; for example, radiation treatment first affects the specialised cells in the body that are fast-replicating, where any change in their rate of formation is noticeable
- continuation of cell division but with DNA that now contains errors; depending on what bonds have been broken and gone undetected, the formation of new cells may lead to cancer. Prolonged exposure of radiations is fatal to most life forms.

Detecting and monitoring radiation exposure

The radiation energy emitted by radioisotopes causes ionisation of the substances that they encounter.

Geiger counter
a device used to detect the presence of radiation (primarily beta radiation)

scintillation counter
a device used to detect multiple forms of radiation energy

film badge
a device worn by workers in the radiation industry to detect and quantify the level of radiation exposure

- A **Geiger counter** is a gas-filled metallic tube that detects radiation (primarily beta radiation) when it penetrates the tube and the gas inside ionises. The audible clicks associated with Geiger counters represent the degree of ionisation that the gas has undergone.
- **Scintillation counters** use a phosphor (a luminescent substance), on a special coated surface, to detect all forms of radiation and produce bright flashes of light, which are then recorded.
- **Film badges** are radiation detectors worn by people who work with radiation. Consisting of several layers of photographic film and black light-proof paper, a film badge monitors ongoing exposure to radioactive sources by the level of development (darkening) of the photographic film at frequent and specific time intervals.

FIGURE 4 Geiger counter for measuring radiation levels.

CASE STUDY 4.2

Radium and polonium

The discovery of the radioactive elements radium and polonium were not an easy feat! Both were discovered in 1898 by Marie and Pierre Curie. The wife and husband extracted the two elements from uranium ore and quickly realised they were highly radioactive metals. They discovered both elements have a large number of isotopes that undergo radioactive decay.

FIGURE 5 Marie Curie (Nobel prize winner)

Marie and Pierre Curie won a joint Nobel Prize in 1903 for Physics, and then in 2011 Marie won her second Nobel Prize in Chemistry for the discovery of the radioactive elements radium and polonium. This made her the first person to win two Nobel Prizes. Pierre Curie died when hit by a horse-drawn cart before he could receive his second Nobel Prize. Marie died in 1934, from prolonged exposure to radiation. It is believed she was exposed to substantial amounts of radiation while transporting test tubes of radioactive isotopes in her pocket!

CHECK YOUR LEARNING 4.2

Describe and explain

1 **Describe** the term 'radioisotope'.
2 **Explain** the term 'half-life'.

Apply, analyse and interpret

3 **Compare** the types of radiation energy emitted by nuclear reactions.
4 **Interpret** the graph of the band of stability for isotopic nuclei (Figure 1) to determine the optimal ratio of protons to neutrons.

Investigate, evaluate and communicate

5 **Investigate** the role of radioisotopes in carbon dating.
6 **Investigate** the role of radioisotopes in medical imaging and treatment.
7 **Evaluate** the significance of the size of the radiation energy, in terms of danger to surrounding living organisms and ability to block radiation.

You can find the following resources for this section on your obook assess:

- » Student book questions
  Check your learning 4.2
- » Weblink
  Radiation health effects
- » Weblink
  Radiation and mutation
- » Weblink
  Radiation safety

CHAPTER 4

Review

Chapter summary

4.1
- Isotopes are atoms of an element that have different numbers of neutrons.
- All atoms of an element have the same atomic number, but the isotopes vary in the mass number (which causes different physical properties).
- The value given in the periodic table is a ratio of the weighted average of the atomic masses of all the isotopes of an element (which considers the natural abundance of each of the isotopes on Earth) to one-twelfth the atomic mass of carbon-12.

4.2
- Radioisotopes are unstable isotopes. Some radioisotopes are naturally occurring, but many are synthesised under laboratory conditions.
- Radioisotopes undergo nuclear reactions to reach a more stable state. In doing so, they release radioactive energy, which has detrimental effects on surrounding substances.
- The penetrative effects of each type of radiation vary depending on factors such as the size of the by-product particle that is being emitted.
- Radioisotopes are often monitored by handheld and attachable devices.

Key terms

- film badge
- Geiger counter
- half-life
- isotopes
- nuclear reaction
- percentage abundance
- radiation energy
- radioactive intermediate
- radioisotope
- relative atomic mass (RAM)
- scintillation counter

Key formulas

Relative atomic mass	$\text{Relative atomic mass} = \frac{(\%_1 \times \text{AM}_1)}{100} + \frac{(\%_2 \times \text{AM}_2)}{100} + \frac{(\%_3 \times \text{AM}_3)}{100} \ldots$

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; ***=high.

Multiple choice

1 Determine which isotope of carbon contains seven neutrons.

- **A** ^{11}C
- **B** ^{12}C
- **C** ^{13}C
- **D** ^{14}C

2 Determine which properties are similar between isotopes.

- **A** Chemical reactivity
- **B** Density and viscosity
- **C** Electron configurations
- **D** Boiling and melting points

3 Relative atomic mass is described as:

- **A** dividing the number of isotopes of an element by 12, relative to carbon
- **B** counting the number of isotopes that exist for an element
- **C** averaging the mass numbers of isotopes that exist for an element
- **D** calculating a weighted average of the atomic masses of the isotopes for an element and expressing it as a ratio to one-twelfth the atomic mass of carbon-12.

4 Which of the following takes place when radioisotopes undergo nuclear reactions?

- **A** Only beta and gamma particles are emitted.
- **B** The unstable nucleus loses energy.
- **C** A new element is always formed.
- **D** Only alpha particles are emitted.

5 When a radioisotope has been decaying for two half-lives, what percentage of the atoms remain?

- **A** 75%
- **B** 50%
- **C** 25%
- **D** 12.5%

Short answer

Describe and explain

★ **6** **Describe** how one radioisotope becomes another isotope.

★ **7** **Explain** why a person working with radiation would require their detection device to record cumulative levels of radiation.

★★ **8** **Construct** balanced equations for:

- **a** alpha particle emission
- **b** beta-negative emission.

Apply, analyse and interpret

★★ **9** 90% of the atoms of an unknown element have an isotope with a mass number of 35. 8% of the atoms have a mass number of 37, and the remainder of the atoms have a mass number of 38. **Calculate** the relative atomic mass of the unknown element.

★★ **10** A radioactive isotope of americium is used in smoke detectors. ^{241}Am undergoes nuclear reactions to produce one neptunium isotope and one alpha particle. **Determine** the mass number of the neptunium product.

Investigate, evaluate and communicate

★★ **11** **Investigate** the role of wood, concrete and lead in protecting substances from radiation.

★★ **12** **Evaluate** the effectiveness of ^{201}Th, ^{131}I and ^{59}Fe in their applicability for medical diagnostics.

★★★ **13** **Investigate** the development and implementation of radiation detectors and cumulative monitors in industries exposed to radiation.

You can find the following resources for this section on your obook assess:

- » **Student book questions**
  Chapter 4 Revision questions
- » **Revision notes**
  Chapter 4
- » **assess quiz**
  Auto-correcting multiple choice quiz
- » **Flashcard glossary**
  Chapter 4

CHAPTER 5

Analytical techniques

Analytical techniques are used to identify or measure the amount of a molecule, atom, element, ion or compound in an unknown substance. The electromagnetic spectrum provides some of the most valuable information about the structure, bonding and composition of molecules.

The earliest demonstration of light being dispersed through a prism was in the form of a rainbow. Light is dispersed through water droplets, which act as a prism, forming the range of visible light that we see as colour. The resulting rainbow was named a spectrum by Sir Isaac Newton in the 1600s. Colours are different types of electromagnetic radiation.

The analytical technique of spectroscopy uses light to determine information about various substances. Molecules interact with electromagnetic radiation depending on the energy of the light used and the structure of the molecule.

OBJECTIVES

- → Understand that mass spectrometry involves the ionisation of substances and the separation and detection of the resulting ions, and that the spectra generated can be analysed to determine the isotopic composition of elements, and interpreted to determine relative atomic mass.
- → Understand that flame tests and atomic absorption spectroscopy (AAS) are analytical techniques that can be used to identify elements; these methods rely on electron transfer between atomic energy levels and are shown by line spectra.
- → Distinguish between absorption and emission spectra and recognise that the emission spectrum of hydrogen provides evidence for the existence of electrons in discrete energy levels (Bohr model), which converge at higher energies. Explain that emission spectra are produced when photons are emitted from atoms when excited electrons return to a lower energy level.
- → Analyse, interpret and evaluate data from flame tests and atomic absorption spectroscopy (AAS) to determine the presence and concentration of metallic ions in solution.
- → Use appropriate mathematical representations to make inferences and to solve problems, including calculating the relative atomic mass of an element and percentage abundances of the isotopes of an element from data.

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PRACTICALS

SUGGESTED PRACTICAL

5.2 Conducting a flame test

FIGURE 1 A continuous spectrum of light passing through a prism

5.1 Spectroscopy

KEY IDEAS

In this section, you will learn about:

- the electromagnetic spectrum
- absorption and emission spectra
- evidence obtained from absorption and emission spectra that supports Bohr's model for electron configuration.

Spectroscopy is the branch of chemistry that investigates the interaction between matter and the electromagnetic spectrum. This interaction, which occurs when molecules interact with light (electromagnetic radiation), will depend not only on the type of molecule, atom, element or compound, but also on the type of electromagnetic radiation that is used.

The electromagnetic spectrum

electromagnetic spectrum
the range of frequencies of electromagnetic radiation and their energies and wavelengths; electromagnetic radiation, or light, is comprised of photons

frequency
the number of waves (referring to wavelength) passing a point each second

wavelength
the distance between two crests or troughs of a wave

photon
a particle and wave of light that carries energy and is absorbed or emitted from atoms

The **electromagnetic spectrum** consists of the full range of **frequencies** of light. It displays the range of electromagnetic radiation, its energies, frequencies and **wavelengths**. Electromagnetic radiation, or light, consists of **photons** (particles lacking mass), which travel at the speed of light in a vacuum, such as space.

A photon behaves as both a particle and a wave, depending on the type of the experiment in which it is used. In terms of the spectra obtained and analysed in chemistry, light behaves like a particle. These particles of light carry energy and are absorbed or emitted from atoms. The energy of a photon is related to its frequency. The higher the frequency (smaller wavelength), the higher the energy of the light.

Figure 1 shows the electromagnetic spectrum of light. It demonstrates the inverse relationship between energy and wavelength: as energy increases, wavelength decreases. Visible light forms the narrowest region of the spectrum.

The different frequencies of electromagnetic radiation have many applications.

For example, radios and televisions use radio waves, microwave ovens use electromagnetic radiation from the microwave region, and X-rays are used to visualise broken bones.

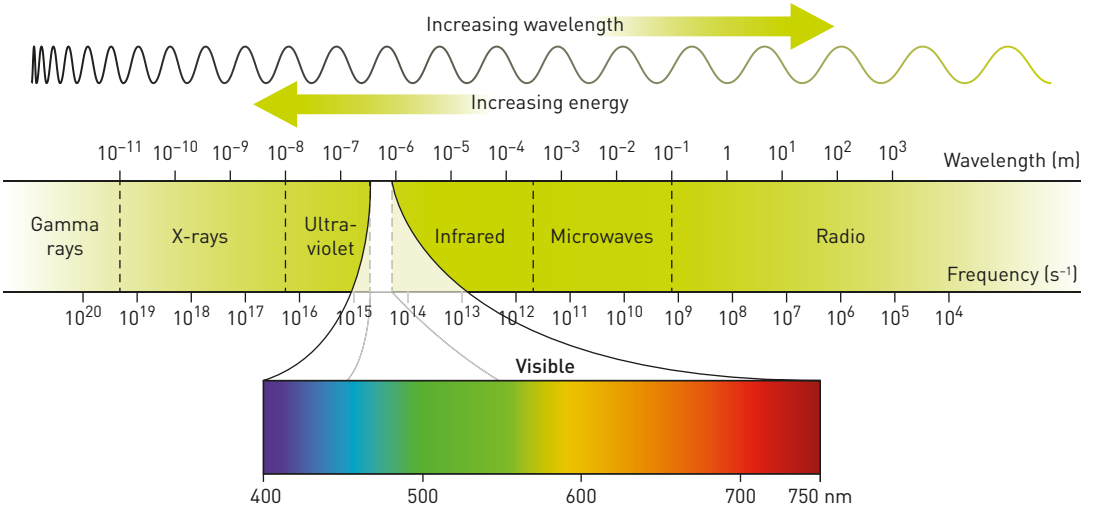

FIGURE 1 The electromagnetic spectrum consists of the full range of frequencies of light.

Study tip

The plural of spectrum is 'spectra'.

Electrons can move between shells

In an atom, negatively charged electrons orbit the positively charged nucleus. Figure 2 shows an electron shell diagram for a hydrogen atom, which contains one positively charged proton in its nucleus, and one negatively charged electron. The atom is in its **ground state**, or most stable state, and the electron is orbiting in the first electron shell ($n = 1$).

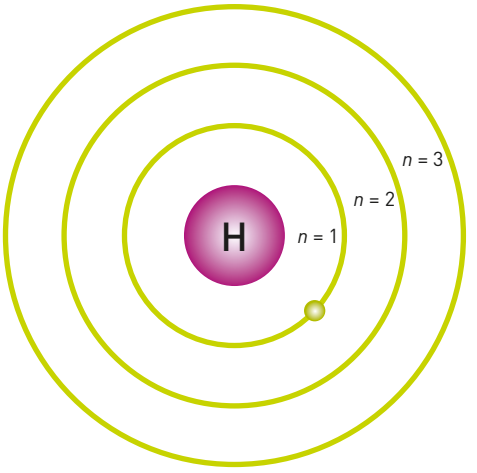

FIGURE 2 Electron shell diagram of the ground state of a hydrogen atom, with its electron orbiting the nucleus in $n = 1$.

When an atom **absorbs** energy, electrons move from a lower energy electron shell to a higher energy electron shell; for example, from $n = 1$ to $n = 2$ or $n = 3$. This is called the **excited state** because electrons contain more energy than they would if they were in their ground state. The more energy that is absorbed, the further away from the nucleus the electrons will move as they jump to a higher energy electron shell.

In order for an electron to move from a lower to a higher energy level, it must absorb energy in the form of electromagnetic radiation. The amount of electromagnetic radiation equals the difference between two energy levels. For example, for an electron to transfer from the third ($n = 3$) to the fifth ($n = 5$) shell, the energy absorbed must be equal to the difference in energy of these shells.

Figure 3 demonstrates the energy levels of electron shells. For an electron to move from the first ($n = 1$) to the second ($n = 2$) shell, it must absorb a single photon containing 10.2 eV (electron volts) of energy with a wavelength of 122 nm (see Figure 4). If the electron were to move from the first ($n = 1$) to the third ($n = 3$) shell, it would need to absorb a photon with a wavelength of 103 nm. An electron its excited state is unstable so will tend to emit the excess energy as a photon of electromagnetic radiation.

ground state
the lowest energy, and most stable, electron shell that an electron can occupy as it orbits the nucleus of an atom

absorbs
the attraction of a substance within a sample to the stationary phase

excited state
an energy state that is higher than the ground state; an electron is unstable in this state and emits energy as light to return to the ground state

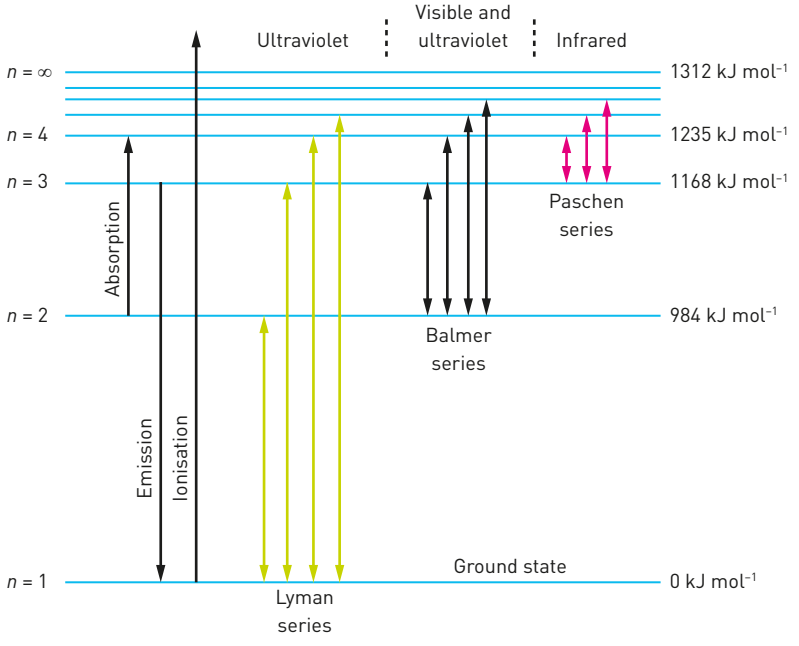

FIGURE 3 Energy levels of electron shells. Transitions from lower to higher energy levels require energy to be absorbed. Transitions from higher to lower energy levels are accompanied by an emission of energy. Different transitions give off radiation in different parts of the electromagnetic spectrum. The Lyman series of transitions occurs in the ultraviolet, the Balmer series occurs in the visible and ultraviolet, and the Paschen series occurs in the infrared region.

Study tip

There is a limit to the amount of energy that an electron can absorb before it is completely removed and the atom becomes an ion. This is called the ionisation energy.

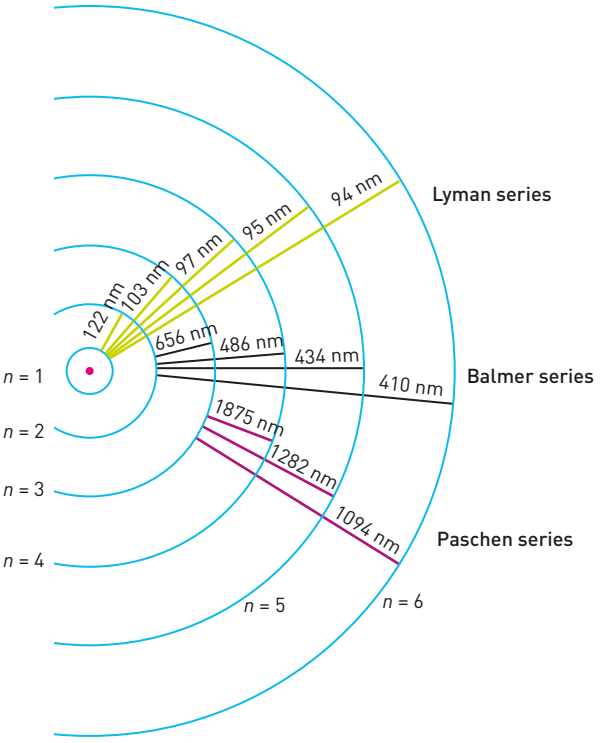

FIGURE 4 The energy emitted by electrons, as a wavelength (nanometres), as they transition from higher energy levels (shells) to lower ones. The pink dot represents the nucleus.

Lyman series

Transitions of electrons between the second electron shell or greater ($n \geq 2$) and the first electron shell ($n = 1$) is called the Lyman series. This was named after Harvard physicist Theodore Lyman, who studied the ultraviolet region of the electromagnetic spectrum. An electron in an excited state in the fourth electron shell emits electromagnetic radiation when it returns to its ground state in the first electron shell. As it transitions between these shells, it emits a single photon with a wavelength of 97 nm (Figure 4), which is in the ultraviolet region of the electromagnetic spectrum (Figure 1).

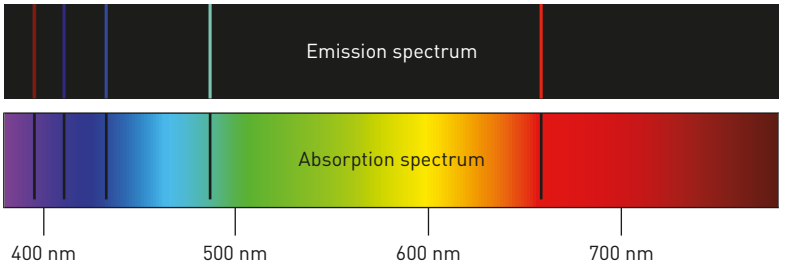

FIGURE 5 The absorption and emission spectra of hydrogen.

Balmer series

The Balmer series represents transitions of electrons in a hydrogen atom from the third electron shell or greater ($n \geq 3$) and the second electron shell ($n = 2$). In 1885, Johann Balmer was researching the hydrogen atom and discovered that, as electrons transition between excited states in the third shell or higher to a lower excited state in the second shell, they emit a photon of light with a wavelength in the visible region of the spectrum (400–750 nm). The electron transition is seen as coloured lines in absorption and emission spectra (Figure 5).

When an electron transitions between the third and second shells, it emits a photon of light of wavelength 656 nm (see Figure 5), which is red. So when transitioning between the third and second shell, the hydrogen electron emits red light. When transitioning between the fourth and second shells, the hydrogen electron emits blue–green light (wavelength 486 nm).

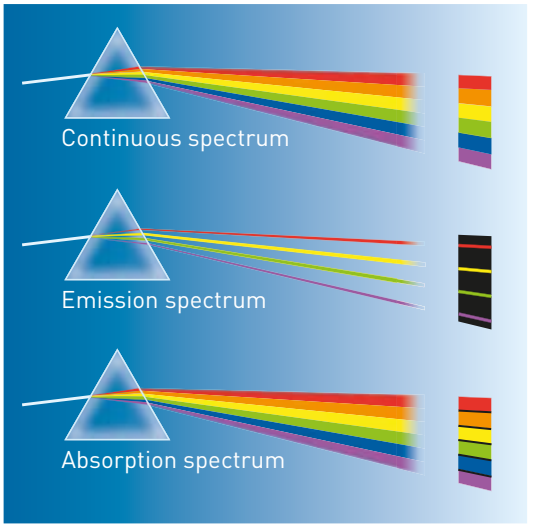

FIGURE 6 When white light passes through a prism, it produces a continuous spectrum – a rainbow of colours. An emission spectrum consists of a set of bright lines against a black background. An absorption spectrum consists of dark lines against a continuous spectrum.

Absorption and emission spectra

As electrons both absorb and emit electromagnetic radiation in the visible spectrum as light, they produce an **absorption spectrum** and an **emission spectrum** (Figure 5).

A continuous spectrum (Figure 6) shows the visible spectrum of light – all of the colours that have wavelengths between 400 and 750 nm.

An emission spectrum (Figure 7) is generated by heating a sample to excite its electrons. The light that the sample emits when the electrons move from an excited state to a lower state is focused through a slit and passed through a prism to disperse the light into a spectrum. The photons of light and their corresponding wavelengths that are emitted when the electrons transition to lower orbitals are seen as coloured lines on a black spectrum (Figure 7).

absorption spectrum
the band of lines observed when an element absorbs light; viewed as a series of black lines on the visible light spectrum

emission spectrum
the band of lines observed when an element emits light; viewed as a series of coloured lines that represent the wavelengths of electromagnetic radiation emitted by a source

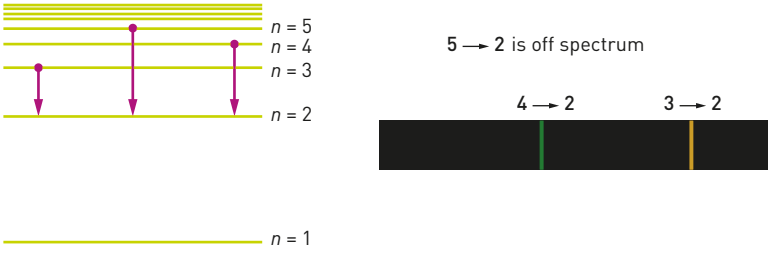

FIGURE 7 The emission spectrum generated by hydrogen when it absorbs photons of light at specific wavelengths on a continuous visible spectrum.

An absorption spectrum is generated when white light is passed through a gas sample. The sample absorbs photons of light of specific wavelengths as electrons transition to higher energy levels. The light that is not absorbed is focused through a slit and passed through a prism to disperse the light into a spectrum. On a continuous spectrum, the wavelengths that have been absorbed are missing and appear as black lines. These black lines correspond to the energy of the photons required to transition an electron from a lower to a higher energy shell (Figure 8).

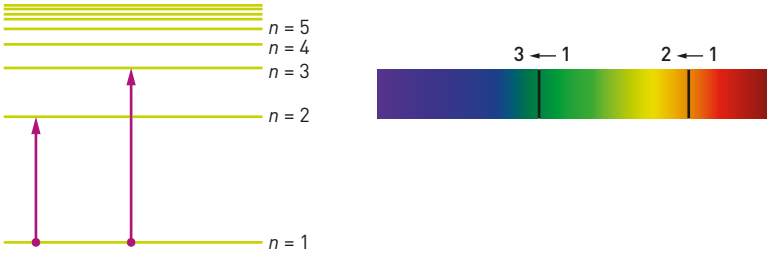

FIGURE 8 The absorption spectrum generated by hydrogen when it absorbs photons of light at specific wavelengths on a continuous visible spectrum.

Paschen series

The Paschen series represents the transition of electrons from the fourth electron shell or greater ($n \geq 4$) to the third electron shell ($n = 3$). The series was named for Friedrich Paschen, who, in 1908, discovered that transitions between these shells emitted photons of light with wavelengths in the infrared region of the electromagnetic spectrum.

Bohr model

The absorption and emission spectra of the element hydrogen provide the evidence for Niels Bohr's theory that electrons orbit the nucleus of an atom in distinct shells.

In 1913, Bohr proposed that:

- electrons orbit an atom's nucleus
- electrons exist in orbitals, which are distinct distances from the nucleus and have definite energy levels
- for an electron to move between orbitals, it must gain energy or lose energy by **absorbing or emitting photons of light with distinct wavelengths.**

CHALLENGE 5.1

Spectra produced by different elements

How would the absorption and emission spectra of iron differ from that of hydrogen? Justify your answer using your knowledge of the electron configuration of both atoms.

CASE STUDY 5.1

The Bunsen burner

Robert Bunsen was a German chemist who studied the electromagnetic spectrum, specifically emission spectra. In 1852, he was hired by the University of Heidelberg, who enticed him into employment by building him a new laboratory with gas inlets.

Bunsen had long grappled with the issue of laboratory burners, which produced soot, generated very little heat and burned a brilliant yellow colour. The yellow colour interfered with emission spectra, and without high temperatures, the electrons of metal atoms would not be excited to higher energy levels. Furthermore, the soot interfered with any data obtained.

Ideally, Bunsen wanted to make a burner to view the emission spectra of metals. He needed a very hot flame (to excite the electrons), which emitted very little to no light, so that it would not interfere with the spectrum, and did not produce soot.

Bunsen approached the university's mechanic, Peter Desaga, to build a prototype that would satisfy all of these criteria. The design included an adjustable air opening, which produced a cooler yellow flame when it was open and a hotter blue flame when it was closed. The blue flame, as well as being very hot, emitted very little light.

The Bunsen burner (Figure 9) is still used in school and professional laboratories and is one of the most effective methods of generating an emission spectrum. Using this technology, Bunsen discovered the elements caesium and rubidium and he received several awards for his achievements.

FIGURE 9 Bunsen burners are widely used to conduct flame tests.

CHECK YOUR LEARNING 5.1

Describe and explain

1 **Explain** the difference between absorption and emission.

2 **Explain** the difference between absorption and emission spectra.

3 **Explain** the difference between the Lyman series and Balmer series.

Apply, analyse and interpret

4 **Analyse** the emission spectrum of hydrogen in Figure 5.

- **a** **Deduce** the number of energy transitions hydrogen has in the spectrum.
- **b** **Determine** what each line on the spectrum represents using its wavelength and colour.
- **c** **Consider** whether it would be possible for hydrogen to have more transitions than those appearing on this spectrum. Explain your answer.

5 **Categorise** the analysis of elements using absorption or emission spectra as a qualitative or quantitative technique.

Investigate, evaluate and communicate

6 **Justify** why an absorption or emission spectrum is referred to as a fingerprint of an element.

7 Based on your knowledge of atomic theory and electron configuration, **discuss** the possibility that there could be more series than just the Lyman, Balmer and Paschen series.

Check your obook assess for these additional resources and more:

- » Student book questions
  Check your learning 5.1
- » Challenge
  5.1 Spectra produced by different elements
- » Video
  The electromagnetic spectrum
- » Video worksheet
  The electromagnetic spectrum

5.2 Techniques for identifying elements

KEY IDEAS

In this section, you will learn about:

- how flame tests can be used to analyse metals
- how atomic absorption spectroscopy can be used to analyse trace levels of metals
- interpreting data obtained from atomic absorption spectroscopy.

flame test
the analysis of metal ions by heating them in a Bunsen burner flame and measuring the light emitted when excited electrons return to their ground state

atomic absorption spectroscopy (AAS)
the determination of the concentration of metal atoms in the gas phase by measuring the amount of light absorbed from a specific lamp

Absorption and emission spectra are essential to the qualitative and quantitative analysis of metals. The **flame test** and **atomic absorption spectroscopy (AAS)** are analytical techniques that can be used to identify elements. They both rely on electron transfer between atomic energy levels.

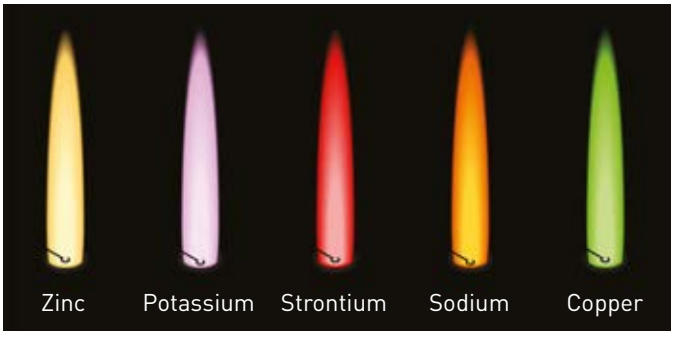

FIGURE 1 The colours produced by some metal ions in a blue Bunsen burner flame.

Flame tests

The flame test is used to demonstrate the characteristic colours that metals emit when placed in a blue Bunsen burner flame (Figure 1). The heat of the flame excites electrons to a higher electron orbital shell. Photons of light are emitted at distinct wavelengths when electrons return to their ground state.

Typically, flame tests are conducted on metal chlorides. A small sample is collected on the end of a wire loop and placed into the blue Bunsen burner flame. However, the colours produced can be quite difficult to distinguish from one another. Some metal ions look very similar and therefore the identification of various metals is impossible with the naked eye. For example, lithium, sodium, calcium and strontium appear yellow, red or orange, and this can make it difficult to tell them apart.

To overcome this issue, a spectrograph is used to generate an emission spectrum of each element (Figure 2). Although they may appear to be the same colour, the photons of light emitted at various wavelengths differ significantly, and therefore identification occurs when the spectrum is generated.

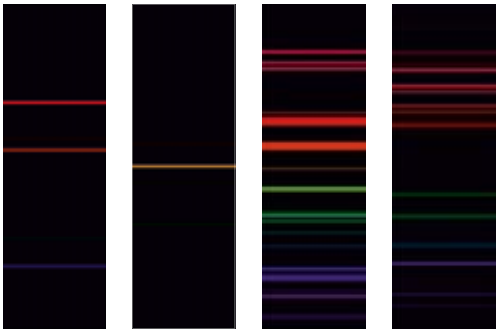

FIGURE 2 Emission spectra of lithium, sodium, calcium and strontium are used to identify the elements.

Atomic absorption spectroscopy

AAS is an analytical technique in which absorption spectra of metals are used to determine the quantity of the metal in an unknown sample. It is essential in the analysis of trace levels of compounds, as low as parts per billion.

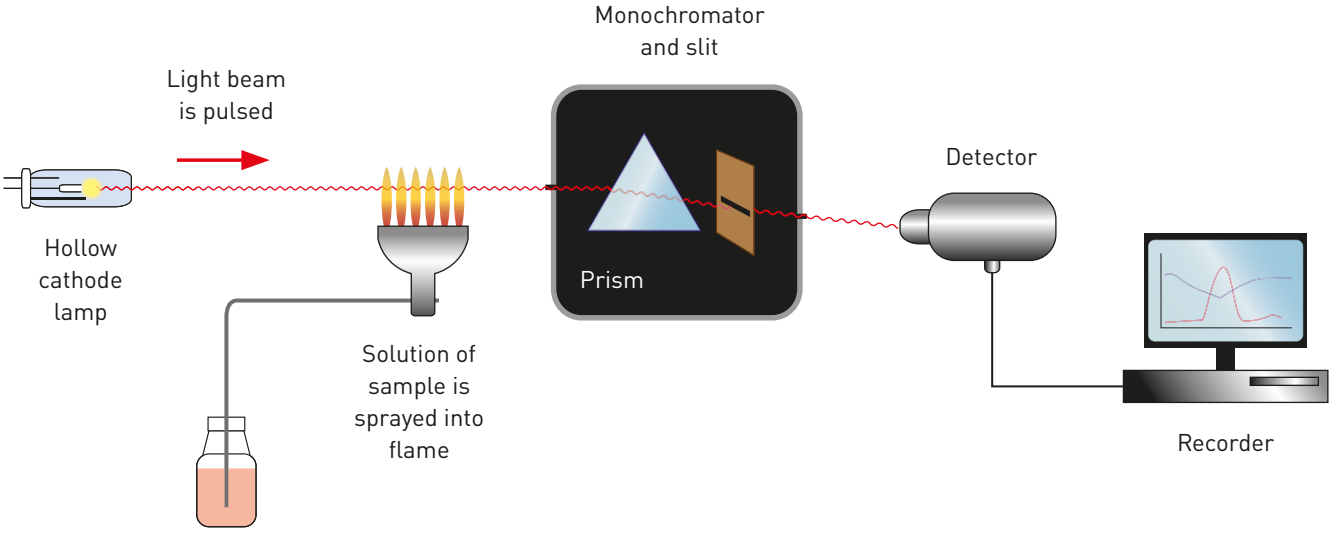

FIGURE 3 Schematic diagram of an atomic absorption spectrometer

An atomic absorption spectrometer (Figure 3) works by spraying a solution of the sample into a flame, which converts the sample into atoms. A lamp is then set to a specific wavelength and the light beam is pulsed at the sample in the flame, to excite the electrons in the atom of interest. As electrons return to their ground state, they emit photons of light with characteristic wavelengths.

The emitted light is passed through a monochromator, which is a prism capable of selecting the wavelength that it will send to the detector. The detector measures the intensity (absorbance) of the light that is emitted relative to the original light source. The greater the intensity (absorbance), the higher the concentration of the metal atoms in the sample.

Data analysis and interpretation

AAS is a quantitative technique because it can measure the concentration of metal atoms in a sample. The sample is compared to a set of **standards** of known concentration of the metal being analysed in the sample. A **calibration curve** of the standards is prepared, and the unknown sample's concentration is calculated from this. The calibration curve is a graph that plots the concentration of the standards on the *x*-axis and the absorbance of the light on the *y*-axis.

standard
a sample with a known concentration of the chemical that is being tested for

calibration curve
the trend associated with plotting the peak area of a set of standards, from a chromatogram, against their concentration

WORKED EXAMPLE 5.2

A sample containing zinc is analysed by AAS. A set of standards is run with the sample and the data obtained is shown in Table 1. Calculate the concentration of zinc atoms in the sample.

TABLE 1 Concentration and absorbance of zinc standards using AAS

Standard/sample	Concentration (ppm)	Absorbance
Standard 1	0	0
Standard 2	15	25
Standard 3	30	50
Standard 4	45	75
Standard 5	60	100
Sample	–	80

SOLUTION

1 Construct a calibration curve by plotting the concentration of the standards against absorbance and draw a linear line of best fit.

2 Draw a horizontal line from the absorbance of the sample (on the *y*-axis) to the line of best fit. Draw a vertical line down to the *x*-axis. Record the concentration value at this point.

The sample has a zinc concentration of 48 ppm.

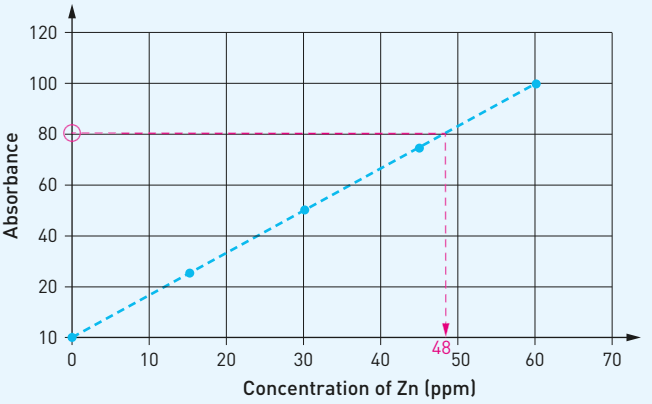

FIGURE 4 Calibration curve for zinc concentration

Study tip

Double check the counting of the squares on the graph axis to ensure your scale is correct. Always use a ruler to ensure the accuracy of the graph.

CHALLENGE 5.2

Determining the Cu^{2+} content of tap water by AAS

A chemist wants to determine the Cu^{2+} content in tap water by AAS. After initial testing, she determines that the sample is too concentrated and dilutes it by a factor of 100. She then reruns the sample with a set of Cu^{2+} standards and obtains the data in Table 2.

TABLE 2 Concentration and absorbance of Cu^{2+} standards using AAS

Standard/sample	Concentration (ppb)	Absorbance
Standard 1	0	0
Standard 2	20	21
Standard 3	40	42
Standard 4	60	63
Standard 5	80	84
Sample	–	72

Determine the concentration of Cu^{2+} in the original sample of tap water.

CASE STUDY 5.2

The development of the atomic absorption spectrometer

The invention of AAS is credited to British–Australian physicist Sir Alan Walsh (1916–1998) (Figure 5). Until Walsh invented the atomic absorption spectrometer, all spectra analysed were emission spectra because scientists found it easier to excite electrons and measure the light emitted when they returned to ground state. Called atomic emission spectroscopy, this method had very little sensitivity and could analyse fewer than ten elements.

Walsh, in a 'Eureka' moment, concluded that scientists had been looking at spectra in the wrong way and that they should be looking at absorption, rather than emission, spectra. Amazingly, it took him less than a day to build a working model of an atomic

absorption spectrometer. Initial testing of the instrument indicated that it could analyse 65 different elements with increased accuracy and sensitivity.

Walsh could not persuade any companies to manufacture the atomic absorption spectrometer commercially. So he separated the instrument into multiple components and sent each component to a separate company for manufacture. The result was a kit version that buyers assembled into an operating spectrometer.

The atomic absorption spectrometer has proven to be essential for analysing elements, specifically metals. It is used in such chemical industries as mining to determine the amount of metal in an ore sample, and it is also used in medicine and environmental analysis. The spectrometer has saved lives by determining unsafe levels of metals in blood or the mercury content in fish and waterways.

FIGURE 5 Sir Alan Walsh – the inventor of atomic absorption spectrometry.

CHECK YOUR LEARNING 5.2

Describe and explain

1 **Explain** the different ways metals are analysed in a flame test versus AAS.

2 **Explain** how AAS relates to absorbance spectra. (Use your knowledge of electron energy levels.)

Apply, analyse and interpret

3 An analysis was conducted by AAS to determine the concentration of cadmium in sea turtle blood. **Determine** the concentration of cadmium in the solution 'Sample 1' from the data in Table 3.

TABLE 3 Cadium concentions of standards

Standard/sample	Concentration (ppb)	Absorbance
Standard 1	0	0
Standard 2	14	42
Standard 3	28	84
Standard 4	42	126
Standard 5	66	168
Sample 1	–	96

4 Two samples of waste from a manufacturing company are analysed to determine whether the concentration of lead(II) is at a safe environmental level. The recommended concentration of lead(II) ions is 5 ppm. **Analyse** the results in Table 4 and provide a recommendation to the company based on your findings.

TABLE 4 Concentration and absorbance of lead standards using AAS

Standard/sample	Concentration ppm	Absorbance
Standard 1	0	0.00
Standard 2	2	0.13
Standard 3	4	0.26
Standard 4	6	0.39
Standard 5	8	0.52
Sample 1	–	0.20
Sample 2	–	0.35

Investigate, evaluate and communicate

5 **Evaluate** the appropriateness of AAS for determining the concentration of glucose ($C_6H_{12}O_6$) in a sample. Justify why it may be appropriate or inappropriate to use the technique.

6 **Investigate** the analytical technique atomic emission spectroscopy. How does it differ from AAS? Explain the differences in processes involved with analysing atoms.

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- » Suggested practical
  5.2 Conducting a flame test
- » Challenge
  5.2 Determining the Cu^{2+} content of tap water by AAS
- » Video
  Flame test

5.3 Mass spectrometry

KEY IDEAS

In this section, you will learn about:

- separating an element into its isotopes
- determining the mass and percentage abundance of isotopes
- calculating relative atomic mass based on the mass and percentage abundance of isotopes
- calculating percentage abundance of isotopes based on relative atomic mass and isotopic mass.

mass spectrometry an analytical technique in which chemical elements are ionised and the ions are sorted according to their mass-to-charge ratio

mass spectrometer an instrument that ionises elements and molecules and produces spectra showing mass-to-charge ratio

ionisation the formation of ions from atoms by the addition or removal of electrons

Mass spectrometry is one of the most useful analytical techniques available to chemists. It is used to determine the isotopic composition of elements and the bonding structure of molecules.

Every element has isotopes that must be considered when determining the element's mass. Isotopes have different masses because they have different numbers of neutrons. Therefore, the mass of an element must be calculated as an average of the masses of an element's isotopes. In order to determine how much of each isotope exists, mass spectrometry is used.

Mass spectrometer

In mass spectrometry, a sample of an element is injected into the ionisation chamber of a **mass spectrometer**.

1. The element is bombarded with electrons from an electron beam. These electrons pair up with one electron from the element to become a pair of electrons, leaving the element positively charged, having lost an electron and formed an ion. This process is called **ionisation**.
2. The ions then pass through an electric field, which accelerates them through a magnetic field.
3. Within the magnetic field, the ions follow a curved path, with the paths of more positively charged or lighter ions bending more than the paths of less positively charged, heavier ions (Figure 1).

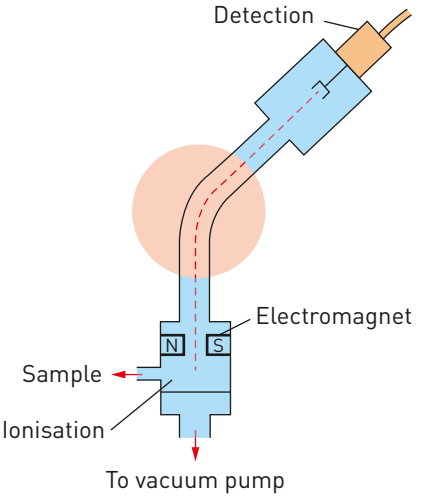

FIGURE 1 The path of particles through a mass spectrometer.

This causes ions to be separated according to their mass-to-charge ratio. The detector collects a **mass spectrum**, which displays the mass-to-charge ratio of each particle on the *x*-axis and its percentage abundance on the *y*-axis.

mass spectrum the spectrum obtained from a mass spectrometer; it displays mass-to-charge ratio on the *x*-axis and percentage abundance of each positively charged ion on the *y*-axis

The spectrum generated shows:

- how many isotopes an element has, because each has a different mass-to-charge ratio
- the relative mass of each isotope, indicated by the mass-to-charge ratio
- how much of each isotope is in the sample, indicated by the percentage abundance.

The mass of each isotope is measured by comparing it to the carbon-12 isotope, which is always assigned a mass of 12. Therefore, isotopic mass is measured relative to ^{12}C and is called **relative isotopic mass (RIM)**.

relative isotopic mass (RIM) the mass of isotopes measured on a mass spectrum, relative to the carbon-12 isotope

Determining relative isotopic mass (and percentage abundance)

In mass spectra, the mass-to-charge ratio represents the RIM of an isotope. The mass spectra for silicon and boron are shown in Figure 2.

- The mass spectrum for silicon reveals that the element has three isotopes. The isotope with a relative isotopic mass of 28 (^{28}Si) has a relative abundance of 92.23%. The remaining two isotopes, which have relative isotopic masses of 29 (^{29}Si) and 30 (^{30}Si), have **percentage abundances** of 4.68% and 3.09% respectively.
- In the mass spectrum of boron, the percentage abundances can be found by reading the heights of the peaks off the *y*-axis. Therefore, ^{11}B, with a relative isotopic mass of 11, has a percentage abundance of 80%, whereas ^{10}B, with a relative isotopic mass of 10, has a percentage abundance of 20%.

percentage abundance the percentage of each isotope contained within a sample of element; it must always add up to 100%

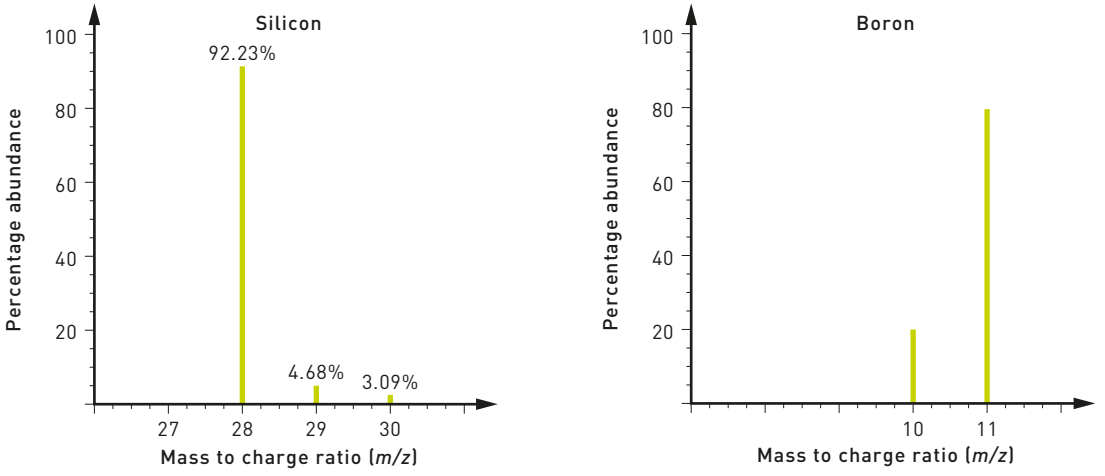

FIGURE 2 The mass spectra of silicon and boron.

Calculating relative atomic mass or mass number

When calculating the mass of an element, you must consider all the isotopes of elements, and their different relative masses and abundances. The mass of an element, called its **relative atomic mass** (A_r), is calculated from the following equation:

relative atomic mass the mass of an atom once the mass and percentage abundance of all isotopes have been considered, relative to the carbon-12 isotope

$$A_r = \frac{(\text{RIM}_1 \times \text{abundance}_1(\%)) + (\text{RIM}_2 \times \text{abundance}_2(\%)) + (\text{RIM}_3 \times \text{abundance}_3(\%)) + \ldots}{100}$$

WORKED EXAMPLE 5.3A

A sample of copper is run through a mass spectrometer and the spectrum in Figure 3 is obtained.

Calculate the relative atomic mass of copper.

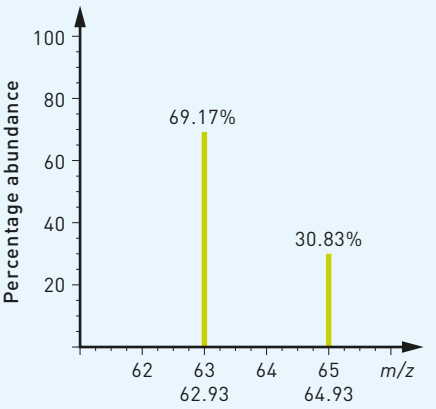

FIGURE 3 The mass spectrum of copper.

SOLUTION

$$A_r = \frac{(\text{RIM}_1 \times \text{abundance}_1(\%)) + (\text{RIM}_2 \times \text{abundance}_2(\%))}{100}$$

$$= \frac{(62.93 \times 69.17) + (64.93 \times 30.83)}{100}$$

$$= \frac{6354.66}{100}$$

$$= 63.54$$

(On the periodic table, the mass number of copper is 63.5.)

WORKED EXAMPLE 5.3B

Magnesium has three isotopes. The data for these isotopes is summarised in Table 1.

TABLE 1 Percentage abundance of magnesium isotopes

Isotope	Mass	Percentage abundance
^{24}Mg	23.985	78.99
^{25}Mg	24.985	10.00
^{26}Mg	25.982	11.01

Determine the relative atomic mass of magnesium.

SOLUTION

$$A_r = \frac{(\text{RIM}_1 \times \text{abundance}_1(\%)) + (\text{RIM}_2 \times \text{abundance}_2(\%)) + (\text{RIM}_3 \times \text{abundance}_3(\%))}{100}$$

$$= \frac{(23.985 \times 78.99) + (24.985 \times 10.00) + (25.982 \times 11.01)}{100}$$

$$= \frac{2430.49}{100}$$

$$= 24.30$$

Backtracking calculations: determining percentage abundance from A_r

If an element has only two isotopes, and the relative atomic mass of the element and the mass of each isotope is known, these values can be used to calculate the percentage abundance of each isotope.

The percentage abundances of the two isotopes must add up to 100% so the abundance of one isotope is denoted as x, and the abundance of the second isotope is $100 - x$.

WORKED EXAMPLE 5.3C

Chlorine has a mass number of 35.5 on the periodic table. Naturally occurring chlorine has two isotopes. One has a relative atomic mass of 34.968 and the other has a relative atomic mass of 36.965. Determine the percentage abundance of the two isotopes.

SOLUTION

The percentage abundances of the two isotopes must add up to 100%.

Let the percentage abundance of chlorine with mass 34.968 = x.

If $^{34.968}Cl = x\%$, then $^{36.965}Cl = 100 - x\%$. Note that $^{34.968}Cl$ has been rounded to ^{35}Cl and $^{36.965}Cl$ has been rounded to ^{37}Cl in Table 2.

TABLE 2 Percent abundance of chlorine isotopes

Isotope	Relative atomic mass	Percentage abundance
^{35}Cl	34.968	x
^{37}Cl	36.965	$100 - x$

1 Substitute these values into the equation:

$$A_r = \frac{(\text{RIM}_1 \times \text{abundance}_1(\%)) + (\text{RIM}_2 \times \text{abundance}_2(\%))}{100}$$

2 Expand this equation.

Remember that, when multiplying x by 36.965, the negative must be included with the x, and so it becomes $-x \times 36.695$.

$$35.5 = \frac{(34.968 \times x) + (36.965 \times (100 - x))}{100}$$

$$35.5 = \frac{34.968x + 3696.5 - 36.965x}{100}$$

3 Multiply both sides by 100.

$$3550 = 34.968x + 3696.5 - 36.965x$$

4 Collect all x-expressions on one side and all numbers on the other.

$$3550 - 3696.5 = 34.968x - 36.965x$$
$$-146.5 = -1.997x$$

5 Divide by the coefficient of x.

$$\frac{-146.5}{-1.997} = \frac{-1.997x}{-1.997}$$
$$73.36 = x$$

Therefore, abundance of 34.968 Cl = 73.36%

Abundance of 36.965 Cl = 100 − 73.36 = 26.64%

Study tip

A_r is measured in amu (atomic mass units). However, typically units are not used.

CHALLENGE 5.3

Determining RIM

Silver has two isotopes. One isotope has an RIM of 106.905 and an abundance of 51.839%.

If the relative atomic mass of silver is 107.9, determine the RIM and abundance of the second isotope.

FIGURE 4 Silver exists as two isotopes.

CASE STUDY 5.3

Proteomics: the study of proteins

Proteomics is the study of the sets of proteins found in a cell, tissue or organism, or their proteome. The name 'proteome' is derived from the words 'protein' and 'genome', meaning the structure and function of proteins. It is also closely related to the human genome project because proteins are derived from DNA. Proteins can act as catalysts by speeding up chemical reactions without taking part in them. They are essential for ensuring that reactions occur in a very small amount of time within the body.

Mass spectrometry is essential to the analysis of proteins. Proteins are analysed in one of two ways (Figure 5).

The 'bottom-up' approach involves digesting the protein with an enzyme to break it down into smaller fragments (peptides). The protein fragments are then injected into the ionisation chamber of the mass spectrometer and bombarded with electrons. This produces positively charged molecular fragments of protein. These fragments provide a snapshot of the composition of the protein.

Proteins are very large on average; a protein polymer contains 400 or more monomer units to form the protein polymer. So the mass spectra produced are very large and contain a lot of information that needs interpreting.

In the 'top-down' approach, the protein is kept intact and analysed as a whole, rather than as fragments.

Mixtures of proteins must be separated before they are run through the mass spectrometer. Scientists use a technique called liquid chromatography to separate the proteins from one another. The liquid chromatograph is connected to the mass spectrometer so that as soon as each protein leaves the liquid chromatograph, it enters the mass spectrometer and is analysed.

These techniques have led to a better understanding of biomolecules and their unique structures and functions within organisms.

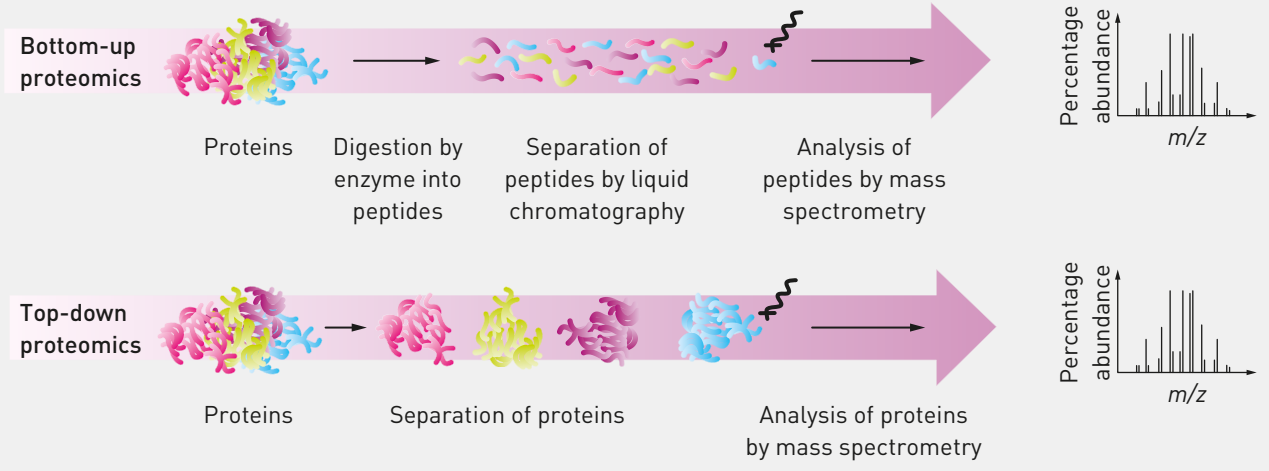

FIGURE 5 Bottom-up and top-down analysis of proteins by mass spectrometry.

CHECK YOUR LEARNING 5.3

Describe and explain

1 **Explain** how a mass spectrometer forms charged ions.

2 **Explain** why an element must become an ion before it can be analysed in a mass spectrometer.

Apply, analyse and interpret

3 **a** Analyse the data obtained from the spectra of silver's two isotopes in Table 3 and **determine** its relative atomic mass.

TABLE 3 Percentage abundances of silver's isotopes

Isotope	Relative atomic mass	Percentage abundance
^{107}Ag	106.9	51.839
^{109}Ag	108.9	48.161

b **Analyse** the mass spectrum in Figure 6 and **determine** the relative atomic mass of silicon.

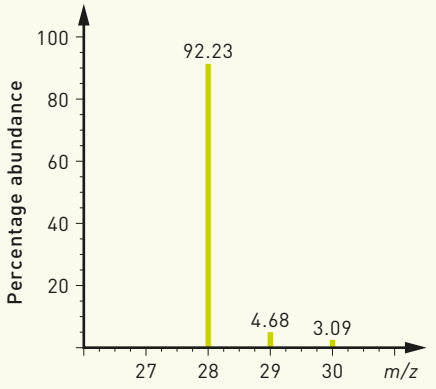

FIGURE 6 Mass spectrum of silicon.

c **Analyse** the mass spectrum in Figure 7 and **determine** the relative atomic mass of boron.

d **Determine** the relative atomic mass of antimony from the spectra of its two isotopes using the data in Table 7.

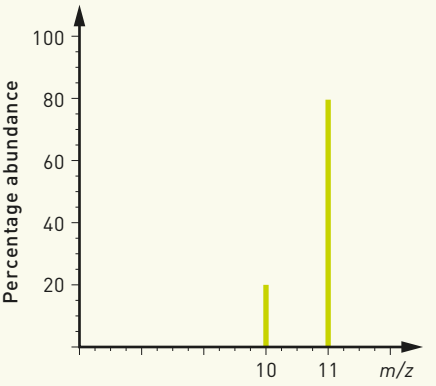

FIGURE 7 Mass spectrum of boron.

TABLE 4 Percentage abundance of antimony's isotopes

Isotope	Relative atomic mass	Percentage abundance
^{121}Sb	120.903818	57.21
^{123}Sb	122.904216	42.79

4 **a** Boron has a mass number of 10.8 on the periodic table. Naturally occurring boron has two isotopes. One has a mass of 10.013 and the second has a mass of 11.009. **Determine** the percentage abundance of the two isotopes.

b Copper has a mass number of 63.54 on the periodic table. Naturally occurring copper has two isotopes. One has a mass of 62.93 and the second has a mass of 64.93. **Determine** the percentage abundance of the two isotopes.

Investigate, evaluate and communicate

5 **Evaluate** why the carbon-12 isotope is a better reference than the carbon-13 isotope for the analysis of isotopes.

6 **Discuss** the meaning of mass-to-charge ratio and how ions are analysed in a mass spectrometer.

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SCIENCE AS A HUMAN ENDEAVOUR

5.4 Distribution of elements in the universe

KEY IDEAS

In this section, you will learn about:

- the analysis of the stars by using electromagnetic spectra, specifically their speed, temperature and composition.

Spectroscopy provides valuable information about the universe. Scientists have long been interested in the speed, composition and temperature of stars, galaxies and various other phenomena, that are well beyond our reach. A scientific understanding of the universe will lead to a better understanding of the life of our own star, the Sun.

Carbon-based organisms

Living organisms are often referred to as 'carbon-based' life forms because most of the molecules that form these organisms are made from elements bonded to carbon atoms. However, this does not mean that carbon is the most abundant element.

Within the human body, the most abundant element (or the most common element by mass) is oxygen. Approximately 65% of the human body is oxygen, whereas carbon only comprises 18.5%. The next two abundant elements are hydrogen (9.5%) and nitrogen (3.2%). Other elements occur in trace (very small) amounts; for example, calcium (1.5%), potassium (0.4%), phosphorus (1.0%) and sulfur (0.3%).

Our planet Earth

Different regions of Earth have different elements. The sea has a different elemental composition from Earth's crust, core and atmosphere.

The most abundant element in the atmosphere is nitrogen. The most abundant element in the oceans and Earth's crust is oxygen. The Earth's core consists mainly of an alloy of nickel and iron. Therefore, it can be misleading to state an overall composition of Earth because it does not reflect the differences between its many layers.

Overall, the most abundant elements on Earth are iron, oxygen, silicon, magnesium, sulfur, nickel, calcium and aluminium. These are all elements found in rock-based formations and it is unsurprising that oxygen has such a high abundance when it is located in rock-based compounds, water and the atmosphere.

Minerals in the Earth's crust

Based on what you have learned about common elements in the Earth's crust. What do you think the most common minerals are in the Earth's crust?

The universe

Scientists have long been interested in the speed, composition and temperature of stars, galaxies and various other phenomena, well beyond our reach.

Doppler effect

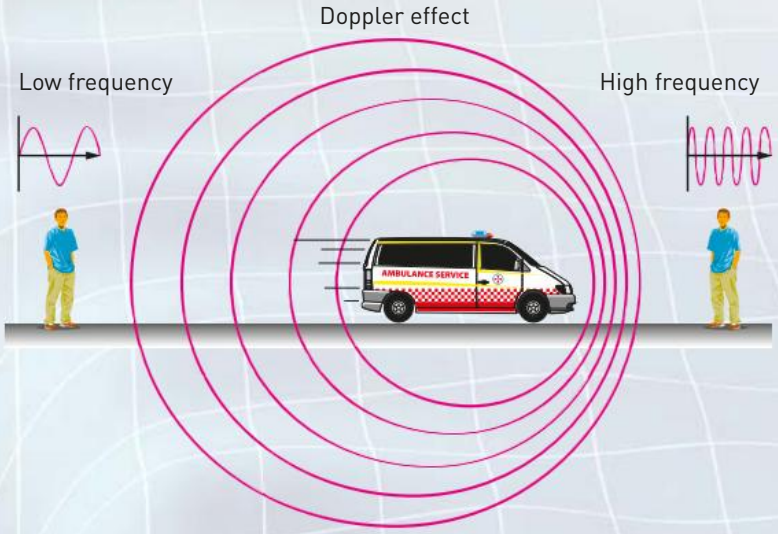

FIGURE 1 The Doppler effect explains why an ambulance siren sounds different to a pedestrian as the ambulance goes past.

The **Doppler effect**, named for Austrian physicist Christian Doppler, is essentially the change in frequency or wavelength observed by a bystander as the source of the wave moves past them. It is best observed when an ambulance passes a pedestrian on the side of the road. As the ambulance approaches, the pedestrian hears the sound of the ambulance siren as a high-pitched (and therefore high-frequency) 'yeee' sound. As the ambulance passes, the sound changes to a low-pitched (and therefore low-frequency) 'owww' sound. Therefore the pitch and frequency of sound shift relative to the motion of the ambulance and the position of the pedestrian.

doppler effect
the change in a frequency of a wave as an object moves towards or away from an observer

The light from moving stars can also be observed to shift in frequency. An object moving towards Earth emits radiation that becomes squeezed. Therefore, the wavelength decreases and the frequency increases. As a result, its absorption spectrum shifts slightly towards the blue end. This is called **blue shift** (Figure 2). Conversely, an object moving away from Earth emits radiation that spreads further apart. The wavelength increases and the frequency decreases. As a result, its absorption spectrum shifts slightly toward the red, resulting in **red shift** (Figure 2).

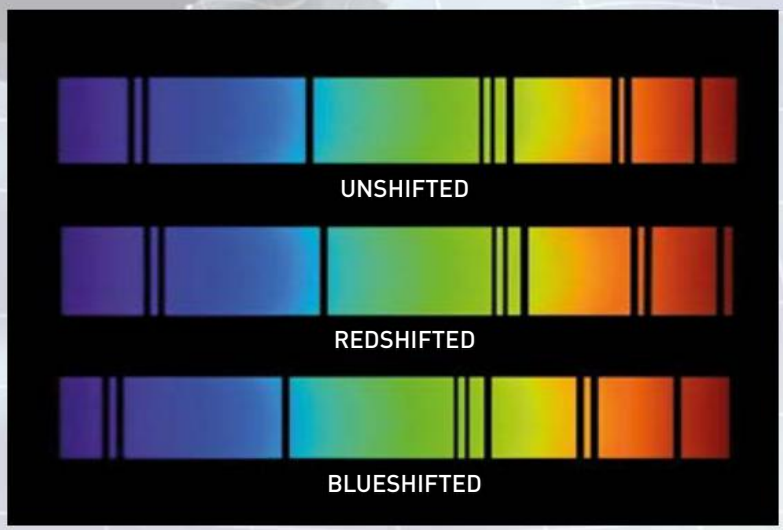

FIGURE 2 The blue shift of galaxies moving towards the Earth and the red shift of galaxies moving away from the Earth.

blue shift
a shift in the spectrum towards the blue region, indicating that a galaxy is moving away from the observer

red shift
a shift in the spectrum towards the red region, indicating that a galaxy is moving away from the observer

These effects are summarised in Table 1.

TABLE 1 The correlation between pitch, frequency and wavelength and its relation to the movement of galaxies

Motion	Sound and the Doppler effect	Light and the Doppler effect
Approaching	The high-pitched 'yeee' shortens the wavelength, increasing the frequency, of the sound of an approaching ambulance.	Lower wavelength light is at the blue end of the visible spectrum. Therefore, light absorbed shifts slightly lower on the spectrum (Figure 2). This indicates that the galaxy is moving towards Earth.
Moving away	The low-pitched 'owww' lengthens the wavelength, decreasing the frequency, of the sound.	Higher wavelength light is in the red region of the visible spectrum. Therefore, light absorbed shifts slightly higher on the spectrum (Figure 2). This indicates that the galaxy is moving away from Earth.

The vast majority of the galaxies in the universe experience red shift, implying that they are moving away from each other and the universe is expanding. This provides evidence for the Big Bang theory.

Temperature of stars

A star's temperature is determined from its colour. A star with a higher temperature will emit more light, and therefore more energy. On the electromagnetic spectrum, shorter wavelengths have more energy; therefore, hotter stars correlate to blue wavelengths.

Colder stars emit less light and therefore less energy. Lower energy on the electromagnetic spectrum results in longer wavelengths; therefore, colder stars appear red.

Chemical composition of stars

To determine the chemical composition of a star, a spectroscope is used. The resultant absorption spectra show dark lines on the continuous spectra. These dark lines are referred to as Fraunhofer lines.

Figure 3 shows a detailed absorption spectrum of the Sun. The Fraunhofer lines represent the electron transitions of each element. By comparing the characteristic fingerprint of absorption lines produced by individual elements to the Sun's absorption spectrum, scientists can identify the elements contained within the Sun.

The most common elements identified in stars are hydrogen, helium, calcium and iron. As the lightest and most common element in the universe, hydrogen makes up most stars. Helium is the next most common.

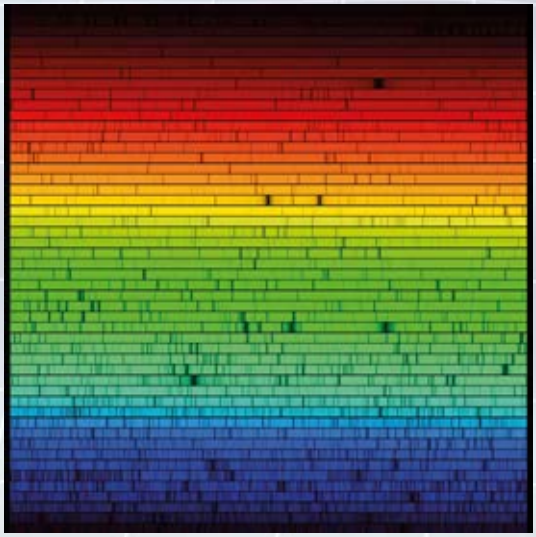

FIGURE 3 The absorption spectrum of the Sun.

Study tip

It is important to link concepts. The electromagnetic spectrum consists of various categories of light, including coloured light, energy, frequency and wavelength. Therefore, any of these concepts can be linked to another to develop deeper understanding.

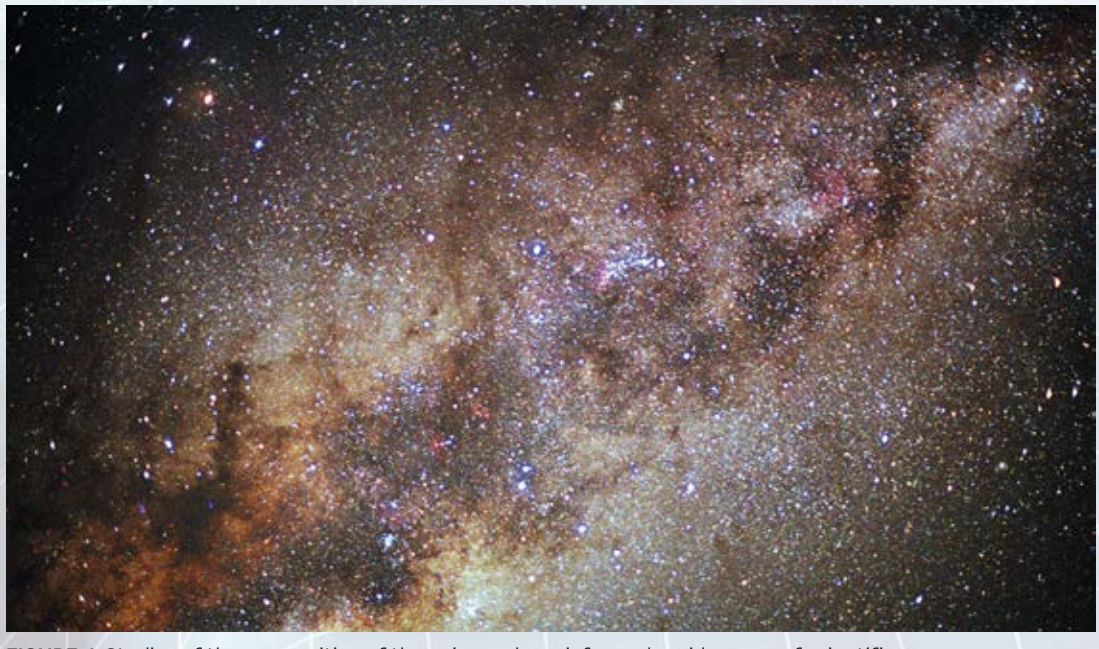

FIGURE 4 Studies of the composition of the universe have informed a wide range of scientific understandings.

CHECK YOUR LEARNING 5.4

Describe and explain

1 **Describe** the Doppler effect.

2 **Explain** how the colour of a star is linked to its temperature.

3 The following questions refer to the Doppler effect.

- a **Describe** how pitch, frequency and wavelength of an ambulance siren changes as it races towards you.
- b **Describe** how pitch, frequency and wavelength of an ambulance siren changes as it races away from you.

4 Figure 5 shows the emission spectrum of carbon.

- a **Sketch** the carbon emission spectrum in red shift.
- b **Sketch** the carbon emission spectrum in blue shift.

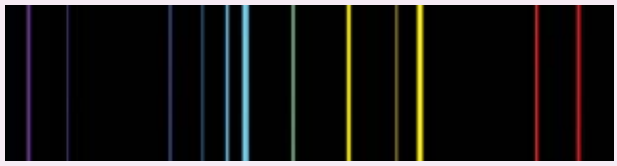

FIGURE 5 The emission spectrum of carbon.

Apply, analyse and interpret

5 **Consider** the ambulance driver in the Doppler effect example. **Infer** whether or not the driver hears the siren's change in pitch.

Investigate, evaluate and communicate

6 **Examine** Figure 3 and determine which Fraunhofer lines contained within the Sun's absorption spectrum represent the element hydrogen.

7 The Doppler effect is a simple principle to demonstrate in a classroom. **Design** an experiment that you could use to demonstrate the Doppler effect to your peers.

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CHAPTER 5

Review

Chapter summary

5.1
- Electrons orbit an atom's nucleus in shells that have discrete energy levels according to the Bohr model of the atom. To transition to a higher energy level, electrons must absorb a photon of light with specific wavelength. Electrons are then in an excited state. An electron will emit a photon of light of specific wavelength when it returns to its ground state. The light both absorbed and emitted can be measured on the electromagnetic spectrum.

5.2
- The flame test is an analytical technique in which electrons are excited by heat from the blue flame of a Bunsen burner. When they return to ground level, the electrons emit a photon of light of a specific wavelength, which is measured on an emission spectrum.
- AAS is a technique that atomises a sample and shines light of a specific wavelength at the atoms. The light excites electrons to a higher electron shell and the electrons emit light when returning to the ground state. The emitted light is measured at a specific wavelength and the resultant absorbance is measured against concentration.
- To determine the concentration of a sample in AAS, a series of standards of known concentration must be analysed. The absorbance is recorded and used to create a calibration curve. The sample's absorbance is plotted on the calibration curve to determine the concentration.

5.3
- Mass spectrometry ionises a sample of an element and creates a spectrum according to the mass and abundance of the sample's isotopes. These are used to calculate the relative atomic mass of an element.
- If the element contains no more than two isotopes and A_r and RIM are known for both, then the percentage composition of the element can be calculated. The isotopes' percentage abundances are represented by x and $100 - x$, as they must add up to 100. Once substituted into the equation, solve for x.

5.4
- The electromagnetic spectrum can be interpreted to determine the speed and direction of galaxy movement, the temperature of stars and the composition of stars.

Key terms

- absorption spectrum
- atomic absorption spectroscopy (AAS)
- blue shift
- calibration curve
- Doppler effect
- electromagnetic spectrum
- emission spectrum
- excited state
- flame test
- frequency
- ground state
- ionisation
- mass spectrometer
- mass spectrometry
- mass spectrum
- percentage abundance
- photon
- red shift
- relative atomic mass (A_r)
- relative isotopic mass (RIM)
- standard
- wavelength

Key formulas

relative atomic mass

$$A_r = \frac{(\text{RIM}_1 \times \text{abundance}_1(\%)) + (\text{RIM}_2 \times \text{abundance}_2(\%)) + (\text{RIM}_3 \times \text{abundance}_3(\%)) + \ldots}{100}$$

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; ***=high.

Multiple choice

1 Which of the following is inconsistent with Bohr's model of electron configuration?

- **A** Electrons in atoms can be at different distances from the nucleus.
- **B** Electrons cannot move into different shells.
- **C** Electrons orbit the nucleus of an atom.
- **D** For an electron to transition between energy levels, it must absorb or emit energy.

2 When an atom absorbs electromagnetic radiation of a specific frequency, electrons:

- **A** spin faster around the nucleus of an atom
- **B** move to an electron shell further from the nucleus
- **C** spin in the opposite direction
- **D** move into an electron shell closer to the nucleus.

3 An emission spectrum is produced when:

- **A** electrons orbiting the nucleus are promoted to a higher energy level
- **B** heat from a Bunsen burner is used to excite electrons
- **C** electrons leave an atom's electron shells and an ion is formed
- **D** electrons shed excess energy by releasing it as light.

4 On the electromagnetic spectrum, visible light is measure between the wavelengths of:

- **A** 400 and 750 nm
- **B** 300 and 700 nm
- **C** 500 and 800 nm
- **D** 200 and 600 nm.

5 What information can a mass spectrometer provide to a chemist?

- **A** the composition of elements in a sample
- **B** the concentration of atoms in a sample
- **C** the properties of molecules within a sample
- **D** the relative mass of atoms within a sample

Short answer

Describe and explain

★ **6** **Explain** how absorption and emission spectra occur.

★ **7** **Explain** the difference and the relationship between relative isotopic mass (RIM), relative atomic mass (A_r) and mass number.

Apply, analyse and interpret

★ **8** **Differentiate** between the Lyman, Balmer and Paschen series.

★★★ **9** **Determine** the atomic mass number of the following elements from the isotope data given.

a Silicon has three isotopes.

TABLE 1

Isotope	RIM	% abundance
^{28}Si	27.977	92.229
^{29}Si	28.976	4.683
^{30}Si	29.974	3.087

b Europium has two isotopes.

TABLE 2

Isotope	RIM	% abundance
^{151}Eu	150.9198	47.81
^{153}Eu	152.9212	52.19

c Zinc has five isotopes.

TABLE 3

Isotope	RIM	% abundance
^{64}Zn	63.929	48.63
^{66}Zn	65.926	27.90
^{67}Zn	66.927	4.10
^{68}Zn	67.925	18.75
^{70}Zn	69.925	0.62

★★★ **10** **Determine** the percentage abundance of the isotopes, given the following information.

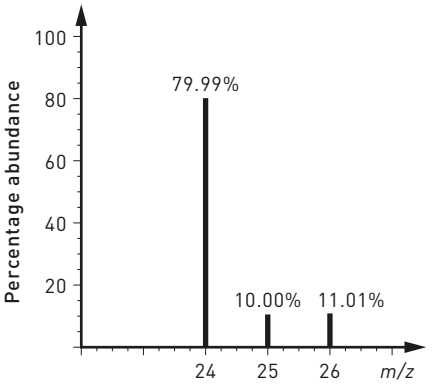
a Bromine has a mass number of 79.9. Its isotopes have relative isotopic masses of 78.918 and 80.916.

b Indium has a mass number of 114.8. Its isotopes have relative isotopic masses of 112.904 and 114.904.

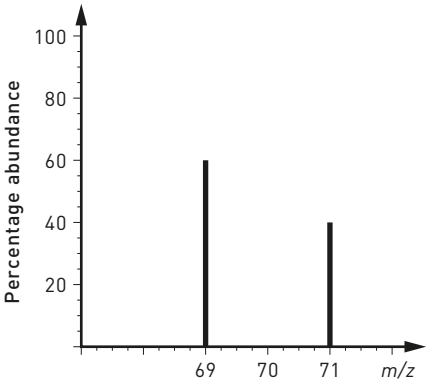
c Rubidium has a mass number of 85.5. Its isotopes have relative isotopic masses of 84.912 and 86.909.

★★★ **11 Determine** the atomic mass number from the following mass spectra.

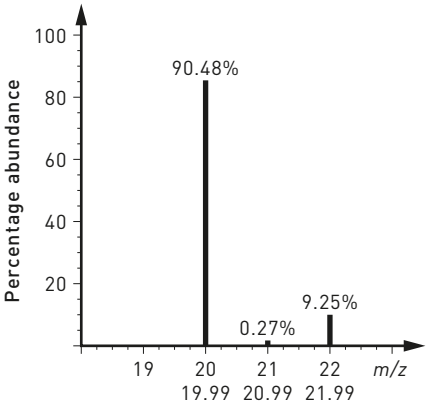
a Mass spectrum for magnesium



b Mass spectrum for gallium



c Mass spectrum for neon



Investigate, evaluate and communicate

★★ **12 Determine** the atomic mass number for the following elements from the isotope data given.

a Isotope data for titanium

TABLE 4

Isotope	RIM	% abundance
^{46}Ti	45.953	8.25
^{47}Ti	46.952	7.44
^{48}Ti	47.948	73.72
^{49}Ti	48.948	5.41
^{50}Ti	49.945	5.18

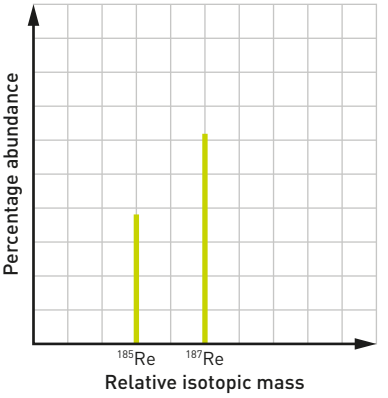
b Isotope data for chromium

TABLE 5

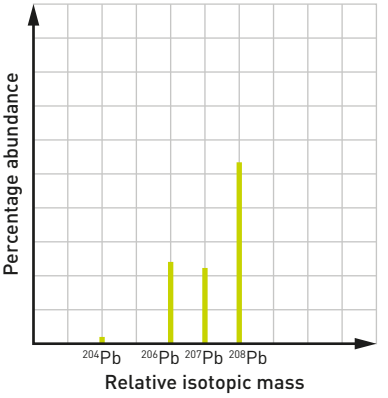
Isotope	RIM	% abundance
^{50}Cr	49.946	4.345
^{52}Cr	51.941	83.789
^{53}Cr	52.941	9.501
^{54}Cr	53.939	2.365

★★ **13 Determine** the relative atomic mass of the following elements using the mass spectra shown. (Hint: For a more accurate scale, use a ruler).

a Mass spectrum for rhenium



b Mass spectrum for lead



★★ **14** Why is AAS primarily used for the analysis of metals and not non-metals? **Investigate** the light source involved with this technique to **explain** your answer.

★★★ **15 a** The magnesium content of blood is analysed against a set of standards and the data in Table 6 is obtained. **Develop** a calibration curve to determine the concentration of magnesium.

TABLE 6

Standard/sample	Concentration (mg L^{-1})	Absorbance
Standard 1	0	0
Standard 2	4	23
Standard 3	8	46
Standard 4	12	69
Standard 5	16	92
Sample	–	53

b The calcium concentration in a sample of milk is analysed against a set of standards and the data in Table 7 is obtained. **Develop** a calibration curve to determine the concentration of calcium

TABLE 7

Standard/sample	Concentration (mmol L^{-1})	Absorbance
Standard 1	0	15
Standard 2	10	30
Standard 3	20	45
Standard 4	30	60
Standard 5	40	75
Sample	–	20

★★★ **16 a** **Evaluate** what information the Lyman, Balmer and Paschen series provide about the structure of an atom.

b **Describe** what the key difference is between the Lyman, Balmer and Paschen series?

c **Determine** how each series relates to the electromagnetic spectrum.

d **Identify** which series emits light in the visible region of the electromagnetic spectrum.

★★★ **17** **Investigate** the use of AAS to measure the concentration of Ca^{2+} in egg shells.

a What steps are required to prepare the egg shell samples?

b How would you quantitatively analyse the samples?

c Assuming that the concentration of Ca^{2+} was lower than expected, what effects may this have on the egg shells?

★★★ **18** AAS is a valuable analytical technique.

a **Discuss** what the technique analyses.

b **Explain** why the hallow cathode lamp must be set to a specific wavelength.

c **Describe** what happens to atoms during AAS and how this is detected.

d Could the same lamp be used for different elements? **Explain** why.

e **Explain** whether, for the same element, the wavelengths for absorption and for emission would be the same.

★★★ **19** **Determine** the percentage abundance of the isotopes, given the following information.

a Lithium has a mass number of 6.9. Its isotopes have relative isotopic masses of 6.015 and 7.016.

b Iridium has a mass number of 192.2. Its isotopes have relative isotopic masses of 190.96 and 102.96.

c Thallium has a mass number of 204.4. Its isotopes have relative isotopic masses of 202.97 and 204.97.

★★★ **20** **Investigate** Planck's constant (6.63×10^{-34} J) and **discuss** the relationship between energy and frequency.

★★★ **21** **Investigate** an occupation that uses the electromagnetic spectrum to analyse elements or molecules. **Explore** what is being analysed and **discuss** why this analysis necessary.

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CHAPTER 6

Compounds and mixtures

All matter is made of atoms. For pure elements, the matter contains the same type of atoms, existing as individual atoms or atoms held together by chemical bonds. When a compound is formed, different types of atoms are held together by chemical bonds. A mixture is a substance that is made from a combination of elements and/or compounds that do not react chemically.

When producing materials, it is essential to know the chemical composition of those materials. Scientists involved in quality control follow strict testing procedures to ensure that the chemical and physical properties of production materials are within quality assurance guidelines.

In recent years, there has been huge growth in the application of nanotechnology into science and engineering fields, facilitating the design of materials at almost a molecular level.

This chapter focuses on the physical properties of materials, and how those properties can be used for identification and separation of substances.

OBJECTIVES

- → Recall that pure substances may be elements or compounds.
- → Recognise that materials are either pure substances with distinct measurable properties (e.g. melting and boiling point, reactivity, strength, density) or mixtures with properties dependent on the identity and relative amounts of the substances that make up the mixture.
- → Distinguish between heterogeneous and homogeneous mixtures.
- → Recognise that nanomaterials are substances that contain particles in the size range 1–100 nm and have specific properties relating to the size of these particles.
- → Analyse and interpret given data to evaluate the physical properties of pure substances and mixtures.

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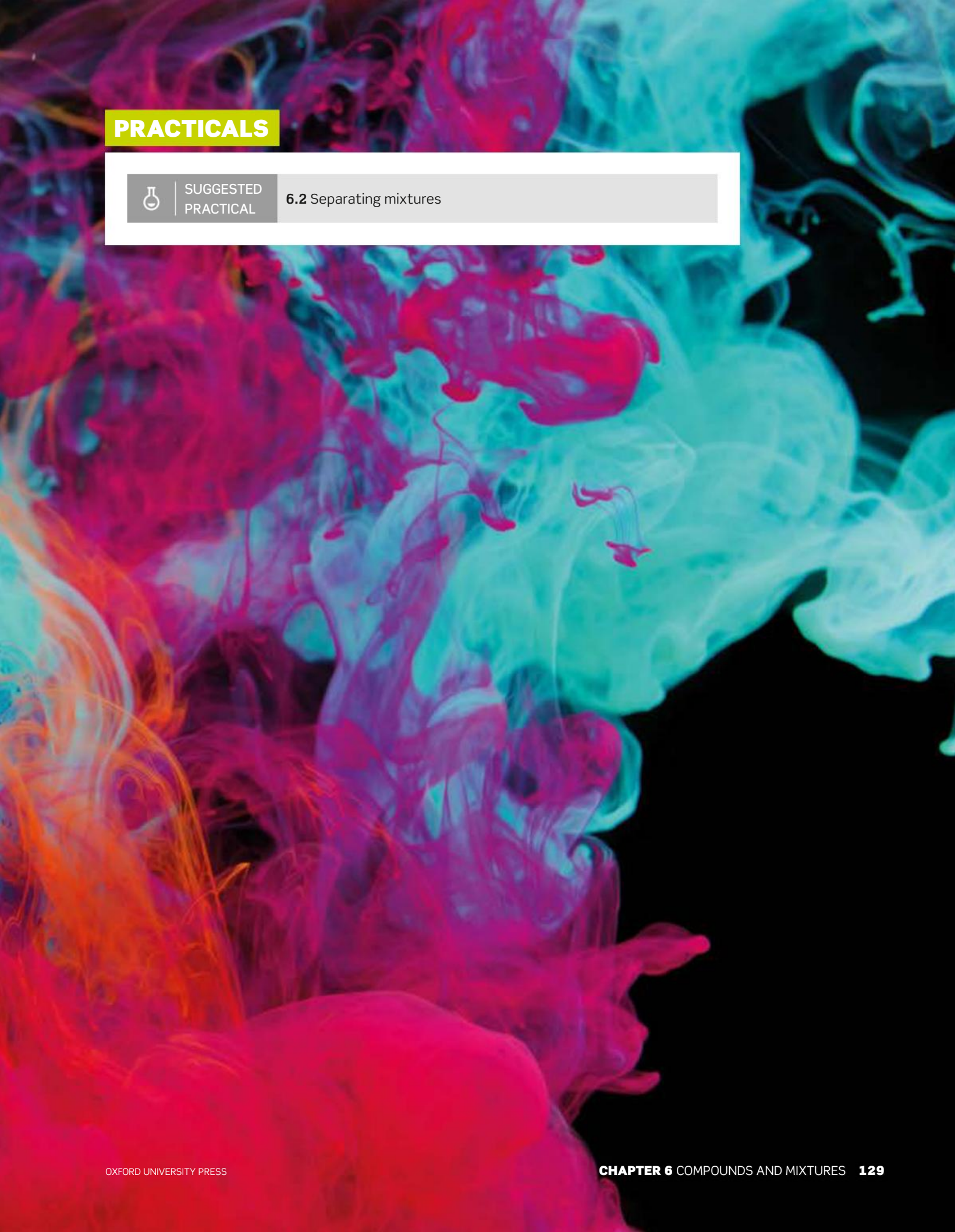
FIGURE 1 This image shows coloured inks as they begin to mix with water. Although it is colourful now, soon the colours will blend together, forming a brown homogeneous mixture with the same composition throughout the mixture.



PRACTICALS

SUGGESTED PRACTICAL

6.2 Separating mixtures

6.1 Pure substances

KEY IDEAS

In this section, you will learn about:

- pure substances, which are made up of elements or compounds
- how pure substances have distinct, measurable properties, including melting point, boiling point, reactivity, strength and density.

mixture
a substance that is made up of two or more substances that are not chemically bonded

pure substance
a substance that is made up of a single type of element or compound

element
a substance that is made up of one type of atom, defined by its atomic number

compound
a substance that is made up of atoms of different elements, joined together by chemical bonds

The world is made up of substances. The glass on a phone screen, the plastic wrapping on food and soil: these items are all made up of **mixtures** of substances.

Pure substances are not mixed with any other type of substance; they are made up of a single type of **element** or **compound** (Figure 1). In other words, a substance that contains only one type of element or one type of compound is considered pure. A pure substance has no impurities and can exist in different phases (i.e. solid, liquid or gas).

For example, water molecules grouped together are considered a pure substance. However, if sodium chloride (NaCl) is added to the water, it is no longer a pure substance because more than one kind of compound is present.

Table 1 summarises the key features of elements and compounds, which were discussed in Chapter 2.

TABLE 1 The key features of elements and compounds

	Description	Examples
Element	Made up of one type of atom	Oxygen (O), hydrogen (H) and carbon (C)
Compound	Made up of atoms of different elements, joined together by chemical bonds	Water (H_2O) and carbon dioxide (CO_2)

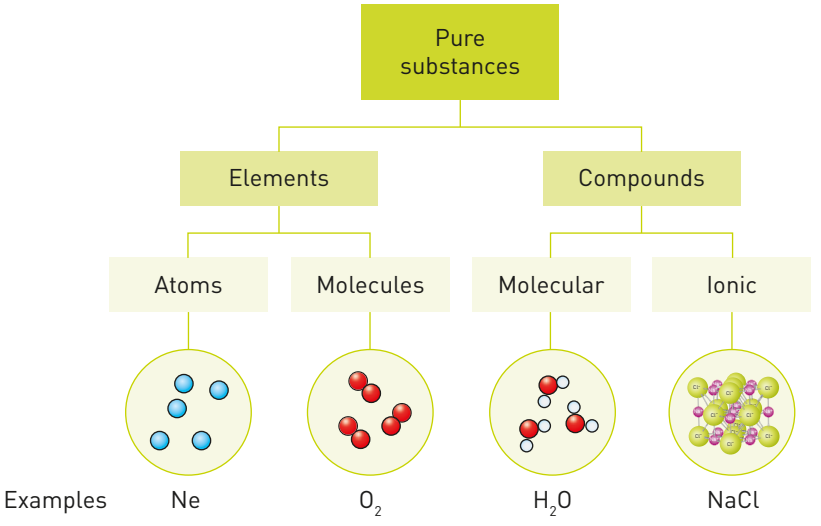


FIGURE 1 Pure substances consist of elements or compounds.

melting point
the temperature at which a solid becomes a liquid

boiling point
the temperature at which a liquid becomes a gas

density
a measure of the mass contained within a given volume

reactivity
a measure of how readily a substance reacts with other substances

Properties of pure substances

Pure substances have distinct, measurable properties. The purity of a substance can be established by collecting or analysing data of its physical and chemical properties, such as **melting point**, **boiling point**, and **density**. Pure substances have characteristic **reactivities** and bond strengths.

TABLE 2 Physical properties of some substances under standard laboratory conditions

Element	Physical state	Melting point (°C)	Boiling point (°C)	Density (g cm^{-3})
Water (H_2O)	Liquid	0	100	0.997
Oxygen	Gas	−218.79	−182.962	1.429×10^{-3}
Nitrogen	Gas	−210	−195.975	1.2504×10^{-3}
Mercury	Liquid	−38.8290	356.73	13.534
Iron	Solid	1538	2862	7.874

Melting point and boiling point

Melting point refers to the temperature at which a solid substance becomes a liquid. A pure solid substance should have a fixed melting point. In other words, if the substance is pure, all of the substance should melt at the same temperature.

Boiling point refers to the temperature at which a liquid substance becomes a gas.

Impurities in a substance typically affect its melting and boiling points because the different compounds may have a range of melting and boiling points.

For example, the boiling point of pure water is 100°C at a pressure of 1 atmosphere. However, any impurities in the water can affect the boiling point (Figure 2). This is referred to as a **colligative property**, and is related to the ratio of the number of solute particles to the number of solvent particles. It does not depend on the nature of the chemicals dissolved.

colligative property
a property of a solution that depends on the ratio of the number of solute particles to the number of solvent molecules, but does not depend on the identity of the solute

Reactivity

Reactivity refers to how readily a substance reacts with other substances. During a chemical reaction, reactant bonds are broken and product bonds are formed. The potential energy changes associated with these processes determine whether the reaction is favourable and can influence the rate at which the chemical reaction occurs.

Bond strength

The chemical bonds that are present in a substance directly influence the properties of that substance. For example, sodium chloride contains strong ionic bonds. Sodium chloride is a hard substance with high melting and boiling points. Untreated rubber is a soft, elastic material, which reflects the nature of the bonding present in the substance.

Density

Density is the mass per unit volume of a substance. The density of a substance is based on the relative atomic masses of the nuclei present and how compact the atoms or ions are. If the mass and volume of a substance are known, the following formula can be used to calculate density:

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

It is important to note that the units of density depend on the units of mass and volume used in the calculation. If mass is measured in grams and the volume is measured in cubic centimetres, the units of density are grams per cubic centimetre (g cm^{-3})

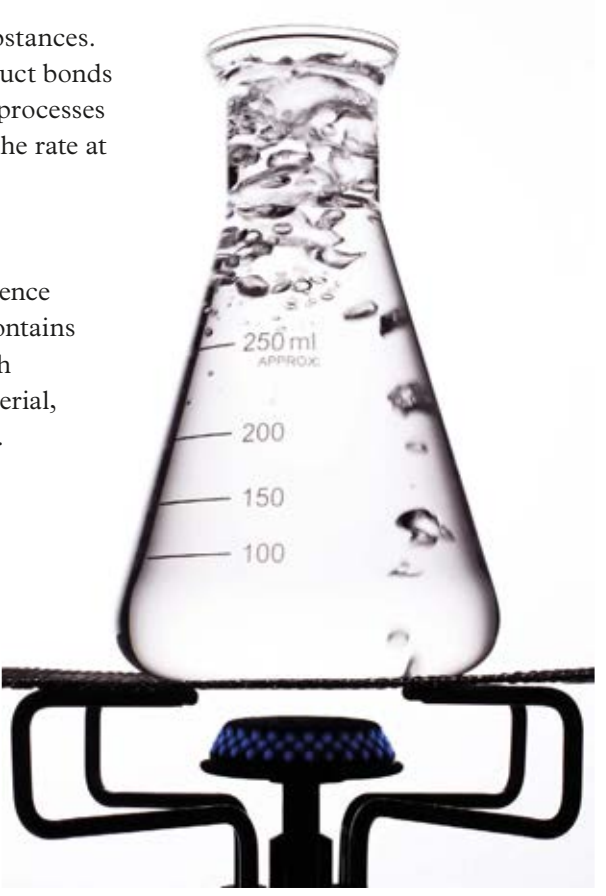


FIGURE 2 The boiling point of water can be affected by impurities in the water.

CHECK YOUR LEARNING 6.1

Describe and explain

1 **Describe** a pure substance and give an example.

2 **Explain** how the density of a known substance such as aspirin can assist to establish its purity.

Apply, analyse and interpret

3 Contrast pure substances comprised of elements with pure substances comprised of compounds. **Provide** an example of each substance.

4 An impure sample melts over a range of temperatures from 101°C to 117°C. **Deduce** the temperature the sample would likely melt at, if it were pure.

5 **Analyse** the melting points of the unknown substances in Table 3.

   a **Determine** which of the substances is most likely to be pure water.

   b **Interpret** the melting point of substance B and compare it with the melting points of the other substances listed.

TABLE 3 Melting points of some unknown substances

Substance	Melting point (°C)
A	–114
B	8–26
C	0
D	6.5

Investigate, evaluate and communicate

6 In 1844, Charles Goodyear patented a process known as vulcanisation. **Investigate** this process and its impact on the production of the pneumatic tyre.

FIGURE 3 Charles Goodyear

7 In Scotland, rock salt is spread on roads during the winter months. **Evaluate** this practice, in terms of the freezing point of water.

FIGURE 4 A frozen road in the Scottish highlands.

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6.2 Mixtures

KEY IDEAS

In this section, you will learn about:

- how the properties of mixtures depend on the identity and relative amounts of their components
- heterogeneous and homogeneous mixtures.

A mixture is made up of two or more substances that are not chemically bonded. The substances in a mixture retain their individual properties. Therefore, the properties of a mixture depend on the identity and amount of each substance that is a component of the mixture. For example, fried rice is a mixture that can be easily separated into its different components. Mixtures are different from pure substances in a number of ways.

- Mixtures can be physically separated.
- Mixtures have not undergone a chemical reaction to combine or bond their components.
- The properties of a mixture depend on the substances that form the mixture.
- The composition of a mixture is not always the same throughout the mixture.

Mixtures are also different from compounds, whose substances are chemically bonded together (Table 1). For example, water is made up of hydrogen and oxygen, but it is not considered a mixture because the elements are chemically bonded together; therefore, water is considered a compound.

TABLE 1 Differences between a mixture and a compound

	Structure	Properties	Separation
Mixture	The amount of each substance in a mixture can vary. The different substances are not chemically bound.	The substances in a mixture retain their own properties.	A mixture can be separated on the basis of the physical properties of its component substances.
Compound	The amount of each element in a compound does not vary. The different elements are chemically bonded.	A compound has properties different from those of its component's elements.	A compound cannot be separated on the basis of its physical properties.

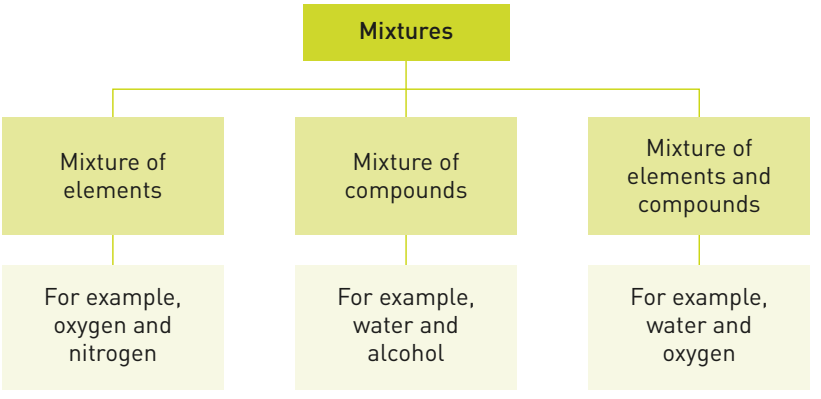

FIGURE 1 Mixtures can contain a mixture of elements, a mixture of compounds or a mixture of elements and compounds

Homogeneous and heterogeneous mixtures

Mixtures can be classified as homogeneous or heterogeneous as shown in Figure 2.

Homogeneous mixtures

homogeneous mixture
a mixture with a uniform composition

A **homogeneous mixture** has a uniform composition. It is made up of substances of the same state (gas, liquid or solid), and appears consistent throughout the mixture. A sample taken from a homogeneous mixture would appear the same and have the same composition, no matter where the sample was taken from in the mixture. For example, brewed tea, salt solution or wine are all mixtures that contain a number of substances but have a uniform composition. A **solution**, such as cordial, is also a homogeneous mixture.

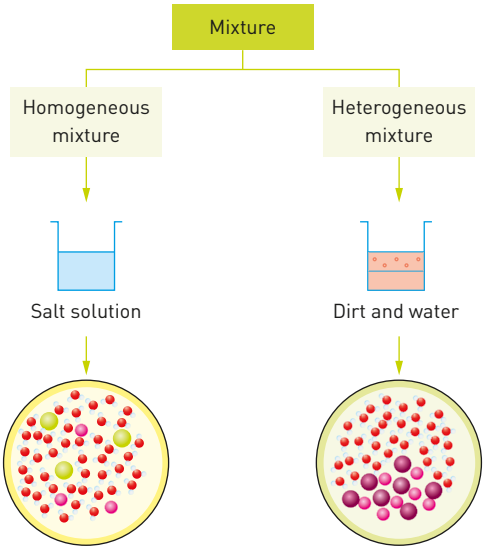

FIGURE 2 A salt solution is an example of a homogeneous mixture, and dirty water is an example of a heterogeneous mixture of soil and water.

solution
a mixture of a solute dissolved in a solvent

Heterogeneous mixtures

heterogeneous mixture
a mixture with an inconsistent composition

A **heterogeneous mixture** has an inconsistent composition. For example, a mixture of oil and water is heterogeneous mixture the two substances do not mix well, and their different physical properties are clearly identifiable (Figure 3a). If a sample was taken from the top of the mixture, it is likely to display the properties of the oil at the top, rather than properties that represent the whole mixture.

Foods such as salads, muesli or a bowl of mixed lollies are heterogeneous mixtures (Figure 3b). In these mixtures, it is clear that there are a number of substances present, which can be easily separated, and one sample of the mixture may have different properties from another sample of the mixture.

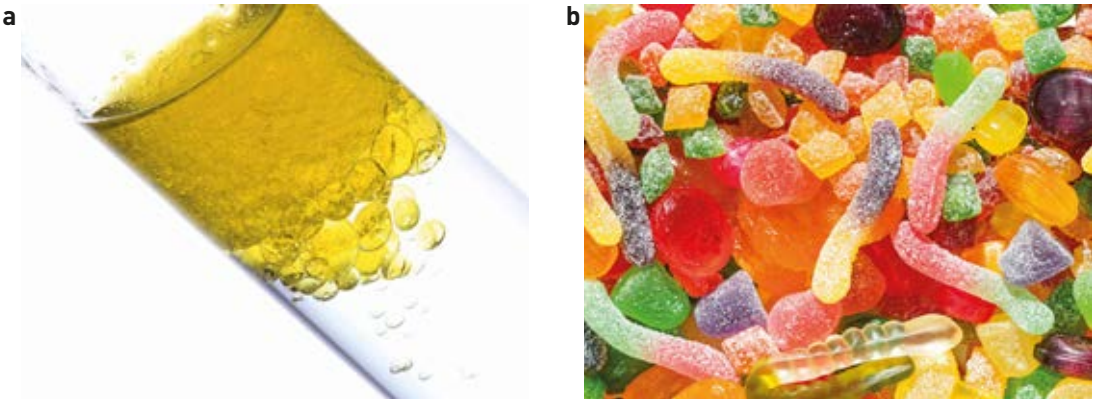

FIGURE 3 Two different heterogeneous mixtures: (a) oil and water and (b) mixed lollies.

colloid
a homogeneous mixture in which dispersed particles do not separate over time

Table 2 shows the different types of mixtures that include solutions, suspensions and **colloids**.

TABLE 2 Different types of mixtures

Type of mixture	Description	Example	Homogeneous or heterogeneous
Solution	A mixture in which a substance is dissolved evenly in another substance	Sugar in water	Homogeneous
Suspension	A mixture in which a substance is dispersed (but not dissolved) in another substance; will separate over time	Oil and sand in water	Heterogeneous
Colloid	A mixture in which dispersed particles do not separate over time	Hair gel	Homogeneous

Using physical properties to separate mixtures

The physical properties of substances, outlined in Section 6.1, can be used to separate substances from mixtures. Therefore, in order to separate mixtures, it is useful to understand the physical properties of each of the individual substances within the mixture.

Different techniques for separation, including **distillation**, **crystallisation by evaporation**, **filtration**, **centrifugation**, **decantation**, **flotation**, **sedimentation** and **physical separation**, have been developed on the basis of these physical properties. The technique chosen for separation will depend on the properties of the substance being separated from the mixture. Some of these techniques are outlined in Table 3.

distillation
the technique of separating a liquid from a liquid mixture by boiling the mixture until the liquid evaporates and is then collected by condensation

crystallisation by evaporation
the process of evaporating a solvent to separate a solid substance that is dissolved in a liquid mixture

filtration
the technique of pouring a liquid mixture through a filter so that the liquid passes through the filter, while the solid substance is caught in the filter

centrifugation
the technique for separating mixtures where tubes are placed into a centrifuge that spins in a circular motion very fast to push solids to the bottom

decantation
the technique by which liquid is carefully poured off the top of a liquid mixture to leave the undisturbed solid substances behind

flotation
the technique for separating different substances in a powdered ore depending on whether they sink in, or float on, a given liquid

sedimentation
the process by which dense solid substances in a liquid mixture sink to the bottom of a container over time because of gravity

physical separation
the technique by which substances in a mixture can be easily removed by physically picking them out of the mixture

TABLE 3 Techniques for separating mixtures on the basis of the physical properties of a substance

Physical property	Techniques for separation	Example
Boiling point	Distillation: A liquid can be separated from a liquid mixture by boiling the mixture until the liquid evaporates and is then collected by condensation (see Figure 4). The liquid substance with the lower boiling point evaporates first.	Separating ethanol from water in the production of spirits.
	Crystallisation by evaporation: Evaporating the solvent can separate a solid substance that is dissolved in a liquid mixture.	Collecting salt from a solution of salt and water by evaporating the water (solvent) to leave the salt crystals behind.
Density	Sedimentation and decantation: Dense solid substances in a liquid mixture will naturally sink to the bottom of a container over time because of gravity (see Figure 5a). The liquid can then be 'decanted' or carefully poured off the top to leave the undisturbed solids behind.	Separating soil from a mixture of soil and water, by allowing soil sink to the bottom of the mixture.
Size	Physical separation: Some substances in mixtures can be easily removed by physically picking them out of the mixture, either by hand or with a tool such as tweezers.	Picking out a chocolate chip from a mixture of cookie dough ingredients.
	Filtration: A solid substance can be separated from a liquid mixture by pouring the mixture through a filter (a porous barrier such as filter paper) (see Figures 5b and 6). The liquid passes through the filter, while the solid is caught in the filter.	Filtering tea leaves by using a strainer.

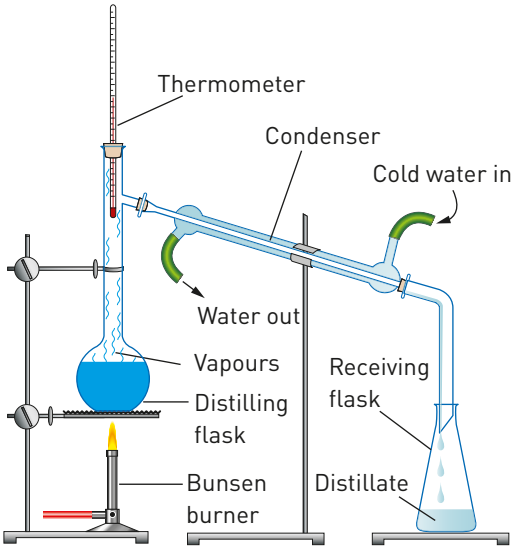

FIGURE 4 Distillation apparatus for boiling and then condensing liquids to separate a liquid mixture

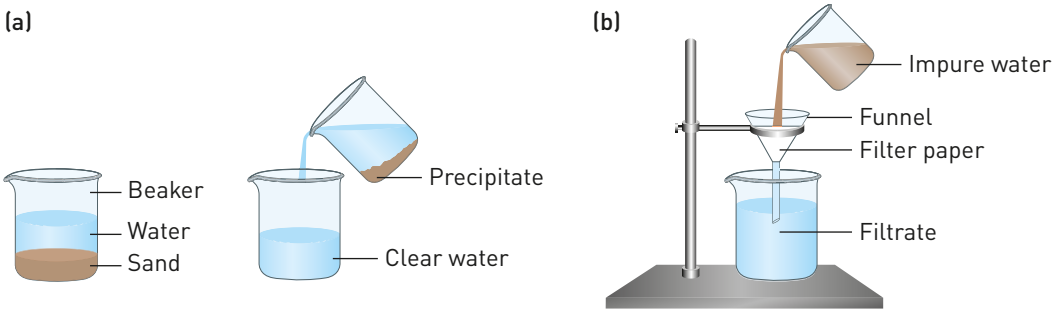

FIGURE 5 Techniques for separating heterogeneous mixtures.

FIGURE 6 A filtration set-up in the laboratory

CHECK YOUR LEARNING 6.2

Describe and explain

1 **Describe** a homogeneous mixture.
2 **Identify** two types of heterogeneous mixtures.
3 **Explain** why a bowl of fruit-and-nut muesli would not be considered a homogeneous mixture.
4 **Explain** why it is important to know the physical properties of substances when trying to separate a mixture.

Apply, analyse and interpret

5 **Contrast** homogeneous and heterogeneous mixtures. Give examples to support your answer.
6 **Consider** two other physical properties of substances (not already outlined in Table 3) that may be useful when separating mixtures.

Investigate, evaluate and communicate

7 **Justify** why air is considered a homogeneous mixture.
8 A student has a mixture of sand and vegetable oil. **Devise** a method that would allow the student to easily separate the two substances.
9 **Propose** a method to separate a mixture of water, sand and table salt.

Check your obook assess for these additional resources and more:

- » Student book questions
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- » Suggested practical
  6.2 Separating mixtures
- » Interactive
  Elements, compounds and mixtures
- » Weblink
  Mixtures

6.3 Nanomaterials

KEY IDEAS

In this section, you will learn about:

- nanomaterials, which are substances that contain particles 1–100 nm in size
- the specific properties of nanomaterials relating to the size of their particles.

nanomaterial
a substance that contains particles with at least one size measurement ≤100 nm

nanotechnology
the study of how the control of individual atoms and molecules in the synthesis of nanomaterials can be applied across science and engineering fields

Nanomaterials are substances that contain particles so small that at least one of the dimensions ranges between 1 and 100 nanometres (nm) (Figure 1). In other words, for a particle to be considered a 'nanomaterial', it should have at least one size measurement (such as length, width or height) ≤100 nm.

Nanomaterials are an important product in developing **nanotechnologies** and are therefore becoming more useful in health care, electronics and cosmetics.

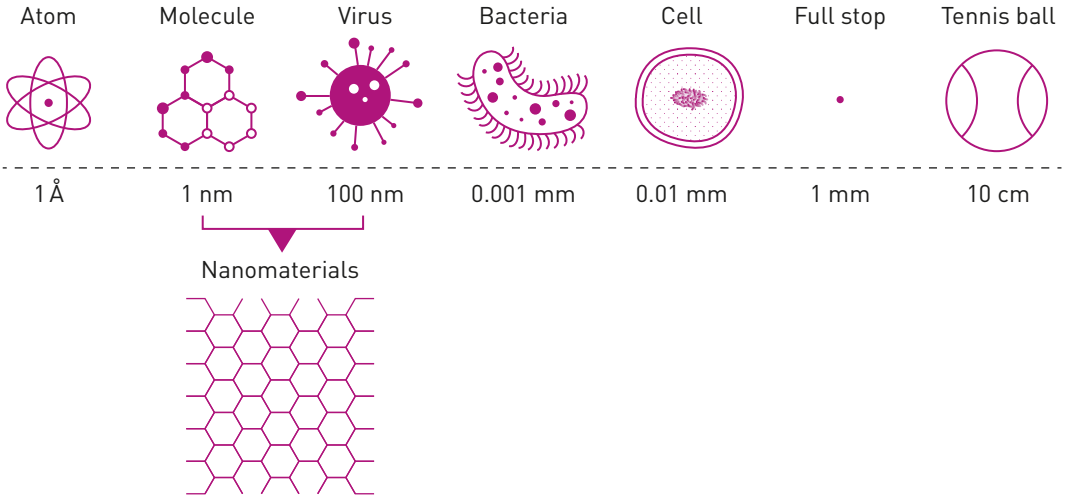


FIGURE 1 Examples of substances across a range of sizes

Properties of nanomaterials

Generally, a substance is considered to be an nanomaterial if 50% or more of the particles are in the size range 1–100 nm. Nanomaterials can have novel electronic, mechanical or optical properties. The ability to engineer nanoparticles, using atoms or simple molecules as building blocks, occurs in natural biological processes. The opportunity for scientists to tailor the synthesis of nanostructures in a similar fashion is a very exciting prospect.

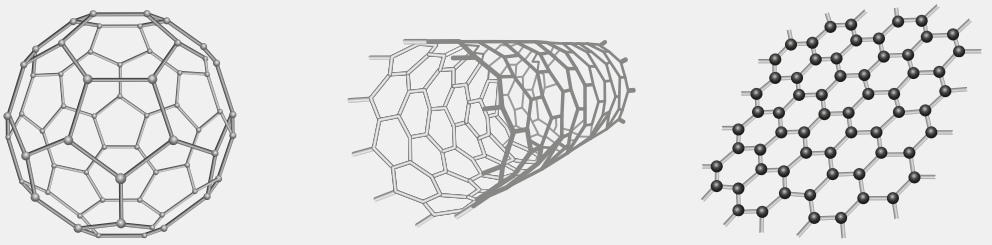
CASE STUDY 6.3

Fullerenes

The 1996 Nobel Prize in Chemistry was awarded to Robert Curl, Harold Kroto and Richard Smalley for their discovery of fullerenes. Prior to their work, there were two known allotropes of carbon – graphite and diamond. Curl, Kroto and Smalley found

that condensing carbon gas produced structures consisting of 60 and 70 carbon atoms. They discovered that the C_{60} structure was chemically inert and did not contain any other elements. Presuming that the 60 carbon atoms were based on a benzene-type bonding arrangement, they attempted to create a molecular structure based on hexagons, but were unsuccessful. It was only when they incorporated carbon pentagons with carbon hexagons that they were able to determine the structure of the molecule. The structure of the C_{60} structure is the same as a soccer ball, made of hexagons and pentagons. The scientists named the compounds 'fullerenes' in honour of Buckminster Fuller and his work on geodesic domes – strong dome-shaped buildings constructed from lattice works of polygons.

Graphite consists of layers of stacked graphene sheets. Individual graphene sheets can be synthesised, rolled into cylinders and capped with fullerenes to form carbon nanotubes, which are essentially elongated fullerenes. They have huge potential in materials science; carbon nanotube-based materials have demonstrated outstanding solid-lubricant properties, electrical conductivities and tensile strengths. There is continuing research into how carbon nanotubes can be incorporated into modern engineering.



Fullerene Nanotube Graphene

FIGURE 2 Fullerene, carbon nanotubes and graphene are just some of the forms that carbon can take.

CHECK YOUR LEARNING 6.3

Describe and explain

1 **Describe** nanomaterials.
2 **Explain** what is meant by 'nanotechnology'.

Apply, analyse and interpret

3 Nanotechnology in chemistry can be described as 'bottom-up' engineering. **Interpret** the meaning of this statement.
4 **Apply** your understanding of fullerene structure and reactivity to suggest how they could be used to safely carry drug molecules around the body.

Investigate, evaluate and communicate

5 **Evaluate** the statement: 'Fullerenes are the first "true" allotropes of carbon'.
6 **Investigate** the impact of a selected nanomaterial in the field of science or engineering.

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» Video
Nanotechnology used to detect bacteria and viruses

» Weblink
Nanoparticles

» Weblink
Nanotechnology

CHAPTER 6

Review

Chapter summary

6.1
- Pure substances are made up of elements or compounds. They have distinct, measurable properties, including melting point, boiling point, reactivity, bond strength and density. These physical properties can be used to identify pure substances.

6.2
- A mixture is made up of two or more substances. Properties of mixtures depend on the identity and relative amounts of their components.
- A homogeneous mixture has a uniform composition. A heterogeneous mixture has an inconsistent composition.
- Substances can be separated from mixtures on the basis of their individual properties by techniques such as distillation, crystallisation by evaporation, filtration, centrifugation, sedimentation and decantation, flotation and physical separation.

6.3
- Nanomaterials contain particles of sizes 1–100 nm. Their properties relate to the size of their particles.

Key terms

- boiling point
- colligative property
- colloid
- compound
- crystallisation by evaporation
- decantation
- density
- distillation
- element
- filtration
- flotation
- heterogeneous mixture
- homogeneous mixture
- melting point
- mixture
- nanomaterial
- nanotechnology
- physical separation
- pure substance
- reactivity
- sedimentation
- solution

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; *** = high.

Multiple choice

1 Pure water is best described as:
- **A** an element
- **B** heterogeneous
- **C** a compound
- **D** a solution.

2 Metals are generally hard substances because:
- **A** metallic bonds are strong
- **B** they are able to form ionic bonds
- **C** metal atoms are heavier than non-metal atoms
- **D** they are found on the left-hand side of the periodic table.

3 A 38.6 g sample of pure gold has a volume of 2.00 cm^3. The density of gold is:
- **A** 77.2 g cm^{-3}
- **B** 38.6 g cm^{-3}
- **C** 0.05g cm^{-3}
- **D** 19.3 g cm^{-3}

FIGURE 1 A gold nugget

4 A beaker of salty water contains undissolved crystals. This is best described as:
- **A** pure
- **B** dilute
- **C** saturated
- **D** heterogeneous

5 Nanoparticles are:
- **A** larger than a virus
- **B** less than 100 nm
- **C** conductors of electricity
- **D** smaller than an atom.

6 A sample of air is best described as:
- **A** an element
- **B** heterogeneous
- **C** a compound
- **D** a solution.

7 When a given volume of water is frozen, the ice formed has a greater volume than the original sample of water. This means that the:
- **A** number of H_2O molecules increases
- **B** sample of water is more concentrated
- **C** strength of the covalent bonds decreases
- **D** density of ice is less than the density of water.

FIGURE 2 Melting ice

8 The density of aluminium is 2.71 g cm^{-3}. If a sample of aluminium has a mass of 14.8 g, the volume of the sample is:
- **A** 0.183 cm^3
- **B** 40.1 cm^3
- **C** 5.46 cm^3
- **D** 12.1 cm^3

9 During a chemical reaction:
- **A** reactant bonds and product bonds are broken
- **B** reactant bonds and product bonds are formed
- **C** reactant bonds are broken and product bonds are formed
- **D** reactant bonds are formed and product bonds are broken.

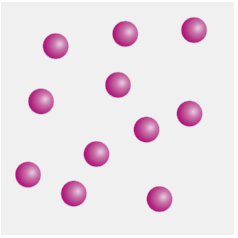
10 A C_{60} fullerene molecule:
- **A** is smaller than an atom
- **B** is an impure substance
- **C** contains only carbon atoms
- **D** is a hexagonal sheet of covalently bonded carbon atoms.

Short answer

Describe and explain

★ **11** **Identify** the following diagrams as showing pure substances, homogeneous mixtures or heterogeneous mixtures.

a



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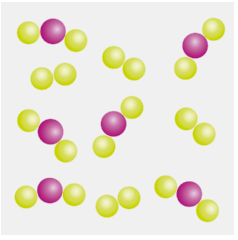

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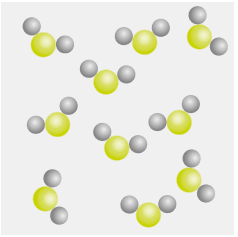

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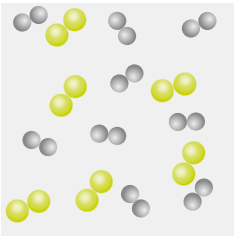
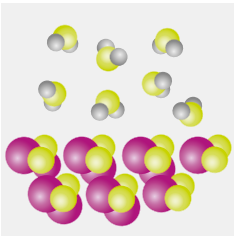


e



★ **12** **Explain** how you could separate a mixture of sand and sodium chloride (salt).

★ **13** **Describe** the structure of a fullerene.

★ **14** A student leaves a beaker containing a solution of saturated copper(II) sulfate on a window ledge in the laboratory. After a week, there are small blue crystals at the bottom of the beaker. **Explain** the appearance of the crystals.

★ **15** From a chemical perspective, **identify** the reason for adding salt to a pan of boiling water that is going to be used to cook pasta.

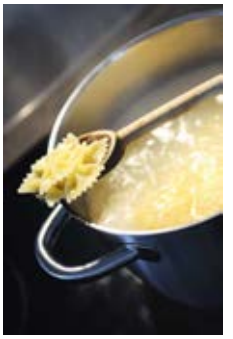


FIGURE 3 Cooking pasta

★ **16** **Explain** the role of the air filter in a combustion engine. Why does this filter have to be changed at regular intervals?

Apply, analyse and interpret

★ **17** **Classify** the following mixtures as homogeneous or heterogeneous.

- **a** A fruit salad
- **b** A glass of diluted cordial
- **c** A milkshake
- **d** Smoke from a fire
- **e** A chocolate chip muffin

★★ **18** Using Table 2 (Section 6.1), **analyse** the melting and boiling temperatures to determine the physical state of water, nitrogen, oxygen, mercury and iron at –200°C.

★★ **19** **Determine** the best sequence of separation techniques to retrieve pure samples of each component in a mixture of:

- **a** vegetable oil and water
- **b** sugar and rice
- **c** alcohol and salty water.

★★ **20** A student pours a bag of pool salt into their pool and leaves it to dissolve naturally. Later that day, they dive into the pool and notice that the water is warmer near the surface and saltier at the bottom. **Interpret** these observations.

FIGURE 4 Swimming pool

★★★ **21**

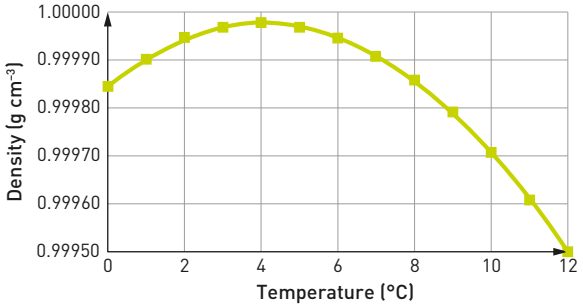


FIGURE 5 Density of water as a function of temperature.

A **Define** what temperature water has the greatest density.

B Assuming that no water is lost through evaporation, **determine** what would happen to a given volume of water as the temperature increases from 6°C to 12°C.

C 1.0000 L of water has a mass of 999.5 g. **Define** the temperature of the water.

D A sample of water has a density of 0.99990 g cm^{-3}. From this information alone, can you determine the temperature of the water? **Explain** your answer.

★★★ **22** **Discuss** why nanoparticles are more reactive than larger particles.

Investigate, evaluate and communicate

★★ **23** **Investigate** the effect of percentage carbon content in steel manufacture. Give examples where high carbon steel is used.

★★★ **24** **Evaluate** the current impact of nanotechnology in society, providing examples of where nanomaterials are being applied.

★★★ **25** Research the structure of C_{60} fullerene. **Construct** a three-dimensional representation of the molecule using hexagons and pentagons of equal side length.

★★★ **26** **Research** whether the Earth's mantle is a heterogeneous or homogeneous mixture of elements. How do geologists investigate this?

FIGURE 6 Earth's mantle

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  Chapter 6

CHAPTER 7

Bonding and properties

On 1 August 2012, trains on the Sunshine Coast line were cancelled because of the theft of copper signalling cable. This was not an isolated incident. The Energex power company estimates they lose around $2 million every year as a result of thefts of copper. What is it about copper that makes it such a tempting target for thieves?

Global demand for copper has increased. More than 50% of copper produced is used for electrical items, cables and wiring (Figure 1). Copper is also a good conductor of heat and is used in thin-walled copper tubing in air-conditioning systems, refrigeration units and car radiators. The Mount Isa mine, as the largest copper producer in Australia, supplies some of this global demand. The copper is present mostly as an ionic compound called chalcopyrite ($CuFeS_2$), which is converted into pure copper metal.

Copper is a metallic element and has high electrical and thermal conductivity because of how the electrons are arranged in the solid. The free electron model of metallic bonding describes metals as a three-dimensional arrangement of metal ions, surrounded by a sea of electrons acting as the 'glue' holding the metal cations together.

This chapter explores metallic bonding further, as well as the types of bonding found in ionic compounds, and introduces the intermolecular forces that occur between molecules. An understanding of chemical bonding and intermolecular forces is needed to know how substances will react and how they can be transformed into new and useful materials.

OBJECTIVES

→ Recognise that the properties of ionic compounds, including high melting point, brittleness and ability to conduct electricity when liquid or an aqueous solution, can be explained by modelling ionic bonding as ions arranged in a crystalline lattice structure with strong electrostatic forces of attraction between oppositely charged ions (metallic lattice, giant covalent networks, allotropes – carbon).

→ Understand that the type of bonding within ionic, metallic and covalent substances explains their physical properties, including melting and boiling point, thermal and electrical conductivity, strength and hardness.

→ Understand that hydrocarbons, including alkanes (saturated), alkenes (unsaturated) and benzene, have different chemical properties that are determined by the nature of the bonding within the molecules.

→ Analyse and interpret given data to evaluate the properties, structure and bonding of ionic, covalent and metallic compounds.

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PRACTICALS

SUGGESTED PRACTICAL

7.4 Testing for saturation

FIGURE 1 Copper wiring is popular because copper is a good conductor of electricity; the copper atoms bond together by metallic bonding.

7.1 Ionic bonding and compounds

KEY IDEAS

In this section, you will learn about:

- the arrangement of ionic compounds in a crystalline lattice structure
- how the main characteristics of ionic compounds are related to their ionic structure and bonding
- the properties of ionic compounds, including high melting point, brittleness and the ability to conduct electricity when a liquid or in an aqueous solution.

ionic compound a compound consisting of positively charged cations and negatively charged anions

crystalline lattice the ordered, repeating three-dimensional arrangement of atoms, molecules or ions in crystalline solids

The ionic model of bonding is a description of how cations (positively charged ions) and anions (negatively charged ions) are held together in **ionic compounds**.

Crystalline lattice structure

An ionic compound is an ordered, repeating three-dimensional arrangement of cations and anions in a **crystalline lattice** structure (Figure 1). In a crystalline lattice, cations are surrounded by anions and vice versa.

For example, in sodium chloride (NaCl), sodium ions (Na^+) are surrounded by chloride ions (Cl^-) in an arrangement that maximises the attraction between the oppositely charged ions. The way that the ions pack together depends on their charges and ionic radii.

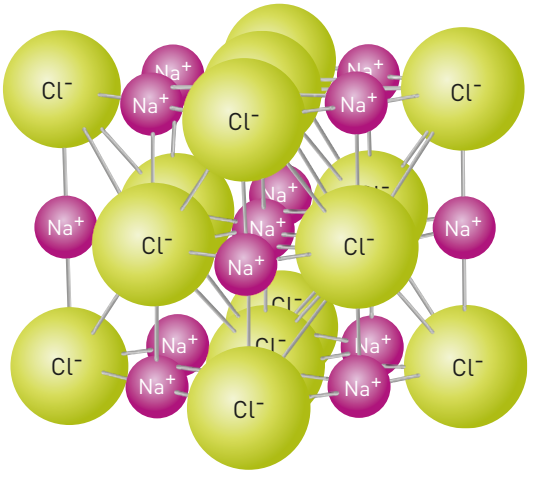


FIGURE 1 Ionic compounds, such as sodium chloride (NaCl), are arranged in a crystalline lattice structure.

Study tip

Ions and bonding were introduced in Chapter 3. It is a good idea to revise this material before advancing.

Ionic compounds are not limited to monatomic ions. Ionic compounds can also consist of polyatomic ions, in which ions have atoms held together through covalent bonds. An example is sodium carbonate (Na_2CO_3), which contains sodium (Na^+) and carbonate (CO_3^{2-}) ions.

Writing formulas of ionic compounds

When writing the name and formula of an ionic compound, write the cation first, before the anion. For example, write sodium chloride, *not* chloride sodium.

The formula for an ionic compound is the formula unit that is used to represent the compound. In an ionic compound, this is the lowest whole-number ratio of ions. Remember though that polyatomic ions are treated as one indivisible unit. For example:

- magnesium nitrate, $Mg(NO_3)_2$ (Note that the brackets around the nitrate ion mean that there are two nitrate ions NO_3^- for every Mg^{2+} ion.)
- ammonium carbonate, $(NH_4)_2CO_3$
- barium nitride Ba_3N_2. (There are three Ba^{2+} ions for every two N^{3-} ions.)

Every ionic compound is neutral overall, so the total number of positive charges must balance the total number of negative charges.

WORKED EXAMPLE 7.1

An ionic compound was analysed and of the ions measured, 649 were found to be M^{n+} cations and 978 were found to be X^{z-} anions, with an experimental error of ±20. Determine the formula unit of the ionic compound.

SOLUTION

First find the lowest whole-number ratio of the two ions:

$649 \pm 20\ M^{n+} : 987 \pm 20\ X^{z-}$

$$\frac{649 \pm 20}{649} M^{n+} : \frac{987 \pm 20}{649} X^{z-}$$

$1.00 \pm 0.03\ M^{n+} : 1.52 \pm 0.03\ X^{z-}$

The formula unit must contain whole numbers, so multiply both sides by 2.

$2\ M^{n+} : 3\ X^{z-}$

Therefore, the formula unit is M_2X_3.

Why do ionic compounds form?

The formation of a compound implies that the compound is more stable than the separated atoms.

Consider the example of sodium chloride to understand why ionic compounds form. The process of making solid sodium chloride can be broken down into three steps (Figure 2). (This is not how the process occurs, but as long as the end result is the same, the overall energy for the process will be the same.)

Step 1: Sodium atoms in the gas phase release electrons to form singly charged cations (Na^+).

Step 2: These electrons are transferred to chlorine atoms in the gas phase to form chloride ions (Cl^-).

Step 3: The resulting ions clump together, attracted by their opposite charges to form solid sodium chloride (NaCl).

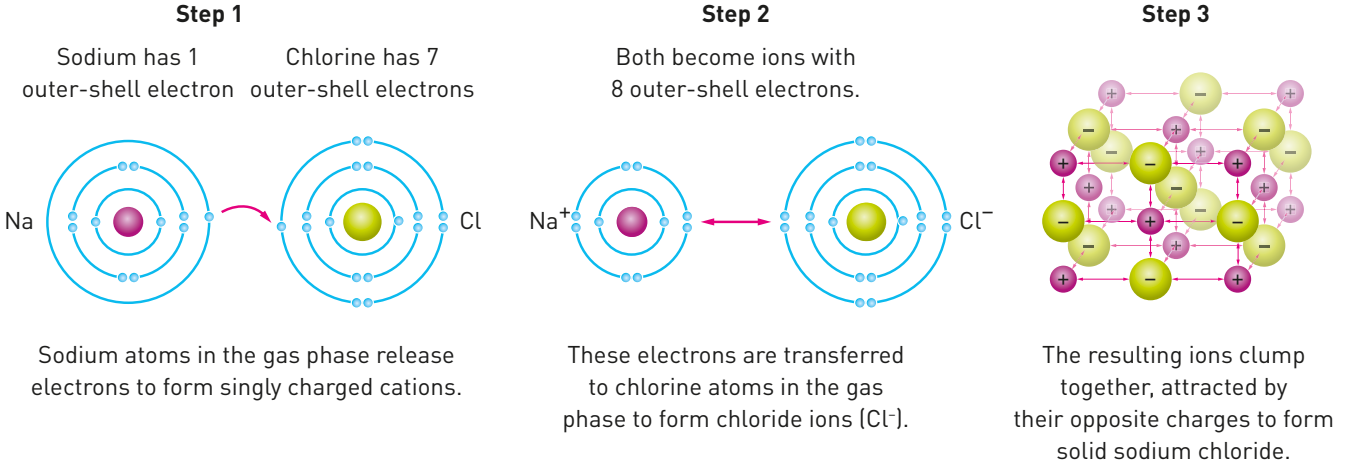


FIGURE 2 The process of making solid sodium chloride in three steps.

ionisation energy
the energy (in kJ mol^{-1}) required by a gaseous atom to remove an electron from that atom

electron gain enthalpy
the energy released or required to add an electron to a neutral atom or species in the gas phase (usually measured in kJ mol^{-1}), with formation of the –1 anion in the gas phase

lattice energy
the energy released when a solid ionic compound forms from ions in the gas phase that are infinitely separated; it is a measure of the stability of an ionic compound

Because sodium is in group 1 of the periodic table, a sodium atom will lose one electron to form a cation (Na^+). The **ionisation energy** of sodium (the energy required to remove electrons from the atoms in the gas phase) is 494 kJ mol^{-1}. Energy needs to be supplied to form the cation, in order to remove the negatively charged electron from the influence of the positively charged nucleus. When chlorine atoms in the gas phase accept an electron, 349 kJ mol^{-1} of energy is released. The energy released or required when an electron is accepted by an atom in the gas phase is called the **electron gain enthalpy**. Combining the first two steps, there has been a net change in energy of:

$$494 \text{ kJ mol}^{-1} + -349 \text{ kJ mol}^{-1} = +145 \text{ kJ mol}^{-1}$$

Therefore, a gas of widely separated Na^+ and Cl^- ions is higher in energy than a gas of sodium and chlorine atoms.

When the separated Na^+ and Cl^- ions are brought closer together, oppositely charged ions are attracted to each other. As the ions are brought closer together, the attractions increase, and the ions come together to form a crystalline solid. The difference in energy between ions in the gaseous state and the same ions packed together in a crystalline solid is called the **lattice energy**. This is the energy released when the solid forms and is very large for NaCl.

$Na^+(g) + Cl^-(g) \rightarrow NaCl(s)$ Energy released = 787 kJ mol^{-1}

Now consider the overall energy for the process: 145 kJ mol^{-1} was needed to form ions from the neutral atoms in the gas phase, and 787 kJ mol^{-1} is released when solid NaCl forms from ions in the gas phase. Therefore, the overall energy change is:

$$145 \text{ kJ mol}^{-1} + -787 \text{ kJ mol}^{-1} = -642 \text{ kJ mol}^{-1}$$

The negative sign indicates that energy is being released.

$Na(g) + Cl^-(g) \rightarrow NaCl(s)$ Energy released = 642 kJ mol^{-1}

Therefore, a solid composed of Na^+ and Cl^- ions is much more stable or has lower energy than a collection of sodium and chlorine atoms in the gas phase.

More generally, when an ionic compound is formed from neutral atoms, energy is required to remove one or more electrons from atoms to form cations. This will be easiest for the elements with the lowest ionisation energies, which are typically metallic elements. The energy released by the formation of the solid ionic compound must be greater than the ionisation energy for this process to occur. The accepting of an electron by an atom to form the anion may require or may release energy.

Study tip

The mole concept is explained in detail in Chapter 8.

Forces between ions

Ionic compounds are held together by the electrostatic attraction between the oppositely charged ions. Charles-Augustin de Coulomb, a French scientist working in the 1700s, developed a formula to describe the forces of attraction between two charged particles. He found that the attractive force depended on the size of the charges on the particles and the distance squared between the particles. A simplified form of Coulomb's law is:

$$\text{Force} = k\frac{q_1 \times q_2}{r^2}$$

where k is a constant, q_1 and q_2 are the charges on the two ions (positive for a cation and negative for an anion) and r is the distance between the two ions.

This formula tells us three main things:

- The attractive force between two oppositely charged ions is greater the more highly charged they are.

- Ions with smaller ionic radii or size have greater attractions between the ions than an ionic compound with larger ions with the same charges, as long as the ions are arranged in the same way.
- Ions of the same charge repel each other.

Characteristics of ionic compounds

The characteristics of ionic compounds can be explained according to what is already known about **ionic bonding**. These properties include the following:

ionic bonding the electrostatic (charge-based) attractive forces holding ions together in ionic compounds

- **Ionic compounds usually have relatively high melting and boiling points.** The melting and boiling points are usually hundreds to thousands of degrees Celsius. The ions in an ionic compound are locked in place by the electrostatic interaction between oppositely charged ions. It takes a large amount of heat energy to overcome the attractions between ions, hence the high melting points. To convert an ionic compound from a liquid to a gas, it is necessary to overcome all of the electrostatic interactions between oppositely charged ions.
- **The solids are relatively hard but brittle.** Ionic compounds are brittle because under pressure, such as when hit with a hammer, ions of the same charge can be brought close together, resulting in repulsion (Figure 3).

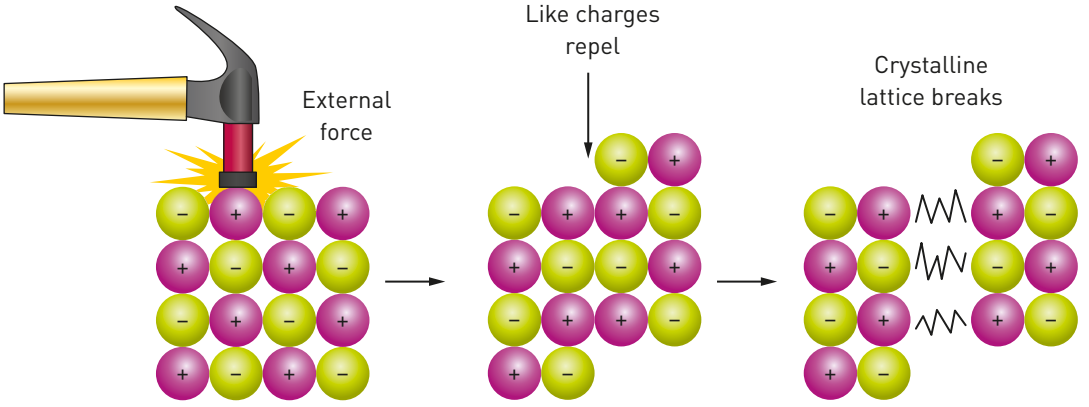


FIGURE 3 Ionic compounds are brittle solids because of their ionic bonds.

- **Ionic compounds do not conduct electricity in the solid state.** In the solid state, the ions are held in a rigid lattice and cannot move freely to carry electrical charge. **However, ionic compounds can conduct electricity as molten liquids.** For a substance to conduct electricity, charged particles need to be able to move to carry current between the electrodes. In the molten state, ions are no longer held in a rigid lattice and are free to move between the electrodes.
- **Aqueous solutions of ionic solids can also conduct electricity.** When dissolved in water, the ions separate. Again, ions move toward the electrode with the opposite charge.
- **Many (but not all) ionic solids are soluble in water.** For an ionic solid to dissolve in water, the lattice energy holding the ions together must be overcome. For example, NaCl and $Mg(NO_3)_2$ are soluble in water, whereas MgO and $PbSO_4$ are insoluble.

The presence of water molecules acts to partially shield the ionic charges from each other, reducing the attraction between oppositely charged ions. Part of the driving force is also provided by an increase in a quantity called entropy as the ordered arrangement of ions in the solid crystalline lattice is replaced by the more random distribution of ions in solution.

Study tip

Review the polyatomic ions that may be encountered in chemistry exams. One way to learn the names and formulas is to put them on either side of flash cards or use a flash card app. Learn the ionic charges as well. Learning the polyatomic ions now will help later with Units 3 & 4.

Study tip

Ionic compounds are more commonly made by double displacement reactions, where the ions already exist. For example:
$AgNO_3(aq) + KBr(aq) \rightarrow AgBr(s) + KNO_3(aq)$
Ag^+ ions were originally present in the silver nitrate solution, and the reaction with potassium bromide does not involve transfer of electrons.

Ionic compounds that are not soluble in water are those that have large lattice energies, where the energy that would be released by hydrating the ions is not enough to overcome the electrostatic interactions between ions in the crystalline lattice (Figure 4).

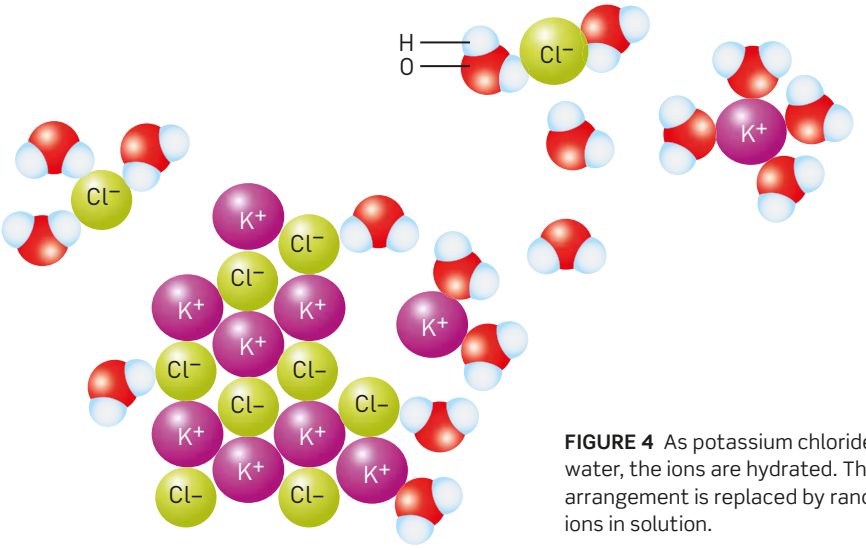


FIGURE 4 As potassium chloride (KCl) dissolves in water, the ions are hydrated. The ordered lattice arrangement is replaced by random distributions of ions in solution.

CHECK YOUR LEARNING 7.1

Describe and explain

1 **Describe** the differences in bonding between potassium iodide and iodine monochloride.

2 **Explain** how the properties of ionic compounds are related to the attractions between ions.

Apply, analyse and interpret

3 **Analyse** the information in Table 1, which shows the melting and boiling points of some ionic compounds that all crystallise in a face-centred cubic pattern. **Explain** the trends observed in the data.

TABLE 1 The melting and boiling points of selected ionic compounds

Compound	Melting point (°C)	Boiling point (°C)
NaF	993	1695
NaCl	801	1413
NaBr	747	1396
KF	770	1420
CsF	682	1251
MgO	2852	3600

4 **Apply** your knowledge of ionic bonding to determine the ions and then write the correct formula for ionic compounds formed from:
- **a** caesium and iodine
- **b** aluminium and fluorine
- **c** gallium and oxygen
- **d** lithium and nitrogen
- **e** sulfur and sodium
- **f** bromine and calcium.

Investigate, evaluate and communicate

5 **Investigate** three different packing arrangements observed for ionic compounds; for example, TiO_2 and CsCl.

You can find the following resources for this section on your obook assess:

- » **Student book questions** Check your learning 7.1
- » **Weblink** Ionic compounds
- » **Weblink** Polyatomic ions
- » **Weblink** Electrical conductivity

7.2 Metallic bonding

KEY IDEAS

In this section, you will learn about:

- the free electron model of metallic bonding
- how the main characteristics of metallic substances are related to their metallic bonds
- the properties of metallic substances, which include high electrical conductivity, high boiling point, malleability and ductility.

malleable
able to be pressed into other shapes

ductile
able to be drawn into a wire

free electron model of metallic bonding
a model of the bonding in metals that presents metals as ions surrounded by one or more valence electrons that are not associated with a particular atomic nucleus

metallic bond
the attractive force between a metal ion and delocalised electrons that hold metals together

Most elements (about 80%) in the periodic table are metals. Metals have a number of common properties, including:

- metallic lustre (Figure 1)
- being solid at room temperature (apart from mercury (melting point –38.83°C) and gallium (melting point 29.8°C)
- being good conductors of electricity and heat
- being easily shaped into sheets (**malleable**) and drawn into wires (**ductile**).

FIGURE 1 The metallic lustre of metals is part of the attraction for metals such as silver and gold in jewellery.

These properties can be explained by the type of bonding in metals. Metallic elements can lose valence electrons relatively easily because they have low ionisation energies. The **free electron model of metallic bonding** considers metals in the solid state as lattices of metal ions surrounded by a sea of valence electrons (Figure 2). These electrons are free to move randomly in the solid and are not associated with any single metal ion but are shared between several neighbouring metal ions. This is called delocalisation. The metal ions can vibrate but are otherwise fixed in a repeating three-dimensional pattern, with the metal ions packed closely together. The forces of attraction between positively charged metal ions and negatively charged electrons can be regarded as the **metallic bonds**, keeping the solid together.

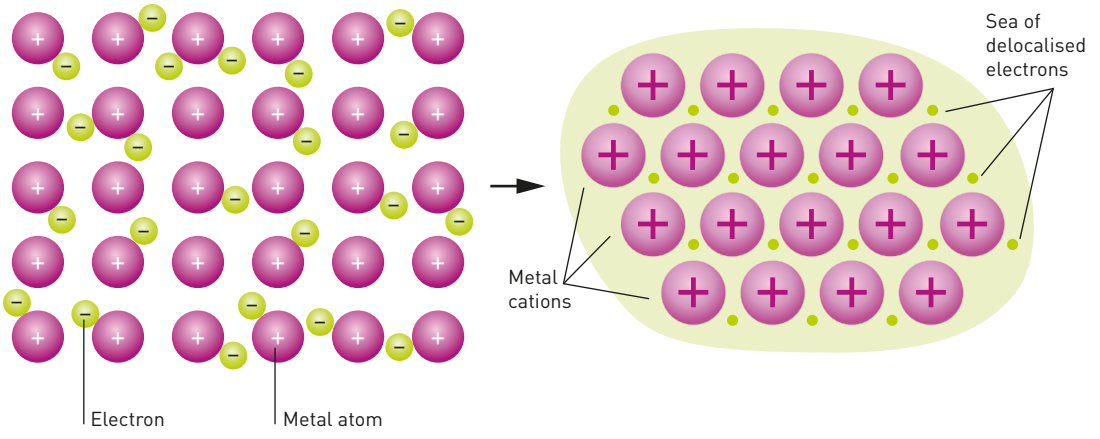


FIGURE 2 Metallic bonding occurs where delocalised electrons surround metal ions in a lattice.

delocalised electron
electron that is not retained in bonds between two atoms or associated with a single nucleus of an atom but is instead spread over several atoms or is highly mobile within a solid

Delocalised electrons have the following effects.

- **High electrical conductivity:** If a piece of metal such as a wire is connected to an electrical source such as a battery and then connected into an electrical circuit, electrons from the battery will enter the wire and displace electrons from the end of the wire closest to the source of the current. Electrons will flow through the solid, finally displacing electrons from the other end of the wire.
- **Malleability and ductility:** When a metal is beaten with a hammer, layers of ions can slip past each other (Figure 3). After the movement, the environment around each metal atom is unchanged because the sea of electrons can easily adapt to the change as a result of the delocalised nature of the electrons.

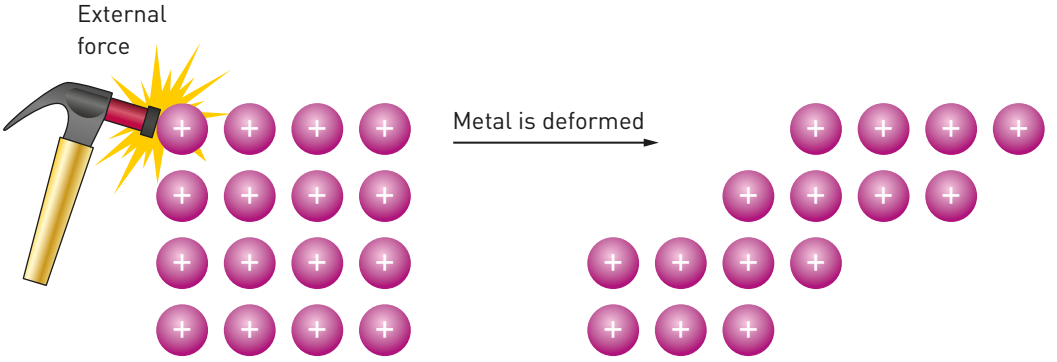


FIGURE 3 When force is applied to metal, layers of metal ions move and the delocalised electrons adapt to the change.

- **High thermal conductivity (of solid metals):** The high thermal conductivity of metals is also due to the free electrons. Some evidence for this is that the thermal conductivity of metals is proportional to how electrically conductive they are and both depend on delocalised electrons. Heat is transferred in metals as electrons move freely from atom to atom, distributing the energy supplied as heat through the solid. The crystalline structure of metals also helps heat transfer, as does their high densities.
- **High boiling points:** In a molten metal, the ordered arrangement no longer exists, but the attractive forces are not completely overcome until enough heat energy is added to reach the boiling point. Therefore, this is a better indicator of the strength of the metallic bonds than the melting point. The strength of the metallic bonds depends on the number of delocalised electrons, the charge of the cation and the size of the cation.

Table 1 shows the melting and boiling points for selected alkali and alkali earth elements.

TABLE 1 Melting and boiling points of selected alkali and alkali earth elements

Alkali	Melting point (°C)	Boiling point (°C)	Alkali earth	Melting point (°C)	Boiling point (°C)
Lithium (Li)	181	1330	Beryllium (Be)	1287	2470
Sodium (Na)	98	883	Magnesium (Mg)	650	1090
Potassium (K)	63	759	Calcium (Ca)	842	1484
Rubidium (Rb)	39	688	Strontium (Sr)	777	1382
Caesium (Cs)	28	671	Barium (Ba)	727	1870

The alkali earths have higher boiling points than those of the alkali elements that precede them in the periodic table. The trends in the boiling points are a better guide to the strength of the metallic bonding than melting points because there are differences in how the metals are arranged in different lattice arrangements in the solid state.

For example, sodium atoms have one valence electron ($3s^1$), whereas magnesium has two valence electrons ($3s^2$). Both of the electrons in magnesium can become delocalised into the sea of electrons, so there is a greater attraction between the 2+ charge from the magnesium and the delocalised electrons.

Magnesium ions are smaller than sodium ions, which also increases the attractive forces. The decrease in melting points and boiling points in the first two groups of the periodic table is due to the increasing atomic radii. As the distance between the atomic nucleus and the delocalised electrons increases, there is a corresponding decrease in electrostatic attraction.

One of the problems with the free electron model of metallic bonding is that it does not explain some of the finer details of the properties of metals or the differences between metals, semiconductors and insulators. It is also unlikely that a metal such as tungsten ([Xe] $4f^{14}6s^25d^4$) with six valence electrons will form a 6+ ion because of the large amount of energy required to remove so many electrons. More advanced models of metallic bonding have been proposed that can explain these phenomena.

CHECK YOUR LEARNING 7.2

Describe and explain

1 **Construct** a labelled diagram that describes the differences and similarities between ionic and metallic bonding.

2 **Explain** why there is an average difference of 793°C between the boiling points of groups 1 and 2 metals.

Apply, analyse and interpret

3 **Analyse** the following *incorrect* statements about metallic bonding. Identify errors in these statements and rewrite them to be correct.

- **a** Metals can conduct electricity in the solid state because the metal atoms can easily slip past each other in order to carry electrical charge through the solid.
- **b** Metals are held together by ionic bonds.
- **c** The strength of metallic bonding in different metals is constant because only a single valence electron from each atom becomes delocalised.
- **d** Metals melt once sufficient heat energy has been provided to boil away the sea of electrons.
- **e** The melting point of a metal is a good indication of the strength of the metallic bonding within that metal.

Investigate, evaluate and communicate

4 **Evaluate** the relationship between whether an element is a metal or non-metal and its electronegativity. Explain this relationship in terms of the free electron model of metallic bonding.

5 **Investigate** why silver and copper are better conductors of electricity than gold is.

You can find the following resources for this section on your obook assess:

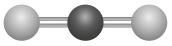
- » Student book questions
  Check your learning 7.2
- » Weblink
  Mercury
- » Weblink
  Alkali elements
- » Weblink
  Electrical conductivity

7.3 Covalently bonded molecules

KEY IDEAS

In this section, you will learn about:

- the main physical characteristics of covalent substances that relate to the relatively weak intermolecular forces that hold them together
- the properties of simple molecules with covalent bonds, including low melting and boiling points and being electrically non-conductive
- the properties of giant covalent networks, including high melting points and variable conductivity because they are held together by covalent bonds



Carbon dioxide

FIGURE 1 The structure of the carbon dioxide molecule. Each oxygen atom shares four electrons with the central carbon atom.

As discussed in Chapter 3, covalent bonds occur by sharing of electrons between atoms. Covalent bonds form when these atoms share electrons. Simple molecules that have covalent bonds are water, methane, ammonia and carbon dioxide (Figure 1).

Giant covalent networks

giant covalent network an element or compound in which each atom is covalently bonded to other atoms to form an extended assembly of covalently bonded atoms

Several substances consist of large assemblies of atoms all covalently bonded together to form **giant covalent networks**, with the network extending throughout the solid. These are also called covalent network solids and may be elements (with only one type of atom present) or contain more than one atom type.

Covalent network solids are usually extremely hard solids, with high melting points, because the atoms are connected together to form one giant molecule. Examples are diamond, silica (silicon dioxide) and boron nitride. To melt diamond, pressures about 99 000 times that of normal atmospheric pressure and a temperature above 4700°C are required because a lot of energy is needed to break the multiple covalent bonds that hold these network solids together.

As the valence electrons in the carbon atoms in diamond are all involved in making covalent bonds, there is no delocalisation of electrons and hence diamond does not conduct electricity. However, it is a good conductor of heat because of the way vibrations can be carried through the solid.

In addition to three-dimensional network solids, there are other solids that form a network, but only in two dimensions or in a single plane. The most well-known of these is graphite.

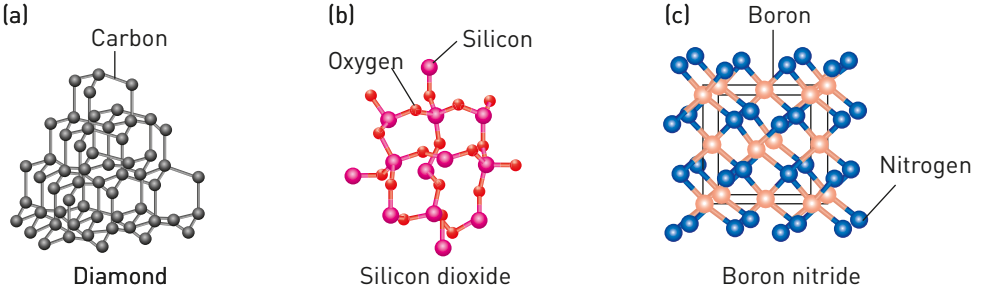


FIGURE 2 Part of the structures of three covalent networks solids. (a) In the structure of diamond, each carbon atom is covalently bonded to four other neighbouring carbon atoms. (b) In quartz (silicon dioxide (SiO_2)), each silicon atom is surrounded by four oxygen atoms in a tetrahedral arrangement and each oxygen atoms bridges two silicon atoms. (c) In cubic boron nitride (BN), boron atoms form four bonds to neighbouring nitrogen atoms.

Each carbon atom in graphite is bonded to three other carbon atoms to form a two-dimensional network. These two-dimensional networks or sheets are stacked on top of each other. There are no covalent bonds between sheets; they are held together by intermolecular forces (discussed in Chapter 12). The purple-dashed lines in Figure 3 represent these dispersion forces, and the grey shading emphasises the flat or planar nature of the layers. Graphite is a strong conductor of energy along its layers, as a result of the delocalisation of electrons in the two-dimensional layers. Graphite is a weak conductor at right angles to its layers.

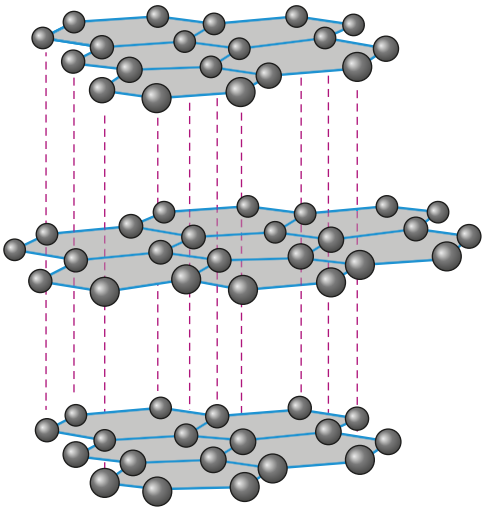


FIGURE 3 Graphite features planes of covalently bonded atoms stacked on top of each other. The dashed lines represent weak intermolecular forces, and the covalent bonds are shown as blue lines.

Allotropes

Graphite and diamond are both solids at room temperature and both contain only carbon atoms. They are called **allotropes**, because they have the same formula of C, but differ in the way that the atoms are bonded to each other. Allotropes are different forms of the same element, existing in the same state (e.g. solid or gas) under the same conditions of temperature and pressure. Carbon is the element with the greatest number of allotropes, as shown in Figure 1. Other forms of the element include amorphous carbon, which does not form a regular arrangement of its carbon atoms. Graphene has a single-layer arrangement of carbon atoms in a hexagonal arrangement. The last type of allotrope of carbon discovered was that of the fullerenes; buckminsterfullerene (or 'buckyballs'), which has the formula C_{60}, was first synthesised in 1985. Other fullerenes have since been found or synthesised, including nanotubes. Fullerenes have been found in nature and detected in outer space (see Case study 6.3).

Another important example of allotropism is the two most common allotropes of oxygen – molecular oxygen (O_2) and ozone (O_3), as shown in Figure 6. O_2 is needed to sustain human life and for respiration. O_3 is a pollutant and lung irritant when found in the lower atmosphere, but helps to absorb UV radiation in the upper atmosphere.

allotrope
one of the different forms in which a particular element may exist in a given physical state; these forms are due to differences in bonding between atoms of that element

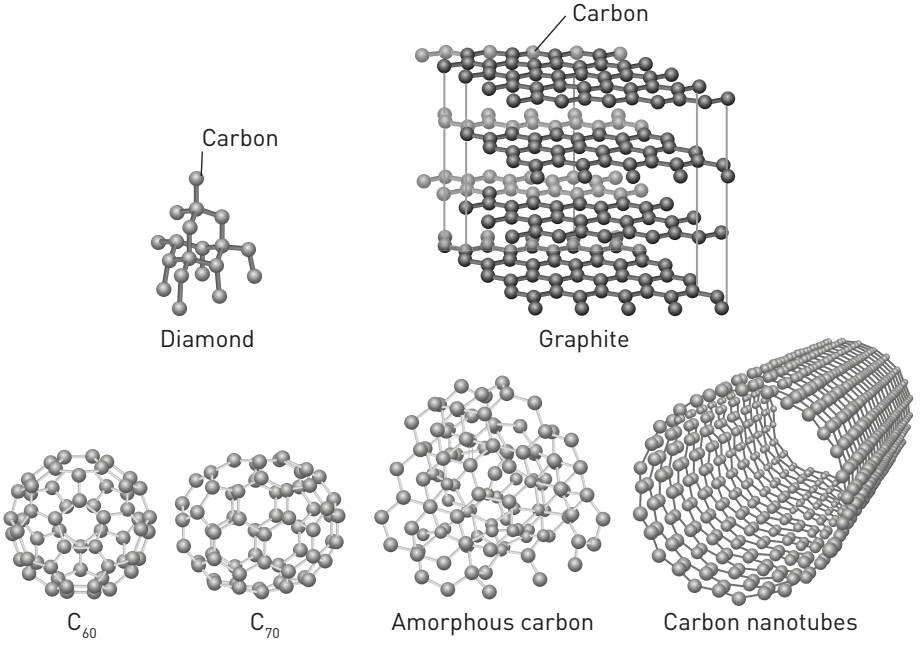


FIGURE 4 Structures of selected allotropes of carbon.

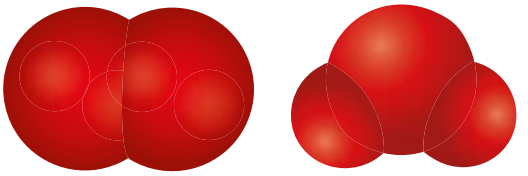
CHALLENGE 7.3

Making diamonds

Can graphite be converted into diamonds. Explain your answers.

FIGURE 5 Diamond

Covalently bonded molecules and intermolecular forces



Oxygen Ozone

FIGURE 6 Two allotropes of oxygen: molecular oxygen (O_2) and ozone (O_3).

Small covalently bonded molecules typically have low melting and boiling points. They do not conduct electricity or heat well and if they are solids at room temperature, they are typically quite soft. Small molecular compounds, such as ammonia and carbon dioxide, are often gases at room temperature and pressure.

Study tip

Intermolecular forces are explained in detail in Chapter 12.

To explain this in terms of the bonding present, a distinction needs to be drawn between the forces that keep atoms in a molecule together and those responsible for the attractions between different molecules.

intramolecular
a force between the atoms in a molecule

Each molecule is held together by the covalent bonds between atoms, resulting from sharing of electrons, and these are **intramolecular** (i.e. between the atoms in a single molecule). These intramolecular interactions are strong and difficult to break. It takes energy to break covalent bonds. For example, to break the O–H covalent bonds in water in the gas phase takes 498.7 kJ for each mole of bonds broken. To break O–H covalent bonds requires heating water to thousands of degrees, or blasting the molecules with high-energy UV radiation. The dissociation energy of covalent bonds varies from around 130 k J mol^{-1} to 1100 kJ mol^{-1}. The strength of bonds increases from single to double to triple bonds, as the number of shared electrons increases.

intermolecular
a force between molecules

In a solid or a liquid made up of many molecules, there are no covalent bonds between distinct molecules. Instead, molecules are attracted weakly to each other through **intermolecular** forces. Intermolecular forces are explained fully in Chapter 12, including dispersion forces, dipole–dipole interactions and hydrogen bonding. When a molecular solid is heated, the thermal energy disrupts some of the intermolecular attractions between different molecules and the solid can go from a solid to a liquid state. These intermolecular forces take much less energy to break than covalent bonds and the molecules remain intact. For example, intermolecular forces between water molecules can be broken simply by heating to temperatures above 100°C at normal pressures. Intermolecular interactions are also constantly breaking and re-forming in the liquid state, unlike covalent bonds. Differences in the boiling points of different molecular compounds are due to differences in the strength of the intermolecular interactions between the molecules.

CHECK YOUR LEARNING 7.3

Describe and explain

1 **Describe** three different allotropes of carbon and explain the differences between them.

2 **Explain** why silicon dioxide (silica) has a high melting point.

Apply, analyse and interpret

3 **Analyse** the structures shown in Figure 2 of diamond and quartz. In what ways are they the same? How are they different?

4 **Interpret** the diagram in Figure 7 of two HCl molecules interacting. Which are intermolecular interactions. Explain which each type of interaction is represented in a different way.

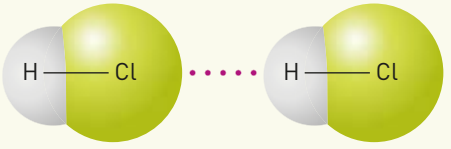


FIGURE 7 HCl molecules

5 **Interpret** the graph in Figure 8. Which compound has the most significant (strongest) intermolecular forces? Which compound has the weakest intermolecular forces?

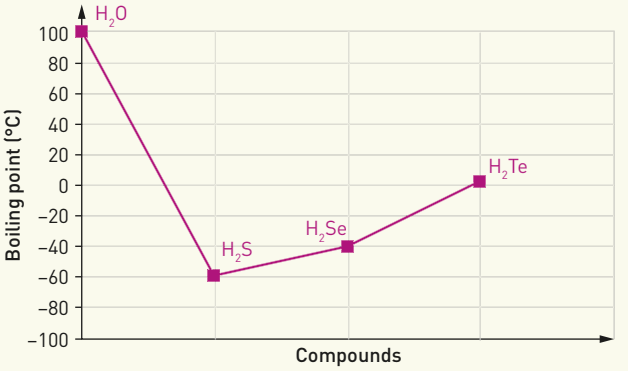


FIGURE 8 Compounds against boiling points

Investigate, evaluate and communicate

6 **Investigate** how different allotropes of sulfur can be prepared. Construct a diagram that communicates in a graphical summary the experimental procedures used.

7 **Investigate** how fullerenes were detected in outer space and where they have been found on Earth.

You can find the following resources for this section on your obook assess:

- » Student book questions
  Check your learning 7.3
- » Challenge
  7.3 Making diamonds
- » Increase your knowledge
  Addition reactions
- » Weblink
  Giant covalent networks

7.4 Hydrocarbons

KEY IDEAS

In this section, you will learn about:

- the different types of hydrocarbons
- how the properties of hydrocarbons relate to the type of carbon–carbon bonds present.

hydrocarbon
a molecular compound that only contains carbon and hydrogen atoms; for example, propane (C_3H_8)

saturated hydrocarbon
a compound of carbon and hydrogen that contains the maximum possible number of hydrogen atoms for the number of carbon atoms present

alkane
a hydrocarbon that contains the maximum number of hydrogen atoms possible for the given number of carbon atoms; for linear or branched alkanes with n carbon atoms, the number of hydrogen atoms is $2n+2$; for a cycloalkane with one ring and n carbon atoms, the number of hydrogen atoms is $2n$

Hydrocarbons are compounds that are composed entirely of carbon and hydrogen atoms. Because of the small electronegativity difference between carbon and hydrogen, all hydrocarbons are non-polar with dispersion forces acting between molecules.

Alkanes

Saturated hydrocarbons contain the maximum number of hydrogen atoms possible for the number of carbon atoms present. Saturated hydrocarbons are also called **alkanes** and only have single bonds between the atoms. The carbon atoms can be arranged in a linear fashion (linear alkanes), or there can be branches coming off the main chain (branched alkanes).

Linear and branched alkanes have the same general formula C_nH_{2n+2}, where n is the number of carbon atoms and $2n + 2$ is the number of hydrogen atoms. If $n = 1, 2, 3, 4 \ldots$, then $2n + 2 = 4, 6, 8, 10 \ldots$, to give the corresponding molecules CH_4, C_2H_6, C_3H_8, C_4H_{10}.

WORKED EXAMPLE 7.4A

What is the formula of the linear alkane with 12 carbon atoms?

SOLUTION

From the general formula C_nH_{2n+2}, when $n = 12$,

$$\begin{aligned} & 2n + 2 \\ &= 2 \times 12 + 2 \\ &= 26 \end{aligned}$$

Therefore, the alkane with 12 carbon atoms has 26 hydrogen atoms, and the molecular formula is $C_{12}H_{26}$. It is named dodecane to indicate that there are 12 carbons in a linear arrangement.

The carbon atoms in alkanes are all sp^3 hybridised, so the groups around them are in an approximate tetrahedral arrangement.

Alkanes can also contain rings in which two ends of a chain are joined together. These are called cyclic alkanes. Cyclic alkanes (or cycloalkanes) with one ring have the general formula C_nH_{2n}; that is, they have two fewer hydrogen atoms than the linear (or acyclic) alkane with the same number of carbon atoms.

Alkanes are largely unreactive, which makes them good solvents for reactions involving otherwise highly reactive chemicals. Hexane, cyclohexane and petroleum spirits (a mixture of low-boiling alkanes) are all used in the laboratory for this purpose. Kerosene is a mixture of about 70% (by volume) of branched- and straight-chain alkanes with between 10 and 16 carbon atoms and also some cycloalkanes, and it is used as an industrial solvent and for fuel. The main reaction of alkanes is combustion, which is the reaction with oxygen to produce carbon dioxide

and water and liberate heat energy. The energy required to break the C–C and C–H bonds is much less than the energy released from the production of C=O and O–H bonds.

The higher molecular-weight alkanes are waxy solids – their larger electron clouds result in more significant intermolecular reactions so they are solids at room temperature (Table 1).

TABLE 1 Formulas, 3D models and boiling points of the first six alkanes

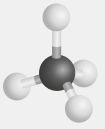
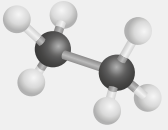
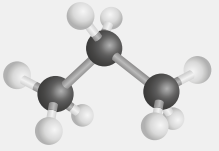
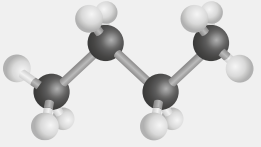
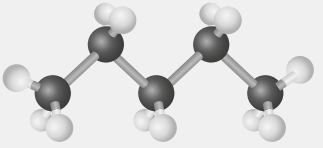
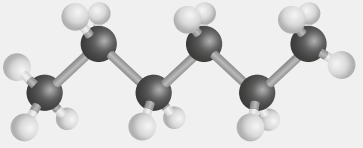
Name	Molecular formula	3D model	Condensed structural formula	Boiling point (°C)
Methane	CH_4		CH_4	–162
Ethane	C_2H_6		CH_3CH_3	–89
Propane	C_3H_8		$CH_3CH_2CH_3$	–42
n-Butane	C_4H_{10}		$CH_3(CH_2)_2CH_3$	–0.5
n-Pentane	C_5H_{12}		$CH_3(CH_2)_3CH_3$	36
n-Hexane	C_6H_{14}		$CH_3(CH_2)_4CH_3$	69

FIGURE 1 LPG is a mixture of gases, including propane and butane.

alkene
a hydrocarbon that contains one or more double bonds between carbon atoms

alkyne
a hydrocarbon that contains one or more triple bonds between carbon atoms

unsaturated hydrocarbon
a compound of carbon and hydrogen that contain double or triple bonds and therefore have fewer than the maximum number of hydrogen atoms for a particular number of carbon atoms

Alkenes and alkynes

Alkenes have one or more double bonds between carbon atoms and their names end in 'ene'. **Alkynes** have one or more triple bonds between carbon atoms and their names end in 'yne'.

Alkenes and alkynes are examples of **unsaturated hydrocarbons** because they contain less than the maximum number of hydrogen atoms possible for the number of carbon atoms present.

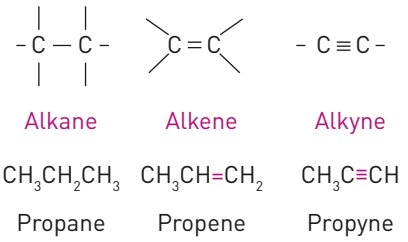


FIGURE 2 The general structures of alkanes, alkenes and alkynes and examples with three carbon atoms.

The general formula for an acyclic alkene is $C_nH_{2n+2-2P}$ where P is the number of double bonds present.

For example, but-1-ene ($H_2C{=}CH{-}CH_2{-}CH_3$) has one double bond and molecular formula C_4H_8, and 1,3-butadiene ($H_2C{=}CH{-}CH{=}CH_2$) has two double bonds and molecular formula C_4H_6 ($n = 4$, so $2 \times 4 + 2 - 2 \times 2 = 6$).

WORKED EXAMPLE 7.4B

Squalene is a polyunsaturated polyene acyclic branched hydrocarbon found in shark liver oil ('a' as a prefix means 'without'). It has the molecular formula of $C_{30}H_{50}$. How many double bonds does it contain?

SOLUTION

This is a branched polyene, so use the formula:

Number of H atoms = $2n + 2 - 2P$, where P is the number of double bonds

$50 = 2 \times 30 + 2 - 2P$

$50 = 62 - 2P$

$2P = 62 - 50$

$2P = 12$

$P = 6$

Therefore, there are six double bonds in squalene.

The structure of squalene (shown in Figure 3) confirms that there are six double bonds.

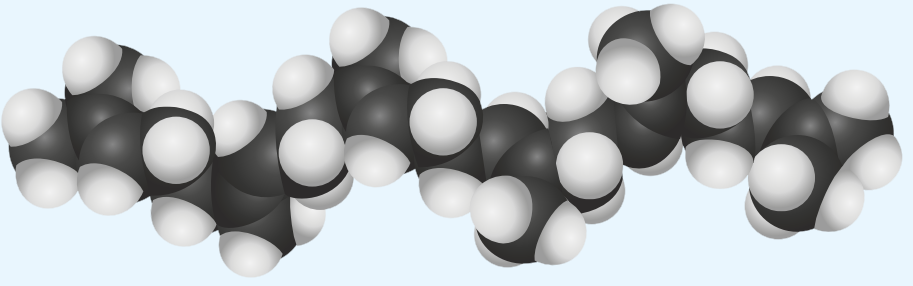


FIGURE 3 The structure of squalene.

Figure 4 shows a type of compound called a terpene, which is built from five-carbon units, and is found in many naturally occurring chemicals. For example, natural rubber consists of long chains made up of many of these five-carbon units.

Monoalkenes

Monoalkenes have boiling points close to but slightly lower than those of the alkanes with the same number of carbon atoms. This feature of monoalkenes is because the number of electrons is two fewer than that for the corresponding alkane as a result of the two fewer hydrogen atoms, which results in slightly weaker interactions between molecules.

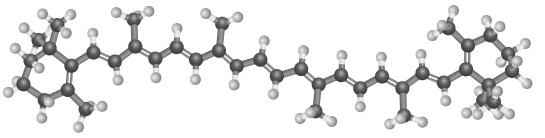
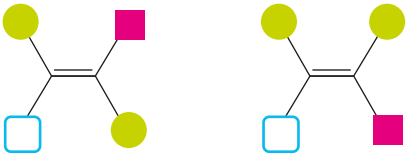


FIGURE 4 Beta-carotene is a terpene found in carrots; it gives them their orange colour.

The presence of the double bond in alkenes limits rotation around the double bond, which can lock groups on either side of the double bond into a fixed geometric relationship (Figure 5). Two compounds containing a double bond may have the same molecular formula and atom connectivity, but have groups that are arranged differently in space. These compounds have different physical properties, including boiling points. If two identical groups are on opposite sides of the double bond, it is called the *trans* isomer. If two identical groups are on the same side of the double bond, it is called the *cis* isomer. The groups being compared are attached to the two carbon atoms in the double bond (not the same carbon atom). A different naming system applies if there are four different groups attached to the double-bond carbon atoms.



trans isomer *cis* isomer

FIGURE 5 The restricted rotation around a double bond results in *trans* and *cis* stereoisomers. Note that the *trans* isomer has the two green groups on opposite sides of the double bond, while in the *cis* isomer, these groups are on the same side of the double bond

Alkenes and alkynes show greater chemical reactivity than alkanes and are able to undergo a type of reaction called addition, in which a reagent adds across the double or triple bond to give a final addition product. These reactions sometimes need catalysts.

Addition reactions have the general form A + B → C, i.e. the addition product is the result of combining two reactants. Several examples are shown in Figure 6.

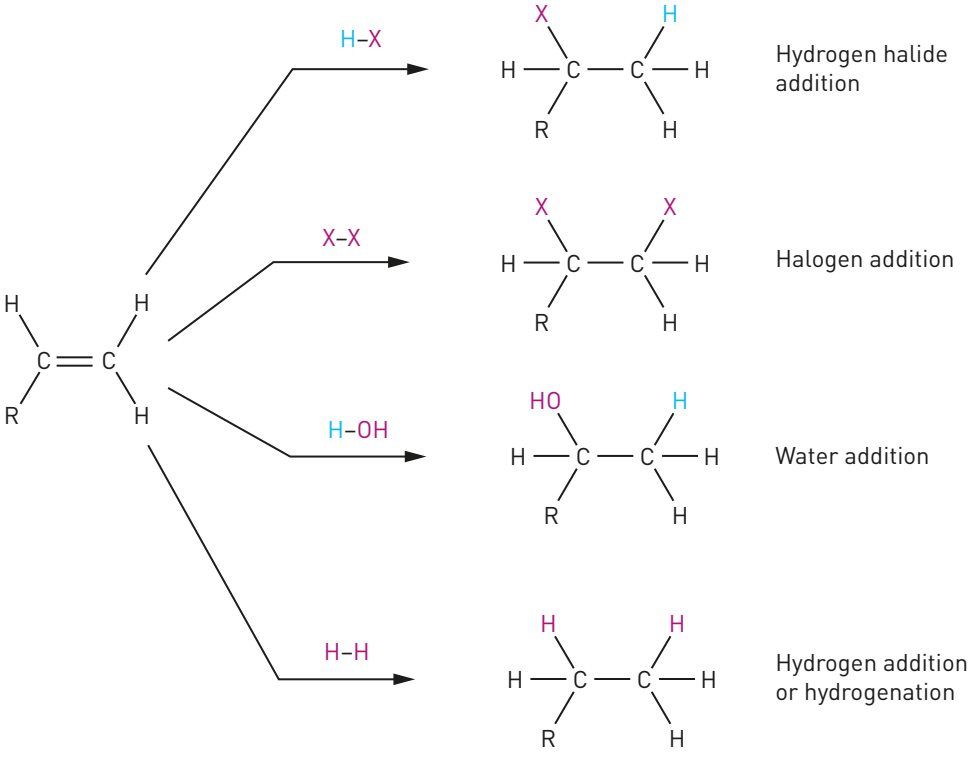


FIGURE 6 Addition reactions of alkenes. R represents any type of carbon-containing group and X is a halogen, such as Cl, Br or I.

Study tip

To help you remember:

- *trans* isomer: the groups have been *trans*ported apart
- *cis* isomer: '*cis*'ters stick together

addition reaction a reaction in which the atoms in a reagent are added to an unsaturated compound

benzene
the simplest aromatic compound with the molecular formula C_6H_6, in which the carbon atoms are arranged in a flat (planar) six-membered ring, with one hydrogen atom attached to each carbon atom

aromatic compound/arene
a cyclic unsaturated hydrocarbon compound that has delocalised clouds of electrons

Benzene

Benzene (Figure 7) is the simplest representative of a class of compounds called **aromatic compounds** (or **arenes**). It is also a hydrocarbon with the formula C_6H_6 and is found in petrol.

In the structures shown in Figure 7, alternating double and single bonds are shown around the ring. The distances between carbon atoms in compounds are known to change with the type of bond present. Representative values for these distances are shown in Table 2.

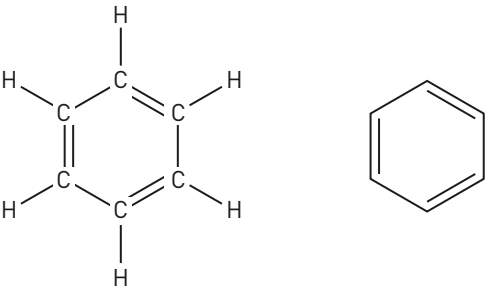


FIGURE 7 Representations of the structure of benzene. The structure on the left shows all of the atoms in benzene. In the structure on the right, the carbon atoms are understood to be located at the corners, and sufficient hydrogen atoms are understood to be attached to each carbon atom to give four bonds

TABLE 2 Carbon–carbon bond types and typical bond distances

Bond type	Bond distance (pm)
C–C single bond	154
C=C double bond	134
C≡C triple bond	109
Bonds between carbon atoms in benzene	140

If benzene actually had alternating single and double bonds, then we would expect that there to be three longer C–C single bonds and three shorter C=C double bonds. When the bond distances are measured, they are the same between all carbon atoms at around 140 pm, which is between the distance expected for single and double C–C bonds.

Chemists explain this with a concept known as resonance. A molecule of benzene can be represented by two Lewis structures shown with a double-headed arrow (Figure 8).

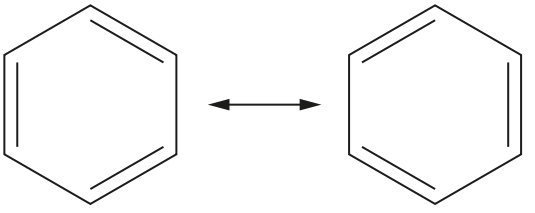


FIGURE 8 Resonance forms of benzenes.

Neither of the two Lewis structures represent the actual benzene molecule. Benzene is not alternating between two different forms; these forms are not in equilibrium. Instead, the real molecule can be described as a hybrid of these two resonance forms, which is more stable than either of them. More accurate descriptions of the bonding in benzene describe the electrons in the double bonds as being delocalised around the ring. These delocalised electrons can be shown as rings or doughnuts above and below the plane that contains the carbon and hydrogen nuclei (Figure 9).

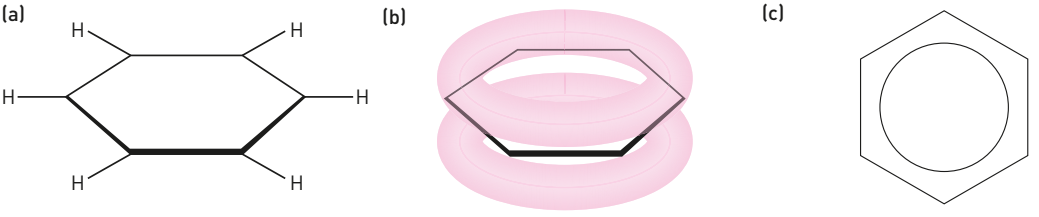


FIGURE 9 (a) The localised bonds in benzene between carbon and hydrogen atoms and (b) the clouds of delocalised electrons. (c) Another way of depicting benzene.

SUGGESTED PRACTICAL

7.4 Testing for saturation
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The extra stability that comes from these delocalised electrons explains why benzene does not readily undergo the addition reactions that are usually seen for alkenes. If benzene were to undergo an addition reaction, then the extra stability that comes from having resonance would be lost. For example, benzene can be hydrogenated just as alkenes can, but higher pressures are required for the reaction.

Instead, benzene typically undergoes substitution reactions, where a hydrogen atom is replaced by another group of atoms (Figure 10). These substitution reactions often require quite reactive reagents or higher temperatures than are required for **substitution reactions** of alkenes; this is because of the high stability of benzene. Benzene is a very important industrial chemical because it can be used to make many other compounds.

substitution reaction
a reaction in which an atom or group of atoms in a molecule is replaced by another atom or group of atoms

Benzene has a boiling point of 80.1°C, which is only 0.65°C lower than that of its fully saturated analogue cyclohexane (80.75°C). The lower number of electrons in benzene (because benzene has six fewer hydrogen atoms than cyclohexane) should lead to weaker intermolecular forces, but this is compensated for by the clouds of delocalised electrons in benzene, which are easily distorted.

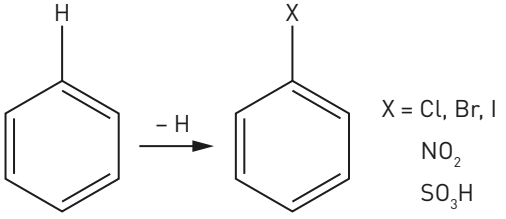

FIGURE 10 Substitution reactions of benzene can be used to replace a hydrogen atom with other atoms such as halogens or nitro or sulfonic acid groups.

CHECK YOUR LEARNING 7.4

Describe and explain

1 **Explain** the differences between linear, branched and cyclic alkanes by drawing structures of molecules of each type, all containing six carbon atoms.

2 **Explain** why alkanes do not react with hydrochloric acid (HCl) but alkenes do.

Apply, analyse and interpret

3 Early organic chemists found that it was possible to isolate only three different isomers of dimethylbenzene, in which two of the H atoms in benzene are replaced by CH_3 groups. **Interpret** the importance of this observation by explaining why the scientists then rejected a structure of benzene with alternating double and single bonds around the ring.

Investigate, evaluate and communicate

4 **Investigate** the historical processes that led to the proposal for the modern structure of benzene. **Construct** a timeline to communicate the main contributions by different chemists.

You can find the following resources for this section on your obook assess:

- » Student book questions
  Check your learning 7.4
- » Suggested practical
  7.4 Testing for saturation
- » Weblink
  Hydrocarbons and climate change
- » Weblink
  Hydrocarbons as a resource

CHAPTER 7

Review

Chapter summary

7.1
- Ionic compounds are held together as solids by the attractions between oppositely charged ions. The ions are held in a repeating three-dimensional arrangement called a crystalline lattice structure.
- Ionic compounds do not conduct electricity in the solid state but can conduct electricity if dissolved in solution or in the molten state.
- Ionic compounds usually have relatively high melting and boiling points, reflecting the strength of the electrostatic interactions between ions. Ionic compounds as solids are relatively hard but brittle. Many ionic compounds are water soluble.

7.2
- Metals in the free electron model of metallic bonding are regarded as positively charged ions surrounded by a sea of delocalised electrons. The electrostatic attractions between the cations and negatively charged electrons hold metallic solids together.
- Metals are malleable, ductile and good conductors of electricity. Generally, they have high boiling points.
- The boiling points of metals decrease going down the first two groups of the periodic table. This is because as the atomic radius increases, the electrostatic attraction decreases.

7.3
- Covalent bonds occur when non-metal atoms share electrons.
- Simple molecules with covalent bonds have low melting and boiling points and are non-conductive. Giant covalent networks consist of many covalent bonds between atoms and generally cannot conduct electricity (although there are exceptions such as graphite). They also have high melting points.
- Allotropes are forms of the same element that have a different structure. For example, diamond and graphite are allotropes of carbon.

7.4
- Hydrocarbons are compounds composed entirely of carbon and hydrogen atoms.
- Alkanes (saturated hydrocarbons) contain the maximum number of hydrogen atoms possible for the given number of carbon atoms. Alkenes (unsaturated hydrocarbons) contain one or more double bonds between carbon atoms.
- Benzene (C_6H_6) is composed of six carbon atoms arranged in ring, with one hydrogen atom attached to each carbon atom.

Key terms

- addition reaction
- alkane
- alkene
- alkyne
- allotrope
- aromatic compound/arene
- benzene
- crystalline lattice
- delocalised electrons
- ductile
- electron gain enthalpy
- free electron model of metallic bonding
- giant covalent network
- hydrocarbon
- intermolecular
- intramolecular
- ionic bonding
- ionic compound
- ionisation energy
- lattice energy
- malleable
- metallic bond
- saturated hydrocarbon
- substitution reaction
- unsaturated hydrocarbon

Key formulas

Molecular formula of a linear or branched alkane	C_nH_{2n+2}
Molecular formula of a cycloalkane with one ring	C_nH_{2n}
Molecular formula of a linear or branched alkene	$C_nH_{2n+2-2P}$ where P is the number of double bonds
Molecular formula of a linear or branched alkyne	C_nH_{2n-2}

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; ***=high.

Multiple choice

1 Which one of the following shows the order of boiling points correctly from lowest to highest?

A $MgCl_2$, Cl_2, Mg
B Mg, Cl_2, $MgCl_2$
C Cl_2, $MgCl_2$, Mg
D $MgCl_2$, Mg, Cl_2

2 Which of the following substances is hard but brittle, has a high melting point and conducts electricity as a liquid but not as a solid?

A Cu
B CaO
C N_2
D SiO_2

3 Which of the following correctly describes metallic bonding?

A interactions between pairs of metal atoms with electrons shared between them
B interactions between positively charged atomic nuclei and outer sphere electrons held in valence orbitals
C interactions between electrons and negatively charged metallic anions
D electrostatic interactions between positively charged metal ions and freely mobile electrons

4 Which of the following is *not* a general characteristic of ionic compounds?

A High boiling points
B High electrical conductivity in the solid state
C Hard but brittle solids
D An ordered arrangement in the solid state

5 Which of the following molecular formulas could represent a linear alkyne?

A C_6H_{12}
B C_6H_{14}
C C_6H_{10}
D C_6H_6

6 Which of the following substances does not contain delocalised electrons?

A silver
B diamond
C benzene
D graphite

7 Which of the following is not a hydrocarbon?

A ethane
B propane
C ethanol
D methane

8 Which of the following is not an allotrope of carbon?

A amorphous carbon
B graphite
C diamond
D carbonate

9 Which of the following is not an ionic compound?

A CO_3^{2-}
B MgCl
C $CaCl_2$
D FeO

Short answer

Describe and explain

★ **10** **Describe** what is occurring at the molecular level as water is heated and changes from a liquid to steam.

★ **11 Explain** why benzene typically undergoes substitution reactions rather than the addition reactions usually seen for alkenes.

★ **12 Explain** the main features of ionic bonding by drawing a labelled diagram.

★ **13 Explain** why methane has the lowest boiling points of the alkanes.

Apply, analyse and interpret

★★ **14** The melting points of magnesium chloride and aluminium chloride differ by more than 1200°C. Aluminium chloride melts at 192.6°C, only 135°C higher than the melting point of silicon tetrachloride. **Interpret** this observation and provide an example of a test that could be used to confirm your interpretation.

★★ **15 Analyse** the densities for the given allotropes of carbon (Table 1) and explain the trend observed.

TABLE 1 Allotropes of carbon

Allotrope	Density (g cm^{-3})
Amorphous carbon	1.8–2.1
Graphite	2.3
Diamond	3.5

★★ **16 Apply** your understanding of the structures of solids to explain the following observation. Borax ($Na_2B_4O_7{\cdot}10H_2O$), naphthalene ($C_{10}H_8$) and camphor ($C_{10}H_{16}O$) are all compounds used to control insects. Both naphthalene and camphor have distinctive odours and are used in mothballs, while borax does not smell.

★★ **17 Analyse** the following information, name the type of reaction described and draw possible structures of both compounds A and B.

Reaction of A (C_6H_{10}) with water in the presence of a trace amount of an acid resulted in a single product B with the molecular formula $C_6H_{12}O$.

★★★ **18** Tests were carried out to determine the identity of the following compounds, each of which have six carbon atoms in their molecular formula: cyclohexane, hex-2-ene and benzene. **Analyse** the test results and assign identities to the three unknowns A–C.

- Test 1: Bromine was dissolved in carbon tetrachloride and added to A, B and C. Only C reacted, with disappearance of the reddish-brown colour of bromine.
- Test 2: A reacted with hydrogen under high temperature and pressure conditions to give a product identical to B.
- Test 3: C was found to be a mixture that could be separated into two compounds with identical molar masses. A and B were found to be pure compounds.
- Test 4: When each of the pure compounds from C was subjected to hydrogenation, the product was identical in both cases.

Investigate, evaluate and communicate

19 Investigate the processes that cause the problem known as 'tin pest'.

FIGURE 1 Deterioration of a tin-based coin

★★ **20 Investigate** a Material Safety Data Sheet for benzene. Chemical suppliers will often have this information available on their websites or you may have access to a database such as ChemWatch. Write a summary of the main hazards of working with benzene.

★★★ **21 Investigate** the most common misconceptions among high school students about ionic bonding. Create a poster to communicate to your fellow students about why a particular misconception is wrong and show the correct information.

★★★ **22 Evaluate** the trend in boiling points for the first 20 linear alkanes. Create a graph to show the trend. Using a spreadsheet program such as MS Excel, fit a curve to the data.

★★★ **23** Concept maps are a useful way to summarise and communicate information about a topic. Concepts are linked by statements. Complete the concept map in Figure 2, and then **create** a separate, new concept map to communicate information about the concepts involved in metallic bonding.

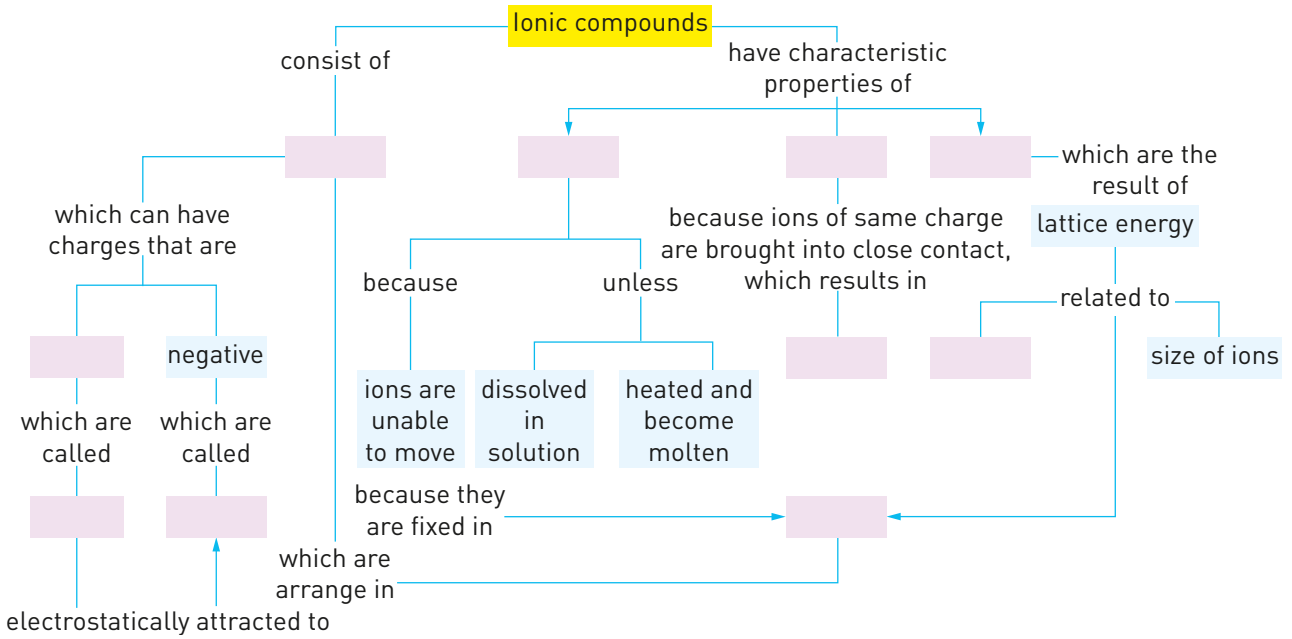

FIGURE 2 Complete this concept map.

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- **Student book questions**
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- **Revision notes**
  Chapter 7
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CHAPTER 8

Mole concept and law of conservation of mass

Until very recently, scientists had not seen an individual atom. This meant that the structure of atoms and the way that they behaved had to be inferred from data gathered from millions of different chemical reactions. Chemists know that atoms are very small, so to make it easier, atoms are counted in large groups, rather than individually. Words such as 'mole' are used to describe and compare the large groups of atoms found in every gram of a substance. When chemists compared the number of moles of atoms in reactants and products, they found that they were the same. This led to the law of conservation of mass, which states that atoms cannot be created or destroyed in a chemical reaction.

OBJECTIVES

- → Recognise that a mole is a precisely defined quantity of matter equal to Avogadro's number of particles.
- → Appreciate the law of conservation of mass and understand that the mole concept relates mass, moles and molar mass.
- → Understand that the empirical formula expresses the simplest whole number ratio of elements in a compound.
- → Use the appropriate stoichiometric ratio to determine that reactants can be limiting.
- → Appreciate that experimental yield can be different from theoretical yield.
- → Use appropriate mathematical representation to solve problems and make predictions, including using the mole concept to calculate the mass of reactants and products; amount of substance in moles; number of representative particles; and molar mass of atoms, ions, molecules and formula units.
- → Use appropriate mathematical representation to solve problems and make predictions, including determining the percentage composition from relative atomic masses; empirical formula of a compound from the percentage composition by mass; and molecular formula of a compound from its empirical formula and molar mass.
- → Calculate percentage yield from experimental or given data.

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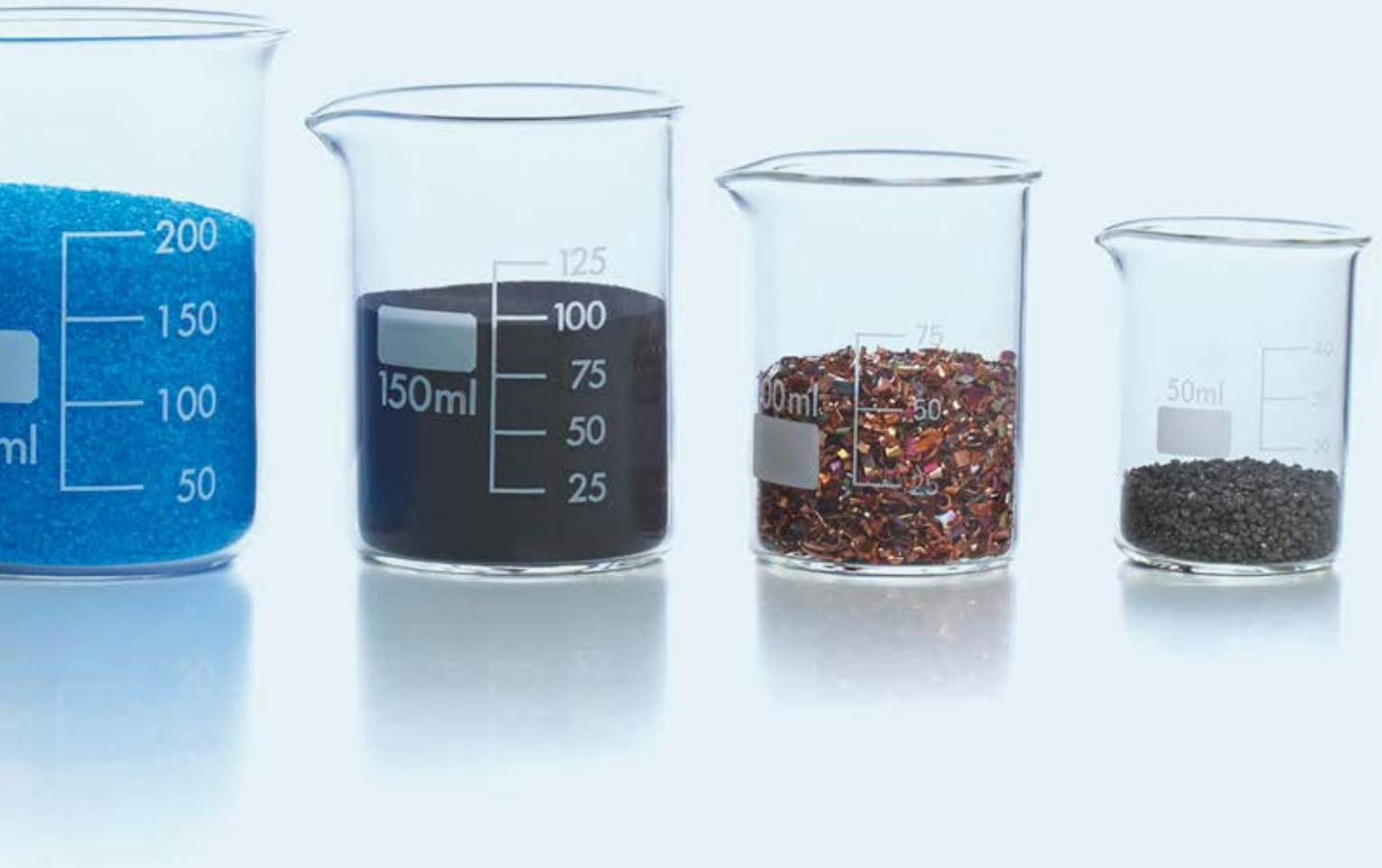
FIGURE 1 Each of these beakers contains one mole of a substance. The mole is one of the base international (SI) units. One mole is equal to 6.02×10^{23} atoms.

PRACTICALS

MANDATORY PRACTICAL	**8.5** Finding the percentage yield of magnesium oxide
SUGGESTED PRACTICAL	**8.4** Investigating limiting reagents



8.1 The mole

KEY IDEAS

In this section, you will learn about:

- Avogadro's number (6.02×10^{23})
- the mole as a quantity of atoms or molecules.

When first learning to count, we use special names to refer to groups of numbers; such as, ones, tens, hundreds, thousands and even millions. The names do not always refer to numbers with zeros. A pair usually means two. Trio means three. One dozen is the same as twelve. As the numbers become larger, they can be called billions (10^9, one thousand million) or even googols (10^{100}).

Because molecules and atoms are so small, it is almost impossible to count them individually. Instead, chemists count the number of atomic or molecular particles in large groups of numbers. The most common of these number groups is 602 000 000 000 000 000 000 000 or 6.02×10^{23}. This number is named after the scientist Amedeo Avogadro, and is often called Avogadro's number, or **Avogadro's constant**.

Avogadro's constant
the number of atoms or molecules in one mole of a substance, equivalent to 6.02×10^{23} particles

Number of molecules = number of moles of substance × Avogadro's number

In the 20th century, scientists linked this large number of atomic particles to the mass of the particles and called the concept *Kilogrammemolekuel* or **mole**. The scientists estimated that there were 6.02×10^{23} atoms (Avogadro's number) of ^{12}C in 12 grams of carbon. As testing has become more advanced, this number of atoms has varied slightly, but the overall concept remains the same. Today, the mole is defined as 'the amount of substance that contains the same number of particles as there are atoms in exactly 12.000 g of ^{12}C atoms'.

mole
the quantity of a substance containing the same number of particles as there are atoms in exactly 12.000 g of carbon-12, which is the same as Avogadro's number (6.02×10^{23})

- One pair of molecules of water = 2 particles of H_2O
- One dozen molecules of water = 12 particles of H_2O
- One billion molecules of water = 10^9 particles of H_2O
- **One mole of water = 6.02×10^{23} particles of H_2O**
- One googol of water = 10^{100} particles of H_2O

When calculating the number of particles, scientists often use symbols instead of words. In chemistry, the mole is represented by the letter n and is abbreviated to 'mol'. Instead of writing 'the number of moles of water', a chemist may write $n(H_2O)$.

Each water molecule has a single oxygen atom and two hydrogen atoms. This means two molecules of water have a total of two oxygen atoms and four hydrogen atoms. If the number of water molecules was one mole, then there would be one mole of oxygen atoms and two moles of hydrogen atoms. Although the number of water molecules has increased (to 6.02×10^{23} molecules), the ratio of hydrogen atoms to oxygen atoms has not changed.

FIGURE 1 One mole of a solid, a gas and a liquid.

WORKED EXAMPLE 8.1

a A sample of dry ice (frozen CO_2) contains 4.5 mol of CO_2 molecules. Calculate the number of CO_2 molecules.

b Calculate the number of oxygen atoms in 4.5 mol of CO_2 molecules.

c If there are 15.0×10^{23} molecules of CO_2 in a cube of dry ice, how many moles does this represent?

SOLUTION

a Number of CO_2 molecules = number of moles of CO_2 × Avogadro's number

$= n(CO_2) \times 6.02 \times 10^{23}$

$= 4.5 \times 6.02 \times 10^{23}$

$= 27.1 \times 10^{23}$

$= 2.71 \times 10^{24}$ molecules of CO_2

b One molecule of CO_2 = contains two atoms of oxygen atoms. Therefore:

The number of particles of oxygen = 2 × the number of particles of carbon dioxide

$n(O) = 2 \times n(CO_2)$ mol

$= 2 \times 4.5$ mol oxygen atoms

$= 9$ mol oxygen atoms

Number of O atoms = $n(O)$ mol $\times 6.02 \times 10^{23}$

$= 9 \times 6.02 \times 10^{23}$

$= 54.2 \times 10^{23}$

$= 5.42 \times 10^{24}$ atoms of oxygen

c 1 mole of $CO_2 = 6.02 \times 10^{23}$ molecules of CO_2

$$n(CO_2) = \frac{15.0 \times 10^{23}}{6.02 \times 10^{23}}$$

$= 2.49$ mol

Study tip

Revise the index laws before completing these questions.

$\frac{10^a}{10^b} = 10^{a-b}$

$10^a \times 10^b = 10^{a+b}$

$10^0 = 1$

CHECK YOUR LEARNING 8.1

Describe and explain

1 **Explain** the concept of a mole. Refer to Avogadro's number in your answer.

2 **Identify** how many atoms are in 12.000 g of ^{12}C.

Apply, analyse and interpret

3 **Determine** the number of hydrogen atoms in:

a 1 mol of hydrogen atoms

b 1 mol of H_2 gas

c 0.5 mol of HCl

d 3.5 mol of H_2O.

4 In one mole of CH_3COOH, **determine** how many molecules there are of:

a carbon **b** hydrogen **c** oxygen

Investigate, evaluate and communicate

5 Imagine you need to explain the mole concept to a new student in your class. **Devise** a series of steps that will help them to calculate the number of molecules in 1.5 moles of water (H_2O).

Check your obook assess for these additional resources and more:

» Student book questions
Check your learning 8.1

» Video weblink
Index laws

» Weblink
The mole

» Weblink
Avogadro's constant

8.2 The mass of a mole

KEY IDEAS

In this section, you will learn about:

- molar mass as the mass of one mole of a substance
- the mole as being equivalent to the mass of a substance divided by its molar mass.

Law of conservation of mass

The law of conservation of mass states that mass is neither created nor destroyed in a chemical reaction. Consider the atoms that made up the carbohydrates in prehistoric plants, were eaten by dinosaurs, digested into individual atoms and rearranged into new molecules in the dinosaur's body. These molecules were later broken down into the same number and type of atoms when the dinosaur died. These same atoms were recycled and still exist around us today.

If the number and type of atoms in a closed system do not change, then the mass of atoms in that closed system will not change. For example, one carbon atom and two oxygen atoms have the same mass as one molecule of carbon dioxide.

$$\text{mass } (CO_2) = \text{mass (1 carbon atom)} + \text{mass (2 oxygen atoms)}$$

This can be extended to larger numbers of carbon and oxygen atoms.

$$\text{mass (1 mol } CO_2) = \text{mass (1 mol carbon atoms)} + \text{mass (2 mol oxygen atoms)}$$

Molar mass

molar mass
the mass of one mole of a substance

The mass of one mole of an element is known as the **molar mass**, and is given the symbol M. The molar mass is measured in g/mol or g mol^{-1}. If the molar mass of carbon-12 is 12 g mol^{-1}, then the mass of one mole of carbon-12 atoms is 12 g. This means the mass of 6.02×10^{23} atoms of carbon-12 (1 mol) is also 12 g.

$$\begin{aligned} M(\text{carbon-12}) &= \text{molar mass of carbon-12} \\ &= \text{mass of one mol of carbon-12} \\ &= \text{mass of } 6.02 \times 10^{23} \text{ atoms of carbon-12} \\ &= 12 \text{ g mol}^{-1} \end{aligned}$$

relative atomic mass
the mass of an atomic once the mass and percentage abundance of all isotopes have been considered, relative to the carbon-12 isotope

It is this number that contributes to the **relative atomic mass** specified near each element on the periodic table (see Chapter 2).

If the molar mass of each element is known, it can be used to determine the mass of a molecule. For example, the mass of one mole of carbon dioxide atoms (6.02×10^{23} molecules of carbon dioxide) can be determined from the molar mass of carbon and oxygen.

$$\begin{aligned} \text{Mass of 1 mol of carbon dioxide} &= \text{mass of 1 mol of carbon} + \text{mass of 2 mol of oxygen} \\ &= 12.01 \text{ g} + 2 \times 16.00 \text{ g} \\ &= 12.01 \text{ g} + 32.00 \text{ g} \\ &= 44 \text{ g} \end{aligned}$$

These calculations can be simplified into a single formula:

$$n = \frac{m}{M}$$

where n is the number of moles, m is the mass of a sample, and M is the molar mass of the substance.

FIGURE 1 These beakers contain one mole of copper(II) sulfate pentahydrate (on the scale) and (from top right clockwise) sucrose, potassium manganate(VII), sodium chloride, iron filings, copper turnings and nickel(II) chloride. The mass of a mole of a substance can be found by using the molar mass of its elements.

WORKED EXAMPLE 8.2

Refer to a periodic table (Appendix 1) for this question. A glass contains 150 g of pure water.

a Calculate the number of moles of H_2O.

b Calculate the number of hydrogen atoms in the glass of water.

c What mass of water contains 2.04×10^{24} molecules?

SOLUTION

a $m(H_2O) = 150\ g$

$M(H_2O) = 2 \times$ relative atomic mass of hydrogen $+ 1 \times$ relative atomic mass of oxygen

$= 2 \times 1.01 + 1 \times 16.00$

$= 2.02 + 16.00$

$= 18.02\ g\ mol^{-1}$

$n = \frac{m}{M}$

$= \frac{150\ g}{18.02\ g\ mol^{-1}}$

$= 8.32\ mol$

b If each molecule of water contains two hydrogen atoms, then:

$n(H) = 2 \times n(H_2O)$

$= 2 \times 8.32\ mol$

$= 16.64\ mol$

Number of hydrogen atoms = $n(H) \times$ Avogadro's number

$= 16.64 \times 6.02 \times 10^{23}$

$= 100.17 \times 10^{23}$

$= 1.00 \times 10^{25}$ atoms

c Before the equation $n = \frac{m}{M}$ can be used, the number of moles of water must be determined.

$n(H_2O) = \frac{2.04 \times 10^{24}\ \text{molecules}}{\text{Avogadro's number}}$

$= \frac{2.04 \times 10^{24}\ \text{molecules}}{6.02 \times 10^{22}}$

$= 3.39\ mol$

The values now available are:

$n(H_2O) = 3.39$ mol

$M(H_2O)$ = 2 × relative atomic mass of hydrogen + 1 × relative atomic mass of oxygen
= 2 × 1.01 + 1 × 16.00
= 2.02 + 16.00
= 18.02 g mol^{-1}

If the equation $n = \frac{m}{M}$ is rearranged, then:

$m(H_2O) = n(H_2O) \times M(H_2O)$
= 3.39 mol × 18.02 g mol^{-1}
= 61.1 g

The mass of water is 61.1 g.

Study tip

The equation $n = \frac{m}{M}$ can be easily rearranged by exchanging the molar mass (M) and the number of moles (n) to form $M = \frac{m}{n}$.

CHECK YOUR LEARNING 8.2

Describe and explain

1 **Describe:**
- **a** mass
- **b** relative atomic mass
- **c** molar mass.

2 **Calculate** the mass of:
- **a** 0.25 mol of water
- **b** 0.2 mol of zinc
- **c** 18×10^{23} atoms of magnesium
- **d** 18×10^{23} molecules of magnesium chloride ($MgCl_2$).

Apply, analyse and interpret

3 **Calculate** the mass of:
- **a** 1 mol of Na
- **b** 1 mol of Cl
- **c** 1 mol of NaCl
- **d** 2.35 mol of NaCl.

4 **Calculate** the amount, in moles, of:
- **a** 24 g of O
- **b** 2.7 g of CO_2
- **c** 2.0 g of NaCl
- **d** 4.6 g of ethanol (C_2H_5OH)

5 Use the periodic table in Appendix 1 to **determine** the:
relative atomic mass of barium
- **b** mass of one mole of barium atoms.

6 **Determine** the molar mass of a pure substance if 3.5 moles of the substance has a mass of 59.584 g.

Investigate, evaluate and communicate

7 Prior to 1961, there were two sets of atomic weights. Chemists used naturally occurring oxygen, which was a mixture of isotopes: 99.76% oxygen-16, 0.04% oxygen-17 and 0.20% oxygen-18. Physicists used the mass of pure oxygen-16 because they made measurements by mass spectrometry.
Evaluate the two measurements by determining the mass of 1 mole of from a chemist's sample and 1 mole from a physicist's sample.

Check your obook assess for these additional resources and more:

- » Student book questions
  Check your learning 8.2
- » Video
  Calculating molar mass
- » Weblink
  Conservation of mass
- » Weblink
  Molar mass

8.3 Empirical and molecular formulas

KEY IDEAS

In this section, you will learn about:

- using the relative atomic mass of a compound to determine its percentage composition
- the empirical formula as the simplest whole-number ratio of elements in a compound
- using the percentage composition by mass to determine the empirical formula
- determining the molecular formula of a compound from its empirical formula and molar mass.

Over 200 years ago, electricity was used to separate the atoms in water to form two different gases, oxygen and hydrogen. This process, known as electrolysis, produced exactly twice as much hydrogen gas as oxygen gas. This suggested that the ratio of the elements was two hydrogen atoms to one oxygen atom.

Although the ratio of atoms is two hydrogen atoms to each oxygen atom, oxygen atoms are much larger and have more mass than hydrogen atoms. This means the single oxygen atom contributes more to the mass of the water than the two hydrogen atoms. The amount of mass each element contributes to a molecule can be expressed as a **percentage composition**. For example, the percentage composition of water is 11.21% hydrogen and 88.81% oxygen (Worked example 8.3A).

percentage composition the percentage that each element contributes to the mass of a compound

The percentage composition of an element (x) in compound can be calculated by the following formula:

$$\%\text{ of } x \text{ in a compound} = \frac{\text{mass of } x \text{ in 1 mol of a compound}}{\text{molar mass of compound}} \times \frac{100}{1}\%$$

WORKED EXAMPLE 8.3A

Calculate the percentage of hydrogen by mass in water.

SOLUTION

$$\%\text{ of hydrogen in water} = \frac{\text{mass of hydrogen in 1 mol of water}}{\text{molar mass of water}} \times \frac{100}{1}\%$$
$$= \frac{2 \times 1.01}{2 \times 1.01 + 16.00} \times \frac{100}{1}\%$$
$$= \frac{2.02}{18.02} \times \frac{100}{1}\%$$
$$= 11.21\%\ H_2O$$

Empirical formula

Many modern medicines are purified versions of traditional treatments; for example, aspirin was traditionally isolated from the bark of the willow tree. Before a new compound can be manufactured, its chemical formula must be determined. First, the percentage composition of all the elements in the compound are determined, and then the relative proportions of each element are identified. This simplest whole-number ratio of elements in a compound is called the **empirical formula**. The empirical formula is like the simplest fraction in mathematics – it cannot be subdivided any further.

empirical formula the simplest whole-number ratio of the elements in a compound

An empirical formula can only be determined through an experiment to establish the ratio of the mass of each element. It can be calculated through a series of steps.

1. Identify all the elements present in the sample.
2. Determine the mass of each element.
3. Use the mass to determine the number of moles of each element.
4. Determine the whole-number ratio by dividing all numbers of moles by the smallest number of moles.
5. If necessary, use a multiplication factor to achieve whole numbers.

WORKED EXAMPLE 8.3B

A compound contains 40.00% carbon, 6.70% hydrogen and 53.30% oxygen by mass. Determine the empirical formula of the compound.

SOLUTION

1. Identify all the elements present in the sample.
   carbon, hydrogen, oxygen
2. Determine the mass of each element.
   Assume there is 100 g of the sample because this makes the use of percentages easier.
   C 40.00g : H 6.70 g : O 53.30 g
3. Use the mass to determine the number of moles of each element.
   $\text{C}\ \frac{40.00}{12.01}\ \text{mol}: \text{H}\ \frac{6.70}{1.01}\ \text{mol}: \text{O}\ \frac{53.30}{16.00}\ \text{mol}$
   C 3.33 mol : H 6.63 mol : O 3.33 mol
4. Determine the whole-number ratio by dividing all numbers of moles by the smallest number of moles.
   $\text{C}\ \frac{3.33}{3.33} : \text{H}\ \frac{6.63}{3.33} : \text{O}\ \frac{3.33}{3.33}$
   C 1 : H 2 : O 1
5. If necessary, use a multiplication factor to achieve whole numbers. In this case, all three numbers are already whole numbers. Therefore, the empirical formula of the compound is CH_2O.
   Remember: This is the simplest ratio of numbers and the actual number of element may be a multiple of this.

Molecular formula

molecular formula
the chemical formula of a compound that indicates the number of each elementary atoms present

The **molecular formula** is the complete chemical formula of a compound that shows how many atoms of each element are in the compound. The empirical formula of water is the same as the molecular formula, H_2O. In other molecules, the molecular formula is a whole-number multiple of the empirical formula. For example, the empirical formula of butene is CH_2, but its molecular formula is C_4H_8 (four times the number of atoms found in the empirical formula).

$$\text{Molecular formula } (C_4H_8) = 4 \times \text{empirical formula } (CH_2)$$

Some compounds have the same empirical formula, but different molecular formulas. If a compound has a different molecular formula, it will have different chemical properties. For example, ethene, propene and butene have the empirical formula CH_2 (Table 1). Ethene is used in the manufacture of plastics, antifreeze and solvents. Propene and butene are useful raw materials for a variety of important products.

TABLE 1 A comparison of empirical and molecular formula

Name	Empirical formula	Molecular formula
Ethene	CH_2	C_2H_4
Propene	CH_2	C_3H_6
Butene	CH_2	C_4H_8

The molecular formula can be obtained from the empirical formula if the compound's molar mass is known (Worked example 8.3C).

WORKED EXAMPLE 8.3C

If the empirical formula of compound is CH_2 and the molar mass is 28.054 g mol^{-1}, find the molecular formula of the compound.

SOLUTION

What is known:

M(C) = molar mass of carbon = 12.01 g mol^{-1}

M(H) = molar mass of hydrogen = 1.01 g mol^{-1}

Empirical formula M (CH_2) = 12.01 + 2 × 1.01

= 14.027 g mol^{-1}

Molar mass of the compound = 28.054 g mol^{-1}

$$\text{Multiple of empirical formula} = \frac{\text{molar mass of compound}}{\text{molar mass of empirical formula}}$$

$$= \frac{28.054}{14.027}$$

$$= 2$$

Therefore, the empirical formula must be multiplied by 2.

Molecular formula = 2 × CH_2

= C_2H_4

Study tip

Always write the calculation you are completing in the equation. For example, distinguishing between the empirical and the molecular formula will help to prevent errors.

CHECK YOUR LEARNING 8.3

Describe and explain

1 **Explain** what is meant by the percentage composition of a compound.

2 **Explain** how the empirical formula of a compound can be different from its molecular formula.

Apply, analyse and interpret

3 **Determine** the percentage composition of each element in glucose ($C_6H_{12}O_6$).

4 A new compound has been identified in a plant in the Amazon rainforest, which may have antibacterial properties. It contains 47.37% carbon, 10.59% hydrogen and 42.04% oxygen. **Apply** the appropriate empirical formula for this new compound.

5 A compound is found to contain only boron and hydrogen. If the compound contains 11.55% hydrogen, **determine** its empirical formula.

6 A compound with a molar mass of 128 g mol^{-1} is known to contain 6.3% of hydrogen and 93.7% of oxygen. **Determine** the empirical and molecular formula for this compound.

Investigate, evaluate and communicate

7 **Investigate** how the method of percentage composition can be used by forensic scientists to build a case against a suspect. What types of compounds or materials could be used in this method?

Check your obook assess for these additional resources and more:

- » Student book questions
  Check your learning 8.3
- » Weblink
  Emperical formula
- » Weblink
  Molecular formula
- » Weblink
  Percentage composition

8.4 Stoichiometric ratios

KEY IDEAS

In this section, you will learn about:

- how stoichiometric ratios can be used to determine limiting or excess reactants
- using mathematical representations to predict the mass or number of moles of reactants required or products produced.

FIGURE 1 The correct ratio of flour, butter, eggs and milk is required for a cake to be edible.

When baking a cake, it is important that there is enough of each ingredient. If there is a lot of flour, but limited amounts of butter, then you cannot produce a full-sized cake. Adding more flour will not help make the cake bigger. Instead, the excess flour will just escape from the mixture. This cake mixture analogy is similar to the limitations that occur in chemical reactions. When one reactant is limited, the amount of product will also be limited.

Study tip

When using the coefficients of a chemical equation to find molar ratio, make sure the chemical equation is balanced.

WORKED EXAMPLE 8.4A

Hydrogen gas can react (combust) with oxygen to form water. Determine the molar ratio of oxygen and water molecules.

SOLUTION

1 Write a balanced equation for the reaction.

$$2H_2 + O_2 \rightarrow 2H_2O$$

2 The coefficients (the numbers before each molecule in the equation) for the reaction indicate the relative number of moles of each molecule that is needed.

2 mol of hydrogen : 1 mol of oxygen : 2 mol of water

Therefore, the molar ratio of oxygen and water is: 1 (O_2) : 2 (H_2O)

stoichiometry the numerical relationship between the amount of reactants used and the products produced in a reaction

limiting reactant the reactant that is totally consumed when a chemical reaction is complete; this limits the amount of product that can be produced

excess reactant the reactant that is not totally consumed by a chemical reaction

Stoichiometry is the use of mathematical calculations to determine the relative quantities of reactants or products formed in a chemical reaction. It is usually presented as a ratio of whole numbers (the stoichiometric ratio). If one reactant is limited and is completely used (the **limiting reactant**), then the reaction will cease and no more product will be produced. Any left-over reactants are called **excess reactants**.

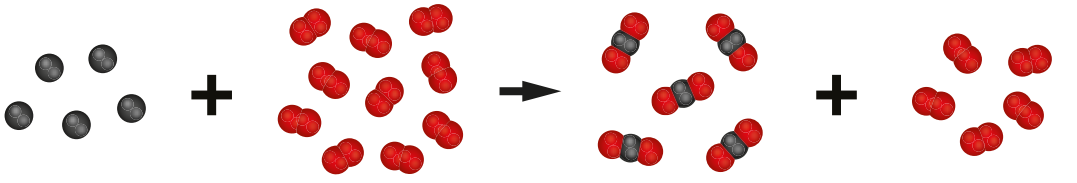


FIGURE 2 The five black C atoms in this reaction are the limiting reactant, because no more product can be produced once the limiting C atoms have all been used in the reaction.

SUGGESTED PRACTICAL

8.4 Investigating limiting reagents
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CASE STUDY 8.4

Formula One: combustion reactions

The chemical reaction that occurs between the fuel and oxygen in a Formula One car's engine demonstrates the idea that if one reactant is limited, the amount of product will also be limited. Too little of either reactant results in a smaller reaction and less available power. The amounts of fuel and oxygen must be in the correct ratio for the chemical reaction to occur efficiently. The balanced combustion reaction of octane can be used as an example.

$$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$$

Although this balanced equation shows two octane molecules reacting with 25 oxygen molecules, in real-life, much larger amounts of reactants (moles) are used. Therefore, the reaction can be described as 2 mol of octane reacting with 25 mol of oxygen, to produce 16 mol of carbon dioxide and 18 mol of water. This is called a **molar ratio** of reactants and products.

molar ratio
the ratio of the amounts in moles of the reactants and products in a reaction

$$2 \text{ mol } C_8H_{18} : 25 \text{ mol } O_2 : 16 \text{ mol } CO_2 : 18 \text{ mol } H_2O$$

If there is a limited amount of one reactant, then only limited amounts of the product can be produced. For example, if there is a lot of oxygen available but only a limited amount of octane (0.5 mol), then only limited amounts of carbon dioxide (4 mol) or water (4.5 mol) can be produced. The limited amount of octane results in less energy being produced in the exothermic reaction. Adding more oxygen into the mixture will not change the amount of carbon dioxide or water being produced or produce more power in the motor. The oxygen is already in excess.

Alternatively, injecting more octane fuel with the oxygen so that both reactants are in the correct ratio will allow the reaction to occur most efficiently.

FIGURE 3 Many Formula One cars have systems that increase in air flow into the engine to ensure the correct ratio of fuel to oxygen for an efficient combustion reaction.

WORKED EXAMPLE 8.4B

Silicon nitride (Si_3N_4) is made by combining silicon (Si) and nitrogen gas (N_2) at a high temperature.

$$3Si(g) + 2N_2(g) \rightarrow Si_3N_4(s)$$

a Determine how many moles of silicon nitride will be produced in this reaction.

b Determine how many moles of silicon are needed to react with 4 mol of nitrogen.

SOLUTION

a Using the molar ratio, 1 mol of Si_3N_4 will be produced.
Molar ratio: 3 Si : 2 N_2 : 1 Si_3N_4

b Double each number in the molar ratio to determine the number of mol of Si needed for 4 mol of N_2.
6 Si : 4 N_2 : 2 Si_3N_4
Therefore, 6 mol of silicon are needed to react with 4 mol of nitrogen.

WORKED EXAMPLE 8.4C

In an experiment, 3.25 g of ammonia is allowed to react with 3.50 g of oxygen according to the reaction:

$$4\,NH_3(g) + 5\,O_2(g) \rightarrow 4\,NO(g) + 6\,H_2O(l)$$

a Identify which reactant is the limiting reagent.

b Determine the mass of nitrogen oxide (NO) produced in this reaction.

c Determine the mass of the excess reactant that remains after the reaction.

SOLUTION

a Determine the number of moles of each reactant and compare it to the molar ratio in the chemical equation.

Molar ratio: $4\,NH_3 : 5\,O_2$

$$n(NH_3) = \frac{m}{M}$$
$$= \frac{3.25\text{ g}}{17.031\text{ g mol}^{-1}}$$
$$= 0.19\text{ mol}$$
$$n(O_2) = \frac{m}{M}$$
$$= \frac{3.50\text{ g}}{31.998\text{ g mol}^{-1}}$$
$$= 0.10\text{ mol}$$

Oxygen is the limiting reagent in this reaction.

b Molar ratio $4\,NH_3 : 5\,O_2 : 4\,NO : 6\,H_2O$

The number of moles of oxygen used is 0.10 mol. Therefore, the amount of nitrogen oxide produced is 0.08 mol.

$n(NO) = 0.08$ mol

$M(NO) = 30.006\text{ g mol}^{-1}$

Rearrange the equation $n = \frac{m}{M}$:

$m(NO) = n(NO) \times M(NO)$
$= 0.08\text{ mol} \times 30.006\text{ g mol}^{-1}$
$= 2.4\text{ g}$

2.4 g of nitrogen oxide is produced.

c If 0.10 mol of oxygen is used, then 0.08 mol of ammonia is used.

Amount of ammonia remaining $= 0.19\text{ mol} - 0.08\text{ mol}$
$= 0.11\text{ mol}$

Convert the amount of ammonia into mass.

$n(NH_3) = 0.11$ mol

$M(NH_3) = 17.031\text{ g mol}^{-1}$

Rearrange the equation $n = \frac{m}{M}$:

$m(NH_3) = n(NH_3) \times M(NH_3)$
$= 0.11\text{ mol} \times 17.031\text{ g mol}^{-1}$
$= 1.87\text{ g}$

1.87 g of ammonia remains after the reaction.

Study tip

Including a statement with the mathematical value allows you to check that you have answered the question that was asked.

CHECK YOUR LEARNING 8.4

Describe and explain

1 **Define**:
- **a** molar ratio
- **b** limiting reactant
- **c** excess reactant.

2 Consider the reaction:

$$Ba_3N_2 + 6H_2O \rightarrow 3Ba(OH)_2 + 2NH_3.$$

Calculate the mass of NH_3 produced if 220 g of barium nitride (Ba_3N_2) reacts with excess water.

3 If 120 g of propane (C_3H_8) is burnt in excess oxygen, **calculate** the mass of water produced.

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

Apply, analyse and interpret

4 90.0 g of iron chloride ($FeCl_3$) reacts with 52 g of hydrogen sulfide (H_2S) to produce iron sulfide and hydrochloric acid.
- **a** Write a balanced chemical equation for the reaction.
- **b** **Determine** the limiting reactant.
- **c** **Calculate** the mass of iron sulfide produced in the reaction.

5 Rescue workers use self-contained breathing apparatus to remove carbon dioxide and water from their exhaled breath. First, potassium oxide removes water from the exhaled breath:

$$KO_2 + H_2O \rightarrow O_2 + KOH$$

The potassium hydroxide (KOH) produced is then available to react with the carbon dioxide:

$$KOH + CO_2 \rightarrow KHCO_3$$

- **a** **Deduce** how much potassium oxide is needed to produce 235 g of oxygen.
- **b** If 123 g of KO_2 is available in the breathing apparatus, **calculate** what mass of carbon dioxide can be removed from the exhaled air.

FIGURE 4 Rescue workers use breathing apparatus

Investigate, evaluate and communicate

6 The chemical equation for photosynthesis is

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2.$$

Synthesise the concepts covered in this chapter to explain why plants in a desert fail to produce enough glucose ($C_6H_{12}O_6$) to grow.

Check your obook assess for these additional resources and more:

- » Student book questions
  Check your learning 8.4
- » Suggested practical
  8.4 Investigating limiting reagents
- » Video
  Balancing a chemical equation
- » Weblink
  Stoichiometry

8.5 Experimental yields

KEY IDEAS

In this section, you will learn about:

- calculating the percentage yield from experimental or given data
- how experimental yield can vary from the theoretical yield.

FIGURE 1 Laboratory scales being used to weigh a product.

Chemical reactions rarely work perfectly to give 100% yield. Chemists measure their effectiveness by comparing the predicted amount of product from their stoichiometric calculations (**theoretical yield**) to the actual amount of product they achieve from the chemical reaction (**experimental yield**). If the theoretical yield and the experimental yield are identical, then the experiment is considered 100% efficient. This is very rare because many factors can affect the reaction and reduce the yield. These factors include impure reactants, other side reactions occurring or lost products through reactants or products remaining on the sides of containers. An added complication is that very few reactions go to completion.

theoretical yield the amount of product predicted from stoichiometric calculations

experimental yield the actual amount of product that is produced in a chemical reaction

The percentage yield can be calculated by the following equation:

$$\text{Percentage yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times \frac{100}{1}\%$$

It is important to remember that the percentage yield can never be higher than 100% because it would mean new atoms had been created in the chemical reaction. This would break the law of conservation of mass.

WORKED EXAMPLE 8.5

When 39.75 g of magnesium carbonate decomposed in an experiment, 15 g of magnesium oxide was formed. Determine the percentage yield of magnesium oxide.

SOLUTION

$$MgCO_3 \rightarrow MgO + CO_2$$

Experimental yield (MgO) = 15 g

Theoretical yield (MgO) = needs to be determined

1 Determine the amount of magnesium carbonate in moles.

$$n(MgCO_3) = \frac{m(MgCO_3)}{M(MgCO_3)}$$

$$= \frac{39.75\ \text{g}}{84.313\ \text{g mol}^{-1}}$$

$$= 0.4715\ \text{mol}$$

2 Determine the molar ratio of MgO to $MgCO_3$.
$1n$ (MgO) : $1n$ ($MgCO_3$)
0.4715 mol : 0.4715 mol
Therefore $n(MgCO_3)$ = 0.4715 mol

3 Calculate the theoretical mass of MgO that should have been produced.
Rearrange the equation $n = \frac{m}{M}$:

$m(MgO) = n(MgO) \times M(MgO)$
$= 0.4717 \text{ mol} \times 40.303 \text{ g mol}^{-1}$
$= 19.00 \text{ g}$

4 Calculate the percentage yield.

$$\text{Percentage yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times \frac{100}{1}\%$$
$$= \frac{15}{19} \times \frac{100}{1}\%$$
$$= 78\%$$

CHECK YOUR LEARNING 8.5

Describe and explain

1 **Explain** the difference between theoretical yield and experimental yield.

2 **Explain** why a percentage yield of 109% is impossible.

Apply, analyse and interpret

3 Magnesium oxide is formed when magnesium ribbon burns in the presence of oxygen:

$$2Mg + O_2 \rightarrow 2MgO$$

a **Determine** the theoretical yield of MgO formed when 12 g of Mg reacts with excess oxygen.

b When a student performed this experiment, they found 18 g of MgO was formed. **Determine** the percentage yield of this reaction.

4 Heptane undergoes a combustion reaction with oxygen to form carbon dioxide and water:

$$C_7H_{16} + 11O_2 \rightarrow 7CO_2 + 8H_2O$$

a **Determine** the theoretical yield of CO_2 formed when 10 g of C_7H_{16} is burnt with excess O_2.

b A chemist analysing this reaction produced 25 g of carbon dioxide. **Determine** the percentage yield for this reaction.

Investigate, evaluate and communicate

5 When a wine bottle has a defective seal, the ethanol (C_2H_5OH) becomes exposed to oxygen. As a result, acetic acid (CH_3COOH) and water are formed.

a **Construct** a balanced chemical equation for this reaction.

b **Decide** if ethanol or oxygen would be the limiting reactant in this reaction. **Justify** your response.

c **Investigate** methods used by wine makers to prevent this reaction from occurring. Present your findings to your peers.

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- » Weblink
  Percentage yield
- » Weblink
  Correct laboratory practices

CHAPTER 8

Review

Chapter summary

8.1
- Avogadro's number is equivalent to 6.02×10^{23}.
- A mole is a quantity of matter that is equal to Avogadro's number of particles.

8.2
- The molar mass of a compound can be determined by adding the individual masses of the elements found in the periodic table.
- The law of conservation of mass states that elements cannot be created or destroyed in a chemical reaction. This suggests that the mass and number of elements (in moles) in the reactants must be equivalent to the mass and number of elements (in moles) in the products.

8.3
- The percentage composition of a compound can be determined from the relative atomic masses.
- The empirical formula is the simplest whole-number ratio of elements in a compound.
- The percentage composition of mass can be used to determine the empirical formula.
- The molecular formula of a compound can be determined from its empirical formula and the molar mass.

8.4
- The limiting reactant is the reactant that is totally consumed when the chemical reaction is complete.
- Stoichiometric ratios can be used to determine the limiting or excess reactants.
- The limiting reactant can be used to predict the mass or number of moles of products produced.

8.5
- The amount of experimental product produced in a chemical reaction will always be less than the predicted theoretical yield. This may be due to impure reactants, other side reactions occurring or lost products through reactants or products remaining on the sides of containers.
- The percentage yield can be determined by dividing the experimental yield by the theoretical yield and multiplying by 100.

Key terms

- Avogadro's constant
- empirical formula
- excess reactant
- experimental yield
- limiting reactant
- molar mass (M)
- molar ratio
- mole
- molecular formula
- percentage composition
- relative atomic mass
- stoichiometry
- theoretical yield

Key formulas

Number of molecules	Number of molecules = number of mole of substance × Avogadro's number
Number of moles	$n = \frac{m}{M}$
Percentage composition	$\%\text{ of } x \text{ in a compound} = \frac{\text{mass of } x \text{ in 1 mol of a compound}}{\text{molar mass of compound}} \times \frac{100}{1}\%$
Multiple of empirical formula	$\text{Multiple of empirical formula} = \frac{\text{molar mass of compound}}{\text{molar mass of empirical formula}}$
Percentage yield	$\%\text{ yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times \frac{100}{1}\%$

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; ***=high.

Multiple choice

1 The number of moles of hydrogen in 1 mol of CH_4O is:

A 1
B 2
C 3
D 4

2 The number of molecules of methane (CH_4) in 0.123 mol of methane is:

A 5
B 2.04×10^{-25}
C 2.46×10^{-2}
D 7.40×10^{22}

3 The combustion of propane (C_3H_8) in excess oxygen produces carbon dioxide and water.

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

If 3.4 mol of oxygen is used in this reaction, the number of moles of carbon dioxide produced would be:

A 2.0
B 3.4
C 0.68
D 2.7

4 The smallest number of molecules would be found in:

A 0.5 g of CH_3OH
B 0.5 g of N_2
C 0.5 g of NH_3
D 0.5 g of C_2H_2.

5 A forensic scientist analyses an unknown compound and finds that it contains 72.7% oxygen and 27.3% carbon. The empirical formula of the compound could be:

A CO
B CO_2
C CO_3
D C_3O_7

Short answer

Describe and explain

★ **6** **Identify** which symbol can be used to write the number of moles of carbon dioxide.

★ **7** **Describe** the factors that can prevent the experimental yield being equivalent to the theoretical yield.

★ **8** **Calculate** the number of atoms of magnesium in 24.3 g of Mg.

★ **9** **Calculate** the molar mass of caffeine ($C_8H_{10}N_4O_2$).

★ **10** **Calculate** the number of hydrogen atoms present in 43.8 g of urea ($(NH_2)_2CO$).

★★ **11** **Identify** an example of a substance that has the same empirical formula and molecular formula. **Explain** your answer.

★★ **12** **Calculate** each of the following found in 3.5 moles of calcium acetate ($Ca(CH_3COO)_2$).

a Moles of calcium atoms
b Moles of carbon atoms
c Moles of hydrogen atoms
d Moles of oxygen atoms
e Number of calcium ions
f Number of carbon atoms
g Mass of hydrogen
h Mass of oxygen

★★ **13** Urea ($(NH_2)_2CO$) is a fertiliser used by many farmers to replace nitrogen in the soil. It can be produced by reacting ammonia (NH_3) with carbon dioxide:

$$2NH_3 + CO_2 \rightarrow (NH_2)_2CO + H_2O$$

If 849.2 g of ammonia is reacted with 1223 g of carbon dioxide, **calculate** how much urea is formed.

FIGURE 1 Farmer applying fertiliser

Apply, analyse and interpret

★ **14** **Calculate** how many molecules there are in:

a one mole of hydrogen

b one mole of H_2O

c one mole of $CuSO_4$.

★ **15** In one mole of CH_3COOH, **determine** how many moles there are of:

a carbon

b hydrogen

c oxygen.

★★ **16** Most living things use glucose ($C_6H_{12}O_6$) and oxygen to produce carbon dioxide, water and energy. This reaction can be represented by the equation:

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$$

a If 968 g of glucose was used, **calculate** the mass of carbon dioxide produced.

b **Calculate** the number of grams of oxygen needed to make 550 g of carbon dioxide.

★★ **17** The amino acid serine has the percentage composition 34.29% carbon, 6.71% hydrogen, 45.67% oxygen and 13.33% nitrogen. If the molar mass is 105.9 g mol^{-1} **determine** the:

a empirical formula of serine

b molecular formula of serine.

★★ **18** Phosphorus reacts with a limited amount of oxygen to form phosphorus trioxide:

$$P_4 + 3O_2 \rightarrow P_4O_6$$

If enough oxygen is available, then the phosphorus trioxide reacts further to produce phosphorus pentoxide.

$$P_4O_6 + 2O_2 \rightarrow P_4O_{10}$$

a If a container holds 5.77 g of phosphorus and 5.77 g of oxygen, **determine** how much oxygen will be in excess in the first reaction.

b Use your answer from part a to **determine** the mass of phosphorus pentoxide that will be produced.

★★ **19** Fool's gold is a common iron ore that contains the mineral iron pyrite (FeS_2). Typically, the percentage by mass of FeS_2 in a sample of fool's gold is 90–95%. The actual percentage in a sample can be determined by gravimetric analysis. The sulfur in FeS_2 is converted to sulfate (SO_4^{2-}), which is then mixed with an excess of barium chloride ($BaCl_2$) to form barium sulfate ($BaSO_4$) according to the equation:

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

FIGURE 2 Fools gold is also known as pyrite

When the reaction has gone to completion, the $BaSO_4$ precipitate is collected in a filter paper and carefully washed. The filter paper and its contents are then transferred to a crucible. The crucible and its contents are heated until constant mass is achieved. The data for an analysis of a mineral sample is shown in Table 1.

TABLE 1 Gravimetric analysis of minerals

Initial mass of mineral sample	14.3 g
Mass of crucible and filter paper	123.40 g
Mass of crucible, filter paper and dry $BaSO_4$	174.99 g
M(FeS_2)	120.0 g mol^{-1}
M($BaCl_2$)	208.3 g mol^{-1}
M($BaSO_4$)	233.4 g mol^{-1}

Analyse the data to **determine** the percentage by mass of FeS_2 in this mineral sample.

★★ **20** The strength of eggshell depends on the amount of calcium carbonate ($CaCO_3$) in the eggshell. The percentage of calcium carbonate in the eggshell can be determined by gravimetric analysis. A total of 0.412 g of clean dry eggshell is completely dissolved in a minimum volume of dilute hydrochloric acid:

$$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(l)$$

An excess of a basic solution of ammonium oxalate ($(NH_4)_2C_2O_4$) is then added to form crystals of calcium oxalate monohydrate ($CaC_2O_4.H_2O$).

The suspension is filtered and the crystals are dried to constant mass. A total of 0.523 g of $CaC_2O_4.H_2O$ is collected.

$$Ca^{2+}(aq) + (NH_4)_2C_2O_4(aq) + H_2O(l) \rightarrow CaC_2O_4.H_2O(s) + 2NH_4^+(aq)$$

Determine the percentage, by mass, of calcium carbonate in the eggshell.

FIGURE 3 Egg shells contain calcium carbonate

★★★ **21 Determine** the:

- **a** empirical formula of a molecule containing 41.40% carbon, 3.40% hydrogen and 55.20% oxygen
- **b** molecular formula of a compound with a molar mass of 136 g mol^{-1} and an empirical formula of C_4H_4O
- **c** empirical formula of acetic acid given that it has a percentage composition of 39.9% C, 6.7% H and 53.4% O
- **d** percentage composition by mass of carbon, nitrogen and oxygen in acetamide (C_2H_5NO).

★★★ **22** Antimony (Sb) is used in the electronics industry to make semiconductor devices, such as infrared detectors. It is obtained by heating stibnite (Sb_2S_3) with scrap iron and drawing off the molten antimony:

$$2Sb_2S_3(s) + 3Fe(s) \rightarrow 2Sb(s) + 3FeS(s)$$

If 300 g of stibnite and 125 g of iron are heated together to give 100 g of Sb, **determine** the:

- **a** limiting reactant
- **b** amount of the excess reactant remaining
- **c** theoretical yield of antimony
- **d** percentage yield of antimony.

FIGURE 4 Infrared dectectors contain antimony

Investigate, evaluate and communicate

★★★ **23 Investigate** Mole Day. When is it and how does it relate to Avogadro's number? Prepare a presentation stating why the school should celebrate Mole Day.

★★★ **24** Caffeine is a chemical found in tea and coffee. The chemical formula for caffeine is $C_8H_{10}N_4O_2$. **Investigate** the caffeine molecule and use the information to write a question that could be used in a chemistry exam. Include the answer to the question and a possible marking scheme.

FIGURE 5 Instant coffee contains caffeine

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CHAPTER 9

Chemical reactions

Chemical reactions involve a change in energy to make new substances. In a chemical reaction, the bonds of reactants are broken, and the atoms rearrange to form new bonds in the products.

The combustion reaction is a chemical reaction that has been used for the benefit of society over the last few hundred years. French chemist Antoine-Laurent de Lavoisier is credited as documenting the first accurate account of a combustion reaction. Lavoisier burned phosphorus and sulfur and noticed that both products had increased in mass. He concluded that both elements must have reacted with a substance in the air.

Electrochemical reactions are also essential chemical reactions. In these chemical reactions, a flow of electrons is produced, which is used to power various processes or appliances. Knowledge of electrochemistry has led to the development of the rechargeable battery, which is used to charge your phone and laptop computer.

Chemical reactions can be manipulated to produce a wide array of useful substances, including plastics, fabrics and medicines.

The ability of chemists to manipulate the energy within chemical reactions has led to the development of such fields as electrochemistry, thermochemistry, equilibrium, reaction rates and industrial chemistry.

OBJECTIVES

- → Recall that chemical reactions and phase changes involve energy changes commonly observable as changes in the temperature of the surroundings and/or the emission of light.
- → Deduce and construct balanced chemical equations when reactants and products are specified and apply state symbols (s), (l), (g) and (aq).
- → Explain how endothermic and exothermic reactions relate to the law of conservation of energy and the breaking and reforming of bonds; understand that heat energy is released or absorbed by the system to or from the surrounds.
- → Understand that heat is a form of energy and that temperature is a measure of the average kinetic energy of the particles.
- → Apply the relationship between temperature and enthalpy changes to identify thermochemical reactions as exothermic or endothermic; deduce from enthalpy level diagrams and thermochemical equations the relative stabilities of reactants and products, and the sign of the enthalpy change (ΔH) for a reaction.

→ Explain, in terms of average bond enthalpies, why reactions are exothermic or endothermic.

→ Construct and use appropriate representations (including chemical symbols and formulas, and chemical and thermochemical equations) to communicate conceptual understanding, solve problems and make predictions.

→ Calculate the heat change for a substance, given the mass, specific heat capacity and temperature change.

→ Use data to calculate the enthalpy change (ΔH) for a reaction.

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PRACTICALS

MANDATORY PRACTICAL	**9.3** Calorimetry: Measuring the enthaly of a reaction
SUGGESTED PRACTICAL	**9.2A** Applying Hess's law
SUGGESTED PRACTICAL	**9.2B** Measuring the molar heat of a chemical reaction

FIGURE 1 The explosive reaction between potassium and water.

9.1 Physical and chemical changes

KEY IDEAS

In this section, you will learn about:

- phase changes involving energy changes
- types of chemical reactions
- balancing chemical equations.

melting
the process of turning a solid into a liquid by applying thermal energy; particles move faster and start to separate

vaporisation
the process of turning a liquid into a gas by applying thermal energy; particles move faster and separate from one another

boiling
vaporisation at a substance's boiling point, which is the temperature at which a liquid becomes a gas

evaporation
the phase change from a liquid to a gas

condensation
the process of turning a gas into a liquid by removing thermal energy; particles move slower and become more attracted to one another

freezing
the process of turning a liquid into a solid by removing thermal energy; particles move slower and become more attracted to one another

Chemical and physical changes are influenced by changes in energy. The energy causes a change within a substance, which results in the formation of a new substance or a phase change. The most common form of energy involved in chemical reactions is thermal energy (heat), which is associated with the movement (kinetic energy) or particles.

Phase changes

Substances change between phases or states upon the addition or removal of thermal energy (heat). Phase changes are physical changes, not chemical changes, because molecules do not break apart to form new molecules. Instead, the molecules separate from one another to change states or phases between solid, liquid and gas.

Changes in thermal energy

Kinetic molecular theory helps to explain phase or state changes when substances are heated or cooled. This theory states that particles are in constant random motion. Changes of state occur with an increase or decrease in temperature. Water is often used to demonstrate state changes because it transforms from solid to liquid to gas over a narrow temperature range, which can be easily observed and measured (Figure 1).

Solid to liquid – melting

Water molecules have relatively strong forces attracting them to one another. In solid water (ice), water molecules form a lattice, which is why they can maintain the solid structure. They can only vibrate within the lattice because they do not have enough energy to overcome the force of attraction between the molecules. As thermal energy is added, in a process called **melting**, the molecules vibrate faster and start to move apart from one another, to become a liquid.

Liquid to gas – vaporisation

In liquid water, the molecules move freely, but still do not have enough energy to overcome the forces that are attracting them to one another. They fill the bottom of the container that they occupy and remain there until more energy is added. As thermal energy is added in a process called **vaporisation**, the molecules move faster and start to overcome the forces attracting the molecules to each other, to become a gas (steam or water vapour). Vaporisation includes both **boiling** and **evaporation**. Boiling occurs at the temperature where thermal energy overcomes the attracting forces between the molecules so that they can form a gas. Evaporation occurs at much lower temperatures and therefore occurs much more slowly. As thermal energy increases, so does the rate of evaporation.

Gas to liquid – condensation

In the gaseous state (steam or water vapour), the movement of the water molecules is greater than the energy that attracts the water molecules to each other. Therefore, they no longer stay together but move about freely in a random motion. When heat is removed, the gaseous water molecules transfer their energy to other gas molecules. Their energy lowers and they begin attract other water molecules again, becoming a liquid. This process, known as **condensation**, occurs in clouds to form rain or precipitation.

Liquid to solid – freezing

As further thermal energy is removed from water molecules, their attraction to one another increases until they align back into the solid ice lattice. This process is known as **freezing**. As discussed earlier, in this state the water molecules cannot move around as vigorously in the lattice because they do not have sufficient energy to break the hydrogen bonds holding them together.

Solid to gas – sublimation

During the process of **sublimation**, molecules change from the solid state to the gaseous state, without going through the liquid phase. Sublimation occurs when snow or ice changes into water vapour without first melting into water. This process depends on the melting point of the solid. Dry ice (solid CO_2) melts at 56.6°C, but only at an atmospheric pressure of 5.1 atm or more. At atmospheric pressure (the pressure that we live under), and temperatures of –78.48°C solid and above, CO_2 is converted to gaseous CO_2, skipping the liquid phase because pressure conditions are not high enough. At room temperature (25°C and 1 atm), dry ice readily converts to gaseous CO_2.

Gas to solid – deposition

Deposition occurs when molecules change straight from a gas to a solid, skipping the liquid phase. It is the opposite of sublimation. On cold mornings, you may notice that frost forms on the windows of a car, when there has been no rain overnight. When the outside temperature falls below the freezing point of water (0°C), water vapour in the air loses thermal energy and converts from a gas to a solid, forming ice on the car (see Chapter 15).

sublimation
the process of turning a solid into a gas by applying thermal energy; particles move faster and separate from one another

deposition
the process of turning a gas into a solid by removing thermal energy; particles move slower and become more attracted to one another

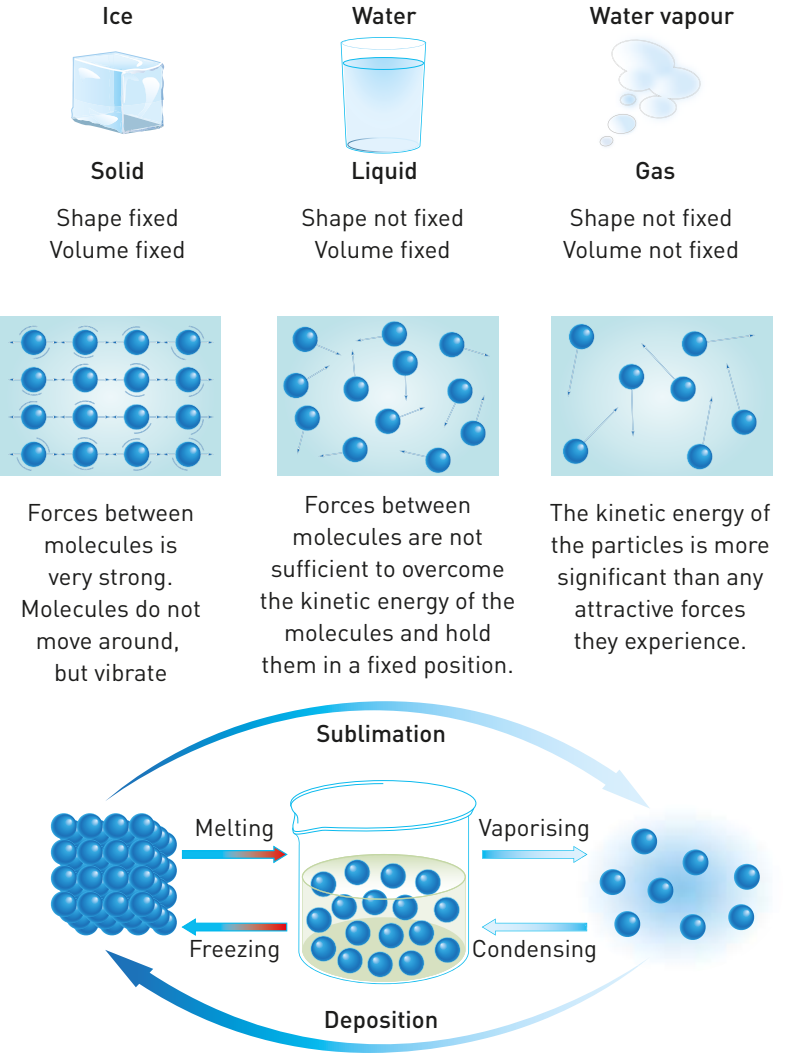


FIGURE 1 The changes of state and movement of water molecules when energy is added or removed.

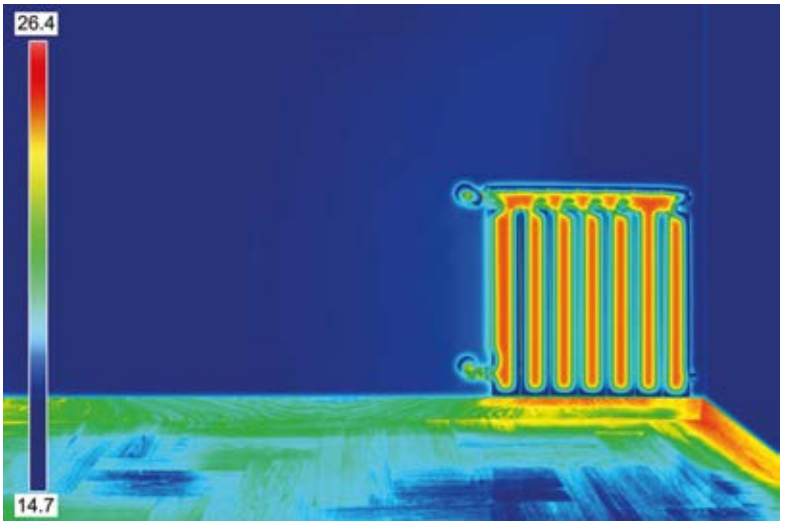


FIGURE 2 Infrared image of the thermal radiation (heat energy) emitted from a radiator heater.

The connection between thermal energy and radiation

Thermal energy is also called thermal radiation. This is electromagnetic radiation that is generated by the kinetic movement of particles and can include visible light and infrared light. For example, when hot metal glows bright yellow-red, it emits visible light. Light bulbs emit infrared light and visible light. Human beings emit infrared radiation, which can only be seen by infrared or thermal imaging. All matter with a temperature greater than absolute zero (0 K) emits electromagnetic radiation.

An example of infrared imaging is shown in Figure 2. When the radiator heater is turned on, it generates thermal energy. This energy is transformed as the air particles within the room gain kinetic energy and their motion increases. This increase in movement causes the particles to spread further away from each other, decreasing their density and causing them to rise, as the cooler, denser particles fall. This is called a convection current.

Study tip

Questions involving balancing chemical equation are typically worth one mark. If the chemical equation is balanced incorrectly or the formulas and states are incorrect, no mark is allocated.

Thermal radiation from a flame is used to heat the contents of a beaker, whereas the Sun's radiation heats Earth. Because heat can be transferred by radiation, and light is a form of radiation heat can be measured on the electromagnetic spectrum. It may be assumed that phase changes involve the emission of light as thermal radiation.

When energy in the form of light or heat is added to an atom, the electrons become excited and jump from their stable, ground orbital shell to a higher energy orbital shell. Electrons are unstable when they have excess energy so they tend to lose the energy by the easiest possible means. The result is that they emit light of characteristic wavelengths and therefore colours to return to their original, stable, ground state.

This concept is explained in detail in Chapter 5.

Chemical changes

Chemical reactions occur when energy breaks apart the bonds within chemical reactants. The atoms then rearrange to form products. The energy that breaks the bonds may be in the form of thermal or light energy. As both are forms of electromagnetic radiation, any form of heat or light may be effective.

As the atoms rearrange, there must be the same number and type of atoms in the products as there are in the reactants. When writing chemical equations, it is essential that they are **balanced** and include the correct formulas and states. This is essential because it will be built on throughout Units 1–4 Chemistry. It also ensures that chemical concepts and principles are correctly communicated.

balanced
having the same numbers of atoms on either side of a reaction arrow

Assigning states and solubility rules

States are used to indicate the nature of reactants or products. They must be assigned to all substances in chemical equations. The four states that are assigned in chemical reactions are solid, liquid, gas and aqueous.

- Solid (s) – all metals, other than mercury (Hg), are solids if they are unreacted. Some non-metals, such as carbon and sulfur, also form solids. In precipitation reactions, a solid is formed. To determine which product has formed the solid, a solubility table can be used. This is a list of ions that indicates whether they form precipitates or remain in solution when combined with other ions. In the table, if a substance is listed as insoluble, it does not dissolve in water and is therefore a solid.

  Some simple rules for solubility of salts in water are:

  1. All nitrate (NO_3^-) salts are soluble.
  2. Most salts of Na^+, K^+ and NH_4^+ are soluble.
  3. Most chloride salts are soluble. Notable exceptions are AgCl and Hg_2Cl_2.
  4. Most sulfate salts are soluble. Notable exceptions are $BaSO_4$, $PbSO_4$ and $CaSO_4$.
  5. Most hydroxides are only slightly soluble. The important soluble hydroxides are NaOH, KOH and $Ca(OH)_2$.
  6. Most sulfide (S^{2-}), carbonate (CO_3^{2-}) and phosphate (PO_4^{3-}) salts are only slightly soluble.

Study tip

To determine if an ionic substance is solid or aqueous, use the 'simple rules for solubility of salts in water' table in the data book.

- Liquid (l) – water is almost always a liquid, unless other stated. For example, water vapour is gaseous water and ice is solid water. Ionic substances can also be liquids, but only when they are melted, or **molten**. For example, sodium chloride melts at approximately 800°C, at which point it becomes a liquid.

molten
a melted substance

- Gas (g) – common gases at room temperature can be found in Table 1. Water can also be a gas at temperature over 100°C and is referred to a water vapour.
- Aqueous (aq) – Solutions formed from water-soluble substances and water are **aqueous** solutions. If a substance is dissolved in water, it is assigned the (aq) state.

TABLE 1 Formulas and names of some common gases at room temperature

Formula	Name
H_2	Hydrogen
O_2	Oxygen
N_2	Nitrogen
Cl_2	Chlorine
CO_2	Carbon dioxide
NO_2	Nitrogen dioxide
NO	Nitrogen monoxide
SO_2	Sulfur dioxide
SO_3	Sulfur trioxide
NH_3	Ammonia

aqueous
a substance that is dissolved in water or the state given to a substance that is dissolved in water

Law of conservation of mass and balancing chemical equations

The law of conservation of mass states that matter cannot be created or destroyed. According to this law, all chemical equations must be balanced. A balanced chemical equation has the same number of each type of atom on either side of the reaction arrow.

An easier method for balancing chemical equations involves **ions**. If a polyatomic ion (an ion that is made up of more than one atom, such as hydroxide (OH^-) or nitrate (NO_3^-)) remains unchanged in a chemical reaction, it can be treated as one entity, rather than breaking it up into individual atoms.

ion
an atom that has gained or lost electrons to become more stable

Single displacement

single displacement a chemical reaction in which a more reactive metal ion replaces a less reactive metal ion in a molecule

Single displacement occurs when a more reactive element replaces a less reactive element in a molecule. The abbreviated reactivity series for metals and non-metals are outlined in Tables 2 and 3.

Consider the reaction of aqueous silver(I) nitrate and copper metal. Copper is more reactive than silver (Table 2), so it displaces Ag^+ ions from solution to form aqueous copper(II) nitrate and silver metal, as the following unbalanced equation shows.

$$Ag(NO_3)(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$$

TABLE 2 Abbreviated reactivity series for metals

K	Most reactive
Na	↓
Li	
Ba	
Sr	
Ca	
Mg	
Al	
(C)	
Mn	
Zn	
Cr	
Fe	
Cd	
Co	
Ni	
Sn	
Pb	
(H)	
Cu	
Ag	Least reactive
Au	

TABLE 3 Abbreviated reactivity series for non-metals

	Most reactive
F	↓
Cl	
Br	
O	
I	
S	Least reactive

WORKED EXAMPLE 9.1A

Balance the following chemical equation between silver nitrate and copper metal.

$$AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$$

SOLUTION

1 The number of atoms are not equal on each side. We can use coloured balls to represent each atom.

2 The nitrate ion (NO_3^-) remains intact, so it is considered as one entity (the grey ball).

$2AgNO_3$	Cu	→	$Cu(NO_3)_2$	Ag
		→		

3 Therefore, the balanced chemical equation is:

$$2AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$

Consider the reaction of sodium bromide with chlorine gas. Chlorine is more reactive than bromine so it displaces bromine from the sodium ion to form sodium chloride and bromine gas, as the following unbalanced equation shows.

$$NaBr(s) + Cl_2(g) \rightarrow NaCl(aq) + Br_2(g)$$

WORKED EXAMPLE 9.1B

Balance the following chemical equation between sodium bromide and chlorine gas.

$$NaBr(s) + Cl_2(g) \rightarrow NaCl(aq) + Br_2(g)$$

SOLUTION

2 NaBr	Cl_2	→	2 NaCl	Br_2
		→		

Therefore, the balanced chemical equation is:

$$2NaBr(aq) + Cl_2(g) \rightarrow 2NaCl(aq) + Br_2(g)$$

Double displacement

A **double displacement** reaction occurs when the **anion** (negatively charged ion) and **cation** (positively charged ion) swap partners. In the following example, lead and sodium swap partners to form new products.

lead(II) nitrate + sodium hydroxide → lead(II) hydroxide + sodium nitrate

According to the solubility table, all nitrates are soluble in water, and are therefore aqueous. Sodium hydroxide is also soluble in water and therefore aqueous, but lead(II) hydroxide is insoluble and therefore a solid. This is an example of a precipitation reaction because a solid forms, as the following unbalanced equation shows.

$$Pb(NO_3)_2(aq) + NaOH(aq) \rightarrow Pb(OH)_2(s) + NaNO_3(aq)$$

double displacement
a chemical reaction in which both the cation and anion of one reactant swaps with the cation and anion of the second reactant, forming new products

anion
a non-metal atom which has gained one or more electrons and has a negative charge

cation
a metal atom that has lost one or more electrons and is positively charged

WORKED EXAMPLE 9.1C

Balance the following chemical equation between lead nitrate and sodium hydroxide.

$$Pb(NO_3)_2(aq) + NaOH(aq) \rightarrow Pb(OH)_2(s) + NaNO_3(aq)$$

SOLUTION

1 As the nitrate ion (NO_3^-) (grey ball) and the hydroxide ion (OH^-) (yellow ball) remain intact, they are each considered a single entity.

$Pb(NO_3)_2$	2 NaOH	→	$Pb(OH)_2$	2 $NaNO_3$
		→		

2 Therefore, the balanced chemical equation is:

$$Pb(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Pb(OH)_2(s) + 2NaNO_3(aq)$$

Acid–base

The reaction of an **acid** with a **base** forms water and a salt. The formulas of some common acids and bases can be found in Table 3. Sulfuric acid reacts with potassium hydroxide (a base) to form a salt (potassium sulfate) and water, as the following unbalanced equation shows.

$$H_2SO_4(aq) + KOH(aq) \rightarrow K_2SO_4(aq) + H_2O(l)$$

acid
a chemical compound that donates a hydrogen ion

base
a chemical compound that accepts a hydrogen ion

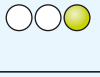
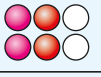
WORKED EXAMPLE 9.1D

Balance the following chemical equation between sulfuric acid and potassium hydroxide.

$$H_2SO_4(aq) + KOH(aq) \rightarrow K_2SO_4(aq) + H_2O(l)$$

SOLUTION

1 The sulfate ion (SO_4^{2-}) (green ball) remains intact, so it is considered as a single entity. The hydroxide ion (OH^-) bonds with the hydrogen ion to form a water molecule and so it must be considered as individual atoms.

H_2SO_4	2 KOH	→	K_2SO_4	2 H_2O
		→		

2 Therefore, the balanced chemical equation is:

$$H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$$

TABLE 4 Some common acids and bases

Acids		Bases	
Formula	Name	Formula	Name
CH_3COOH	Acetic acid (ethanoic acid)	$NaHCO_3$	Sodium hydrogen carbonate
HCl	Hydrochloric acid	KOH	Potassium hydroxide
HNO_3	Nitric acid	NH_3	Ammonia
H_2CO_3	Carbonic acid	MgO	Magnesium oxide
H_2SO_4	Sulfuric acid	$NaOH$	Sodium hydroxide
H_2SO_3	Sulfurous acid	Na_2O	Sodium oxide
H_3PO_4	Phosphoric acid	$Ca(OH)_2$	Calcium hydroxide

Combustion

combustion
a chemical reaction with oxygen to form a metal oxide, covalent compound or carbon dioxide and water

A **combustion** reaction is a reaction between oxygen and an element or a molecule. There are two primary types of combustion reactions – those involving the combustion of metals and non-metals and those involving the combustion of hydrocarbons.

Combustion of metals and non-metals

A metal combusts (reacts with oxygen) to form a metal oxide:

$$\text{metal} + \text{oxygen} \rightarrow \text{metal oxide}$$

For example: $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

A non-metal combusts to form a covalent compound:

$$\text{non-metal} + \text{oxygen} \rightarrow \text{covalent compound}$$

For example: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

Combustion of hydrocarbons

When a hydrocarbon is combusted with sufficient energy, it usually forms carbon dioxide and water. For example, the combustion of the fuel methane produces carbon dioxide and water:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

This chemical equation is a simple one to balance as long as carbon and hydrogen are balanced first. Other combustion reactions of hydrocarbons can be balanced by applying the following rules.

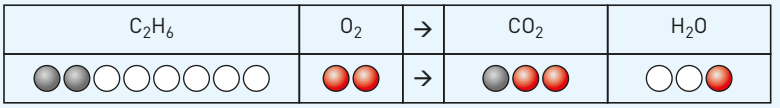
1. Balance oxygen last. Oxygen is in three of the four molecules, so balancing carbon and hydrogen first makes it easier to balance oxygen as the last step.
2. After carbon and hydrogen are balanced, there is sometimes an odd number of oxygen atoms on the product side. If this occurs, double everything. Consider the combustion of ethane (Worked example 9.1E).

WORKED EXAMPLE 9.1E

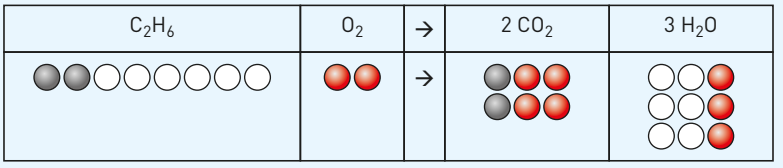
Balance the following chemical equation for the combustion of ethane gas.

$$C_2H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$$

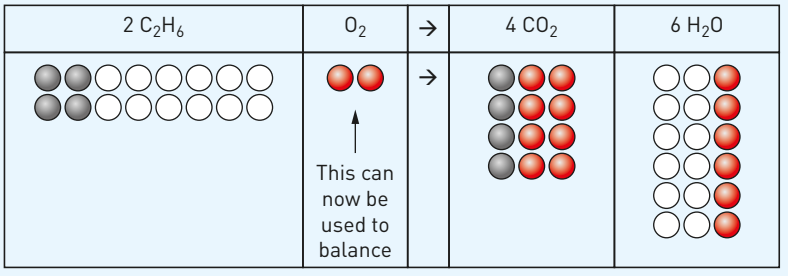
SOLUTION



1 Balance carbon and hydrogen.



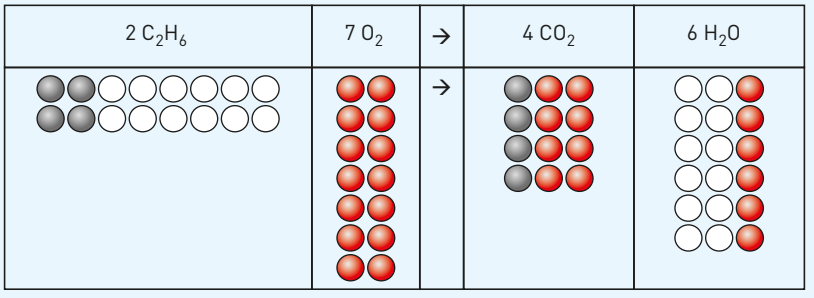
2 When carbon and hydrogen are balanced, there is an odd number of oxygen atoms on the product side. Therefore, double everything, except the O_2 on the reactant side, which will be balanced last. (The formula of oxygen is O_2 and any coefficient of the reactant oxygen always results in an even number. Therefore, there must always be an even number of oxygen atoms on the reactant side. For the combustion of ethane, there are two oxygen atoms on the reactant side and seven on the product side, which cannot be balanced because it is uneven. By doubling everything, an even number is generated on the product side and the oxygen can be balanced.)



3 Balance oxygen.

4 Therefore, the balanced chemical equation is:

$$\mathbf{2}C_2H_6(g) + \mathbf{7}O_2(g) \rightarrow \mathbf{4}CO_2(g) + \mathbf{6}H_2O(l)$$



Exceptions to the formation of carbon dioxide in the combustion of hydrocarbons

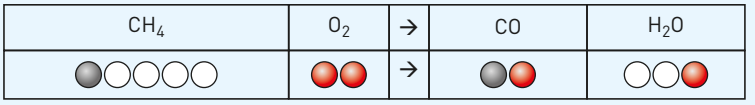
There are two types of combustion reactions where carbon dioxide does not form. These occur when the amount of oxygen is low or limiting. A reduced amount of oxygen means there is less oxygen bonding to the carbon atom and the products are either carbon monoxide (CO) and water (H_2O) or carbon (C) and water (H_2O).

WORKED EXAMPLE 9.1F

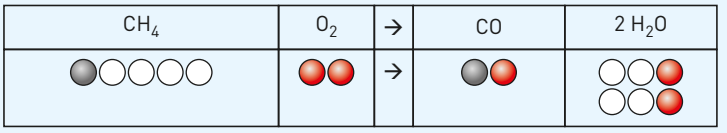
Balance the following chemical equation for the combustion of methane gas in limited oxygen.

$$CH_4(g) + O_2(g) \rightarrow CO(g) + H_2O(l)$$

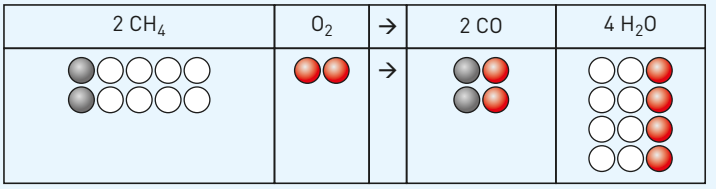
SOLUTION



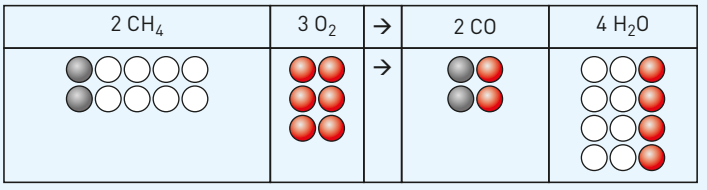
1 Balance carbon and hydrogen.



2 There are an odd number of oxygen atoms on the product side, so double everything except O_2.



3 Balance oxygen.



4 Therefore, the balanced chemical equation is:

$$\mathbf{2}CH_4(g) + \mathbf{3}O_2(g) \rightarrow \mathbf{2}CO(g) + \mathbf{4}H_2O(l)$$

Combination

combination
a chemical reaction in which two atoms or elements react to form a molecule; also called a synthesis or addition reaction

A **combination** reaction, sometimes called synthesis or addition, involves reacting two substances to form one product; for example, the reaction between sodium and chlorine to form sodium chloride.

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl_2(s)$$

The combustion of metals and non-metals are also examples of combination reactions, (Worked examples 9.1A–F).

Decomposition

A **decomposition** reaction occurs when a molecule breaks down into smaller molecules or even atoms. This happens when heat or electricity is applied. For example, water decomposes when electricity is applied using a power pack, wires and platinum electrodes, according to the chemical equation:

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$

Calcium carbonate can also be decomposed by thermal decomposition or heat according to the chemical equation:

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

Dehydration is another example of decomposition. A hydrated salt is a crystalline salt that contains a certain number of water molecules loosely held in the crystal lattice. The number of water molecules is signified by putting '·xH$_2$O' after the salt. For example, $CuSO_4 \cdot 5H_2O(s)$ contains five water molecules for every $CuSO_4$.

$$CuSO_4 \cdot 5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(l)$$

decomposition
a chemical reaction in which a molecule breaks down into smaller molecules on the application of heat or electricity

Simple redox reactions

Simple **redox** reactions involve the transfer of electrons from one reactant to another. These types of reactions encompass several of the examples shown previously. The reaction of zinc metal with hydrochloric acid is shown below. In this single displacement reaction, zinc is more reactive than hydrogen and will therefore react, forming zinc chloride and hydrogen gas. This is a redox reaction because the **oxidation state** (valency or assigned charge) of the two reactants has changed (see Table 5).

$$Zn(s) + HCl \rightarrow ZnCl_2(aq) + H_2(g)$$

redox
a chemical reaction involving the transfer of electrons from one reactant to another

oxidation state
the charge assigned to substances in a chemical reaction

TABLE 5 Oxidation states of reactants and products

Zn	H^+	Cl^-	→	Zn^{2+}	Cl^-	H_2
0	+1	−1	→	+2	−1	0

The oxidation state of zinc increases from 0 to +2, the oxidation state of hydrogen decreases from +1 to 0 and the chloride ion is a **spectator ion** in the reaction because its oxidation state is unchanged. Note that the oxidation state of elements is always 0.

spectator ion
an ion that does not participate in a chemical reaction

WORKED EXAMPLE 9.1G

Balance the following chemical equation between zinc metal and hydrochloric acid.

$$Zn(s) + HCl \rightarrow ZnCl_2(aq) + H_2(g)$$

SOLUTION

Zn	2 HCl	→	$ZnCl_2$	H_2
●	○● ○●	→	●●●	○○

Start with the metal, Zn has 1 on each side of the equation. Move to the non-metal, Cl has 1 on the reactant side and 2 on the product side. So there needs to be 2HCl molecules on the reactant side to balance Cl. Now count the hydrogens. There are 2 on both sides. The equation is balanced.

Therefore, the balanced chemical equation is:

$$Zn(s) + \mathbf{2}HCl \rightarrow ZnCl_2(aq) + H_2(g)$$

Study tip

When balancing redox reactions, start by balancing the metals, then non-metals, then oxygen and lastly hydrogen.

CHALLENGE 9.1

Writing balanced equations

Write a fully balanced chemical equation, including states, for the reaction between sodium sulfate and lead (II) nitrate.

CASE STUDY 9.1

The death of the internal combustion engine

The internal combustion engine was invented in the late 1700s. The impact of the combustion engine on society was immense: every aspect of society has benefited. From the development of the efficient internal combustion engine: from transport to agriculture, industry and power generation.

Study tip

If, when balancing an equation you end up with all even numbers, simplify the equation by halving everything.

Since its first development, the engine has been re-designed to be smaller, lighter, safe to use, and start very quickly. When fuel combusts in the engine, the gaseous products cause an increase in pressure, which in turn causes a turbine blade, piston or rotor to move. Combustion engines have become highly efficient in their design. Although they lose some energy in the form of heat and sound, they lose significantly less than the first models developed.

However, internal combustion engines have negative environmental impacts. They burn hydrocarbons and emit carbon dioxide and water vapour, both of which are greenhouse gases and contribute to climate change.

Newer technologies have resulted in the development of electric cars, which are either battery charged or powered by hydrogen fuel cells. Scientists working on each type are competing to refine the technology and create the best possible, cost-effective, fuel-efficient and environmentally friendly alternative to the internal combustion engine.

FIGURE 3 A hydrogen fuel cell is an alternative to batteries or traditional combustion engines to power cars. This Hyundai Tuscan Hydrogen refuels with hydrogen gas in minutes, and its only emission is pure water vapour.

Battery-powered cars are currently leading this race because they require only a recharge station, which can be built into homes or petrol stations. However, in Australia, the electricity that powers the cars is mostly produced by burning coal, which releases greenhouse gases. So although electric cars are environmentally friendly to run, the electricity they run on may be produced in non-environmentally friendly ways.

Hydrogen fuel cells react within the car's engine to form environmentally friendly products and have the potential to be topped up at a fuel station. However, current infrastructure does not support hydrogen as a commercially available fuel and there are significant risks involved with storing such a highly flammable and explosive gas.

CHECK YOUR LEARNING 9.1

Describe and explain

1 **Explain** why is it necessary to balance chemical equations.

2 **Describe** how the state (aq) differs from the state (l).

Apply, analyse and interpret

3 Correctly **structure** the following reactions by balancing them.

- **a** $C_3H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$
- **b** $Al(s) + O_2(g) \rightarrow Al_2O_3(s)$
- **c** $N(g) + O_2(g) \rightarrow NO(g)$
- **d** $C_4H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$
- **e** $C_6H_{12}O_6(aq) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
- **f** $KOH(aq) + H_3PO_4(aq) \rightarrow K_2PO_4(aq) + H_2O(l)$
- **g** $Zn(s) + Pb(NO_3)_4(aq) \rightarrow Zn(NO_3)_2(aq) + Pb(s)$
- **h** $Na_2SO_4(aq) + BaCl_2(aq) \rightarrow NaCl(aq) + BaSO_4(s)$
- **i** $NaOH(aq) + Cu(NO_3)_2(aq) \rightarrow NaNO_3(aq) + Cu(OH)_2(s)$
- **j** $H_2SO_3(aq) \rightarrow H_2O(l) + SO_2(g)$

4 **Determine** the states of the following substances at room temperature and normal atmospheric pressure.

- **a** silver bromide
- **b** potassium nitrate
- **c** barium sulfate
- **d** silver chloride
- **e** ammonium chloride
- **f** calcium hydroxide
- **g** carbon monoxide (CO)
- **h** calcium carbonate ($CaCO_3$)
- **i** magnesium metal
- **j** molten lithium fluoride
- **k** lithium fluoride
- **l** nitrogen

5 Correctly **structure** the following reaction by balancing the decomposition of nickel (II) oxide.

$$Ni_2O_3(s) \rightarrow Ni(s) + O_2(g)$$

6 Correctly **structure** the following reaction by balancing the combustion of methane gas in very limited oxygen.

$$CH_4(g) + O_2(g) \rightarrow C(s) + H_2O(g)$$

Investigate, evaluate and communicate

7 The simple rules for solubility of salts in water state that some ionic compounds are 'slightly soluble'. **Discuss** this statement. In your response, identify what state should be assigned to these compounds.

8 **Discuss** the terms 'aqueous', 'liquid' and 'molten' with respect to the substances involved and particle movement.

Check your obook assess for these additional resources and more:

- » Student book questions
  Check your learning 9.1
- » Challenge
  9.1 Writing balanced equations
- » Video
  Balancing a chemical equation
- » Increase your knowledge
  Additional worked examples

9.2 Chemical energy and thermochemistry

KEY IDEAS

In this section, you will learn about:

- the law of conservation of energy
- endothermic and exothermic reactions
- bond energy and enthalpy
- enthalpy level diagrams
- thermochemical equations.

law of conservation of energy
a law that states energy cannot be created or destroyed

chemical energy
a form of potential (stored) energy generated by the attraction and repulsion of protons, neutrons and electrons, as well as the electron movement in a chemical substance

exothermic
describes a reaction in which the reactants have more energy than the products; excess energy is released to the environment as heat

endothermic
describes a reaction in which the reactants have less energy than products; the required energy is absorbed from the environment as heat

enthalpy
chemical substances, referred to as its chemical energy or heat content

bond enthalpy
the amount of energy stored within a single bond

The **law of conservation of energy** states that **energy cannot be created or destroyed**. Energy is instead transferred from one place to another or transformed from one type of energy into another. For example, when a light is switched on, electrical energy is transformed into heat (thermal) and light energy. Chemical energy is a type of potential, or stored, energy.

Chemical energy

Chemical energy is the energy stored within the bonds of substances. Different molecules are made up of different atoms bonded in different ways, so the reactants and products in chemical equations have different amounts of potential energy. When a chemical reaction takes place and atoms rearrange to form products, the energy level within the reaction system must change. One of two situations will occur.

1. If the reactants have more energy than the products, excess energy is released from the system (the reaction) into the environment. This is called an **exothermic** reaction.
2. If the products have more energy than the reactants, more energy is needed to form them. The extra energy is absorbed from the environment into the system. This is called an **endothermic** reaction.

Enthalpy

The amount of chemical energy stored within the bonds of a substance is called its **enthalpy**, or heat content, and is represented by the symbol *H*.

Each bond in a molecule has a different amount of energy or enthalpy. **Bond enthalpy** is the amount of energy required to break 1 mol of a bond. Therefore, the enthalpy for a whole molecule is the sum of the bond enthalpies within its structure. Table 1 lists some bond enthalpies.

Calculating enthalpy from bond enthalpy

To calculate enthalpy, you must know the structure of a molecule. Then add the bond enthalpies for each bond.

It is important to note that the values in Table 1 are the average bond enthalpies for these bonds. The values that are calculated experimentally are often very different from these theoretical values because theoretical values:

- are average values and do not take into consideration that the same bond can have different energies depending on their reaction environment

- do not take into consideration any **intermolecular** or **intramolecular forces** (Chapter 12) that may exist in the reaction system
- were measured using gaseous reactants and products and therefore only really apply to gaseous systems.

Therefore it is difficult to measure the energy of a reaction, but easier to measure the change in enthalpy.

intermolecular force
a force between molecules

intramolecular force
a force between the atoms in a molecule

TABLE 1 Some average bond enthalpies (kJ mol^{-1})

Single bonds							
C–H	414	N–H	391	O–H	463	F–F	159
C–C	346	N–N	158	O–O	144		
C–N	286	N–F	278	O–F	191	Cl–F	255
C–O	358	N–O	214	O–Cl	206	Cl–Cl	242
C–F	492	N–Cl	192	O–I	201		
C–Cl	324	N–Br	243			Br–F	249
C–Br	285			S–H	339	Br–Cl	219
C–I	228	H–H	436	S–F	327	Br–Br	193
C–S	289	H–F	567	S–Cl	253		
		H–Cl	431	S–Br	218	I–Cl	211
Si–H	323	H–Br	366	S–S	266	I–Br	178
Si–Si	226	H–I	298			I–I	151
Si–C	301						
Si–O	368						
Multiple bonds							
C=C	614	N=N	470	O=O	498		
C≡C	839	N≡N	945				
C=N	615			S=O	523		
C≡N	890			S=S	418		
C=O	804						
C≡O	1072						

WORKED EXAMPLE 9.2A

Determine the amount of energy required to break 1 mol of propane (Figure 1) into its constituent atoms.

SOLUTION

Determine the enthalpy of propane from its bond enthalpies in Table 1.

Propane contains 2 C–C bonds and 8 C–H bonds.

Enthalpy of C–C bond = 346 × 2
= 692 kJ mol^{-1}

Enthalpy of C–H bond = 414 × 8
= 3308 kJ mol^{-1}

∴ Enthalpy of propane = 692 + 3308
= 4000 kJ mol^{-1}

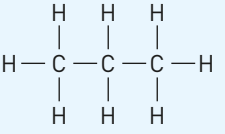


FIGURE 1 Structural formula of propane.

Study tip

Use bond enthalpy values from the data book.

Study tip

Remember to count each individual bond only once. If it helps, cross them out as you go.

Thermochemistry

Thermochemistry is the study of heat that has been absorbed (endothermic) or released (exothermic) from a chemical reaction. This can be represented on an enthalpy level diagram.

Enthalpy level diagrams

An enthalpy level diagram is used to represent the difference in bond energies between reactants and products. Enthalpy (H) is represented on the y-axis, and the progress of the reaction, or time, is represented on the x-axis (Figure 2).

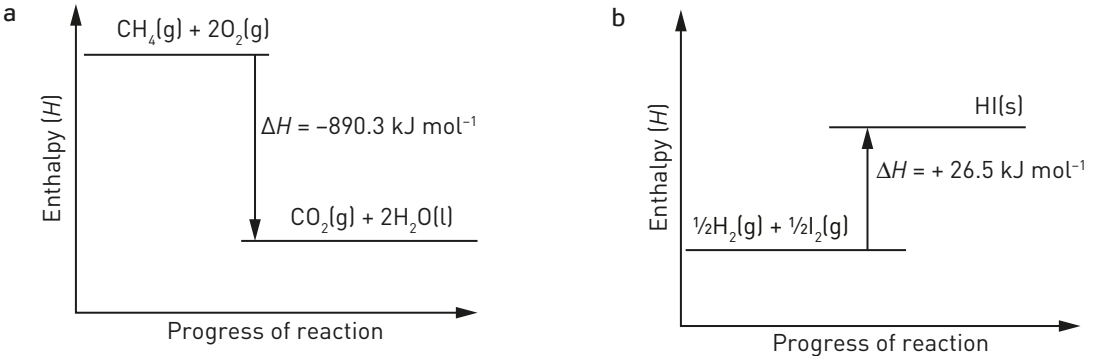


FIGURE 2 Enthalpy level diagrams of an (a) exothermic and (b) endothermic reaction.

Several key pieces of information may be generated from an enthalpy level diagram.

stability
related to bond enthalpy

The **stability** of a particular bond can be related to the bond enthalpy. Higher bond enthalpy values indicate more energy is required to break the bond and therefore that bond is more stable.

The difference in enthalpy values between products and reactants represents the amount of energy that is absorbed or released by the system. This difference indicates whether the reaction is endothermic or exothermic.

In an exothermic reaction (Figure 2a), the reactants have a higher enthalpy than the products. Therefore, when products form, there is residual energy that transforms into heat and moves from the system to the environment, leaving the reaction vessel feeling hot. This agrees with the law of conservation of mass because energy released is the difference between the enthalpies of reactants and products.

In an endothermic reaction (Figure 2b), the reactants have a lower enthalpy than the products. For products to form, energy, as heat, must be absorbed from the environment, leaving the reaction vessel feeling cold. This agrees with the law of conservation of mass because energy absorbed is the difference between the enthalpies of reactants and products.

Determining ΔH

ΔH
the difference in enthalpy between reactants and products in a chemical equation

The difference in enthalpy values between reactants and products is represented as **ΔH**. It is calculated as:

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

As energy must be added to reactants to break their bond and rearrange their atoms to form products, the following formula may also be used:

ΔH = energy required to break bonds – energy released when bonds are formed

$$\Delta H = \Sigma H_{\text{bonds broken}} - \Sigma H_{\text{bonds formed}}$$

Any energy put into the system to break reactant bonds may be released when product bonds form and so this energy cancels out and only the difference in enthalpy between reactants and products determine the size and sign of ΔH.

A **positive** $\Delta \boldsymbol{H}$ value represents an **endothermic** reaction, whereas a **negative** $\Delta \boldsymbol{H}$ value represents an **exothermic** reaction.

WORKED EXAMPLE 9.2B

Calculate the ΔH value for the following reactions and determine whether they are endothermic or exothermic.

a Methane (CH_4) and hydrochloric acid (HCl) form chloromethane (CH_3Cl) and hydrogen gas (H_2); the enthalpy values of reactants is 2083 kJ mol^{-1} and the enthalpy values of products is 2003 kJ mol^{-1}

b Water (H_2O) decomposes to form hydrogen (H_2) and oxygen (O_2); the enthalpy values of reactants is 926 kJ mol^{-1} and the enthalpy values of products is 931 kJ mol^{-1}

SOLUTION

a $\Delta H = H_{products} - H_{reactants} = 2003 - 2083 = -80$ kJ mol^{-1}, which is exothermic

b $\Delta H = H_{products} - H_{reactants} = 931 - 926 = +5$ kJ mol^{-1}, which is endothermic

Determining ΔH from bond enthalpies

A value for ΔH can be calculated by determining the enthalpy of reactants and products from their bond enthalpies in Table 1.

Therefore, an enthalpy level diagram can allow you to conclude whether a reaction is endothermic or exothermic.

WORKED EXAMPLE 9.2C

Ethyne (H–C≡C–H) reacts with hydrogen (H–H) to form ethane (CH_3–CH_3).

Calculate the enthalpies of each molecule and use these to find ΔH of the reaction and conclude whether it is endothermic or exothermic.

SOLUTION

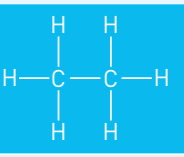
1 The structures of each molecule are known.

2 Balance the chemical equation:

$$C_2H_2(g) + 2H_2(g) \rightarrow C_2H_6(g)$$

3 Use Table 1 to determine bond enthalpies and add them together for each molecule.

TABLE 2 Calculating bond enthalpies

H–C≡C–H	2 H–H	→	
C≡C: 839 kJ mol^{-1} C–H: 2 × 414 kJ mol^{-1}	H–H: 2 × 436 kJ mol^{-1}	→	C–C: 346 kJ mol^{-1} C–H: 6 × 414 kJ mol^{-1}
1667 kJ mol^{-1}	872 kJ mol^{-1}	→	2830 kJ mol^{-1}

4 Calculate ΔH value, using the following formula.

$$\Delta H = \Sigma H_{\text{bonds broken}} - \Sigma H_{\text{bonds formed}}$$
$$= (1667 + 872) - 2830$$
$$= -291 \text{ kJ mol}^{-1}$$

5 Determine whether the reaction is exothermic or endothermic.

ΔH is negative so the reaction is exothermic.

Thermochemical equations

thermochemical equation
a balanced chemical equation, including states, with a ΔH value that is positive or negative

Thermochemical equations are balanced chemical equations that always include a ΔH value. The ΔH value represents the energy difference in kilojoules (kJ) between the reactants and products when 1 mol of the reactant is used.

Writing thermochemical equations

1 Write a balanced chemical equation, including states. These should be given in a question, but they may also be given in a data table. Table 3 lists the heats of combustion and states of some common fuels. The state must be added to the equation. Also, because these are combustion reactions, water will typically vaporise and so should be represented as a gas.

Study tip

If ΔH is written separately, it is negative as exothermic reactions lose energy to their environment.

2 Write the ΔH value next to the equation. This must be negative for an exothermic reaction, which all combustion reactions are. If the coefficient of the fuel in the balanced equation is greater than 1, ΔH must be multiplied by the coefficient of the fuel.

For example, for the combustion of ethanol:

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g) \quad \Delta H = -1360 \text{ kJ mol}^{-1}$$

TABLE 3 Heats of combustion of some common fuels

Fuel	Formula	State	Heat of combustion (kJ g^{-1})	Molar heat of combustion (kJ mol^{-1})
Hydrogen	H_2	Gas	141	282
Methane	CH_4	Gas	55.6	890
Ethane	C_2H_6	Gas	51.9	1560
Propane	C_3H_8	Gas	50.5	2220
Butane	C_4H_{10}	Gas	49.7	2880
Octane	C_8H_{18}	Liquid	47.9	5460
Ethyne (acetylene)	C_2H_2	Gas	49.9	1300
Methanol	CH_3OH	Liquid	22.7	726
Ethanol	C_2H_5OH	Liquid	29.6	1360

Doubling and reversing thermochemical equations

When the coefficients of a chemical equation are doubled, the ΔH value is also doubled because it is a measurement of the energy per mole of reactant.

When a chemical equation is reversed, the sign of the ΔH value is reversed because the reactants and products in the enthalpy equation switch sides.

WORKED EXAMPLE 9.2D

Write the thermochemical equation for the combustion of ethanol. (Refer to Table 3 for ΔH.)

SOLUTION

1 Write a balanced chemical equation, including states. (Remember: combustion is a reaction with oxygen to form carbon dioxide and water.)

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$

2 The balanced chemical equation requires 1 mol of ethanol, therefore ΔH. is as stated in Table 3. ΔH. is negative because this is a combustion reaction.

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g) \qquad \Delta H = -1360 \text{ kJ mol}^{-1}$$

WORKED EXAMPLE 9.2E

Write the thermochemical equation for the reverse reaction of – the combustion of butane. (Refer to Table 3 for ΔH values.)

SOLUTION

1 Write a balanced chemical equation for the combustion of butane, including states. (Remember: combustion is a reaction with oxygen to form carbon dioxide and water.)

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$$

2 The balanced chemical equation requires 2 mol of butane; therefore, multiply ΔH by 2. ΔH must be negative because this is a combustion reaction.

$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$ $\qquad \Delta H = -2880 \times 2 = -5760 \text{ kJ mol}^{-1}$

Alternatively, divide all of the coefficients by 2 and leave ΔH as its original value.

$C_4H_{10}(g) + \frac{13}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)$ $\qquad \Delta H = -2880 \text{ kJ mol}^{-1}$

Note: this is the only time where a fraction in a chemical equation is acceptable because it represents the enthalpy value per mole of fuel burnt.

3 Reverse the chemical equation and change the negative sign on ΔH

$8CO_2(g) + 10H_2O(g) \rightarrow 2C_4H_{10}(g) + 13O_2(g)$ $\qquad \Delta H = +5760 \text{ kJ mol}^{-1}$

WORKED EXAMPLE 9.2F

Write the thermochemical equation for the combustion of butane. (Refer to Table 3 for ΔH.)

SOLUTION

1 Write a balanced chemical equation, including states. (Remember: combustion is a reaction with oxygen to form carbon dioxide and water.)

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$$

2 The balanced chemical equation requires 2 mol of butane; therefore multiply ΔH. by 2. ΔH. is negative because this is a combustion reaction.

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g) \qquad \Delta H = -2880 \times 2 = -5760 \text{ kJ mol}^{-1}$$

Alternatively, divide all of the coefficients by 2 and leave ΔH as its original value.

$$C_4H_{10}(g) + O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g) \qquad \Delta H = -2880 \text{ kJ mol}^{-1}$$

Note: this is the only time where a fraction in a chemical equation is acceptable because it represents the enthalpy value per mole of fuel burnt.

Hess's law

Hess's law
a law that states the enthalpy change for the overall reaction is equal to the sum of the enthalpy changes for the individual steps

Hess's law states that the enthalpy change for the overall reaction is equal to the sum of the individual steps. In the reaction between carbon and oxygen to form carbon dioxide, there is an intermediate step:

Step 1 $C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g)$ $\Delta H = -110.5$ kJ mol^{-1}

Step 2 $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$ $\Delta H = -283$ kJ mol^{-1}

If the overall chemical equation is: $C(s) + O_2(g) \rightarrow CO_2(g)$

$\Delta H = -110.5$ kJ mol^{-1} + -283 kJ mol^{-1}

$= -393.5$ kJ mol^{-1}

WORKED EXAMPLE 9.2G

Calculate ΔH for the overall equation, $4CuO(s) \rightarrow 2Cu_2O(s) + O_2(g)$, from the enthalpy values of the intermediary steps:

Step 1: $2Cu(s) + O_2(g) \rightarrow 2CuO(s)$ $\Delta H = -312$ kJ mol^{-1}

Step 2: $2Cu(s) + \frac{1}{2}O_2(g) \rightarrow Cu_2O(s)$ $\Delta H = -170$ kJ mol^{-1}

SOLUTION

1 Reverse the first equation so that CuO becomes a reactant as per the overall equation. When you reverse the reaction, you reverse the ΔH value

$2CuO(s) \rightarrow 2Cu(s) + O_2(g)$ $\Delta H = +312$ kJ mol^{-1}

2 Because 4 CuO units are required, the coefficients of the first equation must be doubled. When coefficients are doubled, so is the ΔH value.

$4CuO(s) \rightarrow 4Cu(s) + 2O_2(g)$ $\Delta H = +312 \times 2 = +624$ kJ mol^{-1}

3 Because 2 Cu_2O units are required as products, the second equation must be doubled. When coefficients are doubled, so is the ΔH value.

$4Cu(s) + O_2(g) \rightarrow 2Cu_2O(s)$ $\Delta H = -170 \times 2 = -340$ kJ mol^{-1}

4 Cancel out reactants and products, e.g. if 4 Cu(s) atoms form 4 Cu(s) molecules, they can be cancelled out.

1: $4CuO(s) \rightarrow \cancel{4Cu}(s) + 2O_2(g)$ $\Delta H = +312 \times 2 = +624$ kJ mol^{-1}

2: $\cancel{4Cu}(s) + O_2(g) \rightarrow 2Cu_2O(s)$ $\Delta H = -170 \times 2 = -340$ kJ mol^{-1}

Note: because there is 1 oxygen on the reactant side, and 2 on the product side, the reactant oxygen is cancelled out and 1 product oxygen molecule remains

$4CuO(s) \rightarrow 2Cu_2O(s) + O_2(g)$ $\Delta H = +624 + (-340) = +284$ kJ mol^{-1}

Study tip

Remember to multiply the ΔH value by the coefficient of the fuel in a combustion equation. When reversing the equation, *only* reverse the sign.

CHALLENGE 9.2

Manipulating enthalpy values

The combustion of iron has a ΔH value of 3926 kJ for every 1 mol of iron.

Write the combustion equation for the reverse reaction, i.e. the decomposition of iron(III) oxide.

CASE STUDY 9.2

TNT – it's not dynamite

Nitroglycerin was first developed in 1847 by Ascanio Sobrero, who deeply feared it. When he first synthesised the compound, it was so unstable that it exploded in his face, leaving him scarred for the rest of his life. Nitroglycerin is a highly unstable contact explosive. This means it is detonated without a flame or spark, but instead needs only a small physical shock. It combusts according to the following thermochemical equation:

$$4C_3H_5N_3O_9(l) \rightarrow 12CO_2(g) + 10H_2O(g) + 6N_2(g) + O_2(g)$$
$$\Delta H = -1456 \text{ kJ mol}^{-1}$$

In the 1840s, Alfred Nobel, for whom the Nobel Prize is named, started commercially producing the substance as an explosive for blast mining. He turned nitroglycerin into a less reactive paste, which could be moulded like plasticine. He added an ignition source and, therefore, dynamite was born. Dynamite is nitroglycerin with absorbents and stabilisers to ensure it does not combust by the smallest jolt.

TNT (trinitrotoluene) was first synthesised as a yellow dye in 1863 by Julius Wilbrand. It took a very long time for people to realise its destructive powers and therefore was not used as an explosive until some years later. It combusts according to the following thermochemical equation:

$$2C_6H_2(NO_2)_3CH_3(s) \rightarrow 7C(s) + 7CO(g) + 3N_2(g) + 5H_2O(l)$$
$$\Delta H = -10\,153 \text{ kJ mol}^{-1}$$

TNT is primarily used in military, industrial and mining applications. It is more stable than dynamite because it is insensitive to shock or friction. It took many years for scientists to realise that TNT was difficult to react because it needed a lot of energy to be added to the reactants so that their bonds would break and they would form products.

The combustion reaction of TNT has a more negative enthalpy than that of dynamite. This makes TNT preferable to use because it generates more energy and is more stable and easier to handle.

FIGURE 3 Cross-polarised light micrograph of TNT, magnified X10. TNT is more stable and easier to use than dynamite.

SUGGESTED PRACTICAL

9.2A Applying Hess's law
Go to your obook assess »

SUGGESTED PRACTICAL

9.2B Measuring the molar heat of a chemical reaction
Go to your obook assess »

CHECK YOUR LEARNING 9.2

Describe and explain

1 **Explain** enthalpy.

Apply, analyse and interpret

2 **Contrast** an endothermic and an exothermic reaction.

3 Apply your knowledge of enthalpy to the following questions, using the bond enthalpies from Table 1.
- **a** **Determine** the enthalpy of combustion of propene (Figure 4).
- **b** **Construct** a fully labelled enthalpy level diagram.
- **c** **Determine** ΔH for the combustion reaction.
- **d** **Determine** whether the reaction is endothermic or exothermic.

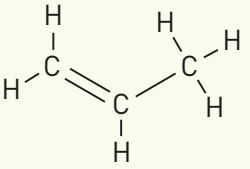


FIGURE 4 The structural formula of propene.

4 **Determine** the thermochemical equations for the combustion reactions of:
- **a** methane
- **b** ethyne
- **c** methanol.

5 **Apply** Hess's law to determine the enthalpy of the overall chemical equation:

$$2C(s) + 4H_2(g) + O_2(g) \rightarrow 2CH_3OH(l)$$

The three intermediary step equations are:

1 $C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H = -394$ kJ mol^{-1}

2 $2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l)$ $\Delta H = -1352$ kJ mol^{-1}

3 $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ $\Delta H = -484$ kJ mol^{-1}

Investigate, evaluate and communicate

6 **Investigate** the statement: 'Bond enthalpies do not take into consideration any intermolecular or intramolecular forces that may exist in the reaction system'. What does 'intermolecular or intramolecular forces' mean? Provide an example of this limitation.

7 **Comment** on the existence of reactions that are not exothermic or endothermic. Investigate this question and communicate your findings.

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- » Student book questions
  Check your learning 9.2
- » Challenge
  9.2 Manipulating enthalpy values
- » Suggested practical
  9.2A Applying Hess' law
- » Suggested practical
  9.2B Measuring the molar heat of a chemical reaction

9.3 Specific heat capacity

KEY IDEAS

In this section, you will learn about:

- the specific heat capacity of water and aqueous solutions
- calculating ΔH in aqueous reactions and combustion reactions
- calorimetry
- the difference between heat and temperature.

In many industries, it essential that the energy released by burning a fuel can be calculated. When the fuel is combusted, it releases energy from the system to the environment. If the energy released to the environment is used to heat a specific volume of water, the change in temperature of the water can be measured and the energy required to heat the water (which is the energy generated by the fuel when it combusts) can be calculated.

Specific heat capacity to calculate enthalpy changes

specific heat capacity the amount of heat energy, in joules (J), that it takes to increase the temperature of a specific volume of water by 1°C (remembering that 1 g of water is 1 mL)

Specific heat capacity (SHC) is a measure of the amount of heat energy, in joules (J), that it takes to increase the temperature of a specific volume of water (remember that 1 g of water is 1 mL) by 1°C. The specific heat capacity of water, under ideal conditions, is 4.2 J g^{-1} K^{-1} (rounded to 2 decimal places) or 75 J mol^{-1} K^{-1}, and can be calculated from the equation:

$$q = m \times c \times \Delta T$$

where q is heat energy in J (not kJ), m is the mass of the water (1 mL = 1 g), c is the specific heat capacity of water (4.2 J g^{-1} K^{-1}) and ΔT is the change in temperature in K (which is equivalent to the change in °C).

Study tip

Remember that $q = m \times c \times \Delta T$ refers specifically to water. q represents the energy required to heat a mass of water by a specific temperature and is calculated using the specific heat capacity of water.

WORKED EXAMPLE 9.3A

200 mL of water was heated in a beaker. The initial temperature of the water was 16°C and the final temperature was 62°C. How much energy was required to heat the water?

SOLUTION

$q = mc\Delta T$

$= 200\ g \times 4.2\ J\ g^{-1}\ K^{-1} \times (62 - 16)\ K$

$= 38\,640\ J$

$= 39\ KJ$

Converting between values of specific heat capacity

It is important that you can convert between the two specific heat capacity values for water. Converting from 4.2 J g^{-1} K^{-1} to 75 J mol^{-1} K^{-1} requires converting from mass to moles using molar mass (M) (Figure 1).

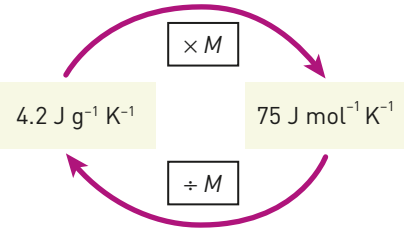

FIGURE 1 Conversion tool for specific heat capacity.

WORKED EXAMPLE 9.3B

Convert the specific heat capacity of water from 4.2 J g^{-1} K^{-1} to J mol^{-1} K^{-1}.

SOLUTION

4.2 J g^{-1} K^{-1} × $M(H_2O)$ = $SHC(H_2O)$ J mol^{-1} K^{-1}

$SHC(H_2O)$ = 4.2 J g^{-1} K^{-1} × 18 g mol^{-1}

= 75.6 J mol^{-1} K^{-1}

= 76 J mol^{-1} K^{-1}

This differs from the value of 75 J mol^{-1} K^{-1} in the databook because we have used a rounded-up value (4.2 J g^{-1} K^{-1}) for the SHC of water (4.184 J g^{-1} K^{-1}):

$SHC(H_2O)$ = 4.184 J g^{-1} K^{-1} × 18 g mol^{-1} = 75.312 J mol^{-1} K^{-1}, which rounds down to 75 J mol^{-1} K^{-1}

Calculating enthalpy changes in aqueous solutions

It can be assumed that aqueous solutions have the same specific heat capacity as water. Therefore, the energy of aqueous solutions can be calculated in the same way.

When a quantity of substance in an aqueous solution is known, a ΔH value can be calculated from the following formula, if the quantity can be converted to moles:

$$\Delta H = \frac{E}{n}$$

This formula becomes obvious when the units of ΔH are considered. If ΔH is measured in kJ mol^{-1}, then it must be calculated as kJ (energy) ÷ moles.

There are two steps involved in completing ΔH calculations.

1 Calculate E the energy that was required to raise the temperature of the water by the given temperature using the equation:

$$q = m \times c \times \Delta T$$

2 The energy calculated will be for the amount of substance in the aqueous solution. Convert this 'amount' to moles n and calculate the ΔH value using the equation:

$$\Delta H = \frac{E}{n}$$

calorimeter
a device that measures the change in temperature of water when an endothermic or exothermic reaction takes place

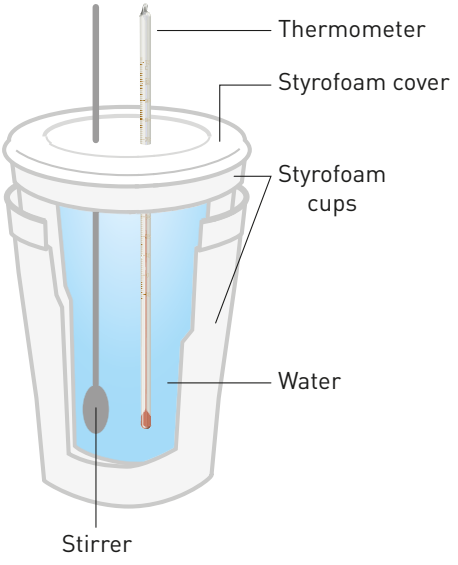

FIGURE 2 The experimental equipment involved in solution calorimetry.

Coffee cup calorimeter

The enthalpy change in an aqueous solution is typically measured in a solution **calorimeter**. A calorimeter is any device that is designed to measure the temperature rise or fall of water when an endothermic or exothermic reaction takes place.

A simple calorimeter can be made from two insulated foam coffee cups and a foam lid, for maximum heat retention (Figure 2). There must be a space in the lid to fit a thermometer, to measure temperature change, and a stirrer for even heat distribution.

However, this experimental technique involves systematic errors because the coffee cups do not retain 100% of the heat of the reaction and some heat is lost to or absorbed from the surroundings.

This is a systematic error as long as the same coffee cup is used for all experiments, and affects all solutions in the same way.

The experimental value for ΔH should always be lower than the theoretical value because the calorimeter does not retain all of the heat energy in the solution.

WORKED EXAMPLE 9.3C

When a 3.85 g sample of caesium fluoride (CsF) is added to 40 mL of water in a coffee cup calorimeter (Figure 2), a temperature rise of 5.4 K is measured. Calculate the enthalpy change of solution for 1 mol of CsF.

SOLUTION

1 Calculate the energy absorbed by the water.

$q = mc\Delta T$

$= 40\ \text{g} \times 4.2\ \text{J g}^{-1}\ \text{K}^{-1} \times 5.4\ \text{K}$

$= 907.2\ \text{J}$

$= 0.91\ \text{kJ}$

2 This is the energy for 3.85 g of CsF. Convert mass to moles and calculate the ΔH for 1 mol of CsF.

$n(\text{CsF}) = \frac{m}{M}$

$= \frac{3.85\ \text{g}}{151.91\ \text{g mol}^{-1}}$

$= 0.0253\ \text{mol}$

$\Delta H = \frac{E}{n}$

$= \frac{0.91\ \text{kJ}}{0.0253\ \text{mol}}$

$= 36\ \text{kJ mol}^{-1}$

Calculating enthalpy changes in combustion reactions

The enthalpy change in a combustion reaction can be calculated in a similar way as for aqueous solutions. However, the methodology is quite different because a flame is used.

Figure 3 shows a spirit burner, which is an effective way of burning a fuel while keeping the flame contained. The mass of the spirit burner is recorded before and after the experiment to determine the mass of the fuel used in the experiment to heat the water. A set volume of water is contained within the beaker or metal can, with a thermometer and stirrer.

As the mass of the fuel and the temperature rise of the water are known, a ΔH value can be calculated.

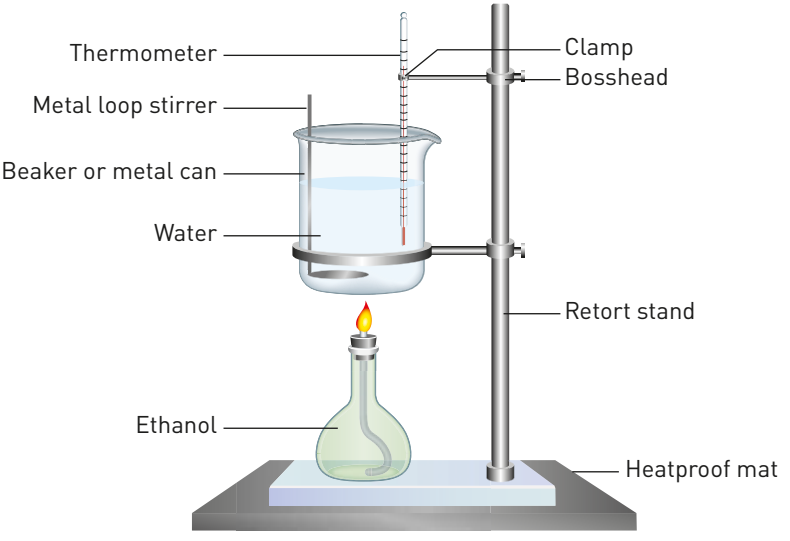

FIGURE 3 A spirit burner set up to record the enthalpy change in a combustion reaction.

WORKED EXAMPLE 9.3D

A 4.5 g sample of propane (C_3H_8) is burnt using a spirit burner. The energy produced by the flame was used to heat a beaker containing 150 mL of water. The starting temperature of the water was 18°C and the final temperature was 68°C. Determine the enthalpy change of solution for 1 mol of propane.

SOLUTION

1 Calculate the energy absorbed by the water.

$$\begin{aligned} q = mc\Delta T &= 150\,\text{g} \times 4.2\,\text{J g}^{-1}\,\text{K}^{-1} \times (68 - 18)\,\text{K} \\ &= 31\,500\,\text{J} \\ &= 31.5\,\text{kJ} \end{aligned}$$

2 This is the energy for 4.5 g of propane. Convert mass to moles and calculate ΔH for 1 mol of propane.

$$\begin{aligned} n(C_3H_3) &= \frac{m}{M} \\ &= \frac{4.5\,\text{g}}{44\,\text{g mol}^{-1}} \\ &= 0.10\,\text{mol} \end{aligned}$$

$$\begin{aligned} \Delta H &= \frac{E}{n} \\ &= \frac{31.5\,\text{kJ}}{0.10\,\text{mol}} \\ &= 310\,\text{kJ mol}^{-1} \end{aligned}$$

Distinguishing between heat and temperature

It is important to understand the distinction between heat and temperature. Heat is a form of energy referred to as 'thermal energy'. Temperature is a measure of the average kinetic energy of the particles.

Therefore, temperature is a measure of thermal (heat) energy.

CASE STUDY 9.3

Measuring energy within food products

Food laboratories work to determine whether the claims a company makes about its food products are accurate.

Food laboratories test for the presence of various biomolecules, including fats and proteins, as well as vitamins and minerals. They also measure the concentration of these nutrients to ensure that food labelling is accurate.

Calorimetry used to be the primary method food chemists used to determine the enthalpy of food products. The simplest method of determining the change in enthalpy due to the combustion of food products is to set the food alight and use the energy generated to heat a volume of water. The specific heat capacity of water is used to determine the energy content of the food. However, only dry food can be burnt and the significant amount of heat loss to the environment limits the usefulness of this approach.

MANDATORY PRACTICAL
9.3 Calorimetry: Measuring the enthalpy of a reaction
Go to page 450 »

The design of calorimeters has improved so that better insulation minimises heat loss in solution (for liquids) or bomb calorimeters (for solids). In a bomb calorimeter, a sample of solid is contained within a reaction capsule (Figure 4). An electric current is used to ignite the sample, which is fed with a constant supply of excess oxygen and the whole sample reacts. The temperature change of the water is recorded and the energy of the food is calculated.

Today, much more sophisticated instrumentation is involved.

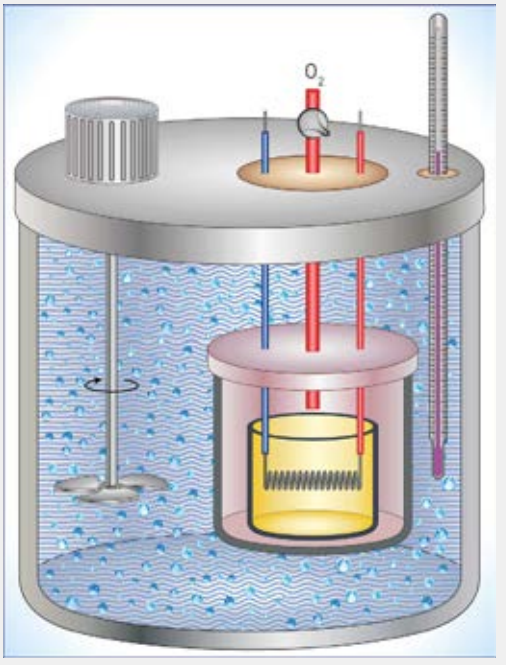

FIGURE 4 A bomb calorimeter

CHECK YOUR LEARNING 9.3

Describe and explain

1 **Identify** what specific heat capacity measures.
2 **Describe** a calorimeter.

Apply, analyse and interpret

3 **Compare** the two techniques for measuring ΔH for aqueous solutions and combustion reactions
4 **Distinguish** between heat and temperature.
5 **Determine** how much energy is required to heat 72 mL of water, if the initial temperature of the water is 273 K and the final temperature is 17°C.
6 **Determine** the enthalpy change of solution of magnesium nitrate ($Mg(NO_3)_2$) when a 0.72 g sample of the salt is added to 123 mL of water in a coffee cup calorimeter and a temperature rise of 0.85°C is measured.
7 A 10.9 g sample of ethanol (C_2H_6O) is burnt to heat a beaker containing 320 mL of water. The initial temperature of the water was 273 K and the final temperature was 72°C. **Determine** the enthalpy change of solution for 1 mol of ethanol.

Investigate, evaluate and communicate

8 **Investigate** and communicate the systematic errors and limitations of the spirit burner experiment when experimentally determining the ΔH of fuels. Hint: refer to Table 3 the heat of combustion table on page 206.
9 **Discuss** the systematic errors involved with coffee cup calorimetry and spirit burner combustion measurements. **Evaluate** both methods and communicate improvements that could be made to each method, based on the errors identified.

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» Student book questions
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» Increase your knowledge
9.3 Experimental combrustion

» Weblink
Calculating enthalpy changes in aqueous solutions

» Weblink
Calculating enthalpy changes in combustion reactions

SCIENCE AS A HUMAN ENDEAVOUR

9.4 Photosynthesis and cellular respiration

KEY IDEAS

In this section, you will learn about:

- chemical energy in the body
- photosynthesis and cellular respiration.

Photosynthesis

autotroph
an organism, such as a plant or alga, that produces its own food by transforming sunlight to chemical energy

producer
an autotroph; it produces its own food by transforming sunlight to chemical energy

photosynthesis
an endothermic reaction that converts CO_2 and H_2O to glucose ($C_6H_{12}O_6$) and O_2 using sunlight; occurs in the chloroplasts of autotrophs

chloroplast
an organelle in plants cells that is the site of photosynthesis

Plants are **autotrophs**. Autotrophs are also called **producers** because they produce their own energy. As the law of conservation of energy states that energy cannot be created or destroyed, the production of energy is actually an energy transformation. An autotroph's energy source is the Sun. Plants use energy from the Sun in the reaction of carbon dioxide (from the atmosphere) with water (from soil via their roots) according to the equation:

$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g) \qquad \Delta H = +2802.7 \text{ kJ mol}^{-1}$$

Photosynthesis is an endothermic reaction. Energy must be absorbed from the environment because the energy of the reactants is lower than the energy of the products.

The enthalpy level diagram for this reaction is shown in Figure 1.

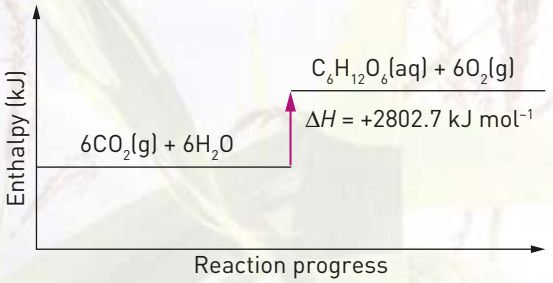

FIGURE 1 The enthalpy level diagram for photosynthesis, an endothermic process.

Photosynthesis occurs in the **chloroplasts**, which are organelles of cells, located in leaves. Chloroplasts contain the green pigment chlorophyll, which is essential for the photosynthesis reaction to occur. The light-dependent part of the photosynthesis reaction occurs on flat disc-like shapes within the chloroplast, called thylakoids (Figure 2).

Once the reaction is complete, the glucose is tranported to other parts of the plant for use as an energy source.

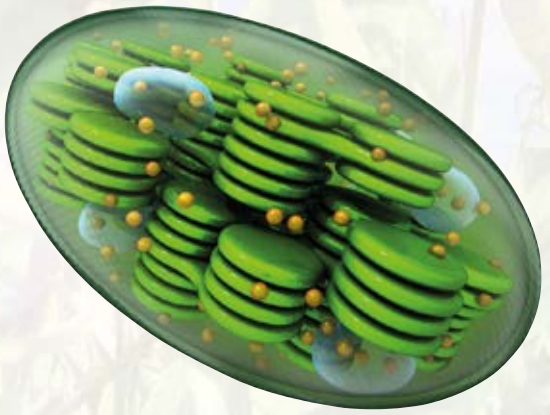

FIGURE 2 The structure of a chloroplast showing the thylakoid discs.

Cellular respiration

Humans are **heterotrophs** because we must consume food for energy. Unlike autotrophs, we cannot transform the Sun's energy into chemical energy. Sugars in our food are ingested and broken down within the body for energy. This process, called **cellular respiration**, occurs according to the chemical equation:

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l) \qquad \Delta H = -2802.7 \text{ kJ mol}^{-1}$$

This is why humans need to inhale O_2 through the lungs and into the bloodstream, and why CO_2 is exhaled as waste.

Cellular respiration is an exothermic reaction. The energy of the reactants is greater than the energy of the products, so energy is released from the system to the surroundings.

The enthalpy level diagram for this reaction is shown in Figure 3.

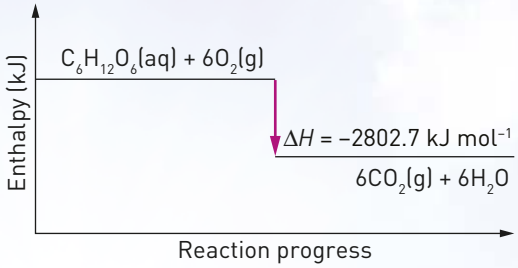

FIGURE 3 The enthalpy level diagram for cellular respiration, an exothermic process.

heterotroph
an organism that must consume an autotroph or other organism to gain energy

cellular respiration
an exothermic reaction that converts glucose ($C_6H_{12}O_6$) and O_2 to CO_2 and H_2O; occurs in the mitochondria of heterotrophs and autotrophs

To put this ΔH value in context, the quantity of water that could be heated to 10°C by this amount of energy is:

$$q = mc\Delta T$$

$$m(H_2O) = \frac{q}{c\Delta T}$$

$$= \frac{(2802.7 \times 1000)\text{J}}{4.2\ \text{J g}^{-1}\text{K}^{-1} \times 10\ \text{K}}$$

$$= 66730.95\ \text{g}$$

$$= 66.7\ \text{kg}$$

Where is energy needed most?

Cellular respiration occurs within **mitochondria**, which are organelles within body cells.

mitochondria
a cell organelle that is the site of cellular respiration

Regular body cells contain 1000 to 2000 mitochondria per cell. Areas of the body with higher energy requirements have cells with more mitochondria. For example, muscles need a great deal of energy for movement, so their cells contain high numbers of mitochondria. Cardiac cells (heart cells) have approximately 5000 mitochondria. Liver cells have 1000 to 2000 mitochondria because they carry out various specialised functions.

Producing more energy without increasing the number of mitochondria

Mitochondria have folds called cristae (Figure 4). Cellular respiration occurs at the surface of these folds. In those parts of the body that require more energy, but cannot increase the number of mitochondria, the mitochondria contain more cristae. Increasing the number of folds in mitochondria increases the surface area available for cellular respiration. Therefore, more cellular respiration occurs and there is an increase in energy production.

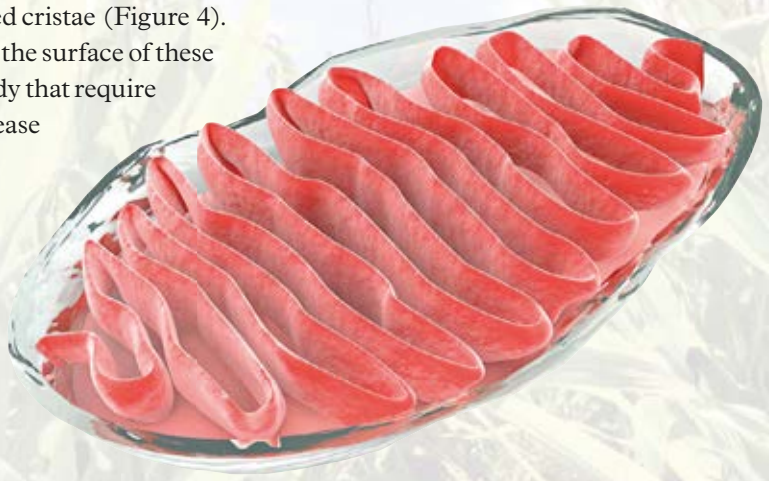

FIGURE 4 Structure of mitochondria, showing the folds or cristae.

CASE STUDY 9.4

What is a calorie?

Chemists and nutritionists use the term 'calorie' to refer to a different amount of energy. This can lead to confusion.

Chemists use 'calorie' (cal) (also known as 'small calorie') for the energy calculated using specific heat capacity. It is defined as the energy needed to raise the temperature of 1 g of water by 1°C at 1 atm.

Nutritionists and the food industry recognise a 'calorie' as a kilocalorie. 1 kilocalorie is the amount of heat required to raise the temperature of 1 kg of water by 1°C.

To determine the difference between the two, food packaging uses a capital 'C' to refer to the dietary 'Calorie', whereas chemists use the abbreviation 'cal'.

Chemists now use the joule (J) as the standard unit of energy, where 1 cal = 4.2 J (SHC).

FIGURE 5 Nutritional calories (C) in food.

CHALLENGE 9.4A

Number of mitochondria

Why does a sperm cell only contain 25 to 50 mitochondria if its biological role is to swim, find an egg and fertilise it?

CHALLENGE 9.4B

Metabolisms of microorganisms

Chemoautotrophs use inorganic energy sources to synthesise organic compounds from carbon dioxide. Photoautotrophs use sunlight to convert carbon dioxide to organic compounds.

Discuss the strengths and weaknesses of both of these metabolisms.

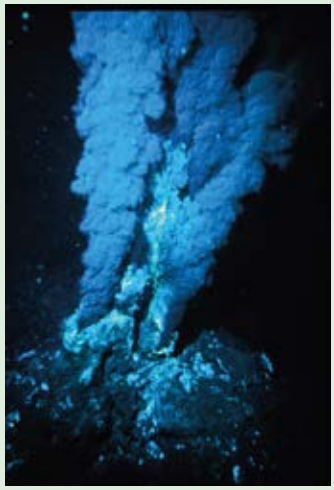

FIGURE 6 Some chemoautotrophs are found at deep sea hydrothermal vents, where there is no sunlight.

CHECK YOUR LEARNING 9.4

Describe and explain

1 **Explain** the difference between an autotroph and a heterotroph.
2 **Explain** the difference between photosynthesis and cellular respiration.
3 **Identify** the organelles in which photosynthesis and cellular respiration occur and **explain** why humans cannot be autotrophs.

Apply, analyse and interpret

4 **Critique** the ΔH value of cellular respiration by calculating bond enthalpies for its reactants and products.
5 **Consider** how a chloroplast could increase the rate of production of glucose.
6 **Discuss** the role of mitochondia within cells.

Investigate, evaluate and communicate

7 **Investigate** whether there are any cells within the body that have no mitochondria. Communicate your findings.
8 **Investigate** whether the brain contains mitochondria. If so, where? If not, why not? Communicate your findings

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- » Challenge
  9.4B Metabolisms of microorganism
- » Weblink
  Photosynthesis

CHAPTER 9

Review

Chapter summary

9.1

- Phase changes occur due to changes in heat and light. They cause physical changes in a reaction system (reversible).
- There are many different types of chemical reactions. All chemical reactions must include the correct formulas and states of reactants and products and be balanced by counting the number of atoms or ions on each side of a chemical reaction.

9.2

- Chemical energy is a form of potential energy due to the arrangement of atoms in elements or molecules. Enthalpy is also the amount of chemical energy stored within the bonds of a substance; therefore, it is a measure of chemical energy. It is sometimes referred to as heat content, and is represented with the symbol *H*.
- Bond enthalpy is the specific amount of energy contained within individual bonds of a substance. By adding the individual bond enthalpies for each bond, an enthalpy value can be calculated for the molecule.
- Enthalpy level diagrams are used to represent the different enthalpy levels of reactants and products. The change in enthalpy between reactants and products (ΔH) determines whether a reaction is exothermic or endothermic. In exothermic reactions, the reactants have more energy than the products, heat energy (enthalpy) is lost from the system to the surroundings, the reaction flask feels hot and ΔH is negative. In endothermic reactions, the products have more energy than the reactants, heat energy (enthalpy) is absorbed from the surroundings to the system, the reaction flask feels cool and ΔH is positive.
- Hess's law states that the enthalpy change for the overall reaction is equal to the sum of the enthalpies for the individual steps.
- Thermochemistry is the study of how heat is absorbed (endothermic) or released (exothermic) from a chemical reaction. Thermochemical equations must include a ΔH value. If the reaction is reversed, so is the sign of the ΔH value. If the reaction coefficients are doubled, so is the ΔH value.

9.3

- Specific heat capacity is a measure of the amount of heat energy (J) it takes to increase the temperature of a specific mass of a substance by 1°C. The specific heat capacity of water, under ideal conditions, is 4.2 J g^{-1} K^{-1} or 75 J mol^{-1} K^{-1}.
- Calorimeters are used to measure the energy changes in a chemical reaction. However, combustion reactions cannot be contained within a coffee cup calorimeter and so a fuel is placed in a pre-weighed spirit burner which is used to heat a beaker of a known volume of water. The temperature increase and the mass of the spirit burnt are measured to determine the mass of fuel used to generate the temperature rise. A ΔH value can be determined by using the specific heat capacity of water, where:

$$q = m \times c \times \Delta T$$

The energy value is converted to kJ and the ΔH value of the fuel is calculated from the equation:

$$\Delta H = \frac{E \text{ (in kJ)}}{n}$$, where n = number of moles of the fuel and E is the energy absorbed and calculated using $q = m \times c \times \Delta T$

9.4

- Photosynthesis occurs in the chloroplasts of plants and other autotrophs, when carbon dioxide and water react to convert energy from the Sun into the chemical energy stored within the bonds of the products.

$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$ $\Delta H = +2802.7$ kJ mol^{-1}

- Cellular respiration is the opposite process, where organisms consume the plants and therefore the glucose. Glucose is metabolised slowly in the body of the organism to generate energy:

  $C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6\ H_2O(l)$ $\Delta H = -2802.7\ kJ\ mol^{-1}$

- Cellular respiration occurs in the mitochondria. The higher the energy need of an organ or structure, the greater the concentration of mitochondria in its cells. Mitochondria that have more cristae folds also produce more energy because they have a greater surface area for cellular respiration to occur.

Key terms

- acid
- anion
- aqueous
- autotroph
- balanced
- base
- boiling
- bond enthalpy
- calorimeter
- cation
- cellular respiration
- chemical energy
- chloroplast
- combination
- combustion
- condensation
- decomposition
- deposition
- double displacement
- endothermic
- enthalpy
- evaporation
- exothermic
- freezing
- ΔH
- Hess's law
- heterotroph
- intermolecular forces
- intramolecular forces
- ion
- law of conservation of energy
- melting
- mitochondria
- molten
- oxidation state
- photosynthesis
- producer
- redox
- single displacement
- specific heat capacity
- spectator ion
- stability
- sublimation
- thermochemical equation
- vaporisation

Key formulas

Enthalpy	$\Delta H = H_{products} - H_{reactants}$
Bond enthalpy	$\Delta H = \Sigma H_{bonds\ broken} - \Sigma H_{bonds\ formed}$
Specific heat capacity	$q = m \times c \times \Delta T$
Enthalpy in aqueous solutions	$\Delta H = \frac{E}{n}$
Number of moles using mass	$n = \frac{m}{M}$

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; ***=high.

Multiple choice

1 The enthalpy for the formation of NO_2 is given by:

$N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$ $\Delta H = +66$ kJ mol^{-1}

Calculate the enthalpy change for the following reaction:

$4NO_2(g) \rightarrow 2N_2(g) + 4O_2(g)$

A +33 kJ mol^{-1}

B +132 kJ mol^{-1}

C –132 kJ mol^{-1}

D –66 kJ mol^{-1}

2 Calculate ΔH for the following reaction.

$2N_2(g) + 5O_2(g) \rightarrow 2N_2O_5(g)$

Use the following information:

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ $\Delta H = -571.6$ kJ mol^{-1}

$N_2O_5(g) + H_2O(l) \rightarrow 2HNO_3(l)$ $\Delta H = -76.6$ kJ mol^{-1}

$N_2(g) + 3O_2(g) + H_2(g) \rightarrow 2HNO_3(l)$ $\Delta H = -348.2$ kJ mol^{-1}

A –29 kJ mol^{-1}

B –1421 kJ mol^{-1}

C +1421 kJ mol^{-1}

D +29 kJ mol^{-1}

3 The heat of combustion of glucose, in kJ g^{-1} is:

A 504 486.0

B 15.6

C 45.2

D 173 767.4

4 Which of the following statements about exothermic reactions is true?

A The reactants have a higher enthalpy than the products; therefore, energy is absorbed from the environment to form products.

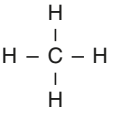
B The products have a higher enthalpy than the reactants; therefore, energy is absorbed from the environment to form products.

C The reactants have a higher enthalpy than the products; therefore, energy is released from the system when products form.

D The products have a higher enthalpy than the reactants; therefore, energy is released from the system when products form.

5 Using the table of bond enthalpies (Table 1, page 203), determine which of the following molecules is the most stable.

A CH_4



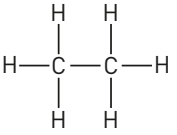
B CO_2

O=C=O

C O_2

O=O

D C_2H_6



Short answer

Describe and explain

★ **6** **Describe** the process of a gas forming a liquid and then becoming a solid.

★ **7** **Explain** why the reactants and products in chemical reactions have different enthalpies.

★ **8** **Describe** why photosynthesis and cellular respiration occur in plants, but only cellular respiration occurs in animals?

★ **9** A 1.2 g sample of octane (C_8H_{18}) is combusted to raise the temperature of 50 mL of water by 30°C. **Calculate** the enthalpy change of octane.

★★ **10** Is the change of state from solid to liquid an endothermic or exothermic process? **Explain** your answer.

★★★ **11** **Construct** an enthalpy level diagram for the following chemical reactions. State the change in enthalpy for each reaction and whether the reaction is endothermic or exothermic.

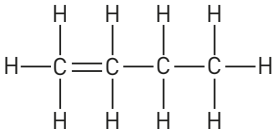
a Sodium hydroxide (NaOH) and hydrochloric acid (HCl) form sodium chloride (NaCl) and water (H_2O), where the enthalpy values of the reactants is 703 kJ mol^{-1} and that of the products is 646.7 kJ mol^{-1}.

b Carbon (C) and oxygen (O_2) form carbon dioxide (CO_2), where enthalpy value of reactants is 0 kJ mol^{-1} and that of the product is –393.5 kJ mol^{-1}.

Apply, analyse and interpret

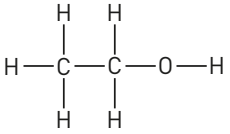
★ **12** **Determine** the enthalpy of propane, from its bond enthalpies in Table 1 (page 203).

★ **13** Using Table 1, section 9.2, **determine** the amount of energy required to break 1 mol of but-1-ene into its constituent atoms.



But-1-ene

★ **14** Using Table 1, section 9.2, **determine** the amount of energy required to break 1 mol of ethanol into its constituent atoms.



Ethanol

★ **15** Use bond enthalpies to **determine** the enthalpy of combustion of ethanol.

★ **16** **Determine** the thermochemical equations for the combustion reactions of:

A octane

B ethane

C propane.

★ **17** **Classify** how much energy is required to raise the temperature of 1 L of water by 20°C.

★ **18** **Consider** why water has a higher specific heat capacity (4.2 J g^{-1} °C) than copper (0.39 J g^{-1} °C).

★★ **19** **Determine** the heat of reaction (ΔH) for the overall equation:

$2C(s) + H_2(g) \rightarrow C_2H_2(g)$

from the enthalpy values of the intermediary steps:

Step 1: $C_2H_2(g) + \frac{5}{2}O_2(g) \rightarrow 2CO_2(g) + H_2O(l)$

$\Delta H = -1299.5$ kJ mol^{-1}

Step 2: $C(s) + O_2(g) \rightarrow CO_2(g)$

$\Delta H = -393.5$ kJ mol^{-1}

Step 3: $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$

$\Delta H = -285.8$ kJ mol^{-1}

★★ **20** Determine the heat of reaction (ΔH) for the overall equation:

$2Fe(s) + \frac{3}{2}O_2(g) \rightarrow Fe_2O_3(s)$

from the enthalpy values of the intermediary steps:

Step 1: $2Fe(s) + 6H_2O(l) \rightarrow 2Fe(OH)_3(s) + 3H_2(g)$

$\Delta H = +322$ kJ mol^{-1}

Step 2: $Fe_2O_3(s) + 3H_2O(l) \rightarrow 2Fe(OH)_3(s)$

$\Delta H = +289$ kJ mol^{-1}

Step 3: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

$\Delta H = -572$ kJ mol^{-1}

★★ **21** **Determine** what would be the final temperature if 100 g of water at 90°C is mixed with 600 g of water at 20°C? (Hint: Heat is transferred from the hotter to the cooler water.)

Investigate, evaluate and communicate

★★ **22** Often, fuels undergo incomplete combustion. **Investigate** how you might determine whether incomplete combustion has occurred in the spirit burner combustion technique. What could you include in the experimental method to ensure that incomplete combustion is avoided.

★★ **23** Ethanol is often used as an alternative to petrol as a fuel for the internal combustion engine. The enthalpy value for ethanol is 1370 kJ mol^{-1} and petrol is 5460 kJ mol^{-1}. **Evaluate** the efficiency of the fuels and determine one advantage and one disadvantage of using each fuel.

★★ **24** **Investigate** the term 'calibration factor'. Why is this concept important in calorimetry? How is it determined experimentally?

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CHAPTER 10

Measurement uncertainty and error

Measurements are a set of numerical information that describes the properties of an object or how it behaves. Every time a measurement is made, there is a possibility of the measured value differing from the true value. This error or uncertainty may be due to problems with the experimental design or to unpredictable random events. Calculating the error of a measurement enables a scientist to evaluate the effectiveness of a procedure and plan to minimise the errors in the future.

OBJECTIVES

- → Distinguish between precision and accuracy and appreciate that all measurements have limits to their precision and accuracy that must be considered when evaluating experimental results.
- → Distinguish between qualitative and quantitative data; appreciate that quantitative data obtained from measurements is always associated with random error/ measurement uncertainties.
- → Communicate measurement uncertainties as a range (±) to an appropriate precision.
- → Understand that propagation of random error in data processing shows the impact of measurement uncertainties on the final result.
- → Calculate the measurement uncertainties in processed data, including the use of absolute uncertainties and percentage uncertainties.
- → Construct and use appropriate graphical representations to communicate understanding, solve problems and make predictions; interpret graphs in terms of the relationship between dependent and independent variables; draw and interpret best-fit lines or curves through data points, including evaluating when it can and cannot be considered as a linear function.
- → Calculate the percentage error when the experimental result can be compared with a theoretical or accepted result (value).
- → Distinguish between random and systematic errors; understand that experimental design and procedure usually leads to systematic errors in measurement, which causes a deviation in a direction; appreciate that repeated trials and measurements will reduce random error but not systematic error.

→ Analyse the impact of random error/measurement uncertainties and systematic errors in experimental work and evaluate how these errors/measurement uncertainties can be reduced.

→ Understand that the number of significant figures in a result is based on the figures given in the data and determine the results of calculations to the appropriate number of significant figures.

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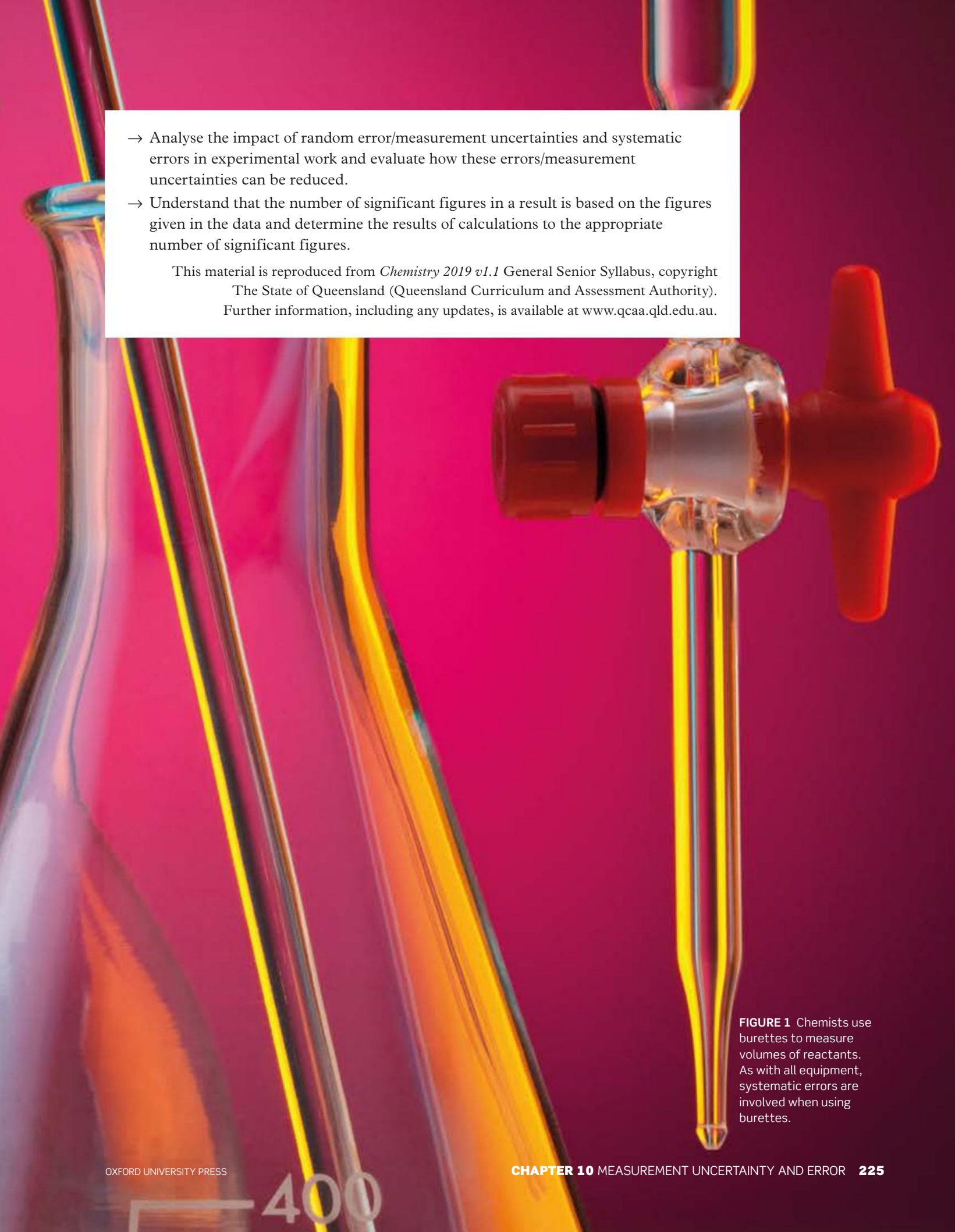

FIGURE 1 Chemists use burettes to measure volumes of reactants. As with all equipment, systematic errors are involved when using burettes.

10.1 Accuracy and precision

KEY IDEAS

In this section, you will learn about:

- the difference between qualitative and quantitative data
- random errors and measurement uncertainties that are always linked to quantitative data
- the difference between precision and accuracy
- the limited precision and accuracy of all measurements that must be considered when evaluating experimental results.

There are many ways of measuring objects. An object may be red, spotted or flat or have the shape of a 3 cm cube or the volume of 5 cm^3. An observer must be able to describe the object so that they can communicate the key features to others. The description and numerical measurements used to analyse an object or attributes are called data.

Qualitative versus quantitative data

When analysing a substance, a chemist first describes the types of elements or grouping of elements that are present in the sample. Non-numerical or **qualitative** analysis of a substance can identify whether an atom, ion, functional group or compound is present, but it does not tell the chemist how much of the substance present.

qualitative
data about types or categories, which may be represented by names, symbols, codes etc.

quantitative
relating to numerical data of a substance or object

Quantitative analysis consists of a series of numbers that describe the amount of each element in the sample. The quantitative analysis of a substance allows a chemist to predict the amount of product that will be produced as well as the purity of a sample. Knowing the amount of an element present in a sample also allows doctors to determine whether a patient is unwell, or a mining company to determine whether a mineral ore is worth mining.

Accuracy vs precision

accuracy
a measure of how close the measured value is to the true or accepted value

precision
the consistency a series of measurements

It is important for a scientist to know whether a measurement is correct or not. The **accuracy** of a measurement relates to how it is close to the true or accepted value. This is similar to playing a ball sport where the ball always lands in the centre of the goal.

A measurement can also be described as precise. **Precision** relates to the consistency of the measurements. To be precise, the series of measurements must be values that are similar to each other. In a ball sport, this is similar to the ball always landing in the same place (that may or may not be close to the goal). Precision is a measure of how repeatable the measurement is (Figure 1). Most sportspeople try to be both accurate and precise.

Study tip

aCcurate is Correct (a bullseye).
pRecise is Repeating (hitting the same spot, but maybe not the correct spot).

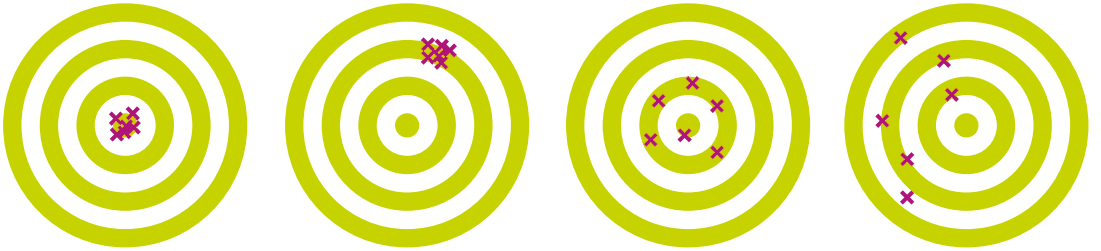

FIGURE 1 Accuracy is a measure of how close the mean result is to the centre 'true' value. Precision is a measure of how closely grouped the results are to each other.

An example of accuracy and precision in a laboratory is when measuring the mass of a 50.0 g standard weight. If an electronic balance obtained the values 48.3, 48.4, 48.3 and 48.5 g, then the balance could be described as precise (all the values are within the range of 48.3–48.5), but not very accurate (they are more than 1 g from the true value). Alternatively, if a second electronic balance gave the values of 50.4, 49.7, 51.0 and 49.3, the mean value would be 50.1 g. This suggests that the balance was accurate (close to the true value) but not precise. In a laboratory, the first, more precise, balance would be better to use if the scientist could adjust for the measurement error.

FIGURE 2 An electronic balance that is precise is often better to use because a scientist can make adjustments for a known measurement error.

WORKED EXAMPLE 10.1

A thermometer is used to measure water at a standard 50°C as 50.4, 49.7, 51.0 and 49.3°C. Calculate the mean of these values and use it to describe the accuracy and precision of the measurements.

SOLUTION

$$\text{Mean} = \frac{\text{sum of all numbers}}{\text{number of values}}$$

$$= \frac{50.4°C + 49.7°C + 51.0°C + 49.3°C}{4}$$

$$= 50.1°C$$

The thermometer is considered accurate because the mean value is 0.1°C from the true value of 50°C.

The thermometer is not considered precise because the values vary from 51.0 to 49.3°C.

Study tip

To calculate the mean, add all the numbers together, then divide by the number of values.

Calibration

Many types of equipment allow for the checking and adjustment of measured values so that they become more accurate. This is known as **calibrating** the equipment. With a set of scales, this may be as simple as pressing the tare button to zero the mass. Other equipment, such as a thermometer, must be compared to a standard. A standard measurement is an accepted value that does not change over time.

calibration
the checking and adjustment of measured values of equipment so that they become more accurate

It is important to calibrate equipment to determine whether it is accurate. For example, a thermometer may be more reliable over a set range. At higher temperatures, the thermometer may become increasingly inaccurate (although it may still be precise). This means the thermometer will make accurate measurements between 15°C and 80°C, but at 10°C it may be 0.2°C above or below the true value. At an even lower temperature, the error may be larger.

FIGURE 3 Digital thermometers need to be regularly calibrated by comparing the value to a standard to ensure accuracy is maintained.

Accuracy can be improved by calibrating equipment, precision can be improved by repeating the measurement because the more values that are obtained, the greater the consistency of those numbers. Repeatedly taking measurements will not improve the accuracy of the results because any error in the equipment will not improve.

CHECK YOUR LEARNING 10.1

Describe and explain

1 **Define** the following terms and describe how a chemist can improve each when measuring a mass.
   - **a** precision
   - **b** accuracy
   - **c** calibration

2 **Describe** the accuracy and precision shown in Figure 4.

FIGURE 4 Representation of a series of data points.

3 A student measures the mass of a substance and finds it to be 12.31 g. They state that their measurement is both accurate and precise. **Identify** errors in this student's statement.

Apply, analyse and interpret

4 **Contrast** qualitative data and quantitative data. Provide an example of each to illustrate your answer.

5 A standard weight (10 g) was measured on an electronic balance and the following measurements were made: 9.9, 10.2, 8.9 and 10.9 g. **Classify** these results as accurate and/or precise. Provide reasoning to support your answer.

6 Three students measured the concentration of a sodium chloride (NaCl) solution. They each made three measurements of the mass of NaCl (in grams) remaining after evaporation of 100 g of solution. Their results are shown in Table 1.

TABLE 1 Results from three trials of measuring a solution of sodium chloride in grams of NaCl/100 g of water

Student	Mass NaCl (g)		
	Trial 1	Trial 2	Trial 3
Student 1	34.7	35.4	36.0
Student 2	41.2	31.9	34.1
Student 3	38.9	39.1	39.2

The true value of sodium chloride is 35.9 g. **Interpret** the results to determine which student was both precise and accurate.

Investigate, evaluate and communicate

7 In chemistry, there are two types of standards – primary standards and secondary standards. **Investigate** the difference between a primary standard solution and a secondary standard solution. Provide a short definition of each for your class.

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- » Increase your knowledge
  Measures of central tendency
- » Weblink
  Calibration of equipment
- » Weblink
  Accuracy vs. precision

10.2 Calculating uncertainties and errors

KEY IDEAS

In this section, you will learn about:

- expressing uncertainties as a range (±) to an appropriate precision
- using the range of random errors when calculating data to show the impact of uncertainties on the final result
- calculating measurement uncertainties in processed data
- calculating percentage error in an experimental data.

When making a measurement with a ruler, it can be difficult to determine which millimetre line is closest to the object being measured. The measurement could be directly between two marked millimetres. This may make the observer uncertain about the measurement and they may question the accuracy of the ruler. Did the ruler manufacturer test each ruler before it was sold?

Uncertainties

uncertainty a numerical measurement that describes the observer's confidence in the accuracy and precision of a measurement

An **uncertainty** is a description of confidence in the accuracy and precision of a measurement. It is typically expressed as a number. All uncertainties are caused by two factors – the accuracy of the measuring instrument (systematic error) and the skill of an experienced person using the instrument (random error). If the person repeatedly uses the equipment the same way, the random error may become a systematic error.

When making a measurement, most people assume that there is a precise value that they should achieve. The amount of absolute uncertainty surrounding the precise value is usually shown as a range above or below the measured value. For example, if a student measured the length of a piece of string as 15 mm, they may not be sure if the end of the string was exactly on the ruler line for that value. This means the piece of string could be between 14.5 and 15.5 mm. Rather than list all the values between these two numbers, the length of string can be expressed as 15.0 ± 0.5 mm. To express this in more general terms:

Measurement = best estimate ± absolute uncertainty

This means a measurement of 2.04 ± 0.05 g suggests that the precise mass is between 1.99 and 2.09 g.

The range of values suggested by the absolute uncertainty depends on the type of equipment being used and the skill of the observer. For example, when measuring a 50 mL volume, a 200 mL beaker may have an uncertainty of 5 mL. In contrast, a measuring cylinder may have an absolute uncertainty of 1 mL. Equipment designed for measuring substances is always more accurate than general containers.

Similarly, equipment used in a laboratory is calibrated to minimise uncertainties. An electronic laboratory balance has a smaller absolute uncertainty than kitchen scales.

FIGURE 1 It is difficult to make an exact measurement with the beaker. The best estimate is 700 mL with an absolute uncertainty of ±10 mL. This means the amount of fluid is between 690 and 710 mL.

parallax error
a random error that results from different values being read due to the angle of the observer

Uncertainties may be further complicated by **parallax error**, where the observer may be measuring a substance at an angle. This can change the value measured in non-digital instruments (Figure 2).

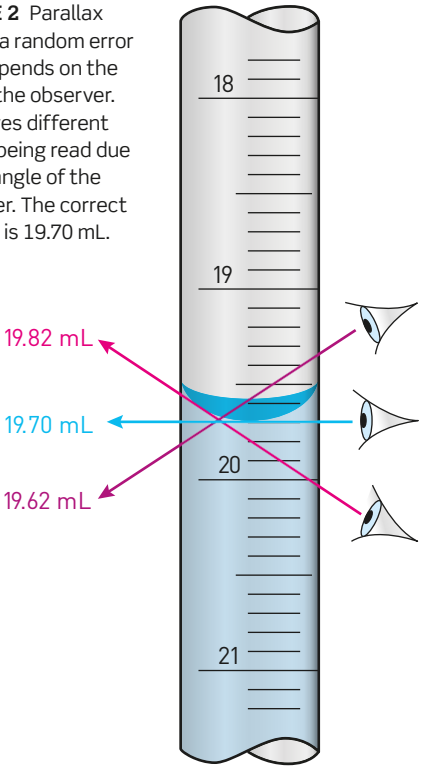

FIGURE 2 Parallax error is a random error that depends on the skill of the observer. It involves different values being read due to the angle of the observer. The correct reading is 19.70 mL.

Calculating absolute uncertainties

When calculating the absolute uncertainty of a measurement, the number of digits used is important. For example, a measurement of 41.2 cm cannot have an absolute uncertainty of 1.2 cm. If the experimenter could measure exactly 0.2 cm, then the absolute uncertainty would be less than 1.2 cm. As a rule, the best estimate should be rounded to the same digit as the absolute uncertainty. This means the absolute uncertainty of the measurement should be 41 ± 1 cm or 41.2 ± 0.1 cm depending on the accuracy of the measuring instrument.

CASE STUDY 10.2

Faster than the speed of light?

In 2011, scientists working on OPERA (Oscillation Project with Emulsion-tracking Apparatus) in Italy announced to the scientific community that they had measured a subatomic particle (a neutrino) travelling faster than the speed of light. This shocked the science world because it challenged Einstein's theory of special relativity, which says nothing can travel faster than the speed of light in a vacuum. Unfortunately, the uncertainty of the speed measurement was found to be much larger than initially calculated. This was due to a wiring error that caused a miscalculation in the timing. When the uncertainty was corrected by OPERA in 2012, the speed range of the neutrino was found to be below the speed of light in a vacuum.

The impact of uncertainties

All scientists want their experimental results to be accurate, but this is not easy to achieve. There are always errors and uncertainties involved in every procedure. Assigning a number to the level of uncertainty can identify whether the value is within an expected range of a known value or whether the experiment can achieve a similar result when repeated.

consistent
when measurements have overlapping uncertainty ranges

discrepant
when there is no overlap between the uncertainties of two measurements

Consider the two measurements of 0.75 seconds and 0.83 seconds. Although the second measurement appears larger than the first, it may be the same measurement if the absolute uncertainty is large enough. If the absolute uncertainty is ±0.08 seconds, then there will be an overlap in the calculated range of the two measurements. Measurements that have overlapping uncertainty ranges are said to be **consistent**. If there is no overlap between the two uncertainties, the measurements are **discrepant**.

WORKED EXAMPLE 10.2A

Describe each of the following measured values as consistent or discrepant.

a 0.75 ± 0.02 s and 0.83 ± 0.02 s

b 1.35 ± 0.06 g and 1.45 ± 0.06 g

SOLUTION

a 0.75 ± 0.02 s = 0.73–0.77 s

0.83 ± 0.02 s = 0.81–0.85 s

There is no number that fits into both intervals, so the measured values are discrepant.

b 1.35 ± 0.06 g = 1.29–1.41 g

1.45 ± 0.06 g = 1.39–1.51 g

The values between 1.39 and 1.41 fit into both intervals, so measured values are consistent.

Percentage uncertainty

The uncertainty of a measured value can also be expressed as a percentage of the value. **Percentage uncertainty** can be calculated by the equation:

$$\text{Percentage uncertainty} = \frac{\text{absolute uncertainty}}{\text{measurement}} \times 100\,\%$$

percentage uncertainty the uncertainty of measurements completed during an experiment expressed as a percentage of the final value

WORKED EXAMPLE 10.2B

Calculate the percentage uncertainty of 1.35 ± 0.06 g.

SOLUTION

Absolute uncertainty = 0.06 g

Measurement value = 1.35 g

$$\text{Percentage uncertainty} = \frac{0.06\text{g}}{1.35\text{g}} \times 100\%$$
$$= 3.33\%$$

Combining uncertainties

When completing calculations involving two or more quantities, the uncertainties of each measurement must be combined. There are some simple rules that can be used to calculate the final uncertainty of the products of the reaction.

1 If the measured values are added or subtracted, the uncertainty ($\pm a$ and $\pm b$) can be calculated from the equation error = $\sqrt{a^2 + b^2}$. For example:

$5.2 \pm 0.2\text{ g} + 3.2 \pm 0.2\text{ g} = 8.4 \pm \text{error}$

$\text{Error} = \sqrt{0.2^2 + 0.2^2} = \pm 0.3\text{ g}$

2 If the measured values are multiplied or divided by a number (with no uncertainty), the uncertainties must also be multiplied. For example:

$3 \times (5.2 \pm 0.2\text{ g}) = 15.6 \pm 0.6\text{ g}$

3 If the measured value with a percentage uncertainty is multiplied or divided by a number (with no uncertainty), the percentage uncertainty will not change. For example:

$3 \times (5.2 \pm 2\%) = 15.6 \pm 2\%$

4 If the measured value is multiplied or divided by a value with a percentage uncertainty, the uncertainties must be added together. For example:

$3\text{ g} + 4\% \times 5.2\text{ g} + 2\% = 15.6 + 6\%$

Percentage error

percentage error the percentage difference between the measured experimental value and the known true value

Percentage error, like percentage uncertainty is a way of expressing the accuracy of an experiment. The percentage error of an experiment is the percentage difference between the measured experimental value and the known true value. Percentage uncertainties are a way of expressing the level of accuracy of measurements made during the experiment, whereas percentage errors describe the accuracy of the final result.

Calculating percentage error

Study tip

Percentage error is how wrong you are. Percentage uncertainty is how confident you are.

When calculating the percentage error of an experiment, it is important to have an accepted true value that is expected. In chemical analysis, it is common for the percentage error to be a negative value. This is because the percentage errors compare the expected yield of a chemical reaction to the expected yield. If the measured experimental yield is larger than the theoretical yield, then the percentage error will be a negative value. Sometimes this negative is sign removed because of mathematical conventions. Removal of positive or negative signs to provide the absolute value is represented by the two horizontal lines in the formula.

The formula used to calculate the percentage error is:

$$\text{Percentage error} = \frac{|\text{measured value} - \text{true value}|}{\text{true value}} \times 100\%$$

WORKED EXAMPLE 10.2C

A student measured the volume and mass of an aluminium cube, and calculated the density of the aluminium to be 2.67 g cm^{-3}. The student researched the density of aluminium and determined the true density value to be 2.70 g cm^{-3}. Calculate the percentage error of the measured density.

SOLUTION

Experimental density of aluminium = 2.67 g cm^{-3}

Accepted true density of aluminium = 2.70 g cm^{-3}

$$\text{Percentage error} = \frac{|2.67 - 2.70|}{2.70} \times 100\%$$
$$= 1.11\%$$

CHECK YOUR LEARNING 10.2

Describe and explain

1 **Describe** the terms 'uncertainty', 'consistent' and 'discrepant'. Use an example to illustrate your answer.

2 **Identify** which measurements are consistent with the rulers in Figure 3.

Ruler A: 4 ± 0.1 cm, 4.3 ± 0.1 cm, 4.00 ± 0.05 cm, 4.3 ± 0.05 cm

Ruler B: 4 ± 0.1 cm, 4.3 ± 0.1 cm, 4.30 ± 0.05 cm, 4.3 ± 0.05 cm

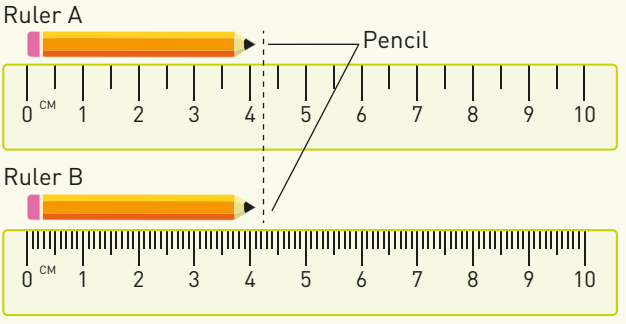

FIGURE 3 Measuring pencil lengths with a ruler

3 **Identify** the range of values suggested by the measurement 10.3 ± 0.8 seconds.

Apply, analyse and interpret

4 **Contrast** percentage uncertainty and percentage error in an experimental result.

5 A student measures the volume of a liquid in a volumetric flask and a measuring cylinder as 16.3 ± 0.3 mL and 15.9 ± 0.2 mL.

- **a** **Determine** the range of values for each container.
- **b** **Classify** the measurements as consistent or discrepant. Use evidence to support your answer.

6 A student combines two liquid reactants they have measured in a measuring cylinder. Reactant A has a volume of 15.3 ± 0.5 mL and reactant B has a volume of 18.5 ± 0.5 mL.

- **a** **Determine** the combined volume of the reactants.
- **b** **Determine** the absolute uncertainty of the combined volume of reactants.
- **c** Use your answer from part b to **determine** the possible range of the combined volume of reactants.

7 A student weighs themselves on electronic scales and reads their weight as 70.5 kg.

- **a** Assuming the scale is accurate, **determine** the absolute uncertainty for the student's weight.
- **b** **Determine** the percentage uncertainty for the student's weight.

8 A student wants to determine the volume of a rectangular container. They use a ruler with an absolute uncertainty of ±0.2 cm to obtain the following measurements: Length 4.5 cm, Width 2.3 cm, Height 1.0 cm.

- **a** **Analyse** the information to determine the percentage uncertainty of each measurement.
- **b** To calculate the volume, the student multiplies the length × width × height of the container. **Determine** the volume.
- **c** **Determine** the percentage uncertainty of the volume of the container.
- **d** If the accepted true volume of the container is 10.7 cm^3, **determine** the percentage error of the experiment.

Investigate, evaluate and communicate

9 **Design** a poster that summarises how to calculate the uncertainties and errors covered in this section. Use different colours and styles of writing to communicate the key concepts.

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- » Weblink
  Percentage error
- » Weblink
  Absolute uncertainty
- » Weblink
  Impacts of uncertainty

10.3 Constructing and using graphs

KEY IDEAS

In this section, you will learn about:

- constructing graphs and using them to communicate, solve problems and make predictions
- interpreting the relationship between the independent and dependent variables on graphs
- drawing and interpreting lines or curves of best fit through data points
- evaluating the shapes of graphs.

An important part of the scientific process is to present data (quantitative or qualitative) as part of communicating results. This process requires a scientist to write up a formal report and present the data in a form that communicates the results effectively. One way of doing this is to draw a visual representation (a graph) of the qualitative data that shows the relationship between the independent variable and the dependent variable.

Graphing features

All scientific graphs have features in common (Figure 1).

- The information on the graph should be easily seen. This includes the heading, axis titles and scales.
- The title should be a descriptive statement that includes the independent variable and the dependent variable.
- The independent variable should be on the horizontal axis and the dependent variable should be on the vertical axis.
- Each axis should start at zero.
- Each axis should be clearly labelled and include the unit measurement.
- The data should not be plotted beyond the end of the axis.
- A pencilled point or circle should be used for plotting each point.
- If there are two sets of data on a single graph, two different symbols should be used.

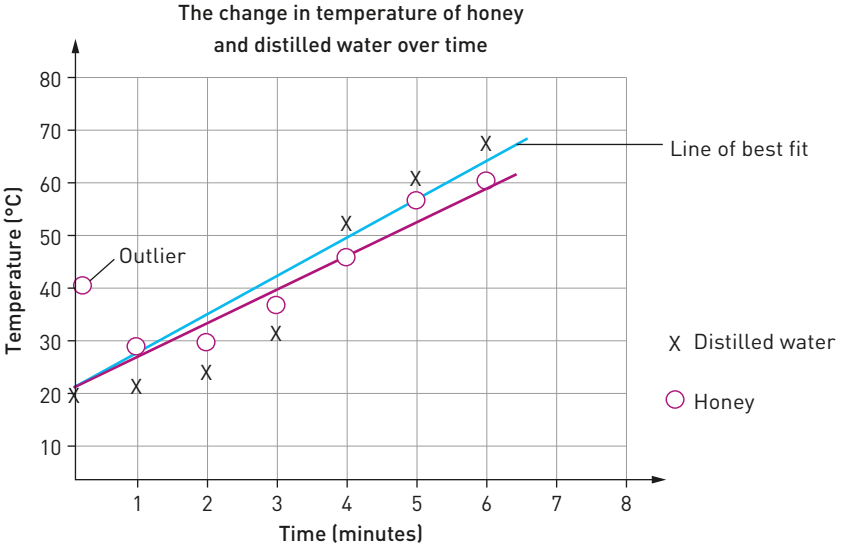

FIGURE 1 All scientific graphs have common features.

Scatterplots

A **scatterplot** (Figure 2) is one of the most common graphs used in chemistry. A scatterplot is a series of data points that are individually plotted on a horizontal and vertical axis. This provides a visual representation of how the independent variable on the horizontal axis affects the dependent variable on the vertical axis. If the absolute uncertainties of each plotted data point are known, then they can also be plotted by adding a small vertical **error bar** to the dot (Figure 2). For example, if the data to be plotted is 2.0 ± 0.5, the main data point should be plotted at 2, and the error bar should link 1.5 to 2.5.

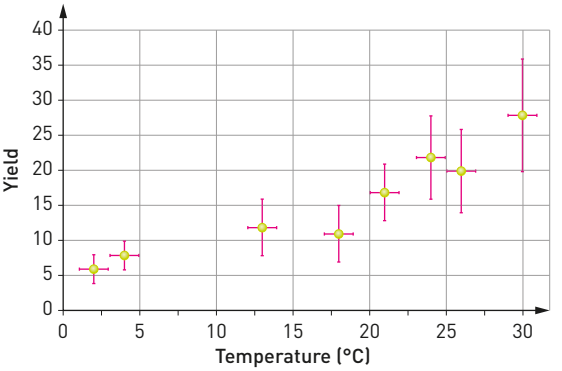

FIGURE 2 The data presented on a scatterplots can show the absolute uncertainty of each measurement with vertical error bars.

scatterplot
a series of data points that are individually plotted on a horizontal and vertical axis

error bar
a small line drawn on a plotted data point to show the uncertainty of the measured value

Correlation

If a change in the independent variable is linked to a change in the dependent variable, then it is described as a **correlation**. This does not mean that the changing independent variable causes the change in the dependent variable – that is **causation**. Causation requires a higher degree of evidence than correlation.

correlation
a link between a change in the independent variable and a change in the dependent variable; this does not mean that the changing independent variable caused the change in the dependent variable

causation
when a change in a single variable effects a change in a second variable

Line of best fit

Sometimes a graph is needed to predict a value between plotted points. A **line of best fit** or trend line (Figure 1) is a line drawn through the middle of the plotted data so that the points are evenly spread on either side of the line. It does not connect all the data points on a graph. If completing this by hand, a transparent plastic ruler allows all the points to be seen. Many computer graphs complete this automatically. The line of best fit can also be extended to predict what would happen if the experiment was extended.

If the line of best fit is a straight line, then the relationship is described as **linear**. If it is a curved line where one variable doubles with every measurement, then the graph is often described as **exponential**.

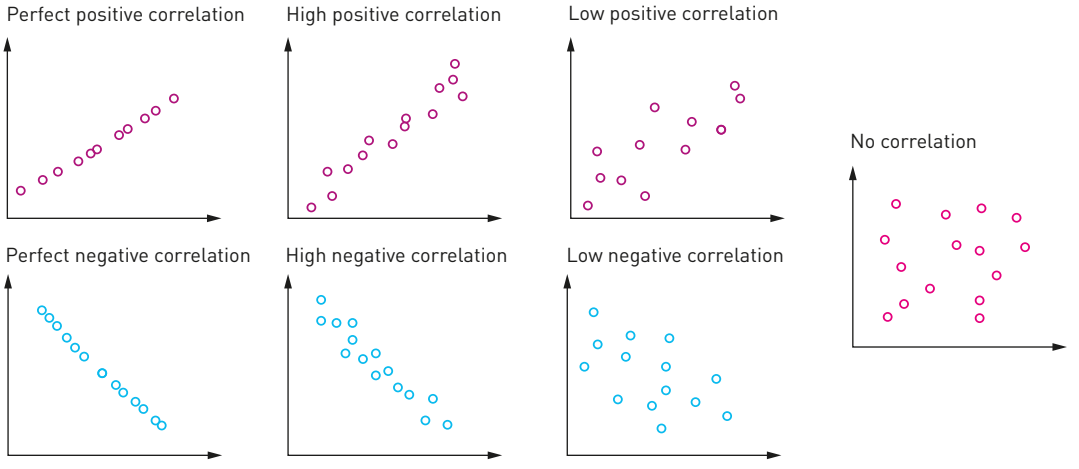

FIGURE 3 There are many different types of correlation that can be used to describe a graph. If is also important to identify when there is no correlation.

line of best fit
a line drawn on a graph joining as many points as possible and showing the general direction of the data; should be drawn with, approximately, an equal number of points above and below the line

linear
a relationship where the *x*-values change proportionally with the *y*-values, leading to a straight line

exponential
a graphed curved line where the dependent variable increases at a greater rate with every measurement

WORKED EXAMPLE 10.3

A student measured the amount of product produced from their chemical reaction and plotted the data on the graph in Figure 4.

a Draw a line of best fit for this data.

b Calculate the amount of product that would have been available at 5 minutes and 30 seconds.

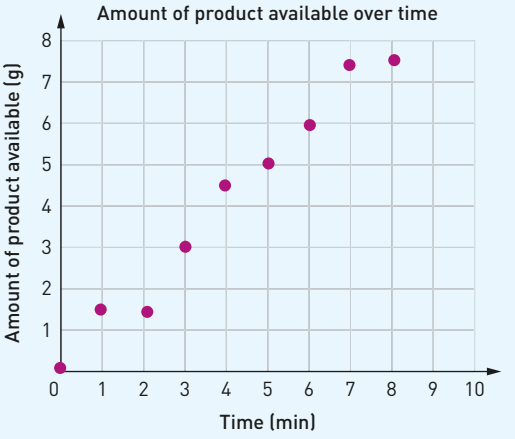

FIGURE 4 Product produced by a chemical reaction over time.

SOLUTION

a The line of best fit goes through as many data points as possible and minimises the total distance of points from the line.

b Locate 5.5 minutes (5 min 30s) on the horizontal axis and draw a vertical line to the line of best fit. Calculate the amount of product on the vertical axis that corresponds to this location.

Amount of product at 5 minutes and 30 seconds = 5.5 g

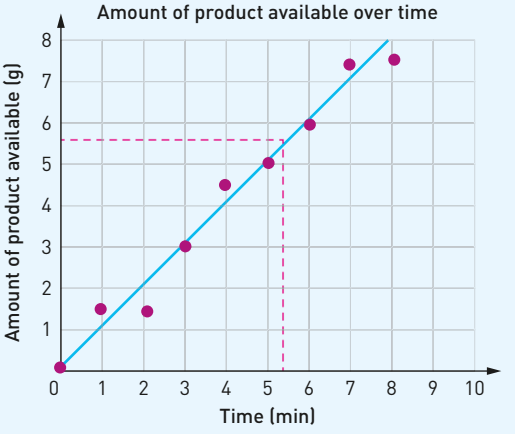

FIGURE 5 Product produced by a chemical reaction over time.

Bar and column graphs

Qualitative data is usually represented by a bar or column graph (Figure 6). The independent variable is usually **discrete** or separated from each other. Examples include colours or numbers of molecules. As it is impossible to add half a molecule to a chemical reaction, this data is described as discrete.

discrete
individual or distinct data; usually used for qualitative data on column or bar graphs

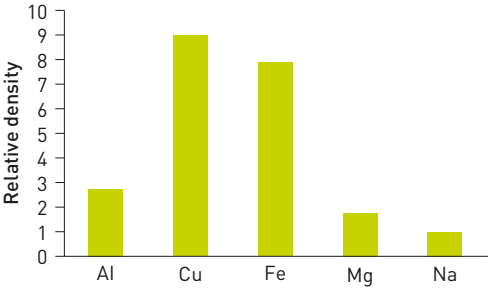

FIGURE 6 A column graph must be used to illustrate this data on elements because it is impossible to have data between the individual elements.

Outliers

It can be very difficult to control every part of an experiment. Errors can result in some data that is very different from the rest of the data. A data point that is outside the rest of the data set is called an **outlier** (Figure 1). This abnormal data point may be caused by mistakes made by the experimenter or equipment during the experiment. Outliers should always be plotted on a graph; however, they may not be included in the line of best fit. Experimenters should attempt to explain the cause of the unexpected data.

outlier
a data point outside the rest of the data set

Interpreting graphs

The ability to describe a graph allows a scientist to describe the relationship between the independent variable and the dependent variable. The description of the graph should include the:

- independent and dependent variables
- type of correlation shown by the graph (Figure 3)
- shape of the graph, i.e. linear or exponential.

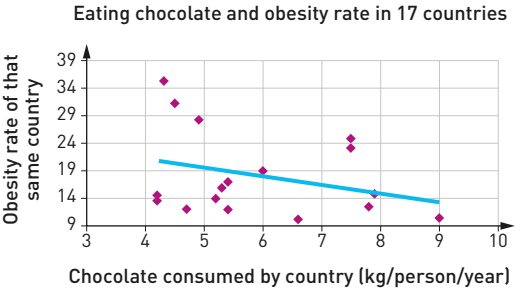

FIGURE 7 Interpreting graphs: the data in this graph shows that there is a weak negative correlation between the amount of chocolate consumed per person and the obesity levels in their country.

When interpreting the data shown on a graph, it is important to remember that correlation does not imply causation. An example of this is shown in Figure 7 where there is a weak negative linear correlation between the amount of chocolate consumed per person and the obesity levels in their country. This is not the same as claiming (as some newspapers did at the time) that eating more chocolate causes the people in the country to become thinner.

Study tip

After describing the relationship between the independent variable and the dependent variable that is shown on a graph, consider how this applied in the experiment. Are the results relevant to the actions performed by the experimenter?

CHECK YOUR LEARNING 10.3

Describe and explain

1 **Define:**
   - **a** causation
   - **b** correlation
   - **c** outlier

2 **Describe** a situation when a bar or column graph should be used.

3 **Explain** why a scientist should not ignore an outlier when graphing or discussing their results.

Apply, analyse and interpret

4 A chemistry class sets up two test tubes, each containing identical amounts of H_2O_2 and MnO_2. A chemical reaction occurs that generates oxygen gas according to the equation:

$$2H_2O_2 + MnO_2 \rightarrow 2H_2O + Mn + 2O_2$$

After one minute, the students measure the amount of oxygen generated in each test tube.

- **a** **Determine** the volume of oxygen produced in the reaction of test tube 1 at 3 minutes and 30 seconds.
- **b** **Construct** a single graph for the results of both test tubes. Draw a line of best fit for each test tube.
- **c** **Identify** any outliers and circle them on the graph.
- **d** The gradient, or slope of the line of best fit, is an indication of the rate of the reaction. **Interpret** the data to **determine** which reaction was producing oxygen at the fastest rate.

TABLE 1 Oxygen produced from hydrogen peroxide reacting with manganese oxide.

Test tube	Amount of oxygen produced (mL) at					
	1 min	2 min	3 min	4 min	5 min	6 min
1	1.4 ± 0.2	2.6 ± 0.2	3.5 ± 0.2	4.2 ± 0.2	4.2 ± 0.2	5.1 ± 0.2
2	1.6 ± 0.2	3.2 ± 0.2	3.9 ± 0.2	3.9 ± 0.2	4.8 ± 0.2	5.9 ± 0.2

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10.4 Random and systematic errors

KEY IDEAS

In this section, you will learn about:

- the difference between random errors and systematic errors
- the systematic errors in experimental design and procedures
- how repeating measurement reduces random errors but not systematic errors
- analysing the impact of random errors and systematic errors in an experiment.

systematic error
a consistent repeatable error that occurs every time a piece of equipment is used

measurement bias
a type of systematic error that occurs when equipment repeatedly makes the same miscalculation

sample selection bias
a type of systematic error where a small sample is taken for testing from a larger amount of substance

Whenever a scientist conducts an experiment, errors will occur. An error is different from a mistake. Mistakes are blunders, such as measuring the wrong amount of a reactant, or not following the procedure correctly. Mistakes should never appear in a scientific report because the experimenter should detect the blunder and repeat the experiment.

There are two main types of errors that can be described in an experimental report: systematic errors and random errors.

Systematic errors

A **systematic error** (or bias) is a consistent repeatable error that occurs every time a piece of equipment is used. This error always results in the measured value changing in the same direction (either more or less). For example, an electronic scale may make a systematic error by measuring every mass as slightly less than the true value. This may be due to the electronic scale being incorrectly calibrated or being used inaccurately. Other equipment, such as a measuring tape, may become stretched over time and consistently measure 102% of the true length. This means the longer the measurement, the greater the measured error.

The size of a systematic error is usually identical because the equipment keeps making the same miscalculation. This type of systematic bias is called **measurement bias**. Measurement bias can be reduced by choosing the equipment carefully, and calibrating the equipment before it is used.

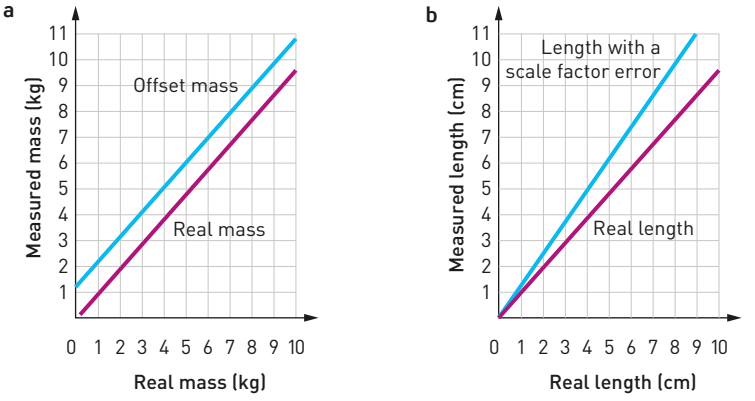

FIGURE 1(a) A systematic error occurs when a set of scales regularly measures the mass of substances above or below the true value. (b) This systematic error is caused by the equipment increasing or decreasing the value by a set proportion; for example, a stretched measuring tape.

Another systematic error is **sample selection bias**, where a small sample is taken for testing from a larger amount of substance. The small sample may not be a true representation of the larger sample if the larger sample is not homogeneous. A 1 mL sample taken from the top surface of a 100 mL solution may be different from a sample taken from the bottom of the solution. Sample selection bias can be reduced by ensuring the solution is thoroughly mixed and selecting samples from random locations in the container

Systematic errors cannot be decreased by repeating measurements or experiments that use the same equipment (Figure 1).

Random error

Random errors are unpredictable and have no detectable pattern. They often occur with measurements or transfers of substances such as when pouring a liquid out of a measuring cylinder, the very small amount of liquid left behind is unpredictable. Although this random error can be expected to some extent, the measured volume may be too large or too small and is unavoidable.

The randomness of the errors, sometimes too high and sometimes too low, means that repeating a measurement many times results in an equal number of misses in each direction. This means repeating a measurement and determining an average improves the accuracy of the value (Figure 2).

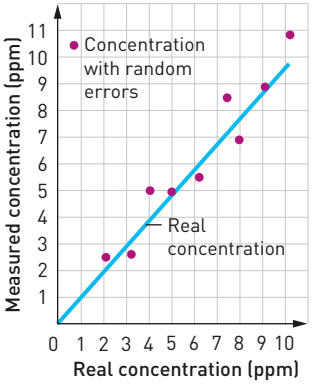

FIGURE 2 Random errors cannot be predicted. Repeating the measurement and determining the average will decrease the impact of the error.

random error
an unpredictable error in measurement or experimental procedure that has no detectable pattern

Study tip

Random errors can be reduced by repeating the measurement. Systematic errors will not be reduced by repeating the measurement.

CHECK YOUR LEARNING 10.4

Describe and explain

1 **Define** and describe examples of the two types of systematic errors.
2 **Explain** random errors. Provide an example.
3 **Identify** the following as random errors or systematic errors.
   - **a** There is a consistent draft pushing down on the electronic balance.
   - **b** There is an air bubble in the liquid of a thermometer.
   - **c** A student takes measurements from the top of the meniscus rather than the bottom.
   - **d** The scale consistently measures 5 units above the true value.
   - **e** A student miscounts the number of drops in the third sample tested.
   - **f** A student does not measure the meniscus at eye level.
   - **g** A student uses a buffer to calibrate a pH meter to 7.0. After completing the experiment, the student discovers the pH buffer was 6.7.

Apply, analyse and interpret

4 An electronic balance is not calibrated effectively so that it consistently measures below the true value. **Identify** an error that would occur and describe how it would affect the results of an experiment.
5 **Consider** how systematic errors could be minimised in an experiment that measures the amount of product produced in a chemical reaction.

Investigate, evaluate and communicate

6 **Propose** a way to minimise the error in an electronic balance caused by the vibrations of people walking past. **Identify** the type of error this is.

Check your obook assess for these additional resources and more:

- » Student book questions
  Check your learning 10.4
- » Weblink
  Bias
- » Weblink
  Systematic error
- » Weblink
  Random error

10.5 Significant figures

KEY IDEAS

In this section, you will learn about:

- significant figures in data
- using the correct number of significant figures in calculations.

Some numbers used in calculations have an impossibly large number of digits. For example, the number pi (π) has an infinite number of decimal places. While it is mathematically interesting, it would be impossible to measure a π amount of a substance. Measuring 3.141 592 6535 … grams of a substance is impossible if the electronic balance is only able to measure 3.142 g. For this reason, the number of figures that can be used in calculations is restricted by the limitations of the measuring equipment.

significant figures the number of numerical figures that can be used to express a measured or calculated quantity pattern

Significant figures are the number of numerical figures that can be used to express a measured or calculated quantity. In this case, the word 'significant' means numbers that have meaning.

In the number 640 724, the most significant figure is 6 because it communicates that the number is slightly more than 600 000 (but less than 700 000). The next most significant figure is the 4 because it communicates that the number is more than 640 000 but less than 650 000. The third figure is 0 because it acts as a place holder so that the 7 is in the correct place.

With much smaller numbers, such as 0.008 507, the 8 is the most significant figure. This is because the 8 communicates that the number is larger than eight thousandths and less than nine thousandths. The next most significant figure is the 5.

Significant figures rules

There are five rules to follow when determining significant figures.

1. All non-zero digits are significant. This means the numbers 1, 2, 3, 4, 5, 6, 7, 8 and 9 are all significant figures.
2. Zeros between non-zero digits are significant. This means 31.407 has five significant figures because the 0 between the 4 and 7 is holding the place and is therefore also significant.
3. Leading zeros to the left of the first non-zero digit are not significant. This means 0.096 only has two significant figures because the zeros at the front do not change the precision when the number is written in scientific notation. For example, 0.096 has the same accuracy as 9.6×10^{-2}. It would not be more precise to write 009.6×10^{-2}.
4. In a number with a decimal point, the trailing zeros to the right are significant. For example, 450.0 can be written as 4.500×10^{2}. Keeping the trailing zeros as part of the four significant figures indicates that the measurement is precise to that number of decimal places.
5. If a number does not have a decimal point, the trailing zeros may or may not be significant. This is because the number is holding a place and may have already been rounded off and the trailing zeros may just be place holders. For example, 250 000 may have six significant figures. Alternatively, if the number is 245 678 and has been rounded up to 250 000, then it is considered to only have two significant figures.
   So the number 0.009 876 543 21 correct to five significant figures is 0.0098765.

Study tip

Remember that when rounding, if the next number is 5 or more, round up. If the next number is 4 or less, round down.

Calculations with significant figures

Occasionally, we need to make calculations with different numbers of significant figures. This may be a result of an electronic balance being able to measure a mass with four significant figures (e.g. 3.642 g) and a measuring flask only measuring to two significant figures (e.g. 4.5 mL).

When adding or subtracting numbers, the calculated result should have the same number of figures after the decimal point as the number with the least significant figures. For example, if 9.4 (one significant figure after the decimal) was added to 5.435 (three significant figures after the decimal), the result should only have one significant figure after the decimal (14.8).

When multiplying or dividing two numbers, the calculated result should have the same number of significant figures as the number with the fewest significant figures.

To prevent errors, the number of significant figures is only determined at the end of the calculation.

Study tip

When calculating significant figures, 23 is not the same as 23.0. 23.0 is ten times more precise than 23. While they have the same value, 23.0 is more trustworthy and has less uncertainty.

WORKED EXAMPLE 10.5

Determine the density of a solution with a mass of 3.640 g and a volume of 4.50 mL.

SOLUTION

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$
$$= \frac{3.640\ \text{g}}{4.50\ \text{mL}}$$
$$= 0.808\,889\ \text{g mL}^{-1}$$

The volume only has three significant figures so the final answer should have three significant figures.

Density = 0.809 g mL^{-1}

CHECK YOUR LEARNING 10.5

Describe and explain

1 **Explain** what is meant by a significant figure.

2 **Explain** why significant figures are important in chemical calculations.

3 **Identify** the number of significant figures in:

a 0.4507 **c** 16×10^2

b 123.00 **d** 5.7×10^{-3}

Apply, analyse and interpret

4 **Apply** your knowledge of significant figures to rewrite the following numbers correct to two significant figures.

a 3.45 **c** 305.01

b 357.03 **d** 6.02×10^{23}

5 **Apply** your knowledge of significant figures to determine the following to the correct number of significant figures.

a 0.0426 – 0.002 **b** 2.45 × 3.0

6 A student calculated the density of an object with a mass of 4.525 g and a volume of 3.0 mL correct to one significant figure. The answer was 1.508 g mL^{-1}. **Identify** errors in this answer. Include the correct answer in your explanation.

Check your obook assess for these additional resources and more:

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Check your learning 10.5

» Increase your knowledge
Significant figures

» Increase your knowledge
Calculations with significant figures

» Weblink
Significant figures

CHAPTER 10

Review

Chapter summary

10.1
- Qualitative data is a non-numerical description of a substance or reaction. Quantitative data is a series of numbers that provide a guideline of the physical properties of a substance.
- A measurement is accurate when it is close to the true or accepted value.
- Precision is a description of the consistency of the measurements. To be precise, the series of measurements must be similar values to each other.
- Calibration involves checking and adjusting measured values so that they become more accurate.

10.2
- An uncertainty is a description of confidence in the accuracy and precision of a measurement. It is typically expressed as a number; for example, best estimate ± absolute uncertainty.
- Measurements that have overlapping uncertainty ranges are said to be consistent. If there is no overlap between the two measurements' uncertainties, they are discrepant.
- Percentage uncertainties are a way of expressing the level of accuracy of measurements made during the experiment.
- The percentage error of an experiment is the percentage difference between the final measured experimental value and the known true value.

10.3
- Graphing data is a way of visually communicating the information.
- Scatterplots are graphs that have two data points plotted on the same set of axes to reveal any correlation.
- A line of best fit is a trend line drawn through as many data points (or their error bars) as possible. It shows the general trend of the data and allows the prediction of values between the data points.
- A column or bar graph is a way of representing discrete data points.
- If two variables are correlated, this does not mean one variable causes a change in the second variable.
- Outliers are values that fall outside the expected set of data values. Outliers should be graphed but may not be included in the line of best fit.

10.4
- A systematic error (or bias) is a consistent repeatable error that occurs every time a piece of equipment is used. This error always results in the measured value moving in the same direction. A systematic error can be improved by calibrating equipment and not by repeating a measurement.
- Measurement bias can be reduced by choosing the equipment carefully, and calibrating the equipment before it is used.
- Sample selection bias occurs when a small sample is not a true representation of the larger sample. This can be reduced by ensuring the solution is thoroughly mixed and selecting samples from random locations in the container.
- Random errors are unpredictable errors that occur in measurement or experimental procedure that have no detectable pattern. Random errors can be reduced by repeating measurements and determining an average.

10.5
- Significant figures are the number of numerical figures that can be used to express a measured or calculated quantity.
- All non-zero digits are significant. Zeros between non-zero digits are significant. Leading zeros to the left of the first non-zero digit are not significant.

- In a number with a decimal point, the trailing zeros to the right are significant.
- If a number does not have a decimal point, the trailing zeros may or may not be significant.
- When adding or subtracting numbers, the result should only include the same number of figures after the decimal point as is in the smallest of the initial numbers
- When multiplying or dividing two numbers, the calculated result should have the same number of significant figures as the number with the fewest significant figures.

Key terms

- accuracy
- calibration
- causation
- consistent
- correlation
- discrepant
- discrete
- error bar
- exponential
- line of best fit
- linear
- measurement bias
- outlier
- parallax error
- percentage error
- percentage uncertainty
- precision
- qualitative
- quantitative
- random error
- sample selection bias
- scatterplot
- significant figures
- systematic error
- uncertainty

Key formulas

Percentage uncertainty	$\text{Percentage uncertainty} = \frac{\text{absolute uncertainty}}{\text{measured value}} \times 100\ \%$
Percentage error	$\text{Percentage error} = \left\lvert\frac{\text{experimental value} - \text{true value}}{\text{true value}}\right\rvert \times 100\%$

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; *** = high.

Multiple choice

1 Students are using a poorly calibrated electronic balance to measure the mass of a watch glass. They repeat the measurement three times to be sure of the answer. Their measurements could be described as having:

A high accuracy and high precision
B high accuracy and low precision
C low accuracy and high precision
D low accuracy and low precision.

2 When a measurement is repeatable and consistent, it can be described as:

A high precision **C** high accuracy
B low precision **D** low accuracy.

3 How close a measurement is to its true value is called:

A reliability **C** precision
B accurate **D** validity.

4 Two students experimentally determined the concentration of a copper (II) sulfate solution. They repeated their experiment three times. Their results are shown in Table 1.

TABLE 1 Concentration of copper (II) sulfate solution

Student 1	Student 2
0.893	0.884
0.897	0.882
0.889	0.883

The true value of the copper (II) sulfate solution was 0.893 M. Which student(s) produced results showing a systematic error?

A Student 1 **C** Both students
B Student 2 **D** Neither student

5 A student adds 4.92 g of sodium chloride to 7.587 g of copper metal. Calculate the final total total mass of the two substances correct to two significant figures.

A 12 g **C** 12.50 g
B 12.5 g **D** 12.51 g

Short answer

Describe and explain

★ **6** **Describe** the correlation (if any) in the following graphs.

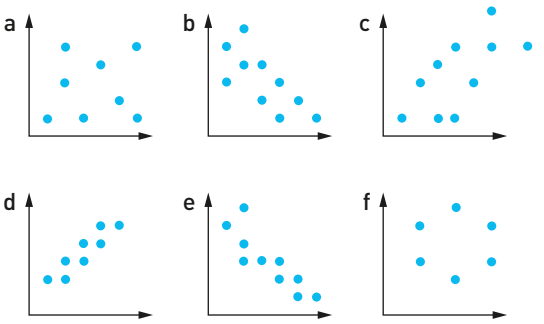

FIGURE 1 Different types of correlation graphs

★ **7** **Calculate** the following to the correct number of significant figures.

a 12.39 + 1.4 **c** 0.043 – 0.03
b 0.4329 ÷ 0.031 **d** 340 × 0.3402

★★ **8** Four students weigh a standard 10 g mass on the same balance. All four students achieve a measurement of 8.9 g for the mass of the standard. **Describe** the accuracy and precision of the measurement.

★★ **9** An experimenter decides to increase the sample size of their experiment. **Identify** the type of error are they trying to reduce. **Explain** your answer.

Apply, analyse and interpret

★ **10** **Consider** why an experimenter should never connect all the data points on a graph. What type of line should be drawn instead?

★★ **11** **Determine** the diameter of the bauble shown Figure 2. **Calculate** the absolute uncertainty of your measurement. Provide reasoning to support the uncertainty you have shown.

FIGURE 2 Measuring a bauble using a mm ruler

★★**12** A scientist used an electronic balance that consistently measured mass as 0.05 g above the true value. Identify an error that would result from using this electronic balance. **Consider** how the scientist could minimise this error in the future.

★★★**13** The graph in Figure 3 shows the radius of an element against its atomic number.

a **Interpret** the shape of the graph for the elements with atomic number 3 –10.

b **Interpret** the shape of the graph for the elements with atomic number 11–18.

c **Consider** what the graph suggests about the structure of different elements as you move across the periodic table.

Investigate, evaluate and communicate

★★**14** Four students obtained the sets of data shown below in Table 2 during a chemistry experiment. The predicted amount of product is 25.11 g. **Evaluate** the accuracy and precision of the data sets.

TABLE 2 Data from a student chemistry experiment

Student 1	Student 2	Student 3	Student 4
26.51 g	24.06 g	25.12 g	23.76 g
25.08 g	28.09 g	25.09 g	23.80 g
23.63 g	29.56 g	25.14 g	23.78 g

★★★**15** A student used a graduated cylinder to measure 11 mL of water and an electronic balance to determine the mass of the water as 11.9200 g. The student calculated that the density of the water is 1.0843 g mL^{-1}. **Decide** which of the rules of significant figures the student did not follow and **justify** your response. Provide the correct answer in your response.

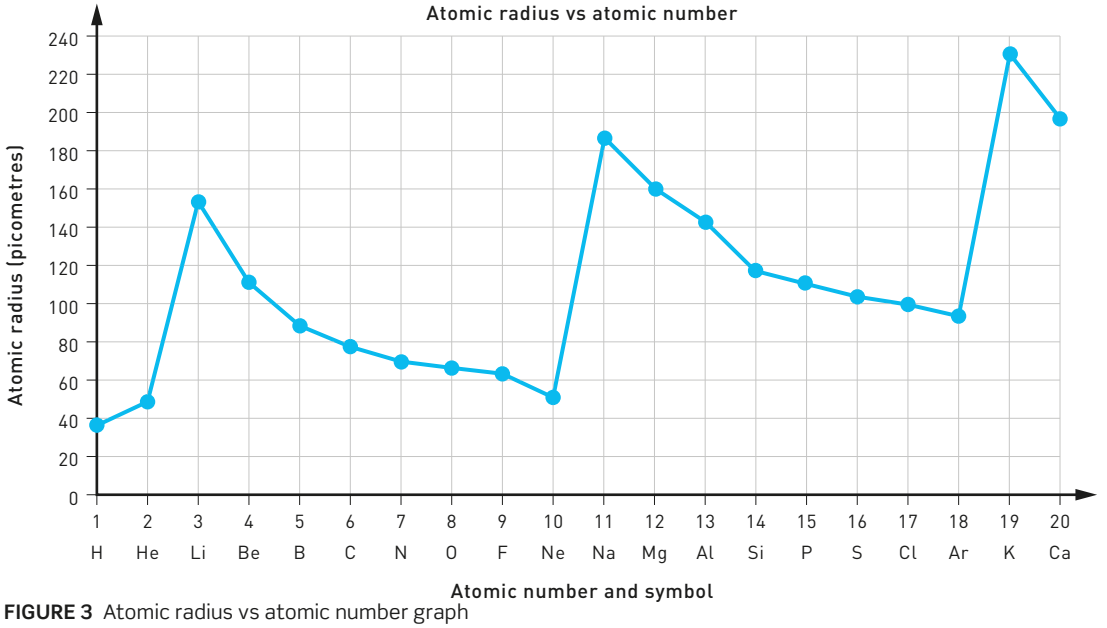

FIGURE 3 Atomic radius vs atomic number graph

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CHAPTER 11

Fuels

Modern society has a high dependency on fuels. We rely on energy to power cars, washing machines, refrigerators, traffic lights, and respirators in hospitals. The simple of act of charging a mobile phone relies on our ability to harness the energy from various fuel sources.

Burning fossil fuels such as coal and crude oil is a common method of generating power. However, these types of fuels can have negative impacts on our environment and may affect our quality of life in the future. There is also a need to look for alternative fuel sources, as fossil fuels are being consumed faster than they are being replenished.

What can be done to meet our energy needs? One solution is biofuels, which could be harnessed to produce the energy that society needs to operate.

OBJECTIVES

→ Compare fuels, including fossil fuels and biofuels, in terms of their energy output, and evaluate their suitability for purpose, and the nature of products of combustion.

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FIGURE 1 Biodiesel can be made from canola oil.

11.1 Fuels and products of combustion

KEY IDEAS

In this section, you will learn about:

- the production of energy in combustion reactions of fuels
- comparing fuels by their energy output
- how the combustion of fuels can result in a number of products, depending on the reactants, which can have varied effects.

fuel
a substance that reacts with oxygen to produce energy

energy
the ability to do work

Fuels are a crucial part of daily life. They are needed for a variety of activities, such as driving cars, cooking food and keeping homes cool.

All substances contain chemical energy. A fuel is a substance that reacts with oxygen to produce useful **energy**. Most fuels burn in a combustion reaction to release energy, so the reaction is considered exothermic. (Combustion reactions were covered in detail in Chapter 8.) Energy is mostly released as heat, but biofuels come from waste plant and animal products. Common examples of fuels are the petrol in cars, the gas in an oven, and the wood used to keep a fire burning.

Study tip

The combustion of fuels is always an exothermic reaction. This means the enthalpy change is negative and ΔH should have a negative sign.

Fuels require enough oxygen to react fully, in a complete combustion reaction. Fuels, such as petrol, that are made of hydrocarbons (compounds of hydrogen and carbon) react according to the following equation:

$$\text{hydrocarbon} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water}$$

In a complete combustion reaction, the carbon oxidises to become carbon dioxide and hydrogen oxidises to become water. For example, propane, which is used in bottle gas, completely combusts to form carbon dioxide and water:

$$\text{propane} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water}$$

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

However, not all fuels are made of hydrocarbons. For example, coal, which is a major fuel, is made up of mainly carbon. The complete combustion reaction of coal is written as:

$$\text{carbon} + \text{oxygen} \rightarrow \text{carbon dioxide}$$

$$C + O_2 \rightarrow CO_2$$

If a combustion reaction does not have enough oxygen, it will be an incomplete combustion. This means that carbon monoxide (CO) may be produced instead of carbon dioxide (CO_2). Carbon monoxide is a toxic gas, so it is a dangerous product of incomplete combustion.

FIGURE 1 The amount of usable energy released from petrol in a car is directly proportional to the amount of petrol consumed in the reaction.

Energy output

The amount of energy resulting from the combustion of a fuel is referred to as its energy output. The amount of energy produced from the fuel's combustion reaction depends on the amount of substance present in the chemical reaction. For example, a drop of petrol in a car does not produce the same amount of energy as a litre of the same type of petrol.

Fuel combustion involves energy transformation, which means energy is converted from one form into another useable form. The amount of energy

present does not change, because energy can neither be created nor destroyed. For example, burning coal transforms chemical energy into heat energy, but it does not create more energy.

The amount of energy produced by fuels can be measured, by using the international system of units (SI units), as **joules** (J). The multiple units for joules are outlined in Table 1.

TABLE 1 SI multiples of the joule (J)

Unit (symbol)	Equivalent number of joules
Nanojoule (nJ)	10^{-9}
Microjoule (μJ)	10^{-6}
Millijoule (mJ)	10^{-3}
Kilojoule (kJ)	10^{3}
Megajoule (MJ)	10^{6}
Gigajoule (GJ)	10^{9}
Terajoule (TJ)	10^{12}

joule
a unit of energy

Specific energy and energy density are measures of energy content.

- Specific energy refers to the amount of energy that is released from a certain mass of fuel (energy per unit mass).
- Energy density refers to the amount of energy that is released from a certain volume of fuel (energy per unit volume).

WORKED EXAMPLE 11.1A

A calorimetry experiment was conducted to investigate the energy released by two fuels, methanol (CH_3OH) and ethanol (CH_3CH_2OH). The fuels were completely combusted with oxygen, producing carbon dioxide and water. Table 2 shows the mass of each fuel used and the heat released during complete combustion.

TABLE 2 Thermochemical data for the complete combustion of methanol and ethanol

Fuel	Mass of fuel used (g)	Energy released (kJ)
Methanol	1.79	36.6
Ethanol	1.86	50.2

a Construct balanced equations for the complete combustion of each fuel.

b Use the thermochemical data to calculate the specific energy (kJ g^{-1}) for each fuel.

c Calculate the molar heat of combustion (kJ mol^{-1}) for each fuel.

SOLUTION

a Complete combustion of methanol:

$$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(g)$$

Complete combustion of ethanol:

$$CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$

b 1.79 g of methanol released 36.6 kJ of energy

Specific energy of methanol:

$$\frac{36.6\text{ kJ}}{1.79\text{ g}} = 20.4\text{ kJ g}^{-1}$$

1.86 g of ethanol released 50.2 kJ of energy

Specific energy of ethanol:

$$\frac{50.2\text{ kJ}}{1.86\text{ g}} = 27.0\text{ kJ g}^{-1}$$

c Molar mass of methanol = 32.04 g mol^{-1}

Molar heat of combustion of methanol:

$$\frac{36.6\text{ kJ}}{1.79\text{ g}} \times 32.04\text{ g mol}^{-1} = 655\text{ kJ mol}^{-1}$$

ΔH (methanol) = −655 kJ mol^{-1}

Note: ΔH is a negative value because the reaction is exothermic.

Molar mass of ethanol = 46.07 g mol^{-1}

Molar heat of combustion of ethanol:

$$\frac{50.2\text{ kJ}}{1.86\text{ g}} \times 46.07\text{ g mol}^{-1} = 1243\text{ kJ mol}^{-1}$$

ΔH (ethanol) = −1243 kJ mol^{-1}

Note: ΔH is a negative value because the reaction is exothermic.

WORKED EXAMPLE 11.1B

Given that ΔH (octanol) = −5281 kJ mol^{-1}, calculate how much energy is released by the complete combustion of 32.57 kg of octanol.

SOLUTION

Molar mass (octanol) = 130.23 g mol^{-1}

Mass of octanol = 32.57 kg

= 32 570 g

$$\text{Number of mol (octanol)} = \frac{32570\text{ g}}{130.23\text{ g mol}^{-1}} = 250.1\text{ mol}$$

$$\text{Energy released} = \frac{32570\text{ g}}{130.23\text{ g mol}^{-1}} \times 5281\text{ kJ mol}^{-1}$$

$= 1.321 \times 10^{6}$ kJ

= 1.321 GJ

Products of combustion

The heat energy produced from combustion reactions can be harnessed by society; other products of combustion are not as useful. For example, carbon dioxide is a product of burning coal. Although carbon dioxide is not harmful in small quantities, it is damaging to the environment in larger quantities.

Common products of combustion include the following.

- Carbon dioxide (CO_2) is produced when carbon in fuel is burnt completely. Carbon makes up most of the mass of fossil fuels, so CO_2 is the most common product of combustion.
- Carbon monoxide (CO) is produced when there is not enough oxygen and the carbon in fuels is not burnt completely. CO is a colourless, odourless gas that is also toxic. It is a product of incomplete combustion.
- Sulfur dioxide (SO_2) is produced when sulfur in fuels is burnt. SO_2 dissolves in water vapour to form a harmful acid (including acid rain) or combines with other gases to form other harmful products.
- Nitrogen oxides (NO_x) are produced when fuels are burnt at very high temperatures. They contain variable amounts of nitrogen and oxygen. Nitrogen oxides are highly reactive gases that are usually colourless and odourless. Nitrogen dioxide, often referred to as smog, has a brown colour and is often seen over cities.
- Water (H_2O) is produced when a hydrocarbon fuel is burnt.
- Particulates are small solids, usually carbon, that result from incomplete combustion. They are released into the air during combustion and often stick to other solid materials.

Products of combustion and climate change

The ability of Earth's atmosphere to retain heat is dramatically affected by the concentrations of carbon dioxide and other polluting emissions released into the atmosphere. A significant portion of these emissions is a result of society's use of fossil fuels to power homes, transport, factories and other industrial activities.

Figure 2 shows how an excess of carbon dioxide produced by burning fuels can affect Earth's atmosphere. Light energy from the Sun is naturally absorbed by Earth's surface, and some heat energy is trapped by the atmosphere. An increase in carbon dioxide concentration in the atmosphere traps more solar radiation, affecting Earth's climate.

FIGURE 2 Excessive polluting emissions produced by burning fossil fuels create a thick layer of greenhouse gases, which trap solar radiation and affect Earth's climate.

CHECK YOUR LEARNING 11.1

Describe and explain

1 **Describe** the difference between complete and incomplete hydrocarbon combustion.

2 **Identify** three fuels that are used in day-to-day life.

Apply, analyse and interpret

3 **Determine** the energy released (kJ) by the complete combustion of 15.02 g of propan-1-ol ($CH_3CH_2CH_2OH$). (ΔH (propan-1-ol) = –2011 kJ mol^{-1})

4 **Consider** the impact of a fuel that produces excessive amounts of:

- **a** carbon dioxide
- **b** carbon monoxide
- **c** particulates.

Investigate, evaluate and communicate

5 **Investigate** the historical levels of CO_2 in Earth's atmosphere. Is it possible to link these levels with burning fossil fuels? **Explain** your answer.

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- » Weblink
  Methane

11.2 Fossil fuels

KEY IDEAS

In this section, you will learn about:

- the use of fossil fuels, including coal and crude oil
- fossil fuels and their suitability for purpose.

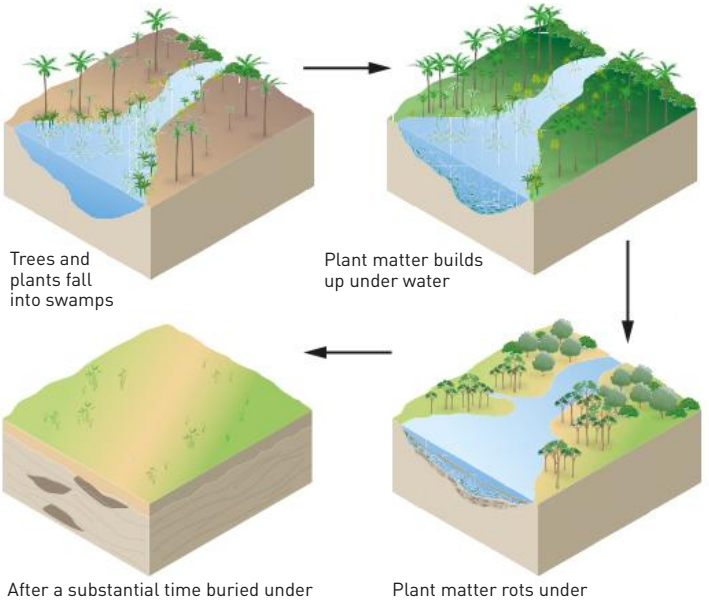


FIGURE 1 The formation of coal, a fossil fuel, occurs over millions of years.

Fossil fuels are sourced from the remains of buried organic matter, such as animals and plants, that has decayed over millions of years. During this time, the heat and pressure of the Earth's crust converts the organic matter into combustible materials (such as coal). This natural process is outlined in Figure 1.

Fossil fuels are being used faster than they can be produced; therefore, they are known as non-renewable resources or, more accurately, **long-term renewable resources**. The term 'long-term renewable' is more appropriate because these resources could be replaced eventually, but it would take millions of years.

However, because the process forming fossil fuels takes so long, there is currently a limited supply of fuels available such as coal, crude oil and natural gas.

The burning of fossil fuels releases a lot of carbon dioxide into the atmosphere. The extraction of fossil fuels can also be damaging to the natural landscape. As a result, there are a number of environmental issues to consider with harnessing fossil fuels.

fossil fuel
a fuel formed from the remains of ancient organisms

long-term renewable resource
a fuel that is only renewable on a long-term timescale, e.g. coal and oil

Coal

Coal is made up of mainly two elements: carbon and hydrogen. It can also include small amounts of other elements in variable quantities, such as sulfur, oxygen and nitrogen, depending on the grade (quality) of the coal.

Coal is comprised of terrestrial (land-based) plant and animal material, typically from old swamps.

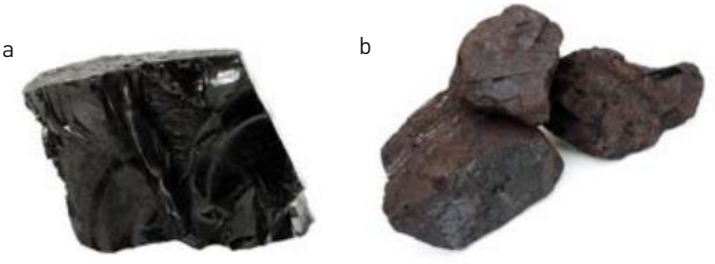


FIGURE 2 (a) Anthracite is the highest quality coal. (b) Brown coal , or lignite, has a lower energy content and is worse for the evironment if burned.

Coal that has formed over a longer period has less moisture and a higher carbon content, making it higher quality.

The different categories of coal are:

- lignite – brown coal, crumbly
- bituminous coal – black coal, hard
- anthracite – hard coal, highest quality.

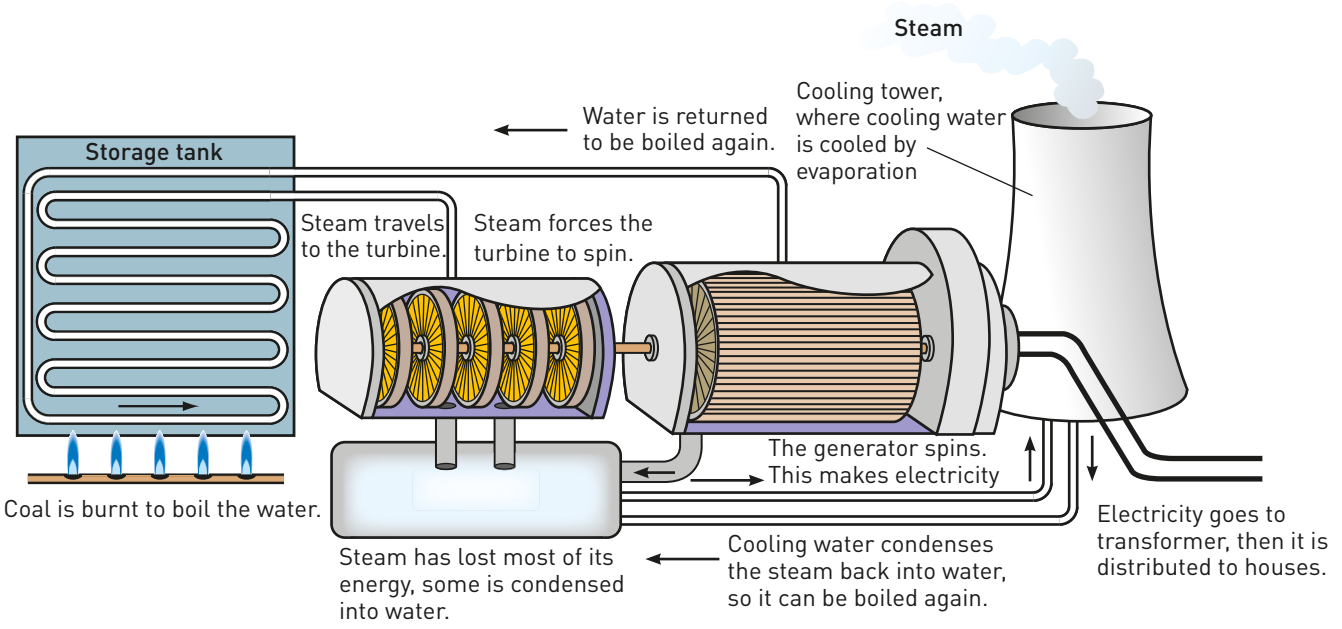


FIGURE 3 The production of electricity in a coal-fired power plant.

Coal is comprised of terrestrial (land-based) plant and animal material, typically from old swamps. Coal used to be burnt in lumps, but now it is ground into a powder first to increase the surface area. This improves the efficiency of combustion.

Energy output

With a specific energy of 29.3–33.5 MJ kg^{-1}, coal does not release a lot of energy per gram compared to other fossil fuels. The combustion of coal also produces the most carbon dioxide for the amount of energy released, approximately 3.59 kg CO_2 per kg of fuel.

Products of combustion

Burning coal produces carbon dioxide, but there are also other products. Combustion can also produce nitrogen oxides and sulfur dioxide, depending on the composition. When it is burnt, coal produces particulates and ash, which are uncombusted organic and inorganic material.

Suitability for purpose

Coal is used mainly to produce electricity (Figure 4). It is considered an affordable option for producing electricity, delivering around 40% of the world's electricity requirements.

Advantages of using coal include:

- it is easy to burn
- it is affordable
- there is a lot of coal still available.

Disadvantages of using coal include:

- it is being used faster than it can be produced
- it produces large quantities of carbon dioxide emissions and is damaging to the environment.

FIGURE 4 An oil rig drilling platform, which extracts oil from the ocean floor.

Crude oil

Crude oil, also known as petroleum, is composed of different types of hydrocarbons sources from marine organisms. As a fossil fuel, it is found trapped in marine sedimentary rocks. Therefore, it can be extracted by drilling. Crude oil can be refined into various types of fuels, including petrol (gasoline), kerosene, diesel oil, fuel oil and asphalt.

CASE STUDY 11.2

Fractional distillation

Crude oil consists of different types of hydrocarbons. Fractional distillation is a method of separating these components into parts or 'fractions'. Each fraction contains similar hydrocarbon molecules.

During fractional distillation, a mixture (in this case, crude oil) is placed below a tall column, which has condensers at different heights. The column is very hot at the bottom and cooler towards the top (Figure 5).

The substances in the crude oil evaporate and condense at different temperatures. The substances with a high boiling point condense at the bottom of the column, and substances with a low boiling point condense at the top of the column. For example, the substances that condense at 200°C will be collected by the condenser for kerosene.

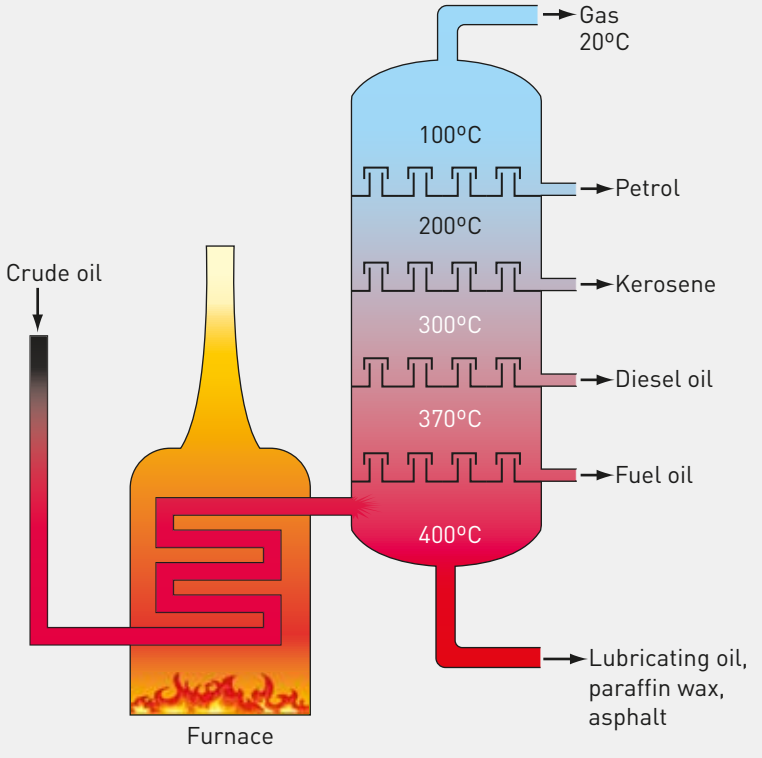


FIGURE 5 Fractional distillation separates the components of crude oil.

Energy output

Crude oil has a high energy density, which means a small amount of oil can produce a lot of energy. In its different refined states, crude oil can have different levels of energy outputs (Table 1).

TABLE 1 Energy outputs of crude oil components

Fuel	Specific energy (MJ kg^{-1})	Mass (kg) CO_2 produced per kg of fuel
Crude oil	42–44	~3.6
Petrol	44–46	~2.3
Diesel	45	~2.7

Products of combustion

When there is enough oxygen present in the reaction, crude oil can react to produce carbon dioxide, water and energy. Other products include nitrogen oxides, carbon monoxide when there is not enough oxygen present, sulfur dioxide when there are impurities in the crude oil, and particulates in the form of soot, which looks like black smoke.

Suitability for purpose

Crude oil has a number of uses. Oil is burnt to create heat and produce electricity.

Crude oil has several advantages. Crude oil:

- has a high energy density
- is easy to extract with the right equipment
- can be easily transported by pipelines
- can be refined and used for a range of purposes, including to manufacture materials such as plastic, medicines, paints and pesticides.

Crude oil also has some disadvantages. Crude oil:

- is a long-term renewable resource, and is being used faster than it can be produced
- produces greenhouse gases, including high amounts of carbon dioxide, which is damaging for the environment
- is more expensive to burn for electricity than coal.

CHECK YOUR LEARNING 11.2

Describe and explain

1 **Summarise** how a fossil fuel is formed.
2 **Explain** why fossil fuels are referred to as 'non-renewable' or 'long-term renewable' resources.
3 **Identify** the main product of coal combustion.

Apply, analyse and interpret

4 **Compare** the use of coal and crude oil in terms of their energy output.
5 Based on the advantages and disadvantages of using crude oil, **judge** whether it is suitable to fulfil its purpose.

Investigate, evaluate and communicate

6 **Investigate** another form of fossil fuel (e.g. natural gas) to determine its energy output and suitability of purpose.
7 **Evaluate** the use of one fossil fuel and discuss whether it has a more positive or negative impact on society.

You can find the following resources for this section on your obook assess:

» Student book questions
Check your learning 11.2

» Weblink
Oil spill

» Weblink
Oil drilling

» Weblink
Coal

11.3 Biofuels

KEY IDEAS

In this section, you will learn about:

- biofuels and their suitability of purpose
- the use of biofuels, including bioethanol and biodiesel.

FIGURE 1 Biofuels come from waste plant and animal products.

Renewable resources are resources that can be produced faster than they are used. Solar energy, wind power and wave power are all examples of renewable energy sources that can be used to power homes and industries.

Biofuels are another form of renewable resource. Biofuels are produced from biomass, or living matter. In other words, biofuels come from waste plant and animal products. Biofuels are often produced by endothermic processes, which means they have high levels of chemical potential energy. These high levels of chemical potential energy can be released and harnessed when burnt.

The two main types of biofuels in Australia are **bioethanol** and **biodiesel**.

renewable resource
a fuel that can be replaced at the same rate or faster than it is being used, e.g. wood or solar energy

biofuel
a fuel that is produced from biomass (living matter)

bioethanol
a biofuel containing ethanol made from sugar case, wheat and corn

biodiesel
a biofuel made from animal fats or vegetable oils

Biofuels are thought to have less impact on the environment than fossil fuels because the plant materials they are made from are produced by photosynthesis. Photosynthesis produces glucose ($C_6H_{12}O_6$) and oxygen:

$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

carbon dioxide + water → glucose + oxygen

The process of photosynthesis removes carbon dioxide from the atmosphere. The combustion of biofuels still produces CO_2, but the biofuel is a product of photosynthesis, so there is less environmental impact with respect to atmospheric CO_2 levels.

Bioethanol

When used as a source of fuel, ethanol (a simple alcohol) is referred to as bioethanol. Bioethanol is a colourless, flammable liquid. It can be produced from biomass such as corn, sugar cane and algae, which undergo a process of fermentation. During this process, yeast digests the sugar in the biomass, and bioethanol and carbon dioxide are released.

Energy output

Ethanol has a lower energy content than fossil fuels such as petrol. The specific energy of bioethanol is 23.4–26.8 MJ kg^{-1}. This means that more ethanol is needed for combustion to produce the same amount of energy.

Products of combustion

In the combustion of bioethanol, ethanol reacts with oxygen to produce carbon dioxide, water and heat energy:

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g) + \text{heat}$$

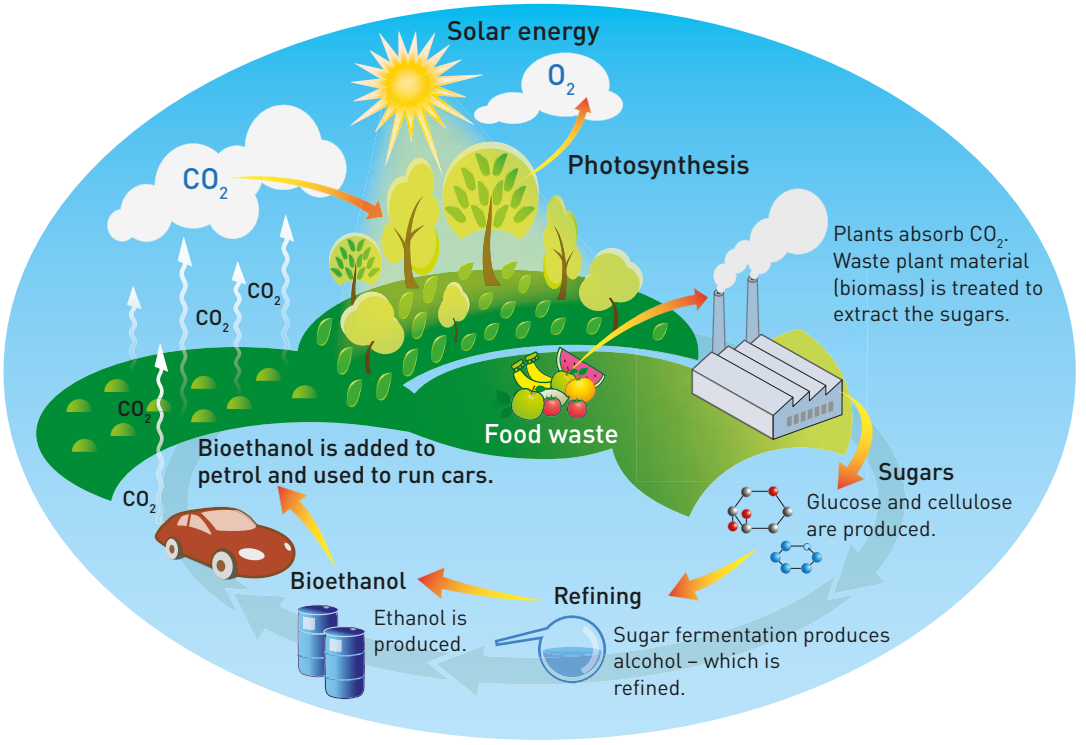


FIGURE 2 Ethanol is produced by plants by fermentation.

The combustion of bioethanol produces carbon dioxide but in lower quantities than other fuels, such as coal or crude oil. Because alcohol contains very little sulfur, bioethanol does not produce much sulfur dioxide or sulfuric acid.

However, because ethanol as a biofuel is derived from plant and animal matter, it also contains large amounts of nitrogen. This nitrogen can produce nitric oxide and other nitrogen compounds when combusted, which can be worse than sulfuric acid for producing acid rain. A lack of oxygen in the combustion reaction can also result in carbon monoxide being produced.

However, bioethanol is often considered a 'cleaner' fuel because it produces no particulates on combustion.

Suitability for purpose

Bioethanol is most often used to power motor vehicles. Usually it is added to petrol; for example, E10 petrol is blended with 10% ethanol.

Bioethanol has several advantages as a fuel. Bioethanol:

- comes from a renewable source
- produces fewer carbon dioxide emissions than other fossil fuels (1.91 kg CO_2 per kg of fuel) and pollutants
- spills are non-toxic compared with crude oil spills.

Bioethanol also has several disadvantages. Bioethanol:

- can be expensive to produce compared to fossil fuels; large areas of land are required to grow crops
- relies on the availability of crops and on good harvests
- yields a lower amount of energy than the same amount of fossil fuels.

Biodiesel

Biodiesel is another renewable fuel. It can be used as a replacement for diesel. Biodiesel is mostly made from vegetable oils, animal fats or recycled greases (such as used cooking oil). It is made by a chemical process called transesterification. During this process, the triglycerides in the animal fats and oils react with methanol or ethanol to form two main products: biodiesel and glycerol (a by-product that can be used in other products such as soaps).

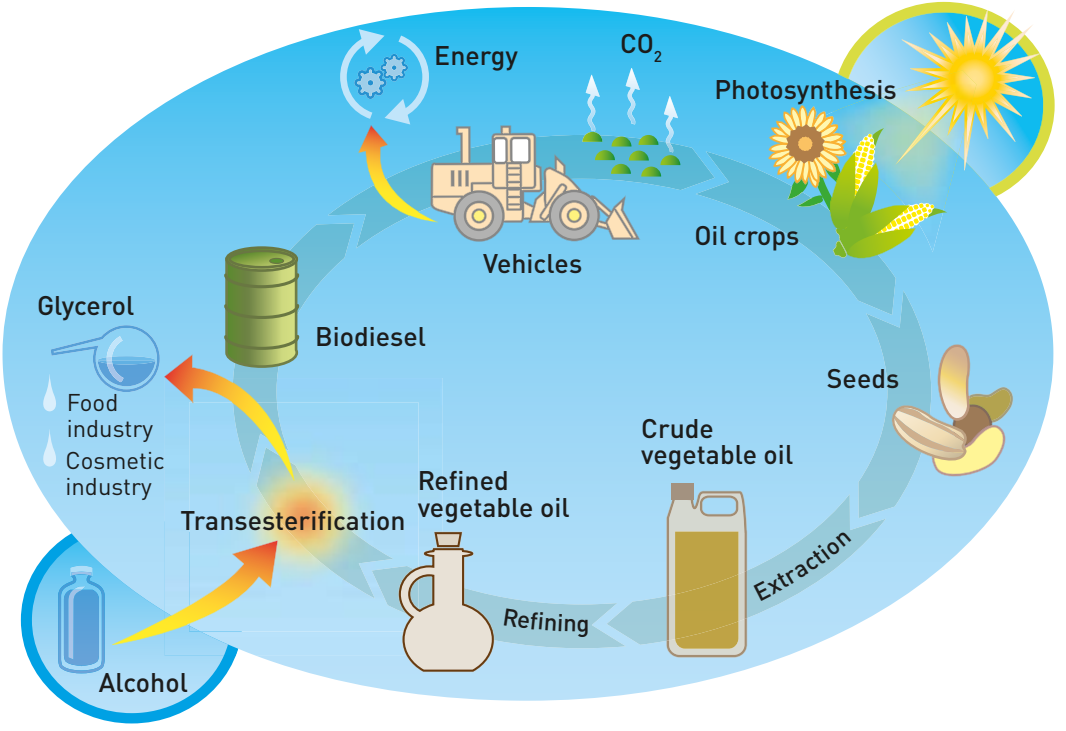


FIGURE 3 Biodiesel is produced from vegetable oils, animal fats or recycled greases.

Energy output

Biodiesel produces less energy than fossil fuels, but it has a higher specific energy (37.8 MJ kg^{-1}) than some other biofuels such as bioethanol. Its energy output is compared with petrodiesel in Case study 11.3.

Products of combustion

There are two types of biodiesel that can be used when examining combustion: C19 and C20. The following equation shows the combustion of the C19 biodiesel methyl linoleate.

$$C_{19}H_{34}O_2 + \frac{53}{2}O_2 \rightarrow 19CO_2 + 17H_2O + \text{heat}$$

Suitability for purpose

Biodiesel can be used as replacement for fuel in diesel engines. It can also be blended with petrodiesel (diesel from crude oil).

Biodiesel has several advantages. Biodiesel:

- comes from a renewable source
- can be used in existing engines and can be blended with other fuels
- produces fewer greenhouse emissions (it produces approximately 2.85 kg CO_2 produced per kilogram of fuel) than other fossil fuels
- is safer to handle because it is less toxic than other fuels.

The disadvantages of using biodiesel are that it:

- is more expensive to use than fossil fuels
- does not flow as easily through fuel lines, and can lose its ability to flow at low temperatures
- relies on the availability of crops and good harvests.

CASE STUDY 11.3

The future of biofuel in Australia

Bioethanol, biodiesel and biogas are currently used as fuels in Australia. Qantas and Virgin Australia are trialling sustainable aviation fuels in order to reduce environmental impact, and the demand for fuel-grade ethanol increased in 2017. The Queensland Government is building a $1 billion sustainable biotechnology and bioproducts sector, and is working with the US Navy as part of their Great Green Fleet initiative.

For the motorist, 10% ethanol-blended petrol (E10) promised cheaper fuel and the ability to reduce the impact on the environment. However, some analysts question whereas this is the reality.

In simple terms, there is less chemical energy locked into a litre of ethanol-blended petrol than there is in a litre of regular unleaded fuel. As a result, motorists use more E10 fuel than regular unleaded petrol on equivalent journeys and have to fill up more often. For E10 to be a viable economic option for the average motorist, the price for E10 petrol needs to be significantly less than that of regular unleaded petrol.

Research is continuing into reducing the cost of producing bioethanol. Designating significant areas of land for the production of the biomass that is used to prepare biofuels reduces the land area that can be used for food and crop production, which may affect the food supply chain.

It is clear that we need to plan for an alternative to fossil fuels; biofuels seem to offer a workable solution, but more research and development is required into making them an economically viable and sustainable component of our fuel supply.

CHECK YOUR LEARNING 11.3

Describe and explain

1 **Identify** the two main types of biofuels used in Australia.
2 **Explain** how one type of biofuel is used by society.
3 What is E10 petrol? **Demonstrate** how it can be used.

Apply, analyse and interpret

4 **Reflect on** why some people prefer the use of biofuels to fossil fuels, and why others may prefer the use of fossil fuels.

Investigate, evaluate and communicate

5 **Investigate** another biofuel (such as biogas) and determine how it is produced, what kind of energy output it produces, and its products of combustion. Compare these characteristics to those of another fuel.
6 'Fossil fuels and biofuels produce carbon dioxide in their combustion reaction.' Given this statement, **evaluate** the impact of fossil fuels and biofuels on the environment.

You can find the following resources for this section on your obook assess:

- » Student book questions
  Check your learning 11.3
- » Weblink
  Biogas
- » Weblink
  Bioethanol
- » Weblink
  Biodiesel

CHAPTER 11

Review

Chapter summary

11.1
- Fuels react with oxygen to produce useful energy in a combustion reaction. The combustion of fuel is always an exothermic reaction and produces heat. This energy is measure in joules (J). Combustion of fuels produces CO_2, a harmful gas for Earth's atmosphere that is contributing to climate change.

11.2
- There are two main types of fossil fuels, coal and crude oil (petroleum).
- Coal is the result of land-based plant and animal matter being buried over millions of years.
- Crude oil is the result of marine organisms being buried under the ocean over million of years.
- Both coal and crude oil are subjected to high temperatures and pressures before they are converted into fossil fuels.
- Fossil fuels are non-renewable energy sources, or long-term renewable resources.

11.3
- Biofuels are sourced from plant and animal matter.
- Biofuels are renewable energy sources meaning they can be produced faster than they are consumed.
- There are two types of biofuels, bioethanol and biodiesel.

Key terms

- biodiesel
- bioethanol
- biofuel
- energy
- fossil fuel
- fuel
- joule
- long-term renewable resource
- renewable resource

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; *** = high.

Multiple choice

1 The burning of a fuel is considered to be:
- **A** damaging to the ozone layer
- **B** a greenhouse reaction
- **C** an inefficient reaction
- **D** an exothermic reaction.

2 Complete combustion of a hydrocarbon fuel always produces:
- **A** heat, carbon and hydrogen
- **B** heat, carbon monoxide and water
- **C** heat, carbon dioxide and hydrogen
- **D** heat, carbon dioxide and water.

3 Identify a fuel that would produce 85% carbon dioxide and small amounts of sulfur dioxide and particulates when combusted.
- **A** ethanol
- **B** coal
- **C** petrol
- **D** hydrogen

4 Biofuels are sometimes referred to as 'carbon neutral'. This is because:
- **A** they do not release carbon dioxide during combustion
- **B** biofuel production from biomass is considered an environmentally friendly process when compared with fossil fuels
- **C** the amount of carbon dioxide absorbed to produce the biomass is equivalent to the

amount of carbon dioxide released during biofuel combustion

D the producers of biofuels receive money for growing biomass, making the fuel production very cheap.

5 Which of the following is a renewable fuel source?

A natural gas

B petrol

C biodiesel

D coal

Short answer

Describe and explain

★**6** **Define** what a fuel is?

★**7** **Identify** a common fuel and describe how it is used.

★**7** **Recall** the chemical equation for the combustion reaction of coal.

★**8** **Explain** why fossil fuels are referred to as 'non-renewable' or 'long-term renewable'.

★**9** **Explain** the formation of carbon monoxide in a combustion reaction.

★**10** **Describe** the process of coal formation.

★**11** **Explain** the link between the combustion of fuels and climate change.

★**12** In a coal-fired power station, **explain** how the chemical energy of the coal is used to produce electricity.

★★**13** **Calculate** how much energy is released by the complete combustion of 148.2 g of butanol (C_4H_9OH). (ΔH (C_4H_9OH) = −2670 kJ mol^{-1})

Apply, analyse and interpret

★**14** **Classify** the following fuel sources as renewable or long-term renewable.

a wood

b corn

c coal

★★**15** **Analyse** the energy and carbon dioxide output of the fuels shown in Table 1.

TABLE 1 Outputs of energy and carbon dioxide from certain fuels

Fuel	Specific energy (MJ kg^{-1})	Mass (kg) of CO_2 produced per kg of fuel
Fuel A	29.3	2.8
Fuel B	43	4
Fuel C	47.6	3.4

Determine which fuel is the most desirable. **Explain** your answer.

★★**16** Petrodiesel and biodiesel release comparable levels energy and carbon dioxide during combustion. **Determine** why is biodiesel considered to be more environmentally friendly?

★★**17** 'More than 70% of Australia's electricity is generated using coal as a fuel source.' **Consider** this statement and comment on the suitability of coal as a significant source of fuel.

★★**18** A 5.50 kg sample of bituminous coal contains 3.30 kg of carbon.

a **Determine** what the mass percentage of carbon in the bituminous coal is.

b **Determine** what mass of carbon dioxide could be produced from the complete combustion of 10.0 kg of bituminous coal.

Investigate, evaluate and communicate

★★**19** **Propose** a suitable fuel source for each of the following situations. **Justify** your answers.

a A company needs heat energy to run the machinery in its factory. The company does not have a lot of money to operate, but needs to produce a lot of steam quickly.

b A car hire company wants to improve its public image by reducing its environmental impact.

★★**20** **Investigate** the global mean surface temperature data for the last 250 years. What conclusions could be made regarding this data and burning fossil fuels?

★★**21** There is much debate about the use of E10 petrol in Australian vehicles. **Investigate** the advantages and disadvantages of using E10 petrol in motor vehicles as a replacement for traditional petrol.

Check your obook assess for these additional resources and more:

- » Student book questions
  Chapter 11 Revision questions
- » Revision notes
  Chapter 11
- » assess quiz
  Auto-correcting multiple choice quiz
- » Flashcard glossary
  Chapter 11

UNIT 1

Practice exam questions

Chemical fundamentals – structure, properties and reactions

Multiple choice

1 What is the electron configuration of a sodium ion?

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

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

C $1s^2 2s^2 2p^6 3s^0$

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

2 Chemical bonds are caused by:

A electrostatic attraction between atoms

B electrostatic attraction between electrons

C electrostatic attraction between protons

D electrostatic attraction between atomic nuclei.

3 Alpha particle radiation is a:

A high-speed electron emitted by an unstable atomic nucleus

B high-speed neutron emitted by an unstable atomic nucleus

C particle of two protons and two neutrons emitted by an unstable nucleus

D high-frequency form of electromagnetic radiation emitted by an unstable atomic nucleus.

4 An unknown metal is poured into a measuring cylinder that holds exactly 22.5 mL. The metal used to fill the container weighs 306.0 g. The metal is:

A lead (density = 11.34 g cm^{-3})

B iron (density = 7.87 g cm^{-3})

C mercury (density = 13.56 g cm^{-3})

D gold (density = 19.32 g cm^{-3})

5 Which of the following comments about 1 mole of Ag and 1 mole of Au is true?

A They are equal in mass.

B They contain the same number of atoms.

C Their molar masses are equal.

D They have the same atomic mass.

6 What is the ionic formula for lithium oxide?

A LiO

B LiO_2

C Li_2O_2

D Li_2O

7 What is the molar mass of magnesium oxide?

A 24.3 g mol^{-1}

B 16.0 g mol^{-1}

C 20.0 g mol^{-1}

D 40.3 g mol^{-1}

8 What is the empirical formula for a compound that is 36.1% Ca and 63.9% Cl?

A CaCl

B Ca_2Cl

C $CaCl_2$

D Ca_2Cl_2

9 What is the percentage of OH in $Ca(OH)_2$?

A 45.9%

B 75%

C 66.6%

D 90.1%

10 The molecular formula for vitamin C is $C_6H_8O_6$. What is its empirical formula?

A CHO

B CH_2O

C $C_3H_4O_3$

D $C_2H_4O_2$

FIGURE 1 Oranges contain a high amount of vitamin C.

Short answer

1 **Construct** the Lewis structure for the following.
   - **a** chlorine atom
   - **b** chloride ion
   - **c** chlorine gas molecule
   - **d** NH_2OH

2 **Describe** how the following properties of elements changes from the left-hand side to the right-hand side of the periodic table.
   - **a** atomic radii
   - **b** valency
   - **c** ionic radii
   - **d** first ionisation energy
   - **e** electronegativity

3 **Calculate** the relative atomic mass of chlorine from these isotopic abundances: ^{35}Cl (75.78%) and ^{37}Cl (24.33%).

4 Atomic absorption spectroscopy (AAS) was used to compare several standard solutions of lead. The results are shown below.

TABLE 1

Lead concentration (mg L^{-1})	2.5	5.0	7.5	10
Absorbance (at 218 nm)	0.17	0.325	0.50	0.27

A 1.0 g sample of soil was dissolved in acid, and then diluted to a volume of 10 mL. The resulting sample was found to have an absorbance of 0.20 at a wavelength of 218 nm. Use a calibration curve to **determine** the concentration of lead (percentage mass lead/mass soil) in the soil.

5 From the table below, **determine** which material has a metallic structure. Provide reasoning to support your answer.

TABLE 2

Material	Melting point (°C)
X	445
Y	1907
Z	-7

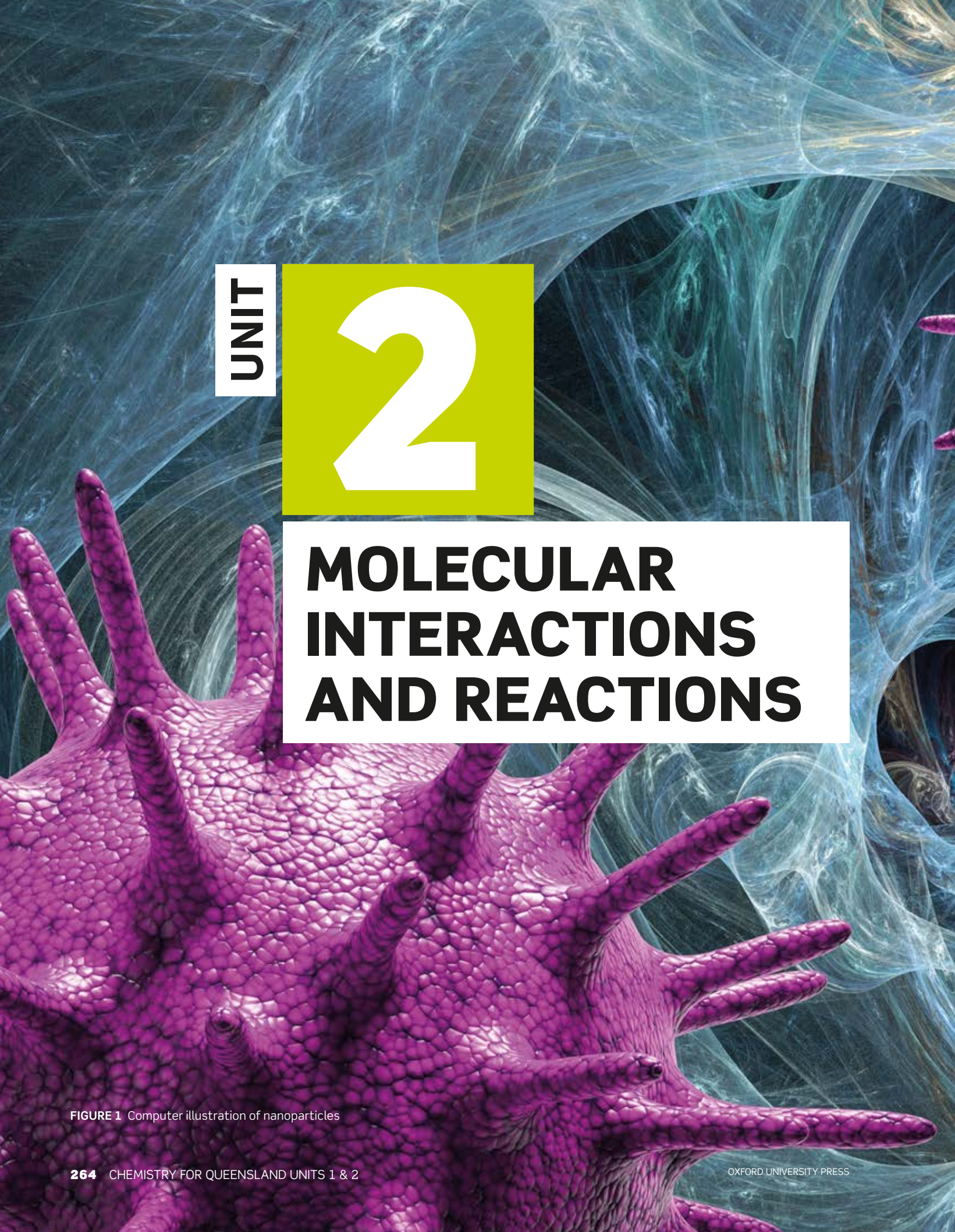
6 **Determine** the number of:
   - **a** molecules in 0.325 mole of sulfur dioxide
   - **b** moles in 5.5×10^{24} molecules of H_2
   - **c** molecules in 150 g of water.

7 Methane (CH_4) burns in oxygen to produce carbon dioxide and water.

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

A student burnt 0.250 mol of methane in 1.25 mol of oxygen in a sealed container.
   - **a** **Identify** the limiting reagent.
   - **b** **Calculate** the theoretical yield (in mass) of water.
   - **c** If 6 grams of water was obtained from the reaction, **calculate** the percentage yield.

8 Many fuels used in homes are hydrocarbons called alkanes. The enthalpy changes that occur in the combustion of alkanes are a source of energy for heating and cooking.
   - **a** **Define** the term 'enthalpy' as it is used in this context.
   - **b** **Evaluate** whether this reaction is exothermic or endothermic. Provide evidence to support your reasoning.
   - **c** **Construct** the equation, including symbols of state, that accompanies the enthalpy change of formation of hexane, $C_6H_{14}(l)$.
   - **d** Use Table 1 in Section 9.2 to **determine** the enthalpy of 1 mole of hexane.
   - **e** If the enthalpy change for the combustion of hexane is -4163 kJ mol^{-1}, **calculate** the change in temperature (°C) that would occur in 100.0 kg of water if 1 mole of hexane was completely combusted to heat the water. (Assume there was no loss of energy in the system.)
   - **f** If the temperature of the water increased by 6.5°C, **calculate** the percentage error of the combustion reaction.



UNIT 2

MOLECULAR INTERACTIONS AND REACTIONS

FIGURE 1 Computer illustration of nanoparticles

Over 70% of Earth's surface is covered in liquid water. Most of this water is saline, mix of soluble ions, dissolved minerals and lifeforms. The chemical reactions of life depend on the solution properties of water. The nature and strength of the intermolecular forces between the water molecules explain its melting point, boiling point and vapour pressure, as well as its ability to dissolve some substances, but not others.

Water as a vapour is subject to the same laws of temperature, volume and pressure as other atmospheric gases. The constantly moving molecules of gas display high levels of kinetic energy. Frequent collisions between high-energy gas molecules increase the rate of chemical reactions. Another way to increase the rate of chemical reactions is through biological catalysts, such as enzymes in blood.

Chapter guide

Topic 1	Intermolecular forces and gases	Chapters 12–14
Topic 2	Aqueous solutions and acidity	Chapters 15–18
Topic 3	Rates of chemical reactions	Chapter 19

Unit objectives

→ Describe and explain intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.
→ Understanding of intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.
→ Analyse evidence about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.
→ Evidence about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.
→ Investigate phenomena associated with intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.
→ Evaluate processes, claims and conclusions about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.
→ Communicate understandings, findings, arguments and conclusions about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.

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CHAPTER 12

Intermolecular forces

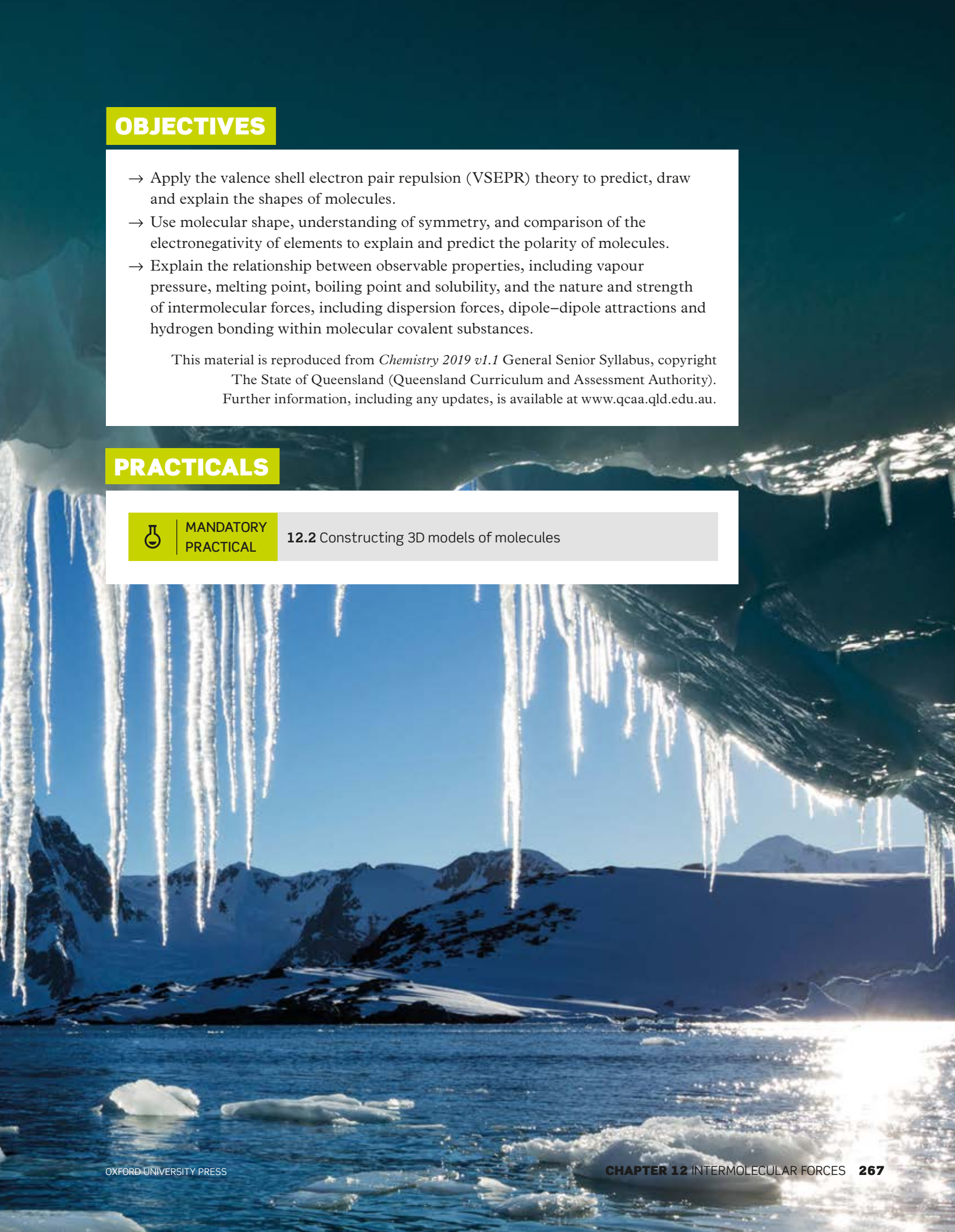
Intermolecular forces are a key concept to understanding most areas of chemistry. No other concept is used as extensively as intermolecular forces.

In organic chemistry, how electrons are shared between atoms in molecules (intramolecular covalent bonds) explains the characteristic reactions of certain chemicals and allows chemists to predict what product will form from two reactants. In analytical chemistry (and forensic analysis), intermolecular forces (i.e. the attractive forces that exist between different molecules and ions) are used to separate substances, based on their properties and attraction to other substances. Biochemists use intermolecular forces to explain the interaction of biomolecules within the body, their functions, their properties and their various health risks.

Inorganic chemists use intermolecular forces to explain how ionic compounds can dissolve in water and certain other organic solvents. Materials chemists use intermolecular forces to determine the properties of plastics, clothing and other various fabrics.

A strong grasp of this topic is essential for future studies in Units 1–4 Chemistry.

FIGURE 1 The properties of water are determined by intermolecular bonding.



OBJECTIVES

→ Apply the valence shell electron pair repulsion (VSEPR) theory to predict, draw and explain the shapes of molecules.

→ Use molecular shape, understanding of symmetry, and comparison of the electronegativity of elements to explain and predict the polarity of molecules.

→ Explain the relationship between observable properties, including vapour pressure, melting point, boiling point and solubility, and the nature and strength of intermolecular forces, including dispersion forces, dipole–dipole attractions and hydrogen bonding within molecular covalent substances.

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PRACTICALS

MANDATORY PRACTICAL	**12.2** Constructing 3D models of molecules

12.1 Shapes of covalent molecules

KEY IDEAS

In this section, you will learn about:

- covalent bonding between non-metals and metalloids
- the valence shell electron pair repulsion (VSEPR) theory
- molecular shapes.

Chapter 3 covered the covalent bonding that occurs between non-metallic elements. The group of molecules formed from these elements is biologically important because it includes DNA, proteins, fats, sugars and medicines.

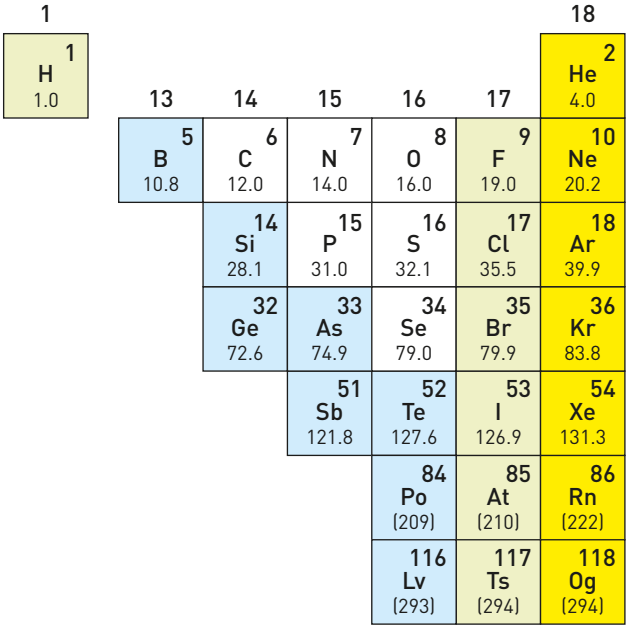


FIGURE 1 The non-metals (yellow, white and green) and the metalloids (blue) on the periodic table form covalent bonds.

Covalently bonded molecules

Covalent bonds occur between two non-metals when they share their electrons (Chapter 3). It can also occur between a non-metal and a metalloid. Non-metals and metalloids are found on the periodic table in the top right-hand side (Figure 1) and include hydrogen. The noble gases are largely unreactive and only form molecules with covalent bonds with great difficulty. No compounds are known for helium and neon.

VSEPR theory, molecular shapes and bond angles

The **valence shell electron pair repulsion (VSEPR) theory** states that electron pairs, whether they are bonded or **non-bonding electrons**, repel each other. Molecules form three-dimensional (3D) shapes, depending on the number of bonds, the type of bond (single, double or triple) and the number of non-bonding pairs of electrons.

valence shell electron pair repulsion (VSEPR) theory a theory that states that electron pairs in bonds and lone electron pairs repel each other around a central atom, forming a 3D shape

non-bonding electrons pairs of electron in an atom's outer shell that are shared because they have already paired up

In order to determine the shape of a simple, small molecule, the central atom must first be identified. All bonds attach to this atom, making it easily identifiable.

Although **lone pairs of electrons** influence **bond angles**, they are not considered when the shape of the molecule is described. Only the position of bonded atoms are considered when describing the molecule's shape.

Valence diagrams and ball and stick diagrams are the best ways to visualise the shape of a molecule.

Linear molecules

A **linear** molecule shape results when the central atom forms two covalent bonds and there are no non-bonded pairs of electrons on the central atom. There are no non-bonding

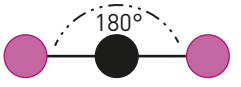


FIGURE 2 Ball and stick model of a linear covalent molecule. The black ball is the central atom and the bonds are at an angle of 180° to one another.

the electrons in the covalent bonds push away from each other. They are as far apart from each other as possible when there is a central bond angle of 180°, placing the terminal atoms on opposite sides of the central atom (Figure 2).

The carbon dioxide molecule is an example of a linear molecule (Figure 3). Carbon is the central atom and forms two double bonds with oxygen atoms with a bond angle of 180°.

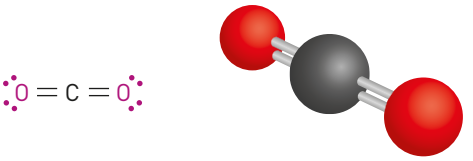


FIGURE 3 Valence and ball and stick models of carbon dioxide, demonstrating a bond angle of 180° and a linear shape.

Trigonal planar molecules

A **trigonal planar** molecule shape results when the central atom forms three covalent bonds and there are no non-bonding pairs of electrons on the central atom, leaving only the **bonding electrons** to push away from each other. The electrons tend to spread as far apart as possible, which results in a bond angle of 120°.

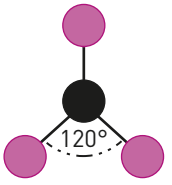


FIGURE 4 A ball and stick model of a trigonal planar covalent molecule. The black ball is the central atom and the bonds are at a bond angle of 120° to one another.

The formaldehyde molecule is an example of a trigonal planar molecule (Figure 5). Carbon is the central atom and forms a double bond with an oxygen atom and two single bonds with hydrogen atoms. The bond angles can be measured and are close to 120°.

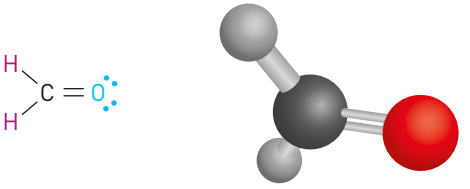


FIGURE 5 The valence and ball and stick models of formaldehyde, demonstrating a bond angle of 120° and a trigonal planar shape.

Tetrahedral molecules

A **tetrahedral** molecule shape results when the central atom forms four covalent bonds and there are no non-bonding pairs of electrons. The electrons in the covalent bonds push away from each other. This means that the bonds are as far away from each other as possible, resulting in a bond angle of 109° (Figure 6).

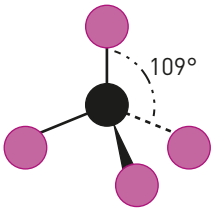


FIGURE 6 A ball and stick model of a tetrahedral covalent molecule. The black ball is the central atom and the bonds are at a bond angle of 109° to one another.

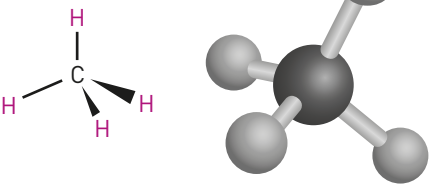


FIGURE 7 The valence and ball and stick model of methane, demonstrating a bond angle of 109° and a tetrahedral shape.

As the bonds in a tetrahedral molecule are not in line with one another (or do not exist in a two-dimensional (2D) plane), a triangular wedge and dotted line are used to represent bond direction. A dotted line represents a bond that is oriented away from the viewer (into the paper) (Figure 6). A solid or triangular line represents a bond that is oriented towards the viewer (coming out of the paper) (Figure 7).

The methane molecule is an example of a tetrahedral molecule (Figure 7). Carbon is the central atom and forms four single bonds with hydrogen atoms. The bond angle is 109°.

lone pair of electrons
valence electrons that are not bonded to another atom

bond angle
the angle between covalent bonds attached to the same central atom

linear
the molecular shape in which the central atom has two covalent bonds and no lone pairs of electrons, forming a straight line

trigonal planar
the molecular shape in which the central atom has three covalent bonds and no lone pairs of electrons

Study tip

Only boron can form a trigonal planar molecule, as the central atom, with single covalent bonds.
All other central atoms must have a double bond and two single bonds to form a trigonal planar molecule.

bonding electrons
the electrons that are shared in a covalent bond

tetrahedral
the molecular shape in which the central atom has four covalent bonds and no lone pairs of electrons, forming a tetrahedral structure

Trigonal pyramidal molecules

trigonal pyramidal
the molecular shape in which the central atom has three covalent bonds and one lone pair of electrons, forming a pyramid structure

A **trigonal pyramidal** molecule shape results when the central atom forms three covalent bonds, with one non-bonding pair of electrons. The non-bonding (lone) pair of electrons and the bonding electrons push the bonds to create as much space between them as possible. However, the lone pair of electrons occupies more space than a bonding pair of electrons. They push the bonds even further, reducing the angle between them. This means that the bonds are further away from the lone pair of electrons, but closer to each other, resulting in a bond angle of 107° (Figure 8), which is slightly less than the 109° bond angle in a tetrahedral shape.

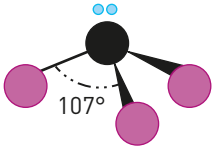


FIGURE 8 A ball and stick model of a pyramidal covalent molecule. The black ball is the central atom and the bonds are at a bond angle of 107° to one another, resulting in a pyramid shape. The smaller blue balls represent the lone, non-bonded, pair of electrons.

FIGURE 9 The valence and ball and stick models of ammonia, demonstrating a bond angle of 107° and a pyramidal shape.

The triangular solid and dashed wedges are also used in this model to represent the bonds that are not in line, but are oriented towards or away from the viewer.

The ammonia molecule is an example of a pyramidal molecule (Figure 9). Nitrogen is the central atom and forms three single bonds with hydrogen atoms, with bond angle 107°.

Bent molecules

bent
the molecular shape in which the central atom has two covalent bonds and two lone pairs of electrons, forming a V-shape

A **bent** molecule shape results when the central atom forms two covalent bonds, and has two non-bonding pairs of electrons. The two non-bonding (lone) pairs of electrons push the bonds even closer, reducing the angle between the bonds. This results in a bond angle of 104.5° (Figure 10).

One example of this orientation is water (Figure 11). Oxygen is the central atom and forms two single bonds with hydrogen atoms and bond angle 104.5°.

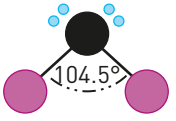


FIGURE 10 A ball and stick model of a pyramidal covalent molecule. The black ball is the central atom and the bonds are at an angle of 104.5° to one another. The smaller blue balls represent the lone, non-bonded, pairs of electrons.

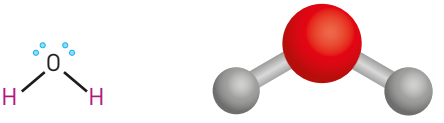


FIGURE 11 The valence and ball and stick models of water, demonstrating a bond angle of 104.5° and a bent shape.

CHALLENGE 12.1A

Bent molecules

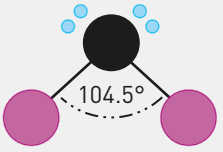
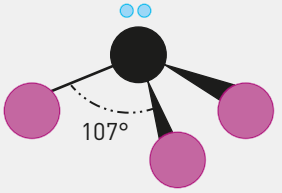
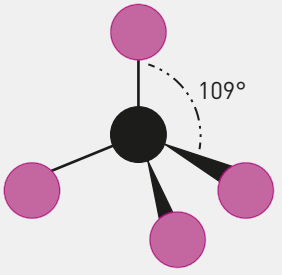
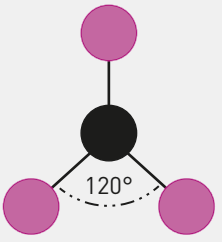
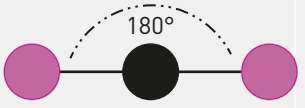
Can you think of other bent molecules? Do they all have the same bond angle?

Covalent molecule shape summary

A summary of covalent **molecular shapes** is provided in Table 1.

molecular shape the shape formed between a central atom and bonded atoms when covalent bonds form

TABLE 1 Bond shapes and angles of covalent molecules

Name	Shape	Number of bonds	Number of non-bonded pairs of electrons	Bond angle	Example
Bent		2	2	104.5°	H_2O
Pyramidal		3	1	107°	NH_3
Tetrahedral		4	0	109°	CH_4
Trigonal planar		3	0	120°	CH_2O
Linear		2	0	180°	CO_2

CHALLENGE 12.1B

Evaluate the shape

Evaluate the structure of ethanoic acid (Figure 12).

How many central atoms does it have? How many shapes does it include and what are these shapes?

Is it fair to say that ethanoic acid has a single shape?

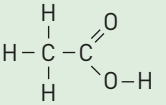


FIGURE 12 The structure of ethanoic acid.

CASE STUDY 12.1

Discovery of a double helix

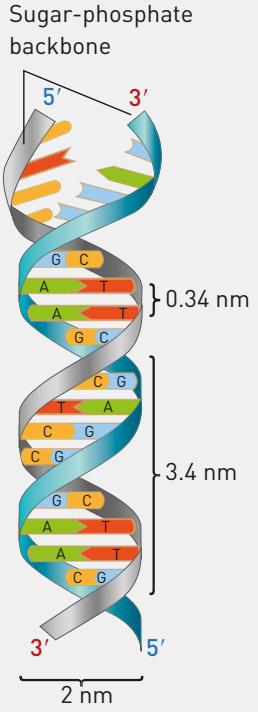


FIGURE 13 The structure of the DNA molecule, showing the pairing of the nitrogen-containing: bases guanine (G) with cytosine (C) and adenine (A) with thymine (T).

In 1953, the journal *Nature* published a ground-breaking paper by James Watson and Francis Crick, with the assistance of Maurice Wilkins, outlining the structure of the DNA molecule and its double helix geometry. The structure consisted of twin strands with pairs of nitrogen-containing units called bases linked by hydrogen bonds. These strands spiralled around each other and could be easily pulled apart for replication (Figure 13).

Biologist Watson and physicist Crick were working at the University of Cambridge in the UK. They already had key pieces of information about the structure and were in a race against other scientists to develop a model for the DNA molecule. They already knew that the structure contained four nitrogen-containing bases (guanine, cytosine, adenine and thymine) and that these bases paired up within the structure by hydrogen bonding.

Watson and Crick used X-ray crystallography, in which an X-ray beam is directed through a crystallised DNA molecule and the resultant image is captured on a photographic glass plate. This image provided information and data for the internal structure of the DNA molecule. But they could not find a geometric shape that matched what they already knew of DNA, and no image that they obtained from X-ray crystallography was clear enough to provide the missing information.

The two scientists discussed their lack of data with colleague Maurice Wilkins, who was working at Kings College in London. Wilkins provided them with an X-ray crystallography image taken by his colleague Rosalind Franklin (Figure 14), and this provided the last piece of the puzzle for them to complete their model. Watson and Crick used Franklin's data and image without her knowledge.

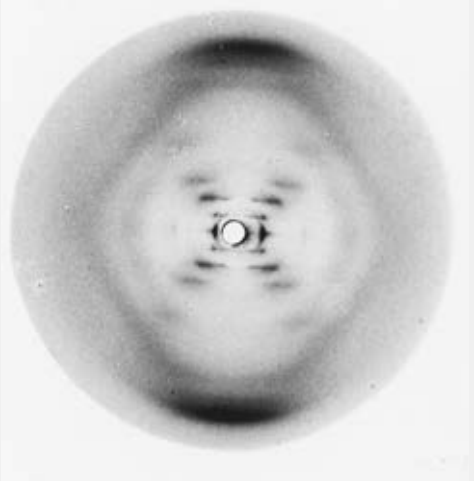


FIGURE 14 The X-ray image taken by Franklin.

In 1958, Rosalind Franklin died of ovarian cancer, which she most likely developed as a result of her exposure to X-rays. In 1962, Watson, Crick and Wilkinson were awarded the Nobel Prize in Physiology or Medicine. The Nobel Prize rules state that the award cannot be received posthumously (after death), so Rosalind Franklin went largely unrecognised for her contribution.

The work of these four scientists gave birth to molecular biology, combining the fields of chemistry and biology.

FIGURE 15 Rosalind Franklin made large contributions to the science of X-rays.

CHECK YOUR LEARNING 12.1

Describe and explain

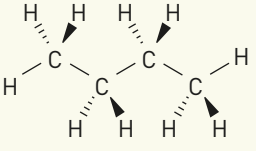
1 **Explain** the difference between tetrahedral, pyramidal and bent molecular shapes.

2 **Describe** the difference between single, double and triple bonds by explaining the behaviour of electrons.

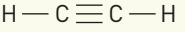
Apply, analyse and interpret

3 **Determine** the shapes the following molecules would form. Use an electron dot diagram and a ball and stick diagram to justify your answer.
- **a** Hydrogen iodide (HI)
- **b** Hydrogen cyanide (HCN)
- **c** Nitrogen triiodide (NI_3)
- **d** Borane (BH_3)
- **e** Dichlorine monoxide (OCl_2)
- **f** Fluoromethane (CH_3F)

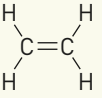
4 **Determine** the shapes of the following molecules. Remember to identify the central atom first.
- **a** Butane



- **b** Acetylene

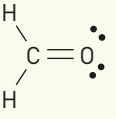


- **c** Ethene

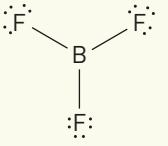


Investigate, evaluate and communicate

5 According to VSEPR theory, electrons repel each other in a covalent molecule. Formaldehyde and boron trifluoride both have a trigonal planar shape (Figure 16).



Formaldehyde



Boron trifluoride

FIGURE 16 Ball and stick representations of formaldehyde and boron trifluoride.

Using your knowledge of the VSEPR theory, **evaluate** both structures to predict whether a double bond affects bond angle. **Justify** your response.

You can find the following resources for this section on your obook assess:

- » Student book questions
  Check your learning 12.1
- » Challenge
  12.1A Bent molecules
- » Challenge
  12.1B Evaluate the shape
- » Weblink
  Trigonal pyramidal molecules

SCIENCE AS A HUMAN ENDEAVOUR

12.2 Development of the VSEPR theory

KEY IDEAS

In this section, you will learn about:

- two- and three-dimensional graphical models that represent and communicate the shapes of molecules.

The molecular geometry of covalent compounds determines every physical, chemical and biological property of molecules. Molecular geometry is a field of study that has expanded quickly since the 1990s.

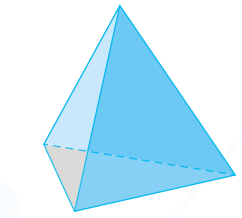
VSEPR theory

VSEPR theory is based on the space requirements of all electron pairs within the valence shell around the central atom. It was first credited to British chemist Professor Ronald James Gillespie and Australian chemist Ronald Sydney Nyholm in 1957. Gillespie specialised in the field of molecular geometry, and Nyholm specialised in inorganic chemistry. Before Gillespie and Nyholm, several other chemists alluded to the idea of molecular geometry; however, it was Gillespie and Nyholm who developed the in-depth model that is referred to today. Historically, VSEPR theory has been referred to as Gillespie–Nyholm theory.

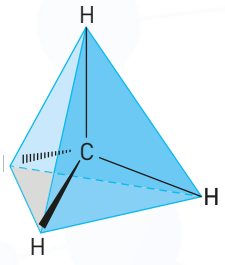
Many scientists made contributions to the field of molecular geometry, even before its development. Friedrich August Kekulé (1829–1896) determined that carbon was tetravalent (had four bonding electrons) and could bond to other carbon atoms to form organic molecules. This was the beginning of the idea of **two-dimensional modelling**. Jacobus Henricus van't Hoff (1852–1911) theorised that these four bonds of carbon were arranged in a tetrahedral shape, forming the basis of **three-dimensional modelling**.

two-dimensional modelling modelling in which all bonds are drawn evenly around the central atom on a 2D plane

three-dimensional modelling modelling in which atoms attached to the central atom are drawn in 3D to maximise the electron repulsion around the central atom



Regular tetrahedron



Methane

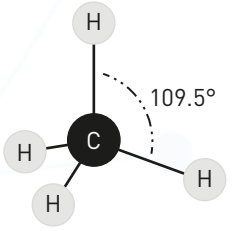


FIGURE 1 The tetrahedral shape proposed by van't Hoff.

Two-dimensional modelling

Two-dimensional structures are drawn with the central atom or atoms in the middle of the structure and any attached atoms spread evenly around it (Figure 2). It is a simple model, but it does not take into consideration the geometry of the molecule and therefore its properties and behaviour.

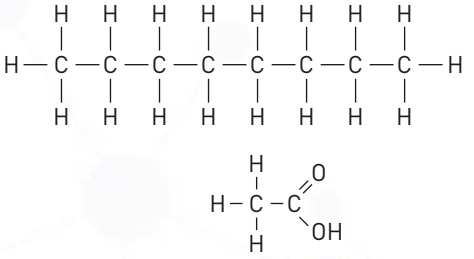


FIGURE 2 Two-dimensional structures of octane and ethanoic acid molecules.

Although the structures of octane and ethanoic acid (Figure 2) show the central atoms, bonds and all atoms in the correct order, they do not reflect molecular geometry or orientation.

Three-dimensional modelling

Molecular geometry outlines the 3D modelling of molecules on the basis of their bonds and lone pairs of electrons. This branch of chemistry is also known as stereochemistry. Chemists noticed that some otherwise identical molecules that had the same behaviours and properties, interacted differently with plane-polarised light. To explain this, it was theorised that, although atoms are attached to the same central atom, they are arranged in a different way in three dimensions. This led to the concept of chiral molecules, which are mirror images of each other (Figure 3) that cannot be placed on top of each other (superimposed).

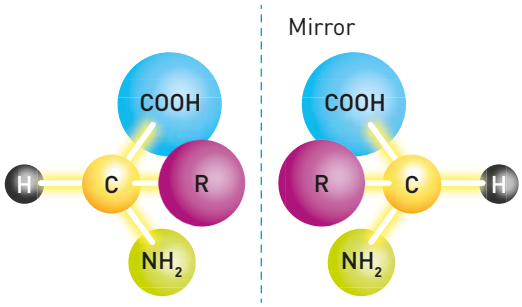


FIGURE 3 Chiral molecules, such as amino acids, are two molecules that are otherwise identical but are mirror images of each other that cannot be placed on top of each other. They have the same physical properties except they interact differently with plane-polarised light. They may have very different biological properties. The R group is a variable side chain.

It became increasingly important to represent molecules three-dimensionally to better determine their physical, chemical and biological properties. Chemists also realised that drawing molecules in their 3D structures was very time consuming so they started developing models that made it unnecessary to illustrate the entire molecule.

Figure 4 represents the tetrahedral **structural formulas**, ball and stick models and **skeletal structures** of octane and ethanol. In a skeletal structure, the lines represents bonds between carbons, but no carbons and hydrogen atoms are usually drawn unless the hydrogen atoms are connected to atoms other than carbon. The bends in the line represent a carbon atom, as do the two ends of the line.

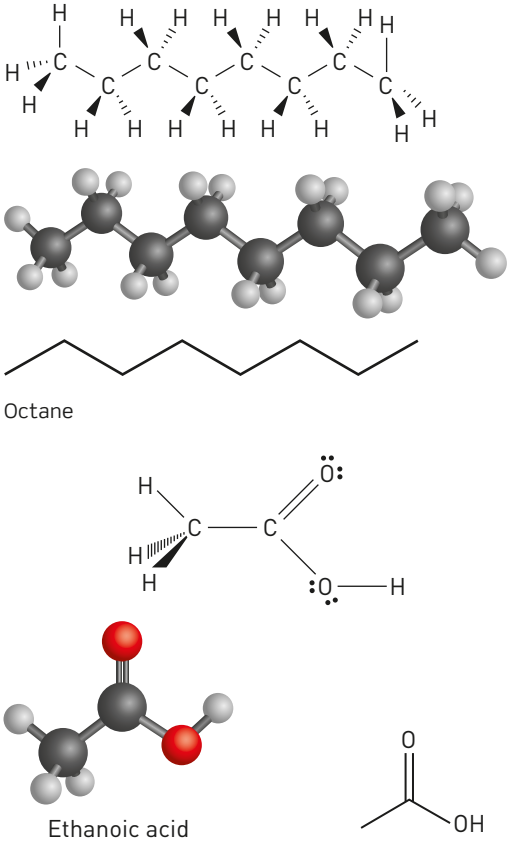


FIGURE 4 The tetrahedral structural formulas ball and stick models and skeletal structures of octane and ethanoic acid. In skeletal structures, every carbon has a total of four bonds, every hydrogen only one bond, and every oxygen two bonds.

Computational chemistry

Computational chemistry can predict the geometric structure of larger molecules, their interactions with other molecules and therefore their behaviour. This is achieved by using various computer modelling programs and is essential in protein design, DNA analysis, forensic science and chemical engineering.

structural formula
a model that shows all atoms and all bonds but not necessarily in 3D, unless solid triangle and dotted lines are used

skeletal structure
a model in which the carbon chain is drawn as a zig-zag line with bends representing carbon atoms; carbon atoms are also at the end of the lines and each line represents a bond between carbons

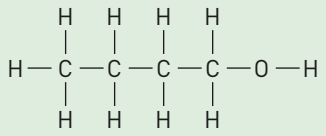
computational chemistry
a branch of chemistry in which the 3D structure of a molecule is generated by technology to assist in the determination of shape, intermolecular bonding and properties

CHALLENGE 12.2

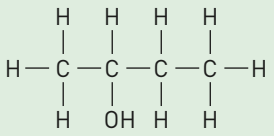
Skeletal structures

Convert the following 2D structures into their skeletal structures.

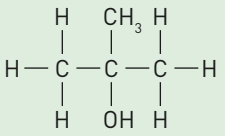
a



b



c

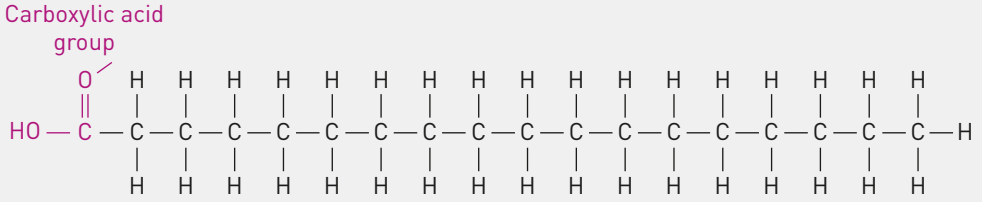


CASE STUDY 12.2

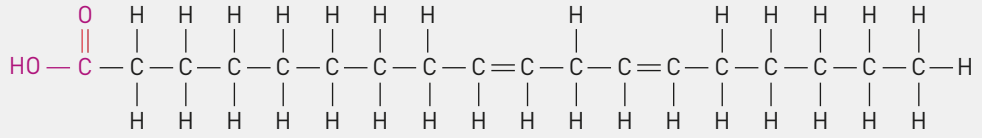
Fatty acid shape

Fatty acids are biomolecules that belong to the carboxylic acid family and form fats within the body.

Fatty acids have the same functional group (group of atoms) located in the same position of each molecule. This is the carboxylic acid functional group and it is highly polar due to the presence of two oxygen atoms (Figure 5). Fatty acids are long carbon chain molecules, sometimes more than 20 carbons long (Figure 5). Carbon chains of each fatty acid may contain a double bond between the carbon atoms. A saturated fatty acid contains only single bonds; an unsaturated fatty acid contains one or more double bonds.



Stearic acid



Linoleic acid

FIGURE 5 Stearic acid is a saturated fatty acid. Linoleic acid is an unsaturated fatty acid that has two double bonds.

The presence of a double bond means that unsaturated fatty acid molecules can form one of two molecular geometries. A *trans* double bond occurs when the hydrogen atoms attached to the double bond are on opposite sides of the double bond. A *cis* double bond occurs when the hydrogen atoms are on the same side (Figure 6). This *cis* configuration is by far the most common in natural fatty acids.

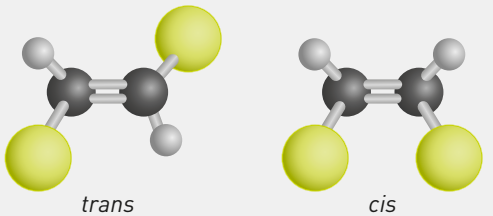


FIGURE 6 On a *trans* double bond, the hydrogen atoms are on opposite sides. On a *cis* double bond, the hydrogen atoms are on the same side

When fatty acids are in the *cis* orientation, they become kinked. In Figure 7, stearic acid contains no double bonds and is a straight molecule. Oleic acid has one double bond so is kinked. Linoleic acid has two *cis* double bonds so has more kinking than oleic acid.

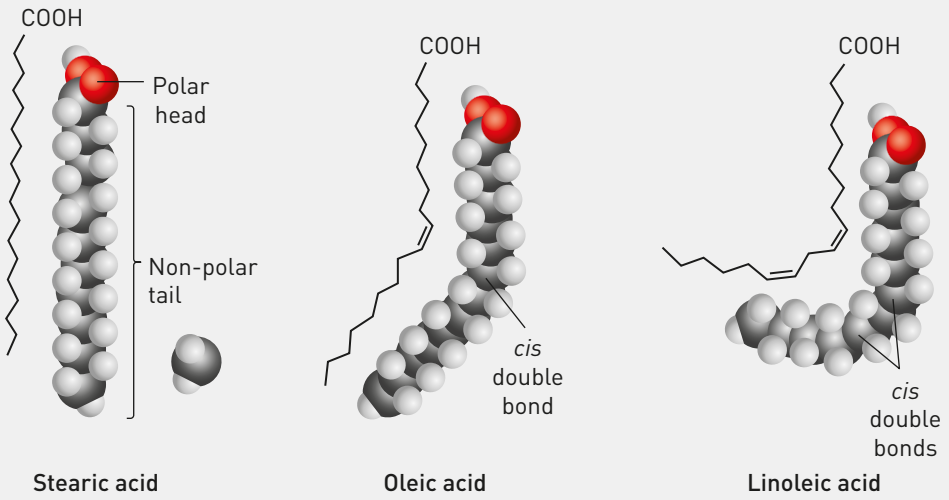


FIGURE 7 The orientation and degree of kinking in fatty acids with *cis* double bonds.

Therefore, fatty acids with *cis* double bonds experience less intermolecular forces because they do not lie flat. A molecule that lies flat (or has a straight chain) can interact more strongly with a neighbouring molecule because of the closeness of the molecules. As *cis* molecules cannot get as close to neighbouring molecules, intermolecular forces are not as strong and they have lower melting points.

CHECK YOUR LEARNING 12.2

Describe and explain

1 **Describe** the advantages and disadvantages of drawing a 2D structure.

2 **Describe** the advantages and disadvantages of drawing a 3D structure.

Investigate, evaluate and communicate

3 A student claims that the VSEPR model works when representing the molecular geometry of all molecules. **Evaluate** this statement and provide a response.

You can find the following resources for this section on your obook assess:

- » Student book questions
  Check your learning 12.2
- » Mandatory practical
  12.2 Constructing 3D models of molecules
- » Challenge
  12.2 Skeletal structures
- » Website
  VSEPR model

12.3 Polarity and intermolecular bonding

KEY IDEAS

In this section, you will learn about:

- electronegativity of elements
- the influence of molecular shape on the polarity of molecules
- the intermolecular bonding, or attraction, between molecules.

polarity
gives the molecule overall positive and negative ends creating a particle charge, the electronegativities of the atoms and the shape of a molecule

polar
having a dipole; for example, a bond or molecule

The shape of a molecule and its ability to form covalent bonds are key to the understanding of the concept of **polarity**. Polarity is an essential concept in chemistry because it helps to explain the properties of covalent compounds and their behaviour in different situations.

Electronegativity and polarity

Covalent bonds are classified as **polar** depending on the electronegativity of the atoms involved in the bond.

Electronegativity

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons within a covalent bond (Chapter 2). Electronegativity increases across the periodic table from group 1 to group 18 and decreases down the periodic table from period 1 to period 7. Fluorine is the most electronegative element of the periodic table (Figure 1) with an electronegativity value of 4.0.

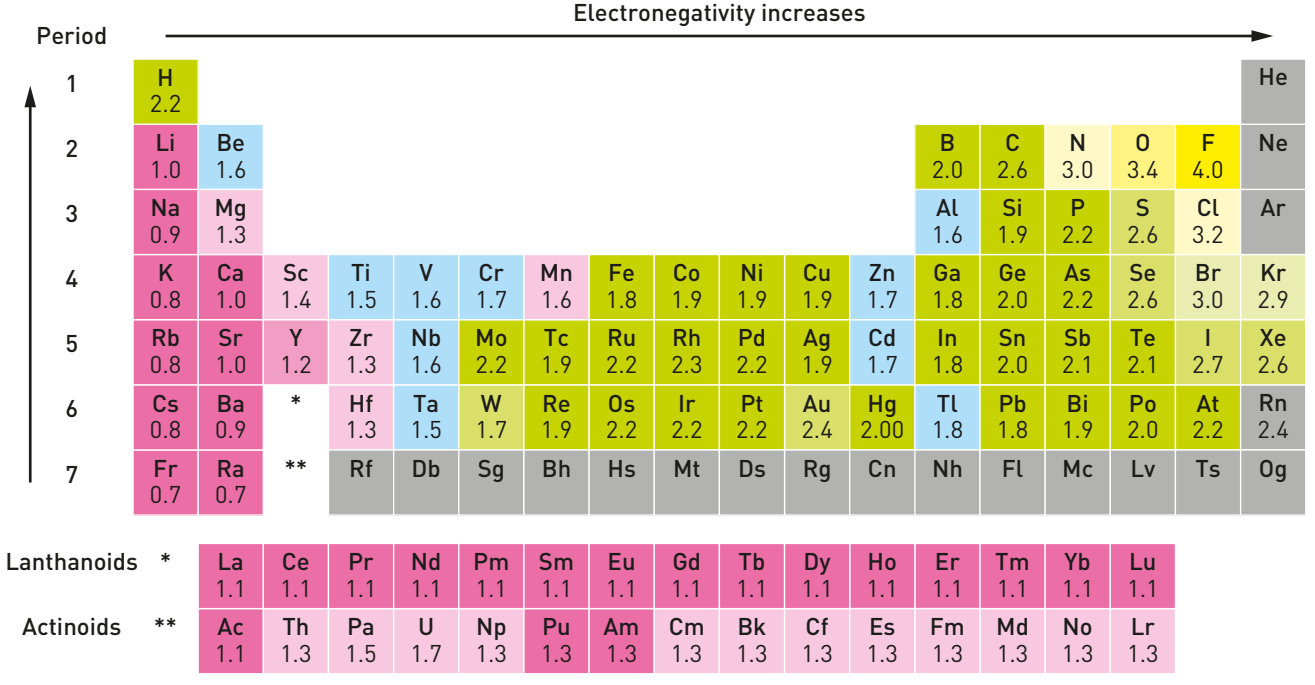


FIGURE 1 Electronegativity values demonstrate a trend in the periodic table.

CHALLENGE 12.3A

Noble gases

Why do the noble gases helium, neon and argon not have electronegativity values reported in Figure 1?

When fluorine is covalently bonded to hydrogen (electronegativity value 2.1), the fluorine atom pulls on the electrons within the bond (Figure 2).

FIGURE 2 Electrons in the H–F bond are not in the centre of the bond but are closer to the fluorine.

As the electrons move closer to the fluorine atom, they give the fluorine a partial negative charge. It is important to note that this is not a full charge (as in ionic bonding). Fluorine is slightly more negative, not because it has gained an electron but because the electrons are closer to the fluorine. Similarly, as the electrons are pulled away from the hydrogen atom, hydrogen gains a partial positive charge. These partial positive and negative charges are represented by δ^+ or δ^- (Figure 3).

FIGURE 3 Representation of partial positive and negative charges on a H–F bond: fluorine is partially negative (δ^-) and hydrogen is partially positive (δ^+).

When two chlorine atoms covalently bond to form a Cl_2 molecule, the atoms have the same electronegativity of 3.0, so they pull on the electrons within the bond equally. The electrons do not move to either side of the molecule and there is no partial positive or negative charge assigned to the molecule (Figure 4).

FIGURE 4 The chlorine (Cl_2) molecule has no overall positive or negative charges.

Polarity

As one side of the hydrogen fluoride molecule is more positive, and the opposite side is more negative, it creates a **dipole**. A permanent dipole occurs when the opposite sides of covalent bonds or molecules have opposite charges. This is also referred to as a polar bond.

The hydrogen fluoride molecule contains a polar bond, whereas chlorine molecules do not.

Effect of molecular shape on polarity

The shape of a molecule also influences its polarity. Carbon tetrachloride is a tetrahedral molecule, with carbon as the central atom bonded to four chlorine atoms (Figure 5). Carbon has an electronegativity value of 2.6 and chlorine has a value of 3.2; therefore, the individual C–Cl bonds in the molecules are slightly polar. However, the carbon tetrachloride molecule is symmetrical, so the partial charges created by the polar bonds cancel each other out and the overall effect is that the carbon tetrachloride molecule is non-polar (Figure 5).

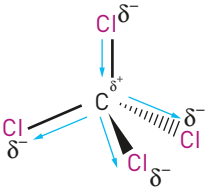


FIGURE 5 Tetrahedral structure of the carbon tetrachloride molecule. The blue arrows show the direction that the electrons are being pulled within the molecule.

The ammonia molecule (Figure 6) has a pyramidal shape. The central nitrogen atom has a larger electronegativity value than hydrogen. The pull of electrons (blue arrows) occurs in one direction in the molecule. This makes the hydrogen end of the molecule electropositive and the nitrogen end of the molecule electronegative, resulting in a polar molecule.

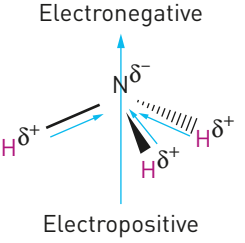


FIGURE 6 The pyramidal structure of the ammonia molecule results in it being polar.

If the shape of a molecule results in the electron pull being cancelled out, such as the tetrahedral carbon tetrachloride molecule (Figure 5), a polar molecule is only formed if one of the atoms attached to the central atom is more electronegative than another.

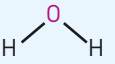
dipole
the separation of opposite charges formed in bond or a molecule so that it has a partial positive charge at one end and a partial negative charge at the opposite end

WORKED EXAMPLE 12.3A

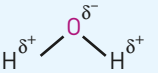
Draw a valence model of a water molecule, indicating which atoms are electropositive and electronegative, the direction of electron pull and the overall effect. Determine whether the water molecule is polar or non-polar.

SOLUTION

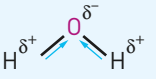
Oxygen has six valence electrons and needs to share two with hydrogen to form water. Therefore, two hydrogen atoms are required to form two covalent bonds and there are two lone pairs of electrons (non-bonding pairs). This gives water a bent shape.



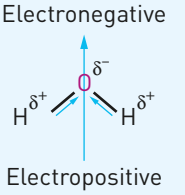
Oxygen has an electronegativity value of 3.4 and hydrogen has an electronegativity value of 2.2. Therefore, oxygen is electronegative (δ^-) and hydrogen is electropositive (δ^+).



The oxygen atom is more electronegative, so it pulls the electrons within the bonds.



Because oxygen pulls the electrons in the same direction, the net effect is that the oxygen end of the molecule is more negative and the hydrogen end is more positive, creating a polar molecule.

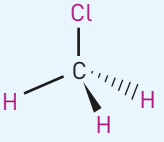


WORKED EXAMPLE 12.3B

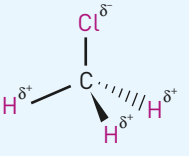
Draw a valence model of a chloromethane molecule (CH_3Cl), indicating which atoms are electropositive and which are electronegative, the direction of electron pull and the overall effect. Determine whether the chloromethane molecule is polar or non-polar.

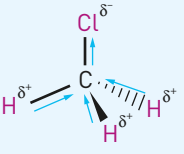
SOLUTION

Carbon has four valence electrons and needs to share all four electrons with the three hydrogen atoms and the chlorine atom, forming single bonds. There are no lone pairs and chloromethane has a tetrahedral shape.



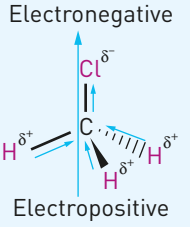
Chlorine has an electronegativity value of 3.2, carbon is 2.6 and hydrogen is 2.2. Although the carbon is more electronegative in the C–H bonds, chlorine is more electronegative in the C–Cl bond. Therefore, hydrogen is electropositive (δ^+) and chlorine is electronegative (δ^-).





Because it is more electronegative, the chlorine pulls the electrons in the C–Cl bond and the carbon pulls the electrons in the C–H bond.

The chlorine pulls the electrons throughout the entire molecule; therefore, the net effect is that the chlorine end of the molecule is more negative and the hydrogen end is more positive, creating a polar molecule.



Intermolecular forces

The polarity of molecules allows them to interact with other polar molecules. This is because the positive end of one molecule interacts with the negative end of a neighbouring molecule, creating an **electrostatic interaction**. The attraction between molecules is called an **intermolecular force**; for example dispersion force. These attractions are not classified as bonds, even though some are called bonding.

electrostatic interaction
the attraction of the negative end of one molecule to the positive end of another molecule

intermolecular force
an attraction between molecules; for example, dispersion forces

Depending on the electronegativities, some polar bonds are stronger than others and result in stronger attraction between molecules.

For example:

- a C–H bond (electronegativities: carbon 2.5, hydrogen 2.1) has a net pull of electrons in carbon's direction to the value of 2.5 – 2.1 = 0.4
- a Cl–H bond (electronegativities: chlorine 3.0, hydrogen 2.1) has a net pull of electrons in chlorine's direction to the value of 3.0 – 2.1 = 0.9.

Therefore, the chlorine–hydrogen bond forms a stronger dipole and results in more interaction between molecules. It is not necessary to calculate these values if you remember the trend in the periodic table (Section 12.1).

There are two main types of intermolecular forces: dispersion forces and dipole–dipole interactions (including hydrogen bonding).

Dispersion forces

Dispersion forces occur between all molecules. Even though methane has four polar C–H bonds (Figure 7), its symmetry cancels out their polarity.

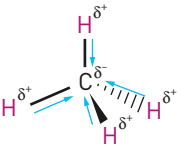


FIGURE 7 A methane molecule has four C–H bonds. The symmetry of the molecule also means it is non-polar.

dispersion force
the weak interaction between non-polar molecules

Non-polar covalent molecules contain valence electrons that constantly move throughout the molecule. At any given time, there may be more electrons on one side of the molecule than on the other, randomly creating a dipole that only exists in this moment, and is not permanent. This 'dipole moment' is temporary and so any electrostatic interaction that it creates is fleeting, disappearing until the next dipole moment occurs.

Therefore, dispersion forces are weak because they only occur for a short moment, and require another molecule to also be experiencing a dipole moment before they are attracted to one another.

Charge cloud diagrams (Figure 8) are used to represent dispersion forces. Charge cloud diagrams demonstrate the charge created at each end of a molecule by electron movement. A partial positive charge is represented in red and a partial negative charge is represented in blue. The black dots within the centre of the charge clouds represent the nuclei of the atoms. Figure 8 represents the intermolecular bonding between two molecules of chlorine (Cl_2) gas. Because the chlorine molecule is non-polar, it experiences dispersion forces.

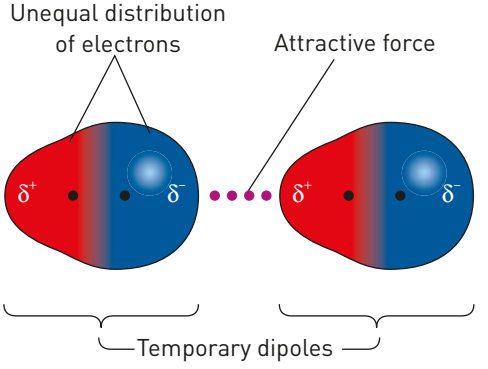


FIGURE 8 The charge cloud diagram for chlorine (Cl_2) gas, showing the temporary dipoles created by the movement of electrons, and the resultant dipole moment in chlorine (Cl_2) gas.

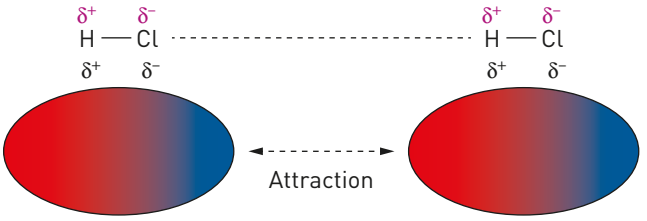


FIGURE 9 The dipole–dipole interaction between neighbouring HCl molecules.

Dipole–dipole attractions

Dipole–dipole attractions occur in molecules that have a permanent dipole. The positive end of the molecule is attracted to the negative end of an adjacent molecule. Repulsion also occurs when the negative ends or positive ends of molecules align.

Figure 9 represents the attraction between two adjacent hydrochloric acid molecules. Both molecules are polar, so they have a permanent dipole. Attraction occurs between the partially negative chlorine of one molecule and the partially positive hydrogen of another molecule. As the dipole within these molecules is permanent, the attraction is stronger than the attraction created by dispersion forces.

Polar molecules can align in various arrangements when they are attracted to one another. This means that they will pack quite closely together when they interact (Figure 10).

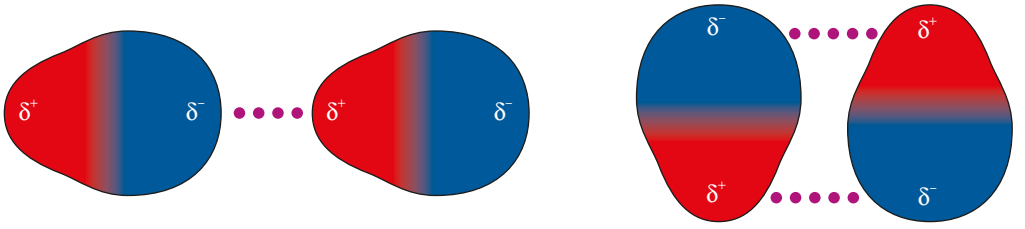


FIGURE 10 Various arrangements of polar molecules, allowing for greater interaction and stronger bonding.

hydrogen bonding
the strongest form of dipole–dipole interaction in which a partially negative atom (typically F, O or N) has an electrostatic interaction with a partially positive hydrogen atom

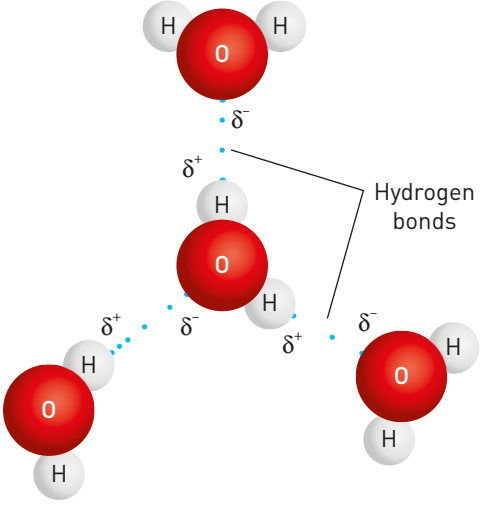
Hydrogen bonding

Hydrogen bonding is a specific type of dipole–dipole interaction, where the partially positive hydrogen atom on one molecule is attracted to the most partially negative atoms from the periodic table. Typically, these atoms are fluorine, oxygen or nitrogen, but hydrogen bonding can occur with other atoms also. This attraction is stronger than a regular dipole–dipole interaction because it involves the most electronegative (F, O and N) and electropositive (H) atoms on the periodic table.

Hydrogen bonding occurs in water. One water molecule forms hydrogen bonds with other water molecules (Figure 11). The partially negative oxygen on one molecule attracts the partially positive hydrogen on a neighbouring molecule, resulting in a hydrogen bond.

Study tip

Electrostatic interactions are always between a positive and negative. This can be a full positive and negative charge, as with ionic substances, or a partial positive and negative charge, as with polar covalent molecules.



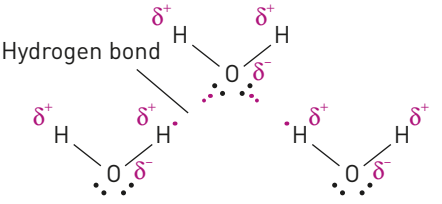


FIGURE 11 Hydrogen bonding exists between the partially positive hydrogen atom of one water molecule and the partially negative oxygen molecule of a neighbouring water molecule.

CHALLENGE 12.3B

The forces in alcohol

Alcohols can be classified as primary, secondary and tertiary depending on the position of the OH group. Figure 12 shows the structure of three different alcohols. Each alcohol contains four carbon atoms and 10 hydrogen atoms. Discuss the intermolecular forces involved with each molecule. Which would have stronger intermolecular forces? Explain why.

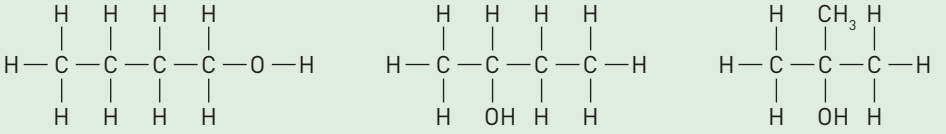


FIGURE 12 A primary alcohol (butan-1-ol), secondary alcohol (butan-2-ol) and tertiary alcohol 2-methyl propan-2-ol.

CASE STUDY 12.3

Hydrogen bonding in DNA

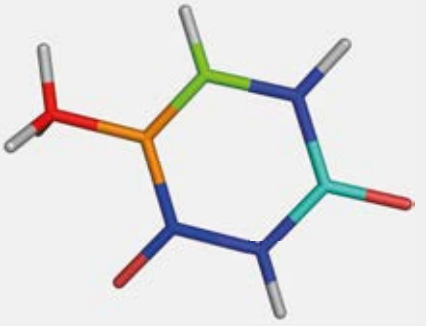


FIGURE 13 Thymine

DNA molecules hold the genetic code that builds all proteins required for the normal functioning of the human body. DNA molecules are shaped like a ladder. The outer bars or handrails of the ladder (light pink in Figure 14a) are made up of alternating sugar and phosphate groups. The sugar groups (dark pink in Figure 14b) connect to the rungs (or steps) of the ladder, which are nitrogen-containing bases. These nitrogen-containing bases are called guanine, cytosine, adenine and thymine (GCAT). In order to connect both sides of the ladder, the bases form hydrogen bonds. Adenine always bonds with thymine and guanine always bonds with cytosine. This is called the base pairing rule. This hydrogen bonding is between a partially negative nitrogen or oxygen and a partially positive hydrogen.

The body makes a copy of DNA by cutting the ladder down the middle, by breaking the hydrogen bonds, and making a copy of both sides. Therefore, it is important that the bonds that hold the nitrogenous bases together are strong, so that the DNA does not break down within the body. However, they cannot be so strong that they cannot be pulled apart when a copy of the DNA is required to make a protein. Hydrogen bonding provides the perfect amount of strength to hold the molecule together. All other bonds within the DNA molecule are covalent.

The DNA ladder forms a double helix because of the hydrogen bonding within the molecule. Guanine and cytosine form three hydrogen bonds, where as adenine and thymine form two hydrogen bonds. This pairing is the most stable arrangement. There is a twofold effect of the number of hydrogen bonds between the nitrogen-containing bases. The first is that cytosine and guanine are harder to break apart because they share three hydrogen bonds, as opposed to two. The second is that three hydrogen bonds hold the two bases closer together than two do. For this reason, the cytosine–guanine base pair is both stronger and held closer together. This is one of the contributing factors to the shape of the double helix. The molecule twists because of the length in the hydrogen bonds.

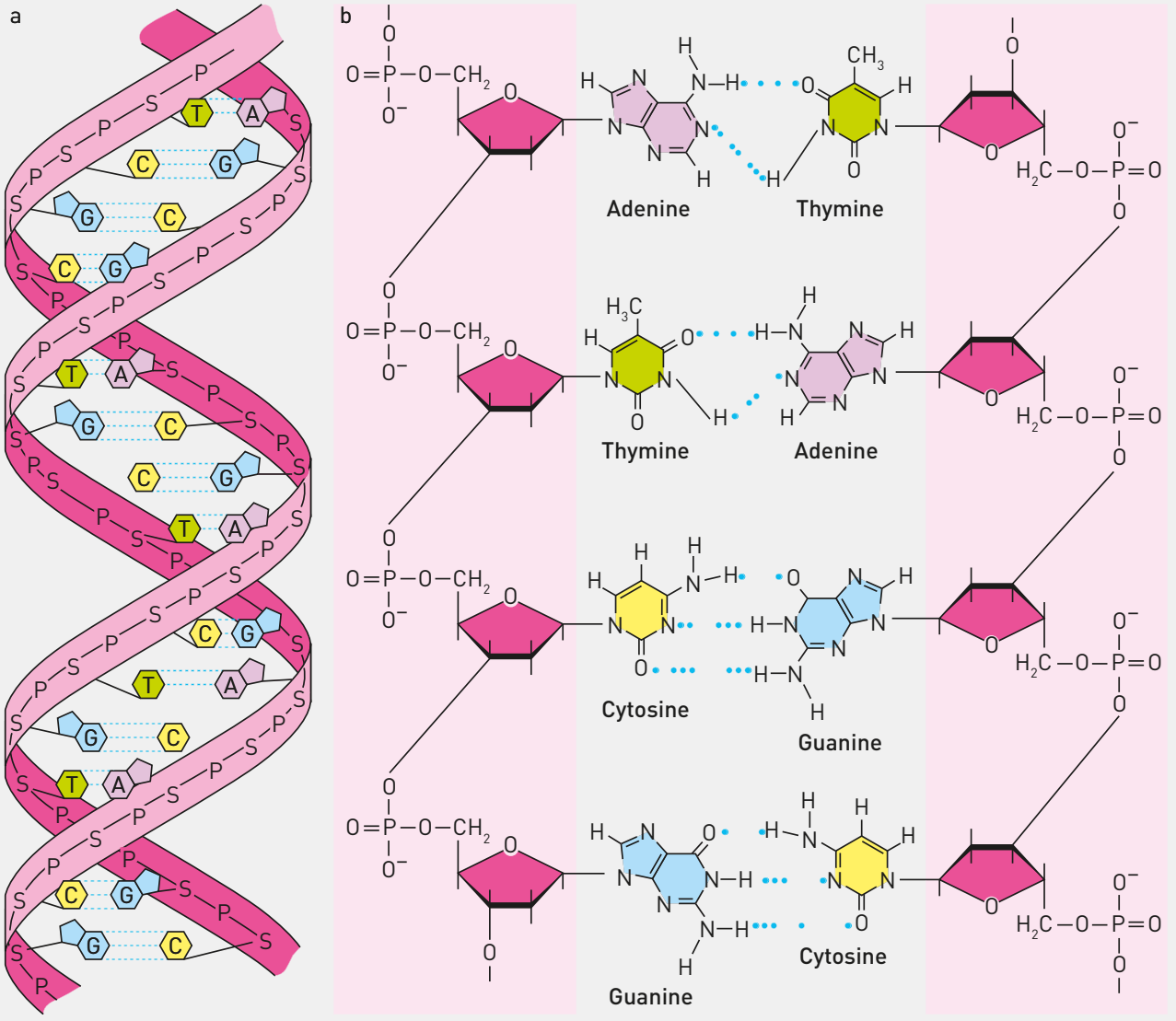


FIGURE 14 (a) The double helix structure of DNA. (b) The structure of DNA, demonstrating hydrogen bonding between nitrogen-containing bases.

CHECK YOUR LEARNING 12.3

Describe and explain

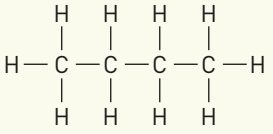
1 **Describe** polarity and explain how a polar molecule gets its charge separation.

2 **Explain** how a hydrogen bond is similar to and different from a dipole–dipole interaction.

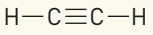
Apply, analyse and interpret

3 **Determine** whether the following molecules interact by dispersion, dipole–dipole or hydrogen bond interactions.

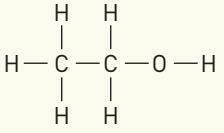
a Butane



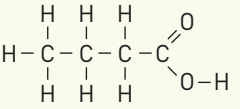
b Ethyne



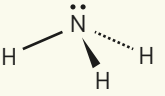
c Ethanol



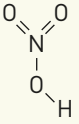
d Butanoic acid



e Ammonia



f Nitric acid



4 **Determine** which intermolecular forces could exist between different pairs of molecules from Question 3.

Investigate, evaluate and communicate

5 **Evaluate** the structure of DNA in Figure 14. What are the various types of bonding and intermolecular forces involved?

6 A student claims that the name 'hydrogen bonding' is misleading and should be changed. **Evaluate** this statement and provide a response.

You can find the following resources for this section on your obook assess:

- **» Student book questions**
  Check your learning 12.3
- **» Challenge**
  12.3A Nobel gases
- **» Challenge**
  12.3B The forces of alcohol
- **» Weblink**
  Intermolecular forces

12.4 Properties of molecules based on shape and polarity

KEY IDEAS

In this section, you will learn about:

- the influence of intermolecular forces on the behaviour of covalent molecules
- the properties of covalent molecules, including melting point, boiling point, vapour pressure and solubility.

The properties of covalently bonded molecules are determined by the intermolecular forces generated between molecules. An increase in the polarity of a molecule means that it is attracted more strongly to other molecules. Therefore, the presence of a polar bond as well as its strength influences how strongly molecules interact with one another.

Properties

The intermolecular bonding and therefore the interaction between molecules determine the properties and behaviours of all covalent molecules. These properties include:

- melting point
- boiling point
- **vapour pressure**
- solubility.

vapour pressure the pressure of a vapour in contact with its solid or liquid form

Melting point and boiling point

All particles have **kinetic energy**. Kinetic energy is a measure of the motion of particles. The greater the kinetic energy of a particle, the faster it moves.

kinetic energy the energy involved with the motion of particles, which can be increased by adding heat energy

When investigating the melting point of a substance, two factors must be considered. These are the:

- kinetic energy of the particles: the more kinetic energy a particle has, the more it moves around and breaks away from other particles
- energy involved with the intermolecular forces holding the molecules together: the greater the intermolecular force holding the particles together, the harder it will be to break them apart.

If the energy of particle movement is greater than the energy of the attraction, the particle will break free.

The melting point of a substance is the temperature at which a solid turns into a liquid (Figure 1). In a solid state, the intermolecular forces between molecules are so great that movement of the molecules cannot break them apart. The molecules are held together in a rigid structure, and their motion is so small that they only vibrate. When heat energy is applied, molecules begin to move faster, and soon have enough energy to start moving away from neighbouring molecules.

FIGURE 1 At the melting point of ice, these cubes become liquid.

The molecules are still attracted to one another but are not now as tightly held. The temperature at which the heat energy is sufficient enough for molecules to break free of their rigid structure, or turn a solid into a liquid, is called the melting point.

The boiling point of a substance is the temperature at which a liquid turns into a gas. In a liquid state, the intermolecular forces between molecules are great enough that molecules are held together and cannot fully break apart from each other. Instead they move over and around each other. Their kinetic energy is not great enough for them to break free of each other. When heat energy is applied, molecules begin to move faster, and soon have enough kinetic energy to start breaking free from neighbouring molecules. The molecules are no longer able to interact with each other because their kinetic energy is greater than the energy involved with their intermolecular bonding (Figure 2). The boiling point is the temperature at which the heat energy is sufficient enough to break molecules free, or turn a liquid into a gas.

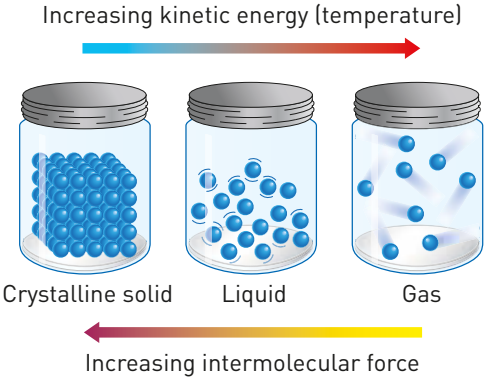


FIGURE 2 The relationship between kinetic energy and intermolecular bonding in solid, liquid and gaseous states.

Trends in melting and boiling points

When the intermolecular bonding between molecules is stronger (dipole–dipole interactions and hydrogen bonding), more kinetic energy, and therefore heat, is required to break them apart into separate molecules. If molecules only interact through dispersion forces, which are very weak, less kinetic energy, and therefore heat, is required to break them apart.

Alkanes are a family of organic compounds that consist only of the elements carbon and hydrogen. The carbons are covalently bonded in a long chain of single bonds, and the remaining bonds are to hydrogen. Some common alkanes are shown in Figure 3.

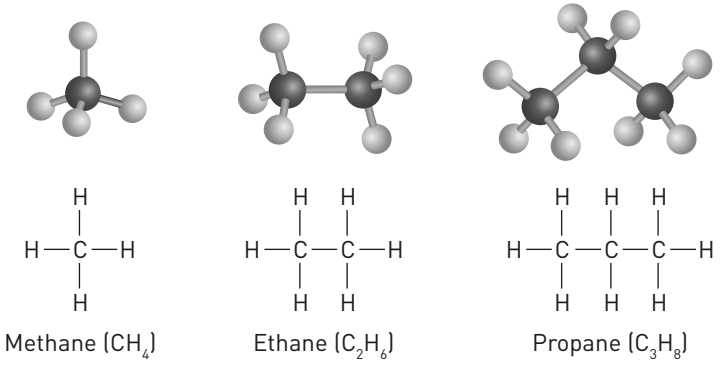


FIGURE 3 Structures of the alkanes methane, ethane and propane.

Alkanes experience only dispersion forces because there is no overall polarity throughout the molecules. All carbons are bonded and there are no lone pairs of electrons. Figure 4 demonstrates the relationship between the number of carbons in an alkane and its boiling point. As the number of carbon atoms increases, the boiling point of the molecule also increases. This is because increasing the number of carbon atoms means an increase in dispersion forces, which means more heat energy is needed to separate the weakly interacting molecules.

Primary alcohols are a family of molecules with an OH group attached to one end (Figure 5).

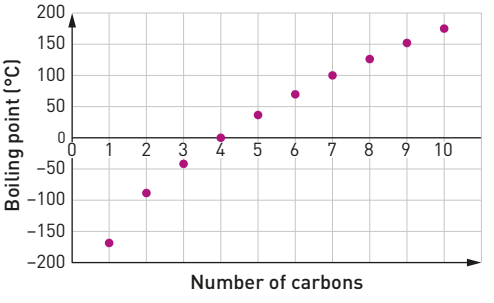


FIGURE 4 Boiling points of straight chain alkanes containing 1–10 carbon atoms.

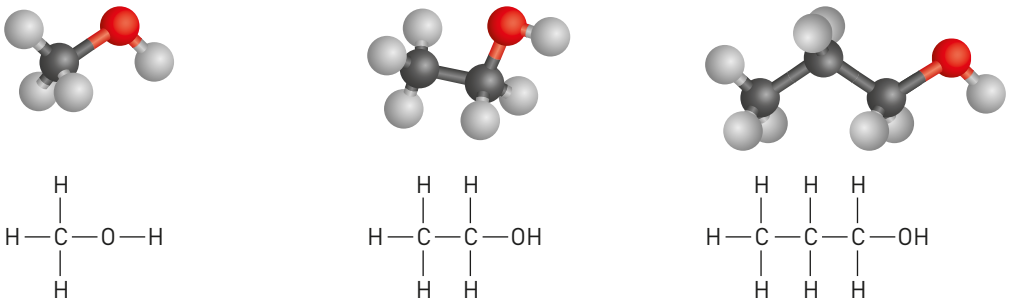


FIGURE 5 Structures of the primary alcohols methanol, ethanol and 1-propanol.

The presence of the oxygen atom and the shape of the molecule make primary alcohols polar and they undergo hydrogen bonding with other alcohol molecules. However, because they contain a carbon chain, one end of the molecule experiences a small amount of dispersion forces. All of the molecules in this family contain the same amount of hydrogen bonding because they have the same polar OH group. Therefore, any differences in boiling points are not due to the polarity of the OH group, but to the length of the carbon chain and therefore the amount of dispersion forces (Figure 6).

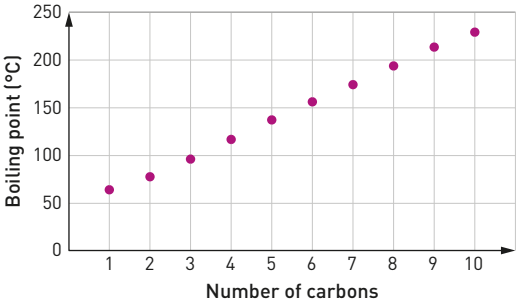


FIGURE 6 Boiling points of primary alcohols containing 1–10 carbon atoms.

When the two sets of data are plotted on the same graph, the trend between molecules that have different intermolecular forces can be seen. Because polar molecules that experience hydrogen bonding are more attracted to one another, they are harder to break apart than molecules that contain dispersion forces. Therefore, converting molecules with dispersion forces (alkanes) from a liquid to a gaseous state requires less energy, so they will have lower boiling points (Figure 7) than primary alcohols.

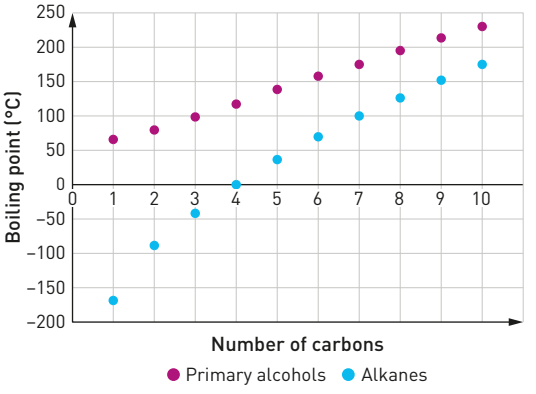


FIGURE 7 Boiling points of straight chain alkanes, containing dispersion forces only, and primary alcohols, containing both dipole-dipole and dispersion forces.

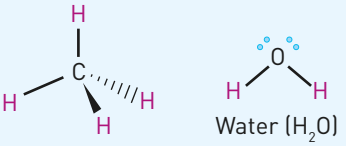
The melting points of covalent molecules display the same trends as boiling points based on the strength of the intermolecular forces between molecules.

WORKED EXAMPLE 12.4A

Determine which molecule has a lower melting point, methane (CH_4) or water (H_2O).

SOLUTION

1 Based on your knowledge of molecular structure, determine the shape of the molecules.

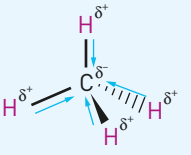


Methane (CH_4) Water (H_2O)

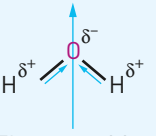
FIGURE 8 Structures of methane and water.

2 Determine whether the molecules are polar. Methane is non-polar because its symmetrical tetrahedral structure cancels out the effect of the four C–H bonds, which only have small electronegativity differences.

Water is polar because oxygen pulls electrons in the two O–H bonds, resulting in a partial negative charge on the oxygen end and a partial positive charge on the hydrogen end.



Electronegative



Electropositive

3 Consider the properties of the molecules and determine which intermolecular bonds are involved between molecules.

Methane experiences dispersion forces because it is non-polar.

Water is polar and has a permanent dipole. The atoms involved in the intermolecular interactions are hydrogen and oxygen, so hydrogen bonding occurs (and dispersion forces).

4 Consider the strength of the intermolecular forces and determine which molecule has the higher melting point.

The dispersion forces between methane molecules are much weaker than the hydrogen bonding between water molecules. Therefore, it takes less energy to break methane molecules apart, resulting in a lower melting point.

Vapour pressure

Vapour pressure is measured as the **pressure** of a **vapour** in contact with its solid or liquid form. The **evaporation** of water is a common example of this phenomenon. Evaporation is the phase change from a liquid to a gas. Water evaporates off the top of liquid water to form water vapour (gaseous water).

Temperature is a measure of the average kinetic energy of a system. To know the energy of every particle or molecule, the **Maxwell–Boltzmann distribution** can be used. This represents the kinetic energies of all particles in a system and forms a normal distribution (Figure 9).

According to the distribution, there will be a small number of particles that, randomly, have a high amount of energy. If this energy is greater than the energy involved with their intermolecular forces, the particles break free of the liquid phase to become a gas. At higher temperatures, the distribution shifts to the right and a greater portion of particles have a higher energy.

pressure
the force exerted, per unit area, by one substance upon another substance

vapour
a substance in the gaseous phase

evaporation
the phase change from a liquid to a gas

temperature
a measure of the average heat energy of the particles within a system

Maxwell–Boltzmann distribution
a representation of the various kinetic energies of each individual particle in a system

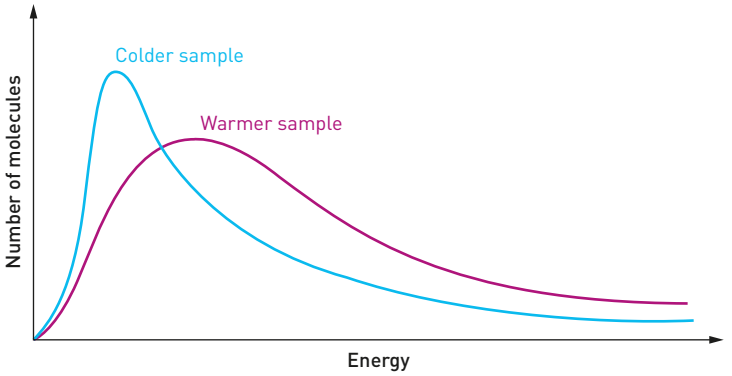


FIGURE 9 The Maxwell–Boltzmann distribution of particles energies within a system.

vaporise
to change from a liquid to a gas

condense
to change from a gas to a liquid

closed system
a system from which gas particles cannot escape

volatile substance
a substance that easily evaporates from a liquid to a gas

dissolve
when a substance dissolves in a solvent

In a container of water, only the particles on the surface of the water can **vaporise**. If a water molecule at the bottom of the container randomly gained enough energy to vaporise, it would still need to move to the surface before it could leave the liquid phase to become a gas. However, the molecules at the surface of the water would experience the same distribution of energies.

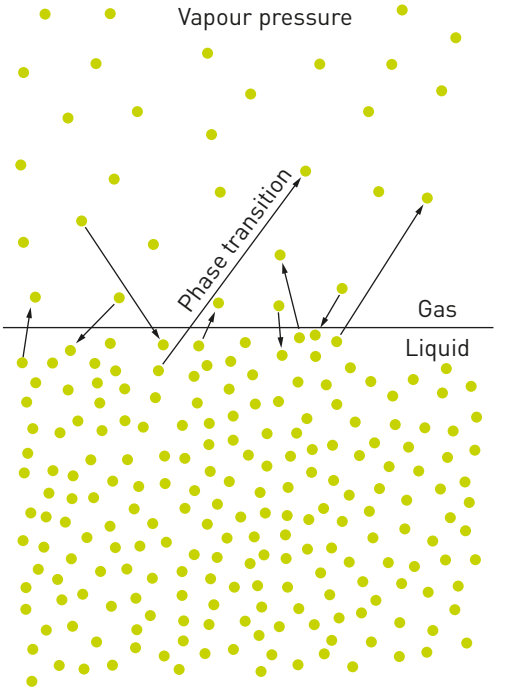


FIGURE 10 Molecules moving from liquid to gas phase and back at the same rate, resulting in vapour pressure.

When a water molecule vaporises, it enters the space just above the water. As more water molecules vaporise, this space contains more and more water molecules.

Chapter 14 explains the properties of gases, one of which is that gases exert pressure. The more water molecules that enter the gas phase, the higher the pressure that they exert.

Molecules within the atmosphere are gaseous and move with high kinetic energy. They can also be represented using the Maxwell–Boltzmann distribution. Low-energy water particles within the atmosphere can **condense** from a gas back into a liquid. This happens more often in a **closed system**, where a lid is placed on the container and none of the water vapour can escape.

Eventually, the rates at which water molecules vaporise (liquid to gaseous) and condense (gaseous to liquid) are equal. The pressure that the water vapour exerts when these rates are equal is called vapour pressure (Figure 10).

The types of covalent molecules that are the most likely to evaporate have:

- high kinetic energy
- weak intermolecular forces.

Dispersion forces are the weakest forces, so the covalent molecules that interact by dispersion forces are more likely to vaporise than molecules that also interact by stronger dipole–dipole interactions or hydrogen bonding.

A substance that readily vaporises is called a **volatile substance**. It has more molecules entering the gas phase and therefore exerts a higher vapour pressure. The more molecules in the gas phase, the higher the pressure that they exert.

Solubility

Water molecules are strongly attracted to other water molecules because of their polarity (partial positive and negative charges). Covalent compounds can **dissolve** in solvents that have similar properties. This is called the 'like dissolves like' rule. A polar substance dissolves in a polar solvent because they both contain partial charges. Therefore, polar substances dissolve in water because their permanent dipoles strongly attract each other.

Non-polar molecules form layers on top of or underneath polar molecules (depending on their relative densities) because the polar molecules experience greater interaction with each other than they do with the non-polar molecules. Figure 11 shows how oil forms a layer on top of water. This happens for two reasons.

- The hydrogen bonding between the water molecules is too strong for any oil molecule to break through and mix into the water.
- Oil, despite being a larger molecule, has a lower density than water and forms the top layer. This is because the hydrogen bonding between water molecules makes then more strongly attracted to each other and therefore they are packed tighter, making water more dense.

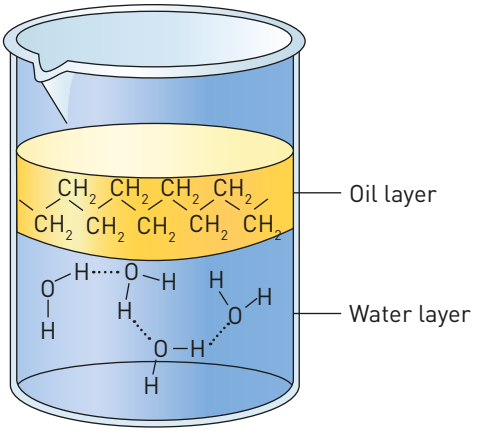


FIGURE 11 Layers form when polar water molecules attract more strongly to themselves than to non-polar oil molecules.

WORKED EXAMPLE 12.4B

Does ethanol (C_2H_5OH) dissolve in water? Explain why.

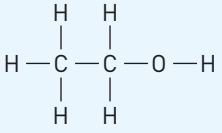


FIGURE 12 Ethanol

SOLUTION

1 Consider the atoms and shape of the molecules to determine the intermolecular forces.

Ethanol is polar because it contains an OH group on one end. It has a molecular dipole and is capable of hydrogen bonding.

Water is polar because it contains two O–H bonds and its shape results in a partial negative charge on oxygen and partial positive charges on the hydrogen atoms. It has a molecular dipole and is capable of hydrogen bonding.

2 If ethanol and water have similar intermolecular properties, they will have higher interaction and ethanol will dissolve in water.

Ethanol (the solute) dissolves in water (the solvent) because they interact by hydrogen bonding to form a solution of diluted ethanol.

Study tip

The like with like, rule or like dissolves like rule, will help you remember that substances dissolve in solvents with similar properties.

WORKED EXAMPLE 12.4C

Does octane (C_8H_{18}) dissolve in water? Explain why.

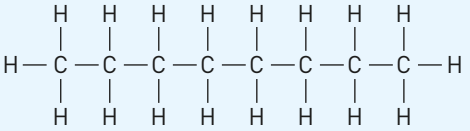


FIGURE 13 Octane

SOLUTION

1 Consider the atoms and shape of the molecules to determine the intermolecular forces.

Octane is non-polar because it contains only C–C and C–H bonds. It experiences dispersion forces.

Water is polar because it contains two O–H bonds and its shape results in a partial negative charge on oxygen and partial positive charges on the hydrogen atoms. It has a molecular dipole and is capable of hydrogen bonding.

2 If octane and water have dissimilar intermolecular properties, they will not interact and will instead form layers with the more dense substance as the bottom layer.

Octane does not dissolve in water because it experiences dispersion forces and has no charge for the water to interact with. It forms a layer on top of water because it is a less dense liquid.

CASE STUDY 12.4

Water only boils at 100°C at sea level

Sometimes water boils below 100°C. This is a result of vapour pressure and atmospheric pressure.

Atmospheric pressure is the collision of air molecules against the surface of Earth. The more collisions there are, the more pressure is exerted. At higher altitudes, pressure decreases because the air is less dense and there are fewer collisions with Earth.

When water molecules evaporate, they push up against the gas molecules in the atmosphere, reducing the amount of pressure pushing down on the water. This makes it easier for the water to form a gas and for the bubble to rise to the surface.

Therefore, boiling is more likely when vapour pressure increases (allowing bubbles to rise) or atmospheric pressure decreases (exerting less pressure on the water). The boiling point occurs when vapour pressure is equal to atmospheric pressure. Different substances have different intermolecular forces and therefore vapour pressures, so they also have different boiling points.

The boiling point of water, 100°C, is only theoretical and occurs when the vapour pressure of water is equal to the atmospheric pressure at sea level, which is 1 atmosphere (1 atm) or 101.325 kilopascals (101.325 kPa). The graph in Figure 14 demonstrates the relationship between vapour pressure and the boiling point of water. If atmospheric pressure is reduced, the boiling point of water also decreases.

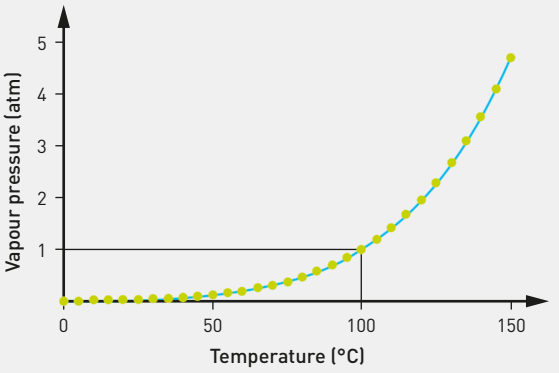


FIGURE 14 The relationship between vapour pressure (atm) and boiling point (°C) of water

CHECK YOUR LEARNING 12.4

Describe and explain

1 **Explain** why kinetic energy affects the ability of a substance to melt or boil.

2 **Explain** how vapour pressure relates to the boiling point of a substance.

Apply, analyse and interpret

3 **Determine** which of the following molecules has a greater solubility in water. Provide reasoning to justify why.

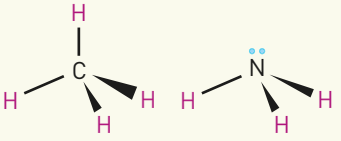


FIGURE 15 Valence models of methane and ammonia

4 Figure 16 shows the structure of the glucose molecule. Use this structure, to **deduce** whether glucose will dissolve in:

a water (H_2O)

b decane ($C_{10}H_{22}$).

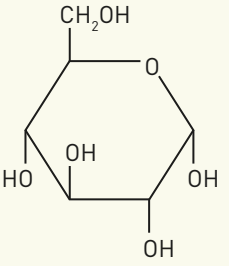


FIGURE 16 The structure of the glucose molecule.

Investigate, evaluate and communicate

5 A student claims that ammonia has a higher solubility than hydrochloric acid in water.

a **Evaluate** the validity of this claim.

b **Investigate** how solubility values for each chemical to support your answer.

6 **Investigate** how the vapour pressure of a solid–liquid differs from that of a liquid–gas.

7 **Investigate** the concept of 'equilibrium' and explain how it relates to vapour pressure.

You can find the following resources for this section on your obook assess:

» Student book questions
Check your learning 12.4

» Weblink
Melting and boiling point

» Weblink
Vapour pressure

» Weblink
Solubility

CHAPTER 12

Review

Chapter summary

12.1
- Covalent bonds are the sharing of electrons between two non-metallic atoms. VSEPR theory states that any bonding and lone pairs of electrons will repel each other to give covalent molecules a three-dimensional shape. This shape depends on the presence of lone pairs of electrons, bonding electrons and whether these bonds are single, double or triple.

12.2
- Electronegative atoms pull on the electrons in a covalent bond, resulting in a partial negative charge at one end of the bond and a partial positive charge at the other end. A molecule's shape may cancel out the polarity of bonds, resulting in a non-polar (uncharged) molecule. If the bonds do not cancel out, then the electrons are pulled to one side of the molecule, resulting in a polar molecule with a dipole (a partial positive and negative charge at opposite ends). Molecules with permanent partial charge separations interact (intermolecular bonding) with dipole–dipole interaction. If a hydrogen atom interacts with fluorine, nitrogen or oxygen, the dipole–dipole interaction is very strong and is called a hydrogen bond. Non-polar molecules interact with dispersion forces.

12.3
- The intermolecular bonding of molecules determines their properties and behaviours. Molecules with stronger intermolecular forces are harder to break apart. Heat energy (kinetic energy) will break these forces. When kinetic energy is greater than the energy in the intermolecular force, molecules move apart.
- Kinetic molecular theory determines the melting point, boiling point and vapour pressure of covalent compounds.
- Solubility is determined by the attraction of the solute to the solvent. A solute that is strongly attracted dissolves more readily. If the solute and solvent do not have similar properties, they will not attract and the solute will not dissolve.

12.4
- The VSEPR theory has been developed from the research of many chemists over more than 200 years. Chemists use two and three-dimensional models of molecules to determine the interaction between molecules, their properties and behaviours in various situations.

Key terms

- bent
- bond angle
- bonding electron
- closed system
- computational chemistry
- condense
- dipole
- dipole–dipole interaction
- dispersion force
- dissolve
- electrostatic interaction
- evaporation
- hydrogen bonding
- intermolecular force
- kinetic energy
- linear
- lone pair of electrons
- Maxwell–Boltzmann distribution
- molecular shape
- non-bonding electron
- polar
- polarity
- pressure
- skeletal structure
- structural formula
- temperature
- tetrahedral
- three-dimensional modelling
- trigonal planar
- trigonal pyramidal
- two-dimensional modelling
- valence shell electron pair repulsion (VSEPR) theory
- vaporise
- vapour
- vapour pressure
- volatile substance

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; ***=high.

Multiple choice

1 The molecule ammonia is:

A tetrahedral and polar
B pyramidal and non-polar
C pyramidal and polar
D tetrahedral and non-polar.

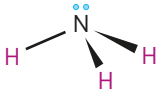


FIGURE 1 The ammonia molecule

2 In liquid hydrogen (H_2), the intermolecular forces present are?

A Dispersion forces
B Dipole–dipole interactions
C Hydrogen bonding
D Ion–dipole interactions.

3 Vapour pressure is greatest when:

A a substance has minimal evaporation
B atmospheric pressure is increased
C a substance has weak intermolecular forces
D a substance has low kinetic energy.

4 To melt a covalent substance, it is best to:

A maximise kinetic energy and intermolecular forces
B maximise kinetic energy and minimise intermolecular forces
C minimise kinetic energy and maximise intermolecular forces
D minimise kinetic energy and minimise intermolecular forces.

5 What is the bond angle in the water molecule?

A 104.5°
B 105.5°
C 106.5°
D 107.5°

6 Which element has the highest electronegativity?

A Carbon
B Iron
C Lithium
D Magnesium

7 Which element has no known compound?

A Helium
B Radium
C Uranium
D Sulfur

8 Which alcohol has two carbon atoms?

A Butanol
B Ethanol
C Methanol
D Propanol

Short answer

Describe and explain

★ **9** **Explain** the difference between melting point and boiling point by using the terms 'kinetic energy' and 'intermolecular forces'.

★ **10** **Identify** and explain which type of intermolecular bonding is the strongest.

★ **11** **Explain** how non-polar molecules are able to interact if they have no overall charges.

★ **12** **Explain** the difference in melting points of sodium chloride (NaCl) 801°C, water (H_2O) 0°C and methane (CH_4) –182°C.

★ **13** Do substances with higher boiling points have stronger or weaker intermolecular forces? **Explain** why.

★ **14** **Explain** the relationship between the molar mass of an alkane and its boiling point.

★ **15** A student claims that all substances experience intermolecular forces. Are they correct? **Explain** your reasoning.

★ **16** **Explain** the intermolecular forces present in the molecule DNA and why it is essential to the replication of the molecule.

★★ **17** **Sketch** a valence model of the following molecules, indicating which atoms are electropositive and which are electronegative, the direction of electron pull and the overall effect. Are the molecules polar or non-polar?

a CCl_4
b HF
c HCN
d BH_3
e CH_3NH_2

★★ **18** Carbon dioxide is a gas at room temperature and a solid (dry ice) at –78.5°C. When dry ice is left on a laboratory bench, it turns into a gas in a process called sublimation. **Explain** why carbon dioxide turns from a solid to a gas, skipping the liquid phase.

FIGURE 2 Dry ice

★★ **19** **Explain** why ethanol has a higher vapour pressure than water at any given temperature.

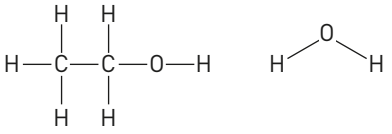


FIGURE 3 Structures of ethanol and water

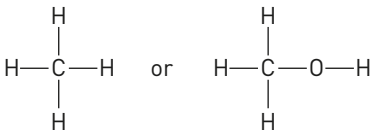
★★ **20** **Explain** why methane (CH_4) has a boiling temperature of –161.5°C and is a gas at room temperature, whereas octane (C_8H_{18}) has a boiling temperature of 125.7°C and is a liquid at room temperature.

Apply, analyse and interpret

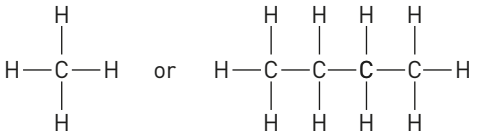
★ **21** **Determine** how a covalent bond is different from an intermolecular force.

★★ **22** **Determine** which of the following pairs of molecules would have a higher boiling point? **Explain** your reasoning.

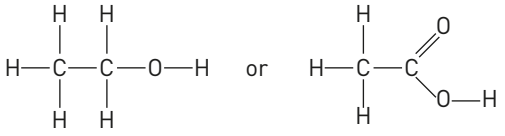
A



B



C



★★ **23** Complete the following table.

TABLE 1 Substances and their properties

Substance	Shape	Polar or non-polar?	Type of inter-molecular bonding
HF			
H_2S			
PF_3			
CO_2			
CCl_4			
BF_3			
O_2			
SO_2			
NH_3			

★★ **24** **Consider** which of the molecular shapes discussed in Section 12.1 result in a polar compound. Use an example for each shape to **justify** your answer.

★★ **25** **Determine** the type of intermolecular bonding that exists between the following pairs of molecules.

a H_2S and H_2O
b CO_2 and N_2
c NH_3 and H_2O
d CH_3OH and CH_3CH_2COOH

★★ **26** Using your knowledge of the structure of water and its intermolecular bonding, **consider** why water expands when it is in solid form and contracts when it is in liquid form.

FIGURE 4 A bottle shattered by ice

★★ **27** **Discuss** why there is little difference in melting points of fluorine (F_2) –220°C, oxygen (O_2) –218.8°C and nitrogen (N_2) –210°C.

★★★ **28** Figure 4 represents the skeletal structure of four fatty acids. Each bend represents a carbon atom; hydrogen atoms are not represented.

- **a** **Explain** how the VSEPR model has helped chemists understand the behaviour of fatty acids.
- **b** **Classify** each of the fatty acids as saturated (no double bonds) or unsaturated (at least one double bond).
- **c** **Classify** each of the unsaturated fatty acids as *cis* or *trans*.
- **d** **Explain** the benefit in drawing a skeletal 3D structure over the 2D structure.

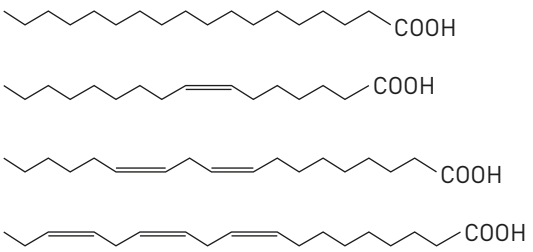


FIGURE 5 Skeletal structures of four fatty acids

Investigate, evaluate and communicate

★★ **29** **Investigate** whether it is possible to form a quadruple bond between two carbon atoms.

★★ **30** **Investigate** why fluorine is the most electronegative element on the periodic table.

★★ **31** The boiling point of 1-propanol is 97°C and that of ethyl methyl ether is 7.4°C (Figure 5). The two molecules have the same number of carbon, oxygen and hydrogen atoms. **Evaluate** the molecules and justify the difference in their boiling points.

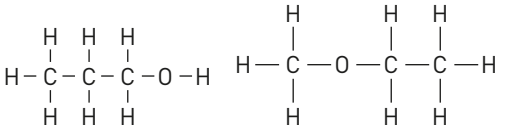


FIGURE 6 Structures of 1-propanol and ethyl methyl ether

★★ **32** Cellulose is a polymer of the molecule glucose.

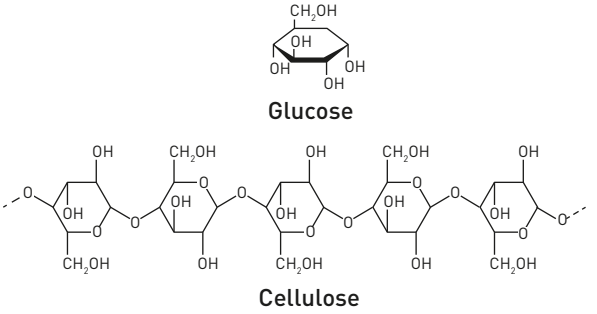


FIGURE 7 Structure of cellulose

Cellulose makes up the cell wall of plants and therefore forms a very rigid structure. This structure is insoluble in water, whereas the glucose molecule is highly soluble in water. **Hypothesise** why is glucose is soluble, but cellulose is insoluble.

FIGURE 8 Plants

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CHAPTER 13

Chromatography techniques

The word *chromatography* is derived from the Greek words for colour and write. It quite literally means 'to write with colour'. Early chromatography was based on these principles.

Chromatography is based on the broad field of separation sciences. Very simply, this field of science focuses on how to separate the chemicals contained within a mixture (Figure 1). Early chromatography was able to separate photosynthetic plant pigments into four coloured components: orange, yellow, blue/green and green.

Not long after this was achieved, it was discovered that chromatography could be applied to most other mixtures of chemicals, regardless of whether they were coloured or not.

Chromatography is essential within the field of chemical analysis. Today, chromatographic techniques focus on identifying the components of a mixture and determining how much of each component exists within a sample.

To fully understand the chromatographic techniques involved with this type of chemistry and the principles that underline them, it is important to have a good knowledge of the intermolecular forces covered in Chapter 12.

OBJECTIVES

→ Recognise that chromatography techniques, including paper, thin-layer, gas and high-performance liquid chromatography, can be used to determine the composition and purity of substances.

→ Describe and explain how variations in the strength of the interactions between atoms, molecules or ions in the mobile and stationary phases can be used to separate components.

→ Analyse, interpret and evaluate data from chromatographs to determine the composition and purity of substances, including calculating R_f values.

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PRACTICAL

SUGGESTED PRACTICAL

13.2 Separating the components of a mixture using paper chromatography

FIGURE 1 Chromotography separates the components of mixtures.

13.1 Principles of chromatography

KEY IDEAS

in this section, you will learn about:

- interactions between atoms, molecules or ions in the mobile and stationary phases of chromatography
- separating atoms, molecules or ions in a mixture
- describing separation processes in chromatography by using the terms 'affinity', 'adsorb' and 'desorb'.

mobile phase
the phase that flows in chromatography, moving the components of a sample at different rates over the stationary phase

solvent
the major component of solution – a substance that can dissolve other substances

stationary phase
the phase to which the components of a chromatographic sample are adsorbed

solute
the minor components of a solution – a substance dissolved in a solvent

affinity
the interaction of a substance within a sample with the mobile or stationary phases

adsorption
the attraction of a substance within a sample to the stationary phase

desorption
the release of a substance within a sample from the stationary phase into the mobile phase

Chromatography is an analytical technique that can be used for qualitative or quantitative analysis. It can both identify components in a mixture and determine how much of each component is present. It is a useful form of separation science that uses intermolecular forces (Chapter 12), to separate the components of a mixture according to their properties.

Chromatography has a wide range of uses, including in:

- pharmaceuticals science – identifying elements and molecules in drugs as well as measuring the purity of medicines
- environmental science – analysing water quality and gases in the air and the impact they may have on Earth
- police work – analysis such as breathalysers and drug detection
- forensic science and archaeology – analysing material such as blood and hair samples
- biomedical research – analysing material such as proteins in cancer research
- food science – analysing the nutritional quality of food and food spoilage.

Stationary and mobile phases

All chromatography uses two phases to separate a mixture. These are the:

- **mobile phase**, which is a **solvent**, that moves over or through the stationary phase
- **stationary phase**, which is either a solid with a high surface area or a liquid coated onto a solid support. It always stays still.

The mobile phase moves through and carries a sample (**solute**) over the stationary phase.

Components of the sample are more attracted to either the mobile or the stationary phase, depending on their intermolecular forces. This makes the components move at different rates and separates them. Several key terms are used to describe this attraction (Figure 1).

- **Affinity**: components of a sample are attracted, by intermolecular forces, to the mobile or stationary phase. High affinity means that the component interacts strongly with the mobile or stationary phase.
- **Adsorption**: components of the sample adsorb onto the stationary phase from the mobile phase if they are attracted to, or have an affinity for, the stationary phase.
- **Desorption**: components of the sample desorb off the stationary phase into the mobile phase if they are attracted to, or have an affinity for, the mobile phase.

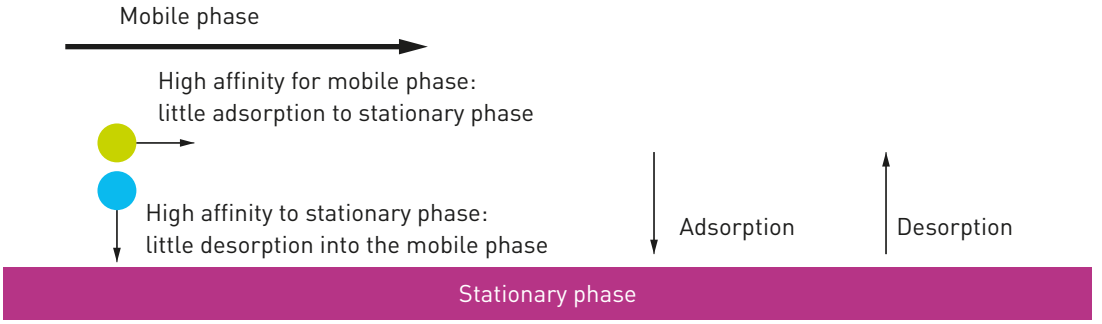

FIGURE 1 The blue and green components in this sample can be separated depending on their affinity for the mobile and stationary phases.

The 'like with like' rule

As a general rule, atoms, molecules and ions are more attracted to the mobile or stationary phase that has the same or similar properties. Therefore, like goes with like. For example, ions contain one or more whole charges. Because of this, they are more attracted to the mobile phase if it contains charged or partially charged molecules or ions, such as water (a polar molecule).

Separation based on carbon chain length or molecular mass

Chemists often need to separate mixtures of molecules that have similar intermolecular forces such as mixtures of alcohols.

In a mixture of methanol, ethanol, propan-1-ol and butan-1-ol (Figure 2), all of the molecules have the same polar –OH group on the terminal, or end, carbon. Therefore, this mixture will not be separated into its components through this functional group. Rather, the mixture is separated by the size of the carbon chain – the longer the carbon chain, the greater the dispersion forces and the higher the interaction with the mobile or stationary phases.

Components are separated according to differences in the relative competition for them by the mobile phase and the stationary phase.

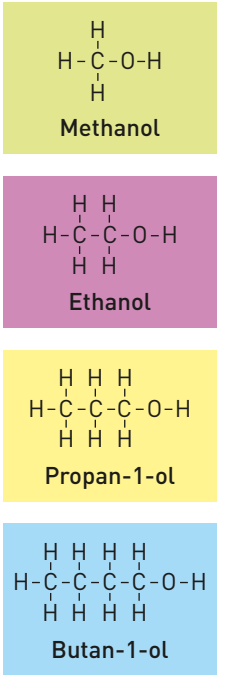

FIGURE 2 These alcohols that have a terminal polar OH group.

Study tip

Chromatography is based on intermolecular forces, so, when answering questions, start by determining which intermolecular forces are involved. Once you have determined the intermolecular forces, explain their affinity to the mobile or stationary phase. Then explain the effect this will have on the data you obtain (data will be covered in Sections 13.2 and 13.3).

Study tip

Adsorption is when a substance sticks to the surface to form a film. Absorption is when a substance penetrates a solid or liquid.

CHECK YOUR LEARNING 13.1

Describe and explain

1 **Explain** the following terms and why they are essential to the separation of mixtures.
- **a** mobile phase
- **b** stationary phase
- **c** affinity
- **d** intermolecular forces

Apply, analyse and interpret

2 **Discuss** which intermolecular forces, studied in Chapter 12, could be used to separate a mixture of substances.

Investigate, evaluate and communicate

3 A mobile phase chosen in an analysis contains only dispersion forces. **Determine** what intermolecular forces the stationary phase should have. **Justify** why.

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13.2 Paper and thin-layer chromatography

KEY IDEAS

In this section, you will learn about:

- separating substances by paper and thin-layer chromatography (TLC)
- calculating R_f values to identify substances within a mixture
- manipulating mobile and stationary phases to optimise separation.

paper chromatography
an analytical technique for separating and identifying mixtures; the stationary phase is a thin strip of absorbent paper

thin-layer chromatography (TLC)
an analytical technique for separating and identifying mixtures; the stationary phase is typically a thin layer of silica gel, aluminium oxide or cellulose supported on a piece of glass or plastic

chromatogram
the pattern of bands, spots or peaks formed on the chromatography paper or TLC plate demonstrating the separation of a mixture

origin
the line applied to a chromatogram to mark the point where the sample or standard was placed

solvent front
the point on a chromatogram where the mobile phase reaches before the analysis is terminated

In **paper chromatography** and **thin-layer chromatography (TLC)**, a sample containing a mixture of substances is applied to the stationary phase as a spot. The spot is carried up the stationary phase by the mobile phase. The resulting **chromatogram** is usually qualitative in nature – it can only be used to identify substances, not determine how much is present.

Paper chromatography

In paper chromatography, the stationary phase is a thin strip of absorbent paper, which is cut to fit inside a container that holds the mobile phase (Figure 1). A pencil line is ruled across the bottom of the paper to mark where the sample will be placed – this is called the **origin**. Pen or ink should not be used to mark the origin because inks will separate and contaminate the sample. The sample is placed on the paper and standards of known identity can also be added to help identify the substances in the sample. In Figure 1 the standards are marked A, B and C.

The mobile phase is added to the container so that it sits *below* the origin line. The paper is then placed into the container and left to adsorb the mobile phase. The sample travels up the piece of paper, separating into its components.

Components of the sample that have a strong affinity for the stationary phase move slowly. On the resulting chromatogram, they have not moved far from the origin. Components that have a higher affinity for the mobile phase move faster. On the resulting chromatogram, they have moved further and are located closer to the **solvent front** (Figure 1).

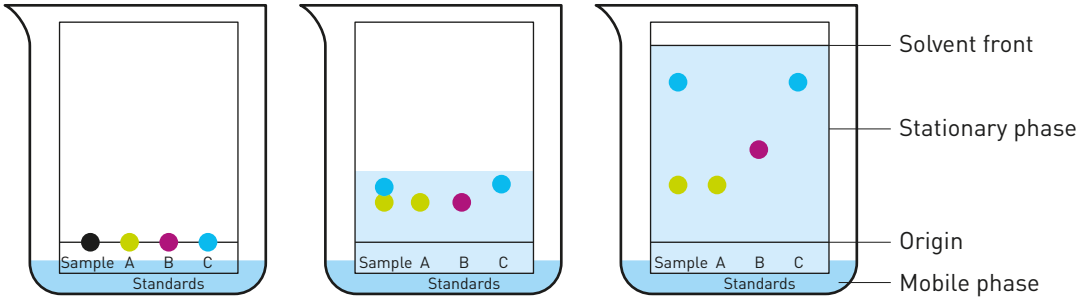


FIGURE 1 The water mobile phase moves up a piece of filter paper (stationary phase) to separate the ink in black marker pen. Components in the sample can be identified by comparing them to known standards A–C.

In Figure 1, the mobile phase is water and the stationary phase is filter paper. Paper is a derivative of cellulose, which contains many polar –OH groups. However, very few of the intermolecular forces of paper extend beyond its network of fibres (and any surface coating).

Water is also a polar molecule, but more polar than the paper. Therefore, any component of the sample that has whole or partial charges, such as ionic or polar substances, will have a higher affinity, or attraction, for the water mobile phase. These components move further from the origin. Any component of the sample that experiences dispersion forces has a higher affinity for the stationary phase and will not move as far from the origin.

FIGURE 2 Paper chromatography can be used to separate the components of photosynthetic pigment (leaf stain). The pigment is placed on the origin, and then the paper strip is placed in a container where it is in contact with the mobile phase, which is drawn up the paper by capillary action.

Retardation factor (R_f) calculations

Once the separation is completed, the resulting chromatogram is analysed to identify the substances in the sample and to determine the purity of the sample. Although we can simply look at a chromatogram to judge whether two substances are identical, it is more precise to calculate the **retardation factor (R_f)** of a substance. This is the ratio of the distance moved by a substance to the distance moved by the solvent, or mobile phase.

retardation factor (R_f) the ratio of the distance travelled by a component of a sample, from the origin, to the distance travelled by the mobile phase

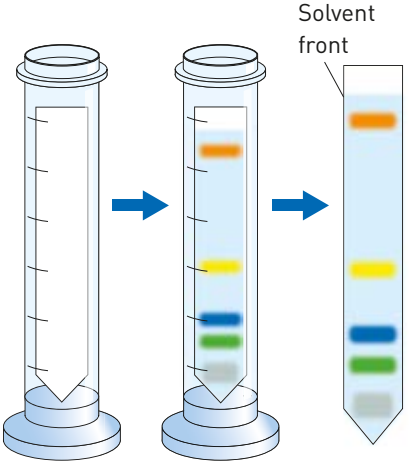
This can be simplified to:

$$R_f = \frac{\text{distance solute moves from origin}}{\text{distance solvent moves from origin}}$$

$$= \frac{\text{distance of sample spot}}{\text{distance of mobile phase}}$$

Each atom, ion or molecule has a unique R_f depending on the properties of the mobile and stationary phases. Increasing the polarity of the mobile phase increases its affinity to charged particles within the sample. This makes the substances move further up the paper.

All R_f values must be expressed as a decimal (never a fraction) and cannot be



Pigment	R_f value
Carotene	0.95
Xanthophyll	0.48
Chlorophyll a	0.30
Chlorophyll b	0.18
Leaf stain	

FIGURE 3 Separation of a photosynthetic pigment into its constituent pigments. The R_f value is calculated from the chromatogram.

greater than or equal to one. An R_f of 1 indicates that the substance has not separated from the mobile phase and therefore cannot be identified. The analysis must be run again under different mobile or stationary phase conditions.

Thin-layer chromatography

TLC works by the same principles as paper chromatography with one major change that makes the process more efficient (faster and more sensitive). In TLC, the support for the stationary phase is a piece of glass or plastic, which is coated in the stationary phase, consisting of silica gel, aluminium oxide or cellulose. This coating is a thin layer spread on the surface of the plastic or glass; hence, the name 'thin-layer' chromatography.

The components within the sample are separated, as with paper chromatography, according to their affinity for the mobile or stationary phase.

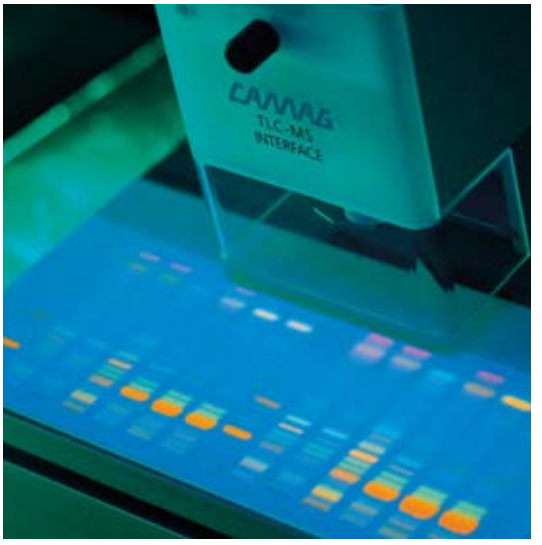


FIGURE 4 TLC performed on fluorescent plates.

It is important to note that both paper chromatography and TLC are not limited to coloured compounds. Fluorescent TLC plates can be used under UV light to see the components of a sample that would not otherwise be visible. These are shown in the darker areas of the plate where the sample blocks the fluorescence of the plate (Figure 4).

Study tip

Never leave R_f as a fraction. This is because it is analysed on a scale of 0–1, with 0 being no affinity to the mobile phase and 1 being no affinity to the stationary phase.

Study tip

Both paper and thin-layer chromatography are named after the stationary phase. This may help with recalling specific information about these techniques.

WORKED EXAMPLE 13.2

A food dye was analysed against three standard dyes A–C to determine which dyes it contained. The resultant chromatogram is shown in Figure 5.

Use the data to determine:

a the retardation factors of the dye standards
b the retardation factors of all dyes within the sample
c the dye that is more strongly attracted to the stationary phase
d the dye that is more strongly attracted to the mobile phase
e which dyes are in the sample
f whether the analysis of this data is valid.

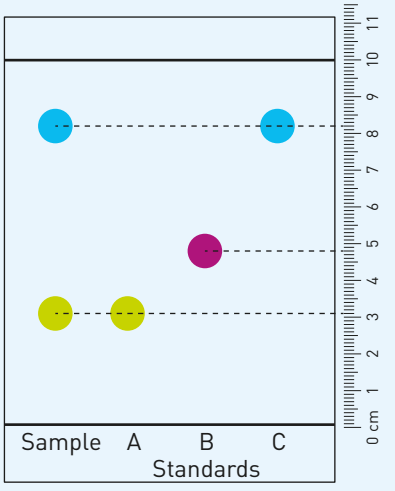


FIGURE 5 The chromatogram from analysing a food dye against three standard dyes.

SOLUTION

a **TABLE 1**

Green dye A:	$R_f = \frac{3.1}{10} = 0.31$
Purple dye B:	$R_f = \frac{4.8}{10} = 0.48$
Blue dye C:	$R_f = \frac{8.2}{10} = 0.82$

b **TABLE 2**

Green dye A:	$R_f = \frac{3.1}{10} = 0.31$
Blue dye C:	$R_f = \frac{8.2}{10} = 0.82$

c Green dye A has the highest affinity to the stationary phase. It has a higher rate of adsorption to the stationary phase because it has moved less distance from the origin.

d Blue dye C has the highest affinity to the mobile phase. It has a higher rate of desorption from the stationary phase because it has moved further from the origin.

e The green dye in the sample and standard A have the same R_f of 0.31. The blue dye in the sample and standard C both have an R_f of 0.82.

f All R_f values are less than 1. This means that all of the substances have separated from the mixture and the data is valid.

Study tip

If $R_f = 1.0$, then no separation from the stationary phase has occurred. A typical exam question may ask how to separate this component in a further analysis. Look at the properties of the stationary phase and determine what you could do to make this substance more attracted to it.

Two-dimensional paper or thin-layer chromatography

Occasionally, a chromatography technique will not separate some components sufficiently within the sample. This is because the components interact the same amount with the mobile and stationary phases. Rather than repeating the analysis, chemists will rotate the chromatogram by 90°, so that the sample is on the bottom. They rule a new origin line and run the analysis again, using a different mobile phase with different properties (Figure 6). The chromatogram is analysed by using R_f calculations for analysis 1 and also for analysis 2.

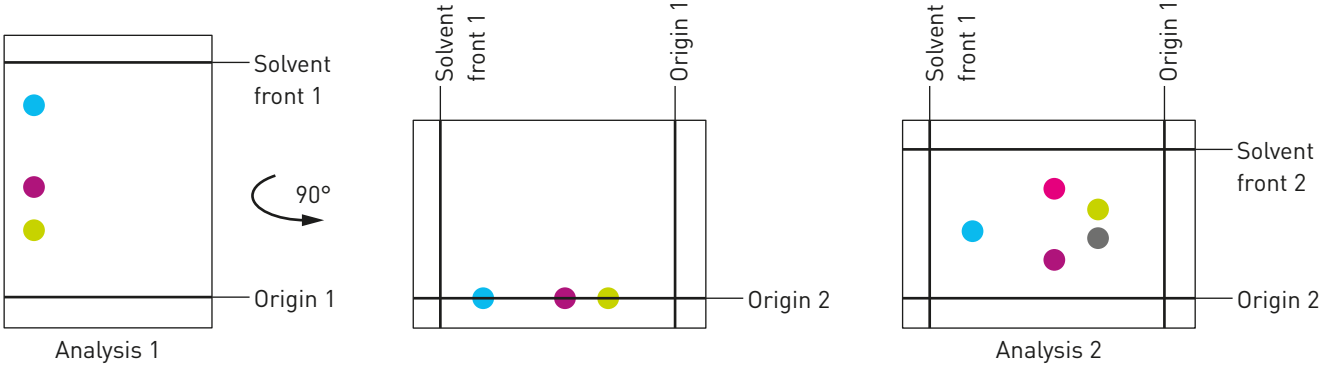


FIGURE 6 Two-dimensional paper chromatography. After the first analysis, the chromatogram is rotated 90°. This gives a new origin line. The second analysis separates the green and purple dots further into their individual components. (Origin 1 and solvent front 1 refer to the first analysis. Origin 2 and solvent front 2 refer to the second analysis.)

CHALLENGE 13.2

Improving separations

A chemist performs a separation under the following conditions.

- Mobile phase: water and 10% ethanol
- Stationary phase: paper

1 What conclusions can you make about the properties of the components of the sample?

2 What alterations would you make to the mobile phase to obtain a better separation of the components?

CASE STUDY 13.2

Separation of amino acids in proteins

Proteins are synthesised by the body when many amino acids are joined into long chains of amino acid residues. Each protein synthesised by the body has a unique function depending on the number and sequence of amino acids in the chain.

In the early 1900s, biochemists were looking for a way to determine the number and sequence of amino acids in proteins. They were largely unsuccessful because of the structure of the amino acid molecules. Amino acids have the same general structure (Figure 7). The only difference between amino acids is the structure of the R group, which gives rise to the different properties of each amino acid and therefore the function of each protein.

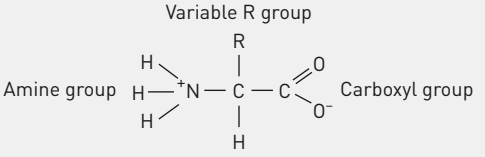


FIGURE 7 The structure of an amino acid.

The problem of separating amino acids was overcome in 1941 when paper chromatography was first developed by British chemist Archer J.P. Martin and biochemist Richard L.M. Synge. Martin and Synge stained the amino acid residues with ninhydrin, a purple dye, and used a piece of filter paper dipped in water as the stationary phase with butan-1-ol as the mobile phase. The characteristic properties of the R group alter the solubility of the amino acid in the mobile phase. Therefore, each amino acid in a protein can be separated by chromatography.

However, some amino acids with similar R groups did not separate. The scientists then decided to rotate the paper by 90° and run it again with a different mobile phase. Therefore not only paper chromatography, but two-dimensional paper chromatography, was born.

Martin and Synge's research became popular throughout the late 1940s and 1950s with major developments in biochemistry being made by using paper chromatography. British chemist Frederick Sanger used this research to determine the sequence of amino acids in insulin, a biochemical that is essential to blood sugar regulation within the body.

Martin and Synge won a Nobel Prize for their work in 1952.

CHECK YOUR LEARNING 13.2

Describe and explain

1 **Describe**, making reference to Figure 5, the steps involved in calculating retardation factor.

2 **Explain** how components of a mixture are identified by paper or thin-layer chromatography.

Apply, analyse and interpret

3 Analyse the chromatogram in Figure 8.

- **a** **Calculate** the R_f values for all components.
- **b** Which component has the highest affinity for the mobile phase? **Explain** why.
- **c** Which component has the highest affinity for the stationary phase? **Explain** why.

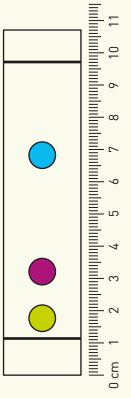


FIGURE 8
A chromatogram

4 If the chromatogram were run with a polar stationary phase, **explain** the types of intermolecular bonding that may be present between each component and the stationary and mobile phases.

5 The R_f values of 10 food dyes are shown in Table 3. These R_f values are specific for TLC using a silica plate and a 1% ethanol mobile phase.

Three foods A–C were tested to **determine** which food dyes were present. The resultant chromatogram is shown in Figure 9.

TABLE 3 R_f values of some food dyes in 1% ethanol

Dye	R_f
Brilliant blue FCF	0.12
Indigotine	0.23
Fast green FCF	0.28
Erythrosine	0.45
Quinoline Yellow	0.50
Carmoisine	0.60
Tartazine	0.65
Green S	0.72
Patent blue V	0.86
Ponceau 4R	0.91

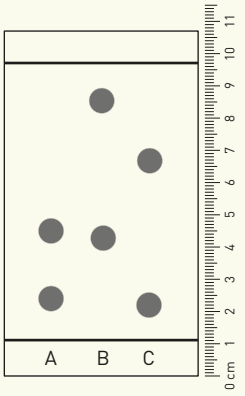


FIGURE 9 A chromatogram of food dyes.

a **Calculate** the R_f values of the components of each sample and determine which food colourings were present in the foods.

b There are two components of these samples that, on visual inspection, appear to be the same. Are these components the same? **Explain** your answer

6 A crime scene note was written in black pen. Two black pens found at the scene were analysed by forensic scientists to determine which pen had written the note. To ensure valid results, five teams were given one of the two pens to test. Each team developed a method that involved using paper chromatography with a water mobile phase. The data in Table 4 was obtained.

TABLE 4 Paper chromatography analysis of pen types.

	Distance (cm) travelled by:			
Team	Solvent front	Component 1	Component 2	Component 3
1	11	6.16	8.69	2.31
2	15	4.80	8.85	13.05
3	9	5.04	7.11	1.89
4	21	11.76	16.59	4.41
5	18	5.76	10.62	15.66
Crime scene note	25	8.00	14.75	21.75

a Use the data to **calculate** the R_f values of components 1–3.

b **Determine** which teams analysed the same pens.

c **Determine** which teams analysed the pen that was responsible for the note at the crime scene.

d Is component 1 the same substance for both pens? **Explain** your answer.

e Is it possible to **identify** the components of the sample from this analysis?

Check your obook assess for these additional resources and more:

» **Student book questions**
Check your learning 13.2

» **Suggested practical**
13.2 Separating the components of a mixture using paper chromatography

» **Challenge**
13.2 Improving separations

» **Weblink**
Thin-layer chromatography

13.3 Column chromatography

KEY IDEAS

In this section, you will learn about:

- the processes involved with separating mixtures by column chromatography
- interpreting gas and high-performance liquid chromatograms
- determining retention times (R_t) to identify substances
- determining concentration using calibration curves.

column chromatography
a technique used to separate and purify individual components from mixtures of compounds

column
the tube-like structure that contains the stationary phase in column chromatography and through which the mobile phase and sample flow

elute
come out of the bottom of a chromatography column

FIGURE 1 Column chromatography in a laboratory.

Column chromatography is a more advanced form of separation science than paper and thin-layer chromatography. It uses the same basic principles. But the stationary phase is contained within a tube, open at both ends.

The simplest form of column chromatography (Figure 1) uses a glass **column**, similar to a burette, which contains a stationary phase of aluminium oxide or silica. The stationary phase resembles finely ground sand. The stationary phase must be packed tightly to minimise pore spaces (gaps). Otherwise, the sample will move too fast and will not interact with the stationary phase. This means that less separation will occur.

The mobile phase is poured into the top of the column and allowed to run slowly down through the stationary phase until it reaches the tap at the bottom. Once the column is soaked in the mobile phase, the sample is added to the top of the column and the tap at the bottom is opened. This allows the mobile phase to run through the column with the sample.

The sample separates into its various components depending on their affinities for the mobile or stationary phases. A higher affinity for the mobile phase means that a component spends longer interacting with it. As the mobile phase flows through the column, the component moves with it and it is collected earlier (Figure 2).

The various components of the sample are collected in separate beakers as they **elute**, or come out of, the column. Once the components are separated, further techniques are used to identify and quantify them.

There are several types of column chromatography. The two most commonly used are **gas chromatography** and **high-performance liquid chromatography (HPLC)**.

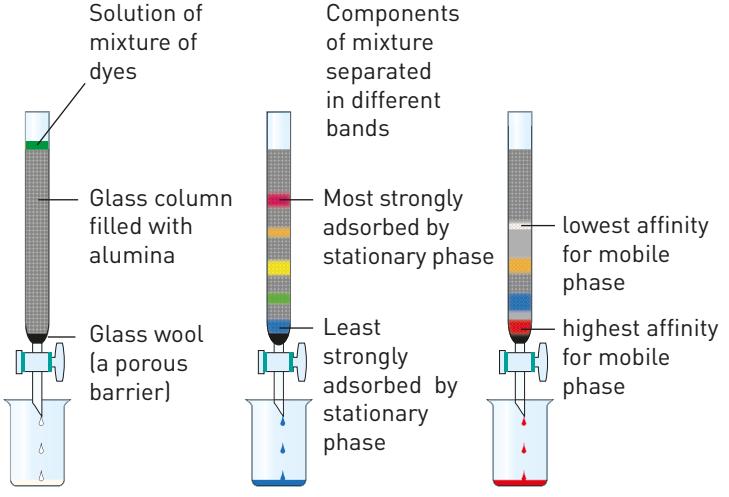


FIGURE 2 Separation of the components of a sample.

Gas chromatography

In gas chromatography, the mobile phase is an inert gas (carrier gas), such as nitrogen (N_2). The sample is **vaporised** and injected into the mobile phase, which carries the sample through a very small (1–2 mm diameter) and very long (up to 60 m or longer) column containing the stationary phase. The stationary phase can be either a solid, packed into the column, or a liquid that coats the inside of the column (Figure 3). Both stationary phases must be able to withstand very high temperatures because the column is contained within an oven to ensure that the sample remains vaporised throughout the separation.

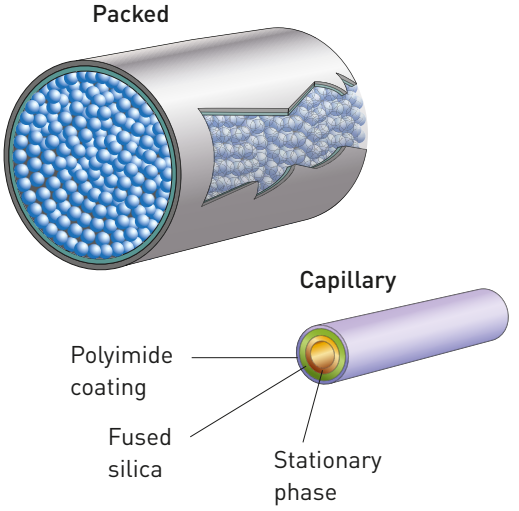


FIGURE 3 The two types of columns used in column chromatography. One is a tube packed with small particles; the other is a much smaller capillary-sized tube, which has a coating on the inside.

The sample separates as each component interacts with the stationary phase. There is no attraction to the mobile phase because it is an inert gas; its purpose is to move the sample through the column. The components with a low affinity for the stationary phase elute from the column first – they are not retained on the column for long. The components with a high affinity for the stationary phase are retained on the column for a longer period of time.

The resulting chromatogram is obtained when the detector recognises the components as they elute from the column. Each component forms a peak in the chromatogram as it elutes (Figure 4).

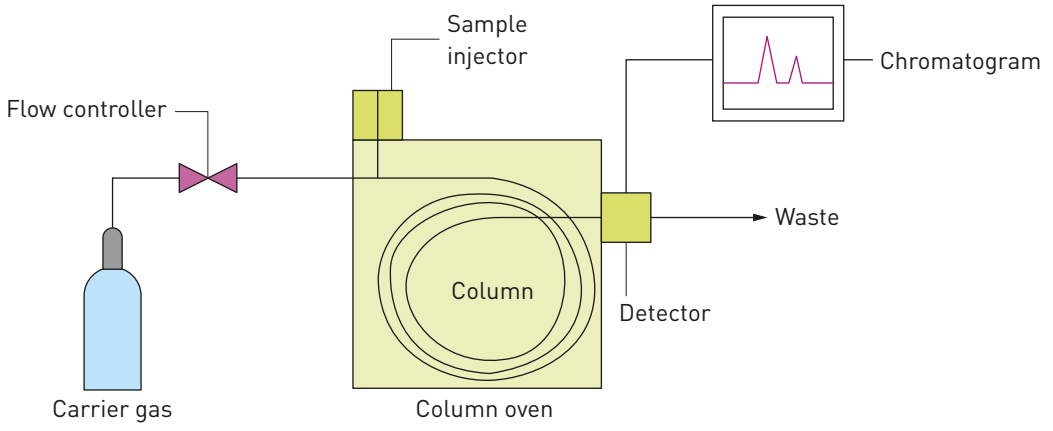


FIGURE 4 A simplified diagram of a gas chromatograph.

Interpreting a gas chromatogram

Figure 5 depicts a gas chromatogram that has been obtained from the separation of a mixture of alcohols. The detector is started at time zero, when the mixture is injected into the column. At 39 seconds, methanol has eluted from the column. Therefore, methanol must have the lowest affinity for the stationary phase. The chromatogram ends with hexanol, which elutes from the column at 162 seconds. Hexanol has the highest affinity for the stationary phase, interacting with it more and eluting last.

gas chromatography
a technique used to separate and purify individual components from mixtures of compounds that can be vaporised

high-performance liquid chromatography (HPLC)
a technique used to separate and purify individual components from mixtures of compounds by pumping a pressurised liquid containing the sample through a column of adsorbent material

vaporise
to change from a liquid to a gas

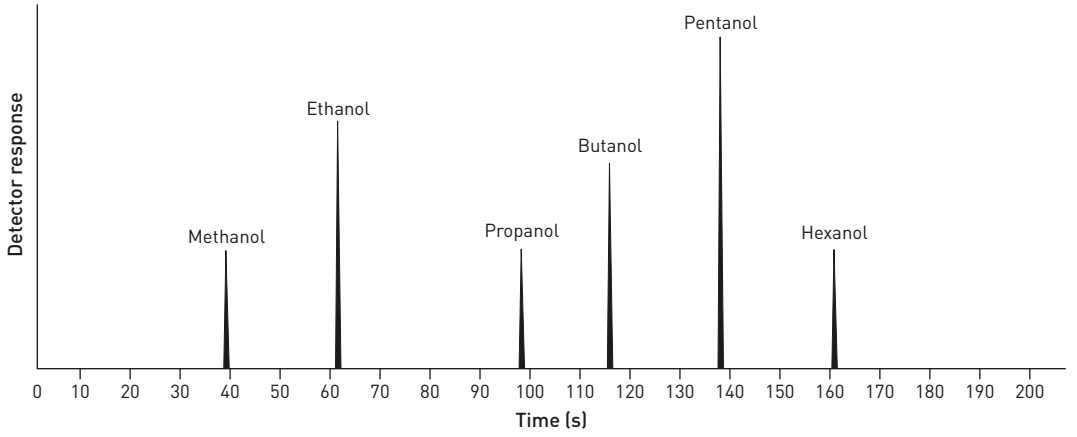


FIGURE 5 A gas chromatogram obtained from the separation of a mixture of alcohols.

Retention time

retention time (R_t) the time that each component is retained by the column

The time that it takes a sample to move through the column, from injection to detection, is called the **retention time (R_t)**. It is the time that each component is retained by the column. Therefore, it indicates the amount of interaction that has occurred between each component of the sample and the stationary phase. Table 1 lists the retention times for the components within the alcohol separation of Figure 5. Figure 6 shows how to measure R_t values.

Study tip

Do not confuse R_t with R_f. In column chromatography, a low R_t means the substance has travelled through the column fast and has a low affinity for the stationary phase. In paper or thin-layer chromatography, a low R_f means the substance has travelled up the stationary phase slowly and has a high affinity for the stationary phase.

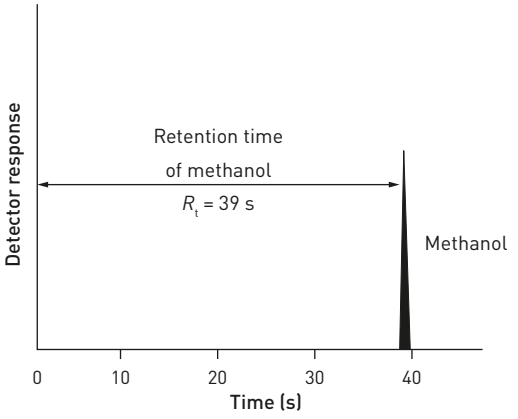


FIGURE 6 The retention time is the length of time that a component is retained by the column.

TABLE 1 Retention times of components of a mixture of alcohols in Figure 5

Component of the sample	R_t (s)
Methanol	39
Ethanol	63
Propanol	98
Butanol	115
Pentanol	138
Hexanol	162

Identifying components

As with paper chromatography and TLC, standards are used in gas chromatography to identify the components within a mixture qualitatively. Figure 7 shows a chromatogram of a propanol standard run under the same conditions as the chromatogram in Figure 5.

Propanol has an R_t of 98 seconds. If the standards are run under the same conditions as the sample, they will have identical retention times to the components of the sample. Therefore, standards that are run under the same conditions can be used to identify components of a sample.

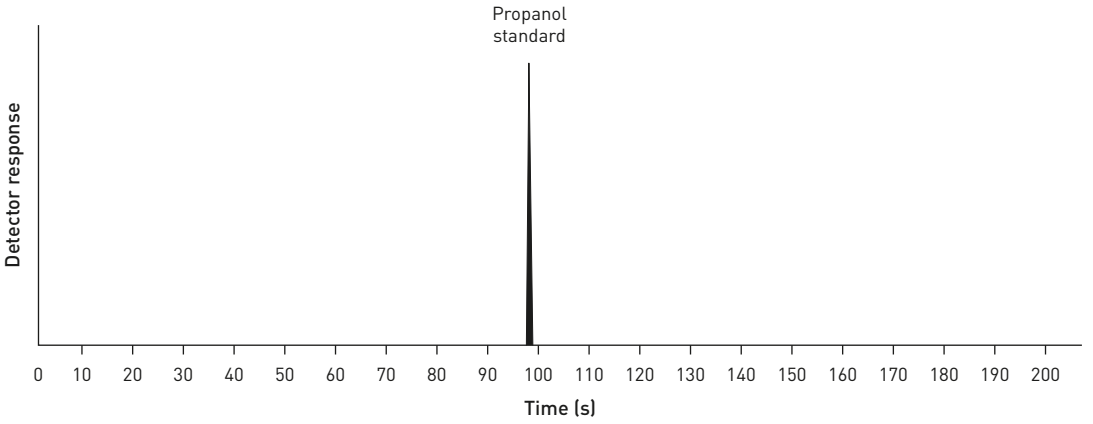


FIGURE 7 A gas chromatogram of a propanol standard.

Concentration of components

Column chromatography has the advantage of measuring the components of a sample quantitatively. This is done by measuring the height of the peak or the area of the peak. Fortunately, most chromatographs are computerised and programmed to complete this calculation. A computerised gas chromatography system usually measures peak area rather than height. The *y*-axis of the chromatogram measures peak area in cm^2. The greater peak height is due to more of that compound eluting over the same very short interval.

Figure 8 demonstrates how this technique is performed. In this example, five pure samples of ethanol were run on a gas chromatograph under identical conditions. Each sample had an accurately known concentration. A sample of ethanol of unknown concentration was also run. As well as the resultant chromatograms, a table of peak area data was generated (Table 2).

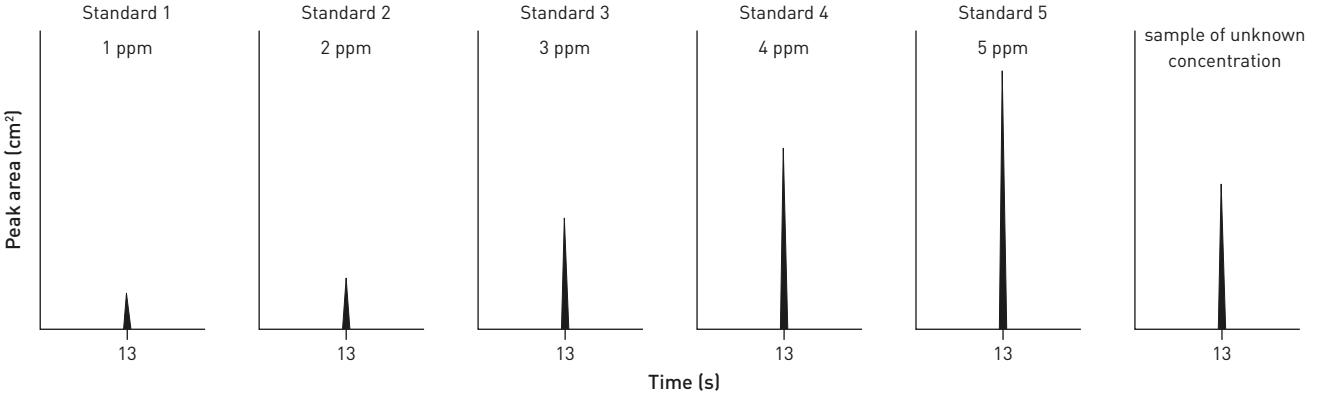


FIGURE 8 The chromatograms of five standards of ethanol and an ethanol sample of unknown concentration.

TABLE 2 Peak areas of ethanol standards and an unknown sample in Figure 8.

Sample	Concentration (ppm)	Peak area (cm^2)
Standard 1	1	5
Standard 2	2	10
Standard 3	3	15
Standard 4	4	20
Standard 5	5	25
Unknown sample	–	19

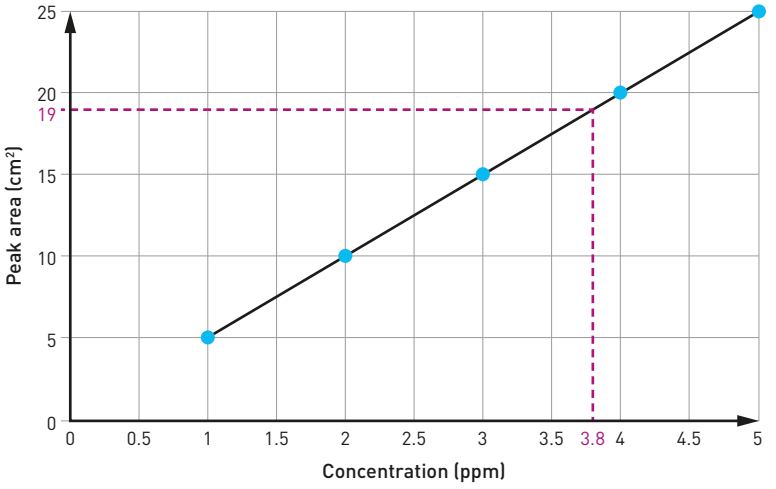


FIGURE 9 The calibration curve of a set of five standards of ethanol.

A graph can be constructed from the data in Table 2, using only the standards (Figure 9). This plots peak area against concentration.

The graph demonstrates a relationship between peak area and the concentration of the standards. This is called a **calibration curve**. If the peak area of the unknown sample is within the range of the standards (Table 2), the calibration curve can be used to determine its concentration. Figure 9 shows that a peak area of 19 cm^2 corresponds to a concentration of 3.8 ppm. This can be determined by ruling lines to the linear calibration line (shown as purple dashed lines on the graph).

calibration curve the trend associated with plotting the peak area of a set of standards, from a chromatogram, against their concentration

WORKED EXAMPLE 13.3

An athlete's urine was tested for a specific drug. Initial testing determined that there was some of this drug in the athlete's system, so gas chromatography was used to determine its concentration. The data in Table 3 was obtained.

TABLE 3 Peak areas of drug standards and an unknown sample

Sample	Concentration (mg L^{-1})	Peak area (cm^2)
Standard 1	7	0.6
Standard 2	14	1.5
Standard 3	21	2.4
Standard 4	28	3.3
Standard 5	35	4.2
Unknown	–	2.9

a Create a calibration curve for the data.

b Determine the concentration of the drug in the urine sample.

SOLUTION

a

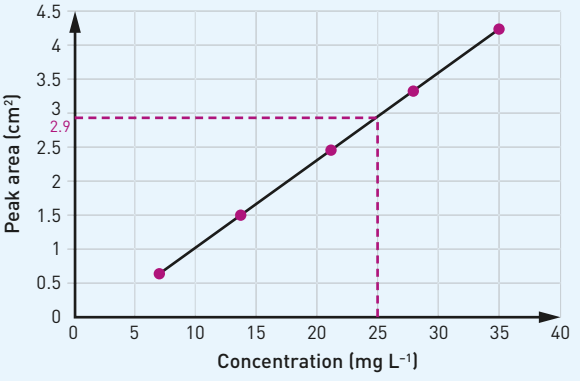


FIGURE 10 The calibration curve of a set of five standards of a drug.

b Reading off the graph the concentration of the drug in the urine sample is 25 mg L^{-1}.

Study tip

Remember that the position of the peaks on a chromatogram is what is used to identify the components in a sample. Their peak area is used to determine how much (concentration) of the component is in the sample.

Problems with chromatograms

Sometimes, when a sample is run through a chromatograph, the first chromatogram includes peaks that overlap. This may make it impossible to measure retention time. Figure 11 shows a chromatogram with two overlapping peaks. The tops of the peaks can be easily observed, so retention time could be measured, even though the most accurate results are gained when there is complete separation of the peaks. However, the peak area cannot be measured because not all of the peak can be observed.

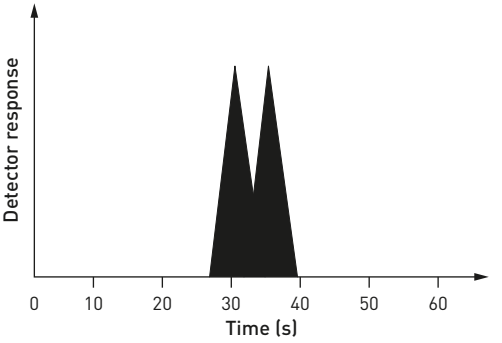


FIGURE 11
Overlapping peaks on a chromatogram can make it impossible to measure retention time and peak area.

To avoid overlapping peaks in a chromatogram, a chemist can change the material in the stationary phase, the flow rate of the mobile phase or the temperature of the column.

High-performance liquid chromatography

HPLC operates on the same principles as other chromatographic techniques. HPLC uses a liquid or aqueous mobile phase, which is pumped through an instrument (Figure 12). The sample is injected, and the mobile phase moves the sample through the column at a specified flow rate. This flow rate can be altered to pump the mobile phase with the sample through faster or slower.

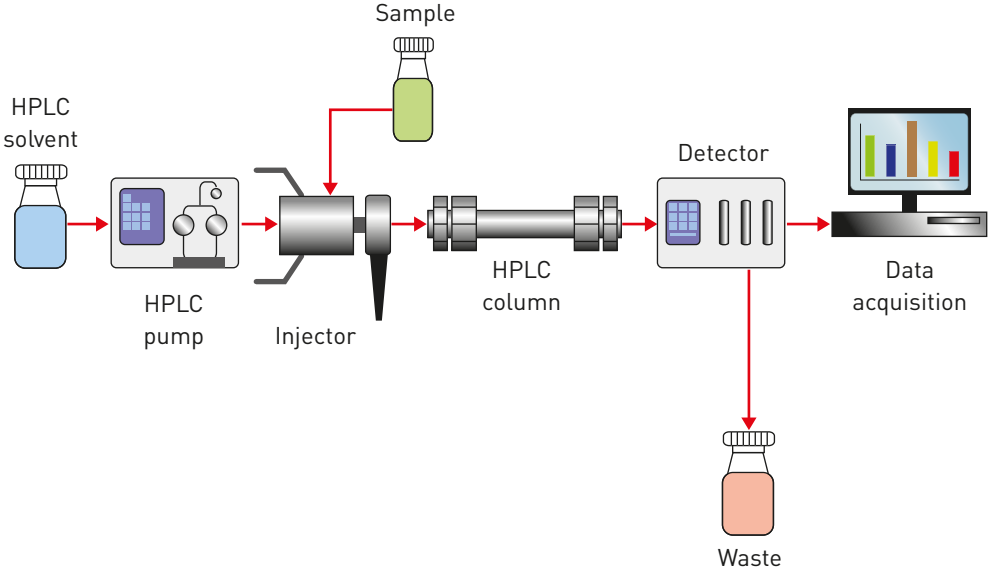


FIGURE 12 A simplified diagram of a high-performance liquid chromatograph.

The column in a high performance liquid chromatography is a solid with a high surface area, packed at very high pressure (about 2700 kPa) into a plastic piece of tubing. It is very similar to the column in chromatography, but particles are much smaller and more tightly packed. As the sample moves through the column, it interacts with both the mobile and stationary phases to separate into its various components.

The detector recognises the components as they elute from the column and generates a chromatogram similar to that generated by gas chromatography.

The temperature of the column can be altered, but not as a much as with a gas chromatograph, because the column is made of plastic, which melts at high temperatures.

Interpreting a high-performance liquid chromatogram

High-performance liquid chromatograms are analysed in the same way as gas chromatograms are. The peaks are identified by running standards under identical conditions and analysing retention times. Concentration is determined by using a set of standards of accurately known concentration, plotting a calibration curve and measuring an unknown sample on the calibration curve.

CHALLENGE 13.3

Overlapping peaks on a high-performance liquid chromatogram

A chemist runs a sample through a high-performance liquid chromatograph under the following conditions.

- Flow rate: 5 mL s^{-1}
- Sample concentration: both substances are 10 ppm
- Mobile phase: highly non-polar
- Stationary phase: highly polar

The chromatogram shown in Figure 13 is obtained.

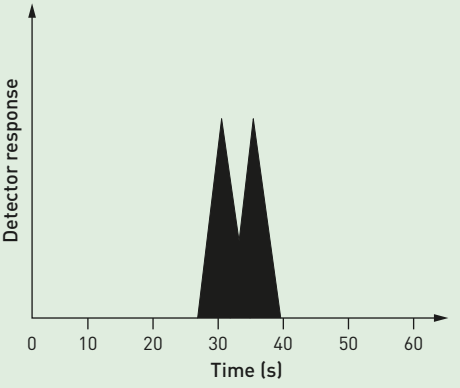


FIGURE 13 The chromatogram run by the chemist.

1. Is this data valid? Explain why or why not.
2. What could be done to improve the separation?
3. Which analytical technique, outlined in Chapter 5, could be used to confirm the identity of the components within the mixture once they have been separated?

CASE STUDY 13.3

The horsemeat scandal of 2013

In 2013, one of the biggest food scandals across Europe was uncovered. Routine food testing in British and Irish supermarkets revealed horse DNA in beef burgers supplied by the ABP Food Group.

HPLC combined with mass spectrometry was used to analyse DNA in the burgers, which were compared with regular beef products. Scientists found inconsistencies between the burgers and the regular beef products. Of the beef products tested, 37% contained horse DNA and 85% contained pig DNA.

FIGURE 14 In 2013, HPLC was used to show that some beef burgers in the UK contained horse and pig meat.

Although these meats are not harmful to consumers and no health issues were raised, it created major problems of trust for consumers.

HPLC showed that some burgers or other meat products contained 100% horse meat, whereas other tests revealed it to be present in only small percentages.

The ABP Food Group supplied beef patties to Burger King, Tesco, Lidl and Aldi. The scandal caused nearly 10 million beef patties to be recalled and removed from supermarket shelves, while Burger King immediately announced a change in their supplier.

CHECK YOUR LEARNING 13.3

Describe and explain

1 **Describe** the steps involved in the creation of a calibration curve.

2 **Explain** the differences between the mobile and stationary phases in gas chromatography and HPLC.

3 **Explain** how the interaction of the components within a sample with the mobile phase and stationary phase differs in gas chromatography and HPLC.

Apply, analyse and interpret

4 A sample of coffee was analysed by HPLC to determine the amount of caffeine present. The data in Table 4 was obtained.

TABLE 4 Peak areas of caffeine standards and an unknown sample

Sample	Concentration (mg L^{-1})	Peak area (cm^2)
Standard 1	4	0.10
Standard 2	8	0.18
Standard 3	12	0.26
Standard 4	16	0.34
Standard 5	20	0.42
Unknown	–	0.23

a **Construct** a calibration curve for the data.

b **Determine** the concentration of caffeine in the sample.

c **Consider** whether gas chromatography could be used to analyse the same sample.

Investigate, evaluate and communicate

5 A calibration curve is developed to measure the ethanol content in a bottle of wine. The resulting graph is shown in Figure 15.

a The sample generates a peak area of 30 cm^2. **Determine** the concentration of the sample.

b Is this concentration valid? **Explain** why.

c **Discuss** what would need to be done to improve the analysis.

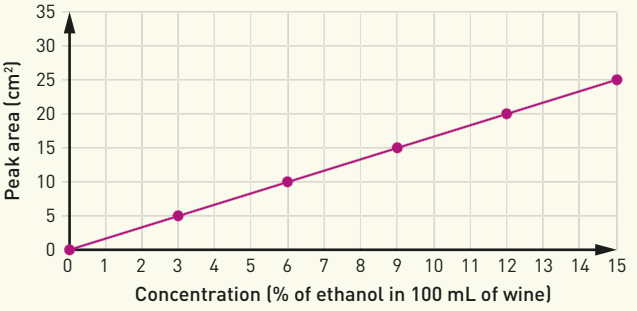


FIGURE 15 A calibration curve for the concentration of ethanol in wine.

You can find the following resources for this section on your obook assess:

» Student book questions
Check your learning 13.3

» Challenge
13.3 Overlapping peakes on a high-performance liquid chromatograph

» Increase your knowledge
Column chromatography

» Weblink
Gas chromatography

SCIENCE AS A HUMAN ENDEAVOUR

13.4 Applications of chromatography in chemical industry

KEY IDEAS

In this section, you will learn about:

+ various applications of chromatographic techniques within the field of chemistry
+ the types of substances analysed by each chromatographic technique.

Chromatographic techniques have a wide variety of uses in the chemical industry. Primarily, these are based on forensic chemistry, food science and pharmaceuticals science.

Applications of chromatography

The type of chromatography technique chosen to separate, identify and quantify a substance depends on the properties of the sample. It also depends on the limitations of each technique in analysing various substances.

Paper chromatography and thin-layer chromatography

Paper chromatography is cheapest of all chromatography techniques because it requires no specialised equipment and is most commonly used for samples of compounds that can be separated by water (a polar molecule). Therefore, it is used for separating molecules by their polarity.

TLC is more sensitive (smaller concentrations of sample can be used) and efficient (faster) than paper chromatography. However, the two techniques analyse similar types of mixtures.

Common analyses include:

- determining the types of food colourings present in inks, dyes and additives in food chemistry
- determining the amino acids in proteins, for various biochemical and medicinal applications; for example, cancer research
- identifying the reactants and products in a reaction system, to determine how far a reaction has proceeded in drug synthesis
- identifying plant extracts, which may be used to develop various pharmaceutical drugs
- determining the purity of a substance; for example, pharmaceutical drugs.

Gas chromatography

Substances analysed by gas chromatography must be volatile (i.e. easily evaporated into the gas phase). Non-polar substances, or molecules with only dispersion forces, are more volatile and therefore tend to be analysed by gas chromatography rather than HPLC.

Gas chromatography is limited by the set-up of the instrument and its requirement for vaporised samples.

- Gas chromatography will only analyse volatile compounds; it cannot analyse semivolatile or non-volatile compounds.
- The instrument must be calibrated regularly because of its high sensitivity.
- Gas chromatography can typically analyse samples with a concentration in mg L^{-1} (or **ppm**). Some gas chromatographs can analyse at lower concentrations, but this is not common.
- Water vapour in the atmosphere can contaminate the analysis.
- The detector is located outside the oven that contains the column. The variation in temperature experienced by the samples moving from the column to the detector can cause errors.
- The instrument must be attached to a gas cylinder, which dispenses gas to act as the mobile phase. However, the cylinder is capable of storing large quantities of gas, so this is not often a problem.

ppm
parts per million; a unit of concentration that measures the amount of substance in milligrams per litre

Applications of gas chromatography include:

- determining the alcohol content in blood
- determining air quality; for example, the amount of gases released as pollutants by cars, power plants and various other fossil fuel burners
- determining the nutritional content of foods; for example, fatty acids, after esterification to make them volatile
- analysing pesticides that may have leached into soils, waterways and ecosystems
- identifying and quantifying cancer-causing biomolecules
- separating mixtures so that individual components can be collected and further analysed by other analytical techniques.

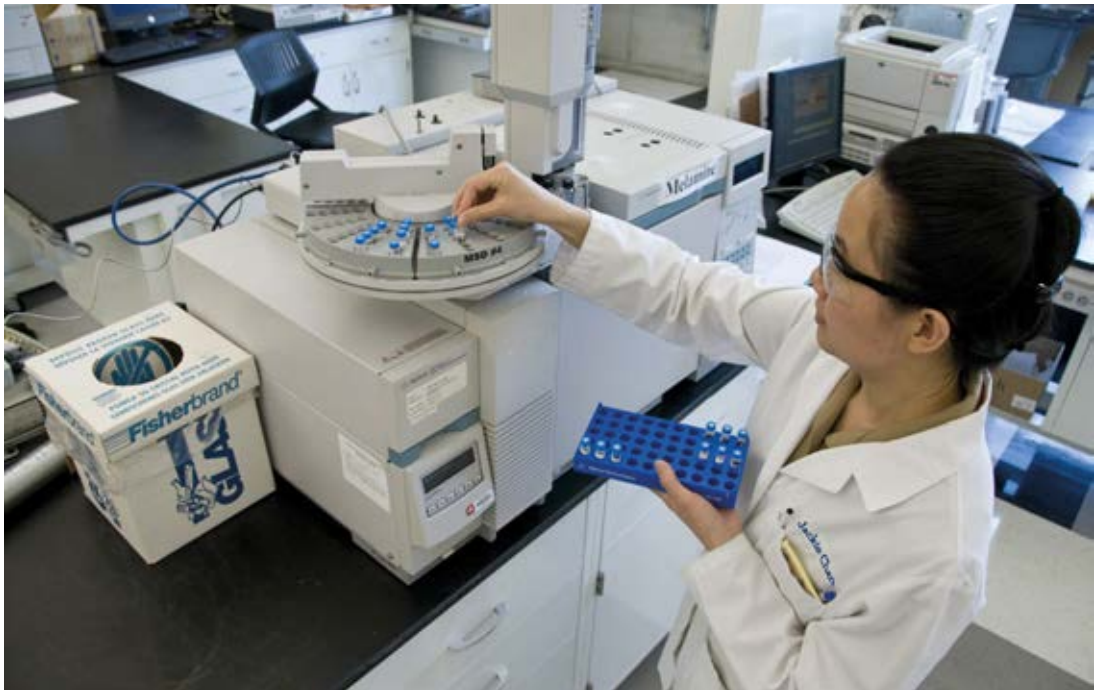


FIGURE 1 Gas chromatography can be used to test the nutritional quality of food.

High-performance liquid chromatography

HPLC is used to analyse non-volatile molecules. There are many types of HPLC that can separate mixtures according to size, charge, molar mass and even hydrophilic properties.

HPLC is limited by the set-up of the instrument and its requirement for liquid or aqueous samples.

- HPLC will only analyse compounds that are soluble in the liquid or aqueous mobile phase. HPLC cannot analyse gaseous or insoluble compounds, which can damage or block the instrument.
- The instrument must be calibrated regularly because of its high sensitivity.
- HPLC can typically analyse samples with a concentration in μg L^{-1} (or **ppb**). Some high-performance chromatographs can analyse lower concentrations but it is not common.
- The mobile phase must be kept clean to prevent impurities contaminating the sample.
- The column must be cleaned by flushing thoroughly with the mobile phase to ensure that no substances are stuck to the stationary phase, blocking its interaction with other molecules.
- The small particle size of the column results in significant resistance to the movement of the mobile phase. Therefore, the mobile phase must be pumped through at very high pressures (sometimes more than 48 000 kPa).
- The instrument must be attached to an eluent reservoir (i.e. the bottle that stores the mobile phase). This dispenses the mobile phase into the instrument. Depending on the rate that the mobile phase is pumped through the instrument and the size of the reservoir, it must be topped up frequently to ensure that the HPLC does not run out of mobile phase.

ppb
parts per billion; a unit of concentration that measures the amount of substance in micrograms per litre

Study tip

When answering questions about chromatographic technique, if possible determine the properties of the sample first. Otherwise, look for the intermolecular properties in the analysis and explain interactions and applications on the basis of these.

FIGURE 2 A scientist working with a high-performance liquid chromatograph checks a sample.

Applications of HPLC include:

- analysing the purity of various drugs in pharmaceutical manufacturing
- determining and quantifying illicit drugs, for use in charging and convicting persons of interest
- analysing food and blood samples to determine the presence and concentration of various nutrients
- immunoassays, which determine the presence and concentration of various molecules within the blood using antibodies
- identifying and quantifying substances in post-blast bomb residues.

FIGURE 3 High-performance liquid chromatography can be used to determine types of nutrients in blood.

CHECK YOUR LEARNING 13.4

Describe and explain

1 **Describe** the main differences between the samples that can be analysed by gas chromatography and by HPLC.

2 **Explain** whether samples that are analysed by TLC can also be analysed by paper chromatography.

Apply, analyse and interpret

3 **Determine** which analytical techniques could be used to analyse the following samples.

a Drug content in blood

b Alcohol content in wine

4 For each of the following three analyses, **deduce** which technique would be used to analyse the sample and why it was chosen.

- Analysis of a crude oil sample to determine the compounds present and their concentrations
- Analysis of a water sample to determine the compounds present
- Analysis of a brown Smartie to determine the food colourings present

You can find the following resources for this section on your obook assess:

» Student book questions
Check your learning 13.4

» Video weblink
Paper chromatography

» Weblink
High-performance liquid chromatography

» Weblink
Applications of chromatography

CHAPTER 13

Review

Chapter summary

13.1
- All chromatography is based on intermolecular bonding. The type of intermolecular bonding determines how strongly the components of a mixture interact with chemicals in the mobile and stationary phases and therefore how effectively they separate.
- Components of the sample with a high affinity for the mobile or stationary phase interact more with it. Components with a lower affinity for the mobile or stationary phase do not interact as much.
- The components of a sample separate when they interact with the mobile or stationary phase as they adsorb to the stationary phase and desorb into the mobile phase.

13.2
- Paper and thin-layer chromatography are the simplest techniques for separating a sample. The resultant chromatogram is analysed by calculating the retardation factor (R_f). This determines the ratio of the movement by components of the sample to the movement of the mobile phase. Both techniques are qualitative.

$$R_f = \frac{\text{distance solute moves from origin}}{\text{distance solvent moves from origin}}$$

13.3
- Gas chromatography and HPLC involve a column that is packed with a stationary phase. The sample and mobile phase are pumped through the column. In gas chromatography, the sample interacts with the stationary phase. In HPLC, the sample interacts with the stationary phase and mobile phase. This interaction separates the sample into its components. The resultant chromatogram displays peaks, with one peak for each component of the sample.
- The chromatogram is analysed in two ways.
  - The retention time (R_t) (time is has spent on the column) identifies each component qualitatively.
  - The areas under the peaks measures the concentration (quantitative) of each component when analysed on a calibration curve.

13.4
- Chromatographic techniques have many applications. To determine the best technique to use, it is important to consider the properties of the components within the sample as well as the limitations and capabilities of each instrument. The mobile and stationary phases must be selected carefully according to the properties of the sample.

Key terms

- adsorption
- affinity
- calibration curve
- chromatogram
- column
- column chromatography
- desorption
- elute
- gas chromatography
- high-performance liquid chromatography (HPLC)
- mobile phase
- origin
- paper chromatography
- ppb
- ppm
- retardation factor (R_f)
- retention time (R_t)
- solute
- solvent
- solvent front
- stationary phase
- thin-layer chromatography (TLC)
- vaporise

Key formulas

Retardation factor	$R_f = \frac{\text{distance solute moves from origin}}{\text{distance solvent moves from origin}}$

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; ***=high.

Multiple choice

1 The measurement used to calculate the concentration of a substance from a series of standards in gas chromatography is called:

- **A** retention time
- **B** peak area
- **C** peak height
- **D** retardation factor.

2 Column chromatography was performed to determine the pigments within a plant sample. The chromatogram shown in Figure 1 was analysed. What conclusions could be made about the blue pigment?

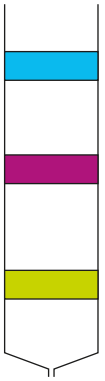


FIGURE 1
A chromatogram of plant pigments.

- **A** It has a high R_t and a high affinity for the stationary phase.
- **B** It has a low R_t and a high affinity for the stationary phase.
- **C** It has a high R_t and a high affinity for the mobile phase.
- **D** It has a low R_t and a high affinity for the mobile phase.

3 Which of the following would result in the highest R_t value?

	Mobile phase	Stationary phase	Substance
A	polar	non-polar	polar
B	non-polar	polar	polar
C	polar	non-polar	non-polar
D	non-polar	polar	polar

4 In HPLC, R_t can be decreased by:

- **A** lowering the temperature
- **B** decreasing the pressure
- **C** increasing the flow rate
- **D** decreasing the flow rate.

5 Which chromatographic techniques are qualitative as well as quantitative?

- **A** Paper chromatography, TLC, gas chromatography and HPLC
- **B** TLC and gas chromatography
- **C** Gas chromatography and HPLC
- **D** HPLC only

Short answer

Describe and explain

★ **6** **Explain** how chromatography separates substances from a mixture. Use the terms 'mobile phase', 'stationary phase', 'affinity', 'adsorb' and 'desorb'.

★ **7** Describe the difference between qualitative and quantitative analysis in chromatography. Use examples in your answer.

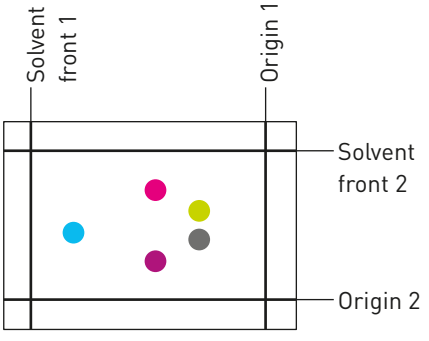
★ **8** **Explain** why it is necessary to prepare a set of standards when performing a quantitative analysis in chromatography.

★★ **9** **Explain** four ways that the retention time of a sample could be increased in an HPLC column.

★★ **10** **Explain** three ways that the retention time of a sample can be decreased in a gas chromatography column.

★★★ **11** **a** **Describe** briefly how this chromatogram would have been produced.

b **Explain** the difference between the mobile phase (solvent) in both separations.



c **Explain** why the chromatography technique was performed twice.

d Using two-dimensional chromatography, you cannot compare the dyes in the mixture directly with known dyes. Instead you must make use of known R_f values.
  a **Explain** why this is true.
  b **Explain** how an R_f value is measured.

e Using a ruler, **calculate** both R_f values for all five chemical substances.

f From the R_f values calculated, **determine** which spot has the:
  a highest R_f value in solvent 1
  b highest R_f value in solvent 2
  c lowest R_f value in solvent 1
  d lowest R_f value in solvent 2
  e highest affinity for solvent 1
  f highest affinity for solvent 2
  g highest affinity to the stationary phase.

★★★ **12** A mixture is separated by paper chromatography. The mobile phase is a 1% ethanol solution. After the mixture is separated, the components of the sample have R_f values of 0.89, 0.62 and 0.27. **Determine**:
  a which component has the highest affiliation for the mobile phase
  b which component has a higher affinity for the stationary phase
  c what can be concluded about the properties of the components in the mixture.

Apply, analyse and interpret

Questions 13–16 are based on the following information.

Two disinfectant sprays were analysed by HPLC to determine whether the ethanol values specified on the labels were accurate.

FIGURE 3 Disinfectant sprays

The data in Table 1 was obtained from a series of standards of ethanol solution.

TABLE 1 Standards of ethanol

Standard	Concentration of ethanol (% per 100 mL)	Peak area (cm^2)
0	0	0
1	3	2
2	6	4
3	9	6
4	12	8
5	15	10

The two disinfectant samples were run through the HPLC instrument and the data in Table 2 was obtained.

TABLE 2 Disinfectant samples

Sample	Label ethanol content (% per 100 mL)	Peak area (cm^2)
Disinfectant 1	1–3	3
Disinfectant 2	6–8	5

★ **13** **Sketch** a calibration curve for the ethanol standards.

★ **14** **Determine** the concentration of ethanol in the two samples.

★★ **15** **Infer** the accuracy of the product labels based on the calculations of concentration from the question 14.

★★ **16** The standards that were used varied in concentration between 0 and 15% per 100 mL. **Consider** why this range was selected.

★★ **17** A mobile phase primarily consists of water. **Consider** why water is chosen as a mobile phase.

★★★ **18** An alkane is an organic molecule that contains only carbon and hydrogen atoms joined by single bonds. A sample of five alkanes is run through a gas chromatograph. The stationary phase is non-polar. Table 3 outlines the number of carbons in each alkane's chain.
  a **Explain** the intermolecular forces between the alkanes and the stationary phase.
  b Which alkane would you expect to elute first from the column? **Explain** why.

c Which alkane would you expect to elute last? **Explain** why.

d **Sketch** a chromatogram of the resulting separation of molecules.

TABLE 3 Alkane samples

Alkane	Number of carbons
Ethane	2
Butane	4
Pentane	5
Heptane	7
Decane	10

Investigate, evaluate and communicate

★ **19** **Investigate** an example of where the 'like with like' rule works.

★★★ **20** An HPLC analysis was performed to determine the R_t of a polar substance by using a polar mobile phase and non-polar stationary phase.

a **Determine** the effects of decreasing the polarity of the mobile phase on the R_t of the substance. Justify your answer.

b **Construct** a chromatogram that demonstrates this effect.

★★★ **21** A student wanted to determine the concentration of methane in natural gas by gas chromatography. After running a set of standards with a flow rate of 1 mL min^{-1}, he determined that this was too slow and increased the flow rate for the samples to 5 mL min^{-1}.

a **Determine** whether the student can generate a calibration curve.

b **Determine** whether it is possible to indicate the concentration of methane.

c **Determine** whether it is possible to determine the identity of methane.

d **Discuss** why the student's data is or isn't valid.

★★★ **22** **Synthesise** the chromatographic techniques into a table such as Table 4.

TABLE 4 Chromatography techniques

Chromatographic technique	Mobile phase	Stationary phase	How the data is collected and analysed	What types of analyses this is used for	Qualitative or quantitative?
Paper chromatography					
TLC					
Gas chromatography					
HPLC					

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CHAPTER 14

Gases

The earliest philosopher in ancient Greece proposed five elements that make up our world: earth, water, fire, aether and air. It was not until the 17th century that scientists developed an alternative, evidence-based theory of gases, which we use today.

Gases are one of the five states of matter, and the first to be thoroughly investigated by early scientists. Early experiments focused on the relationships between the pressure, volume and temperature of gases. The result of these experiments was the formulation of the gas laws, the earliest of which was developed in 1662 by Robert Boyle and is still used in modern chemistry.

A thorough understanding of kinetic molecular theory is essential to understanding the behaviour of gases. This theory describes the way that gas particles move, which gives an insight into their properties. Kinetic molecular theory is used to explain the early research of scientists, such as Boyle, and the relationship between pressure, volume and temperature.

Gas laws can be applied to various phenomena and occupations. Scuba divers trust the gas laws to help them to avoid decompression sickness when under water. Similarly, aeroplane travel relies on a knowledge of gases in the atmosphere.

This chapter focuses on the understanding of kinetic molecular theory, and its ability to explain the gas laws and their relationship with the number of particles contained within a gaseous system.

OBJECTIVES

- → Consider the relationship between the volume, number of moles and molar volume at standard temperature and pressure (STP).
- → Use the kinetic theory of gases to describe and explain the behaviour of gases, including the qualitative relationships between pressure, temperature and volume.
- → Appreciate that the kinetic theory of gases applies to ideal gases and solve problems related to the ideal gas equation.
- → Use appropriate mathematical representation to solve problems and make predictions, including the mole concept, to calculate the mass of chemicals and/or the volume of a gas (at standard temperature and pressure) involved in a chemical reaction.

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PRACTICALS

MANDATORY PRACTICAL	**14.1** Investigating the properties of gases
SUGGESTED PRACTICAL	**14.2** Boyle's law

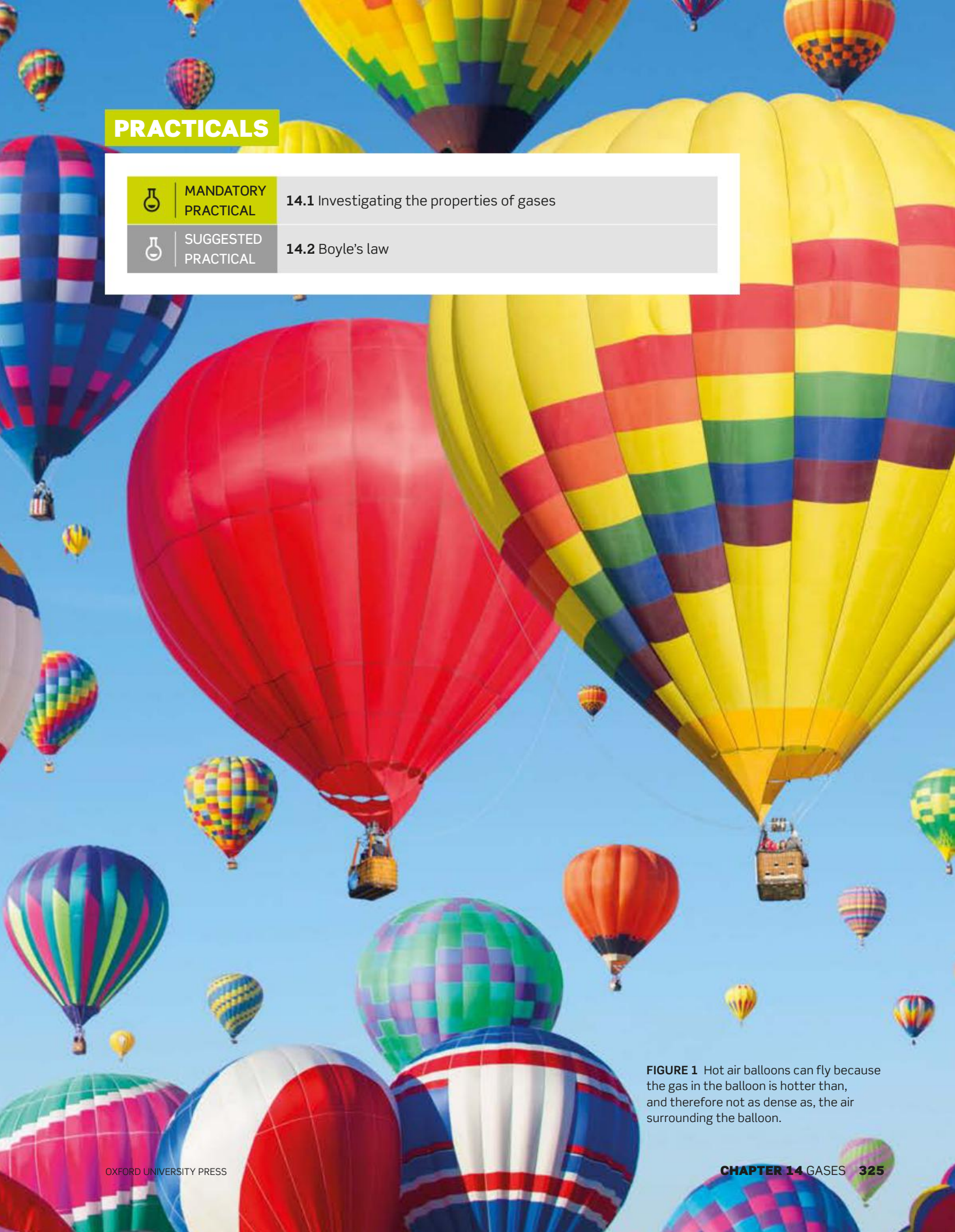

FIGURE 1 Hot air balloons can fly because the gas in the balloon is hotter than, and therefore not as dense as, the air surrounding the balloon.

14.1 Gas properties

KEY IDEAS

In this section, you will learn about:

- how the kinetic theory of gases can be used to explain gas behaviour
- the properties of gases based on kinetic molecular theory
- the units of pressure, temperature and volume
- converting between units of pressure, temperature and volume.

A gas can be described by the particle model of matter. In the gaseous state, particles do not remain together. They have a high level of energy so they separate and move freely. The behaviour of gases can be explained using **kinetic molecular theory**.

kinetic molecular theory
the theory that states that all particles are in constant random motion

kinetic energy
the energy associated with objects based on their movement or motion

Kinetic theory of gases

Kinetic molecular theory explains the movement and behaviour of particles in solids, liquids and gases. **Kinetic energy** is the energy associated with the motion of particles. The more kinetic energy they have, the more they will move.

The kinetic molecular theory of gases can be summarised into the following key points.

- Gas particles are in constant, random, straight-line motion.
- Gas particles have very little attraction to one another (Figure 1) because the energy of their movement is greater than the energy of attraction between particles.
- The space between gas particles is much larger than the space that the particles themselves occupy (Figure 1). This means that the gas can be compressed and has a low density.
- Kinetic energy is conserved in gas particle collisions. This means that when gas particles collide, no energy is transferred from one particle to another particle or to the sides of the container.
- The average kinetic energy of a gaseous system depends on its temperature (see page 328). The higher the temperature, the higher the average kinetic energy.

Solid Liquid Gas

FIGURE 1 Gases expand to occupy the whole space of a container, unlike solids and liquids.

Properties of gases

The behaviour of gases is affected by:

- pressure
- temperature
- volume.

Pressure

Pressure is defined as the force exerted, per unit area, by one substance on another substance. For gases, this is the pressure exerted by gas particles when they collide with the walls of a container. This relationship can be represented mathematically as:

$$\text{pressure} = \frac{\text{force}}{\text{area}} \text{ or } P = \frac{F}{A}$$

A good analogy to help you visualise pressure involves a punching bag. If a punching bag is hit only once, a small force is exerted, and the bag may move a little. If a punching bag is hit continually, with many hits, a larger force is exerted.

Gas particles that have more kinetic energy will move more, increasing their chances of colliding with the walls of the container as well as increasing the average force of each collision. An increase in the number of collisions with the walls of the container will increase the pressure exerted on the container.

The unit for pressure is the pascal (Pa). This is the force of one newton per square metre of substance (N m^{-2}).

FIGURE 2 Pressure is generated when air molecules collide with the internal walls of a balloon.

pressure
the force exerted, per unit area, by one substance upon another substance

Normal atmospheric pressure

At sea level, atmospheric pressure is measured as one atmosphere (1 atm). This is the equivalent of 101 325 Pa, 101.3 kPa (kilopascals) and 760 mmHg (millimetres of mercury). A pressure of 1 mmHg is that needed to increase the volume of mercury so that it travels 1 mm higher up the column of a manometer (a device for measuring pressure).

Figure 3 demonstrates how to convert between pressure units.

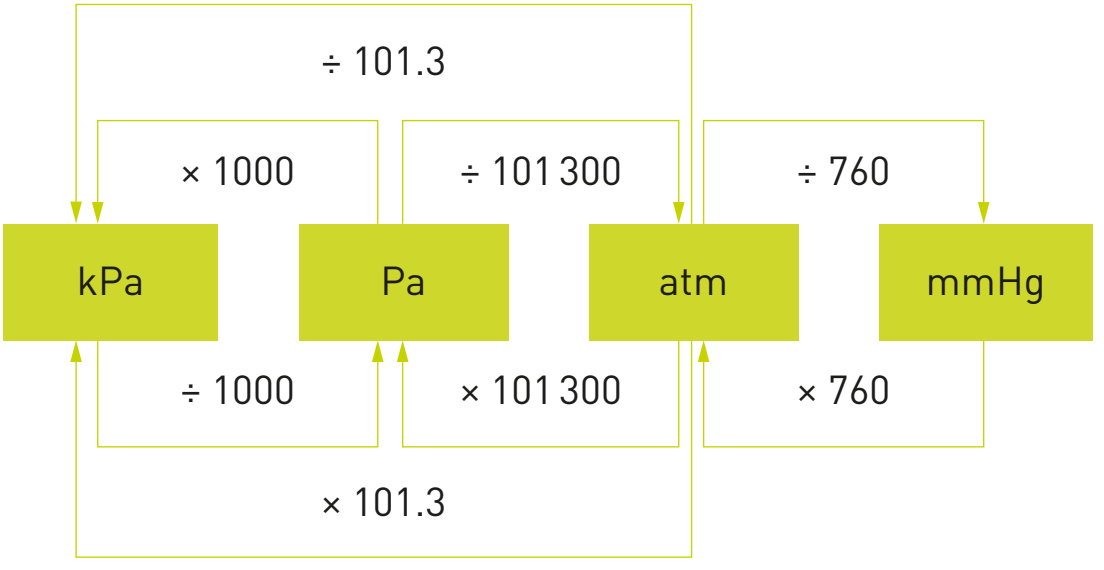

FIGURE 3 The conversion tool for converting between units of pressure.

WORKED EXAMPLE 14.1A

Convert the following units.

a 0.872 atm to kPa

b 0.672 atm to mmHg

c 635 kPa to atm

d 92 572 kPa to mmHg

e 824 mmHg to kPa

f 7362 mmHg to atm

SOLUTION

a $1\text{ atm} = 100.0\text{ kPa}$

$$\frac{0.872\text{ atm}}{1} \times \frac{100.0\text{ kPa}}{1\text{ atm}} = 87.2\text{ kPa}$$

b $1\text{ atm} = 760\text{ mmHg}$

$$\frac{0.672\text{ atm}}{1} \times \frac{760\text{ mmHg}}{1}\text{ atm} = 510.72\text{ mmHg}$$

c $1\text{ kPa} = \frac{1}{100.0}\text{ atm}$

$$\frac{635\text{ kPa}}{1} \times \frac{1\text{ atm}}{100.0\text{ kPa}} = 6.35\text{ atm}$$

d $100.0\text{ kPa} = 760\text{ mmHg}$

$$\frac{92\,572\text{ kPa}}{1} \times \frac{760\text{ mmHg}}{100.0\text{ kPa}} = 703\,547\text{ mmHg}$$

e $1\text{ mmHg} = \frac{100.0}{760}\text{ kPa}$

$$\frac{824\text{ mmHg}}{1} \times \frac{100.0\text{ kPa}}{760\text{ mmHg}} = 108.42\text{ kPa}$$

f $1\text{ mmHg} = \frac{1}{760}\text{ atm}$

$$\frac{7362\text{ mmHg}}{1} \times \frac{1\text{ atm}}{760\text{ mmHg}} = 9.69\text{ atm}$$

Study tip

The databook states the pressure conversion.
Normal atmospheric pressure:
1 atm = 100.0 kPa
= 760 mmHg

temperature
a measure of the average heat energy of the particles within a system

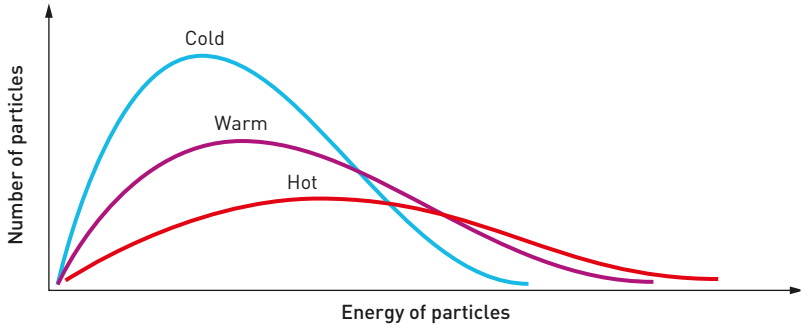

FIGURE 4 The Maxwell–Boltzmann representation of the amount of energy distributed in a reaction system at different temperatures. The curves represent the energy distributions at different temperatures.

Temperature

All particles in a reaction vessel have different amounts of heat energy. **Temperature** is a measure of the average heat energy of the particles within a system (see Chapter 8). This can be represented by a Maxwell–Boltzmann distribution curve (Figure 4). Particles with more heat energy will move more in a reaction vessel. Therefore, they have an increased amount of kinetic energy.

As temperature increases, the heat energy of a system increases. This causes the kinetic energy and movement of the particles to increase (Figure 4).

Kinetic energy is measured as:

$$\text{KE} = \frac{1}{2} m v^2$$

where KE is kinetic energy (J), m is the mass of particles (kg) and v is the average velocity of particles (m s^{-1}).

The kinetic energy of any gaseous system is the same at a specified temperature (Figure 4). It can be concluded that for kinetic energy to remain constant, when the mass of particles increases, their velocity must decrease. This means that larger particles move slower and smaller ones move faster on average, at the same temperature.

The SI unit for temperature is the kelvin (K). One degree on the Celsius scale is equivalent to one kelvin. So:

0°C = 273.15 K

1°C = 273.15 K + 1 = 274.15 K

Figure 5 demonstrates how to convert between temperature units.

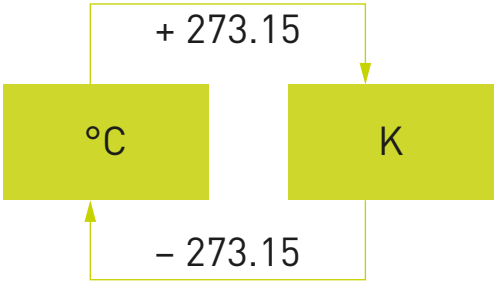

FIGURE 5 The conversion tool for converting between units of temperature.

WORKED EXAMPLE 14.1B

Convert the following units.

a 100°C to K
b 25°C to K
c –124°C to K
d 345 K to °C
e 123 K to °C
f 526 K to °C

SOLUTION

a 100 + 273.15 = 373.15 K
b 25 + 273.15 = 298.15 K
c –124 + 273.15 = 149.15 K
d 345 – 273.15 = 71.85°C
e 123 – 273.15 = –150.15°C
f 526 – 273.15 = 252.85°C

Volume

Gases occupy the whole container. Therefore, the **volume** of a container is the equivalent of the volume of the gas inside that container. Two types of gaseous systems are considered when discussing gas behaviour.

FIGURE 6 A balloon is an example of a variable volume system because it can adjust to the volume of gas.

- **Fixed volume or rigid wall systems** are usually contained within gas cylinders because their walls will not expand or collapse. This means that the volume that the gas occupies within the cylinder is always constant.
- **Variable volume systems** have walls that can adjust to the volume of the gas such as a balloon or a syringe. As more gas is added into the balloon (Figure 6), the elastic of the balloon stretches to accommodate the gas. When the gas leaves the balloon, the elastic compresses to a smaller size.

volume
a measure of the space occupied by a substance

The SI unit for volume is cubic metres (m^3) or litres (L), where:
1 m^3 = 1000 L = 1 000 000 mL
0.001 m^3 = 1 L = 1000 mL
Figure 7 demonstrates how to convert between volume units.

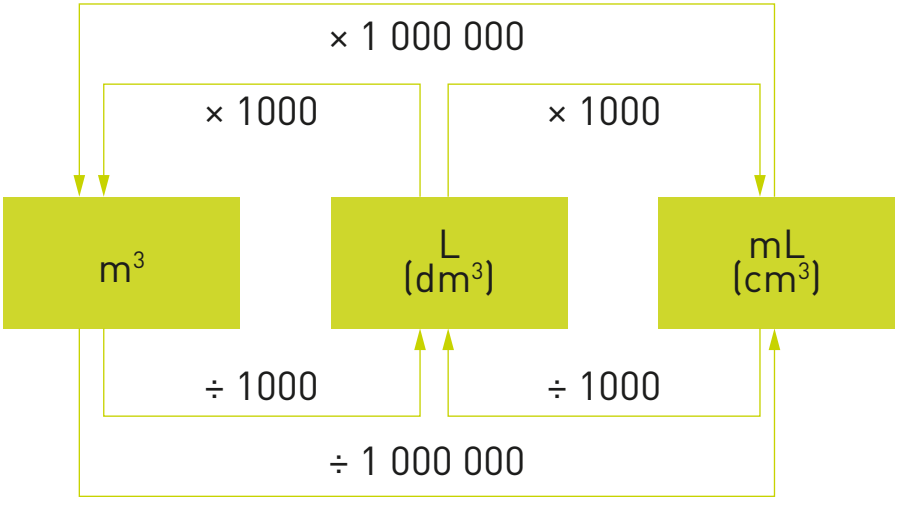

FIGURE 7 The conversion tool for converting between units of volume.

WORKED EXAMPLE 14.1C

Convert the following units.

a 25 mL to m^3

b 45.672 L to m^3

c 2.5 m^3 to L

SOLUTION

a $\frac{25\ mL}{1\ 000\ 000\ m^3} = 2.5 \times 10^{-5}\ m^3$

b $\frac{450.672\ L}{1\ 000\ m^3} = 0.045\ m^3$

c $2.5 \times 1000 = 2500\ L$

CHALLENGE 14.1

Movement of gases

If gases moved in a circular pattern, rather than in straight lines, what impact would this have on the pressure of a system?

CASE STUDY 14.1

Hot air balloons

Hot air balloons date back to ancient China (200–300 CE) when military personnel would release paper lanterns into the air as a form of communication. The first air balloon was developed by French brothers Joseph-Michel and Jacques-Etienne Montgolfier in 1783 (Figure 8). In order to inflate the balloon, the brothers used a basket of hot coals to heat the air. The hot air, being much less dense than the cooler air on the outside of the balloon, was able to lift it off the ground. It flew for 10 minutes, unmanned, and laid the foundation for the design of future manned hot air balloons.

FIGURE 8 A model of the first hot air balloon developed by the Montgolfier brothers.

The first manned hot-air balloon flight followed shortly after and also used hot air to generate lift in the balloon. Around the same time, the discovery of hydrogen (H_2) gas and the hypothesis that this light gas could also lift a balloon was tested by French inventor Jacques Charles. Charles employed the help of French engineers and brothers Anne-Jean Robert and Nicolas-Louis Robert to build the balloon, which was filled with hydrogen gas rather than hot air.

The balloon was built with a release valve, capable of expelling gas from the balloon into the atmosphere and was weighed down with sand bags. In order to generate the hydrogen required, Charles used the reaction between an acid and a metal to produce salt and hydrogen. He used 250 kg of dilute sulfuric acid (H_2SO_4) and 500 kg of scrap iron:

FIGURE 9 One of the early designs of Meusnier's dirigible.

$$H_2SO_4(aq) + Fe(s) \rightarrow FeSO_4(aq) + H_2(g)$$

This reaction is exothermic so the hydrogen gas was hot when it was produced. Filling a balloon with this hot gas was problematic. As the gas cooled, the particles packed closed together and the volume of the balloon decreased. Charles would need to produce more hydrogen to fill the space. This led to the development of Charles' law (section 14.2), which outlines the relationship between the temperature of a gas and the volume of space that it occupies.

This flight was highly successful, leading Jacques Charles and the Robert brothers to build an elongated balloon, filled with hydrogen gas and modelled off French mathematician, Jean Baptiste Meusnier's design of the first dirigible.

CHECK YOUR LEARNING 14.1

Describe and explain

1 **Describe** kinetic energy and **explain** why it has an essential role in the particle model of matter.

2 **Use** kinetic molecular theory to explain why:
- **a** there are spaces between gas particles
- **b** gas particles have more energy than particles in a solid or liquid
- **c** gases occupy the space of a container.

3 **Convert** the following units.
- **a** 315 K to °C
- **b** 74°C to K
- **c** 273 mm Hg to kPa
- **d** 50k Pa to atm
- **e** 25 mL to m^3
- **f** 0.45 m^3 to L

Apply, analyse and interpret

4 The kinetic energy of a gas with a mass of 2.50 g is calculated to be 62.5 J. What is the average velocity of the gas particles? **Determine** a calculation of average velocity? **Explain** why this is.

Investigate, evaluate and communicate

5 At 100°C, some particles still have the same amount of energy as they did at 0°C. **Discuss** how this is possible.

6 When a scented candle is lit at one end of a room, the scent slowly diffuses to the opposite side of the room. **Propose** how the kinetic molecular theory explains this phenomenon.

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- » Challenge
  14.1 Movement of gases
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  Gas and temperature

14.2 Gas laws

KEY IDEAS

In this section, you will learn about:

- the mathematical relationships between pressure, temperature and volume
- molar calculations at standard temperature and pressure (STP).

gas laws
the laws that explain the properties of gases and how they behave at different temperatures, volumes and pressures

The relationship between pressure, volume and temperature is summarised into three main **gas laws** that are named after the scientists who developed them.

Temperature versus pressure: Gay-Lussac's law

Gay-Lussac's law reflects the relationship between the temperature and pressure of a gas. As temperature increases, the kinetic energy of particles within a fixed volume system also increases. This means that particles move more and collide with the walls of the container more frequently and with more force. As the walls of the container cannot move, and the container cannot expand to accommodate these excess collisions, the pressure within the container increases. This occurs when volume is constant; that is, a fixed volume system.

Numerical representation of Gay-Lussac's law

The relationship between temperature and pressure can be represented numerically. Figure 1 depicts a round-bottom flask filled with a gas and attached to a barometer (used to measure pressure). The same flask is placed in three water baths at different temperatures to determine the effect of temperature on gas pressure. Table 1 summarises the data gathered in the experiment.

The relationship between temperature and pressure is shown in the graph in Figure 2. The graph demonstrates that there is a linear relationship between temperature and pressure. Therefore, pressure is proportional to temperature in Kelvin ($P \propto T$). This relationship can therefore be represented mathematically as:

$$\frac{P}{T} = \text{a constant}$$

Furthermore, if a change in temperature or pressure to a gaseous system is made, new values can be calculated from the equation:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

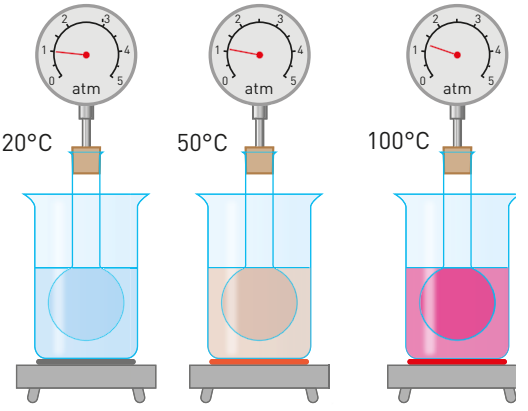

FIGURE 1 The effect of temperature on pressure in a gaseous system.

Study tip

Gas laws often use Kelvin (K) as a measurement of temperature. You need to know how to convert degrees Celsius (°C) to Kelvin (K). Use the formula K = °C + 273.15

TABLE 1 Data obtained from the experiment in Figure 1

Temperature (°C)	Temperature (K)	Pressure (atm)	P/T (atm °C^{-1})	P/T (atm K^{-1})
20	293.15	1.00	0.050	0.0034
50	323.15	1.10	0.022	0.0034
100	373.15	1.27	0.013	0.0034

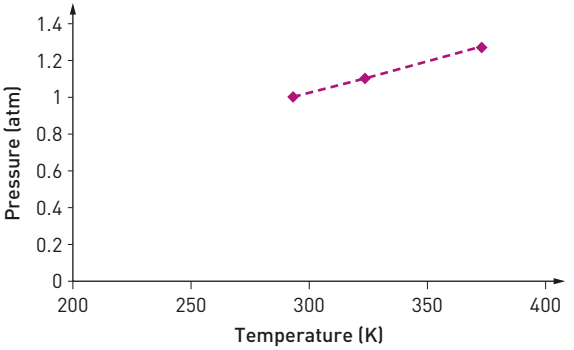

FIGURE 2 The relationship between temperature (K) and pressure in a gaseous system.

WORKED EXAMPLE 14.2A

At 293.15 K the pressure in a reaction vessel is 1.00 atm. If the temperature is raised to 308.15 K, what is the pressure of the vessel?

SOLUTION

1 Substitute values into the equation

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\frac{1\text{ atm}}{293.15\text{ K}} = \frac{P_2}{308.15\text{ K}}$$

2 Rearrange to solve for P_2.

$$P_2 = \frac{1\text{atm}}{293.15\text{ K}} \times 35$$

$$= 1.05\text{ atm}$$

Temperature versus volume: Charles' law

Charles' law states that at a constant temperature, the volume of a gas is proportional to the temperature of the system. This means that when the volume of a container is variable (able to change), temperature will have an effect. As temperature increases, the kinetic energy (movement) of particles increases and they collide with the walls of the container more frequently and with more force. However, as the volume of the container can vary, these collisions will push on the walls of the container, causing the container to expand, rather than increasing the pressure. In this system, pressure is constant, i.e. the container expands, and in so doing, maintains the pressure of the system.

A balloon is one example of this phenomenon. When inflated and placed in a beaker filled with liquid nitrogen (–196°C), the low temperature causes the volume of the balloon to decrease, seen in Figure 3.

When the balloon is removed from the beaker and the temperature starts to increase, the air inside the balloon expands back to its original volume (Figure 4).

FIGURE 3 A balloon decreases in volume when placed in a beaker filled with liquid nitrogen (–196°C).

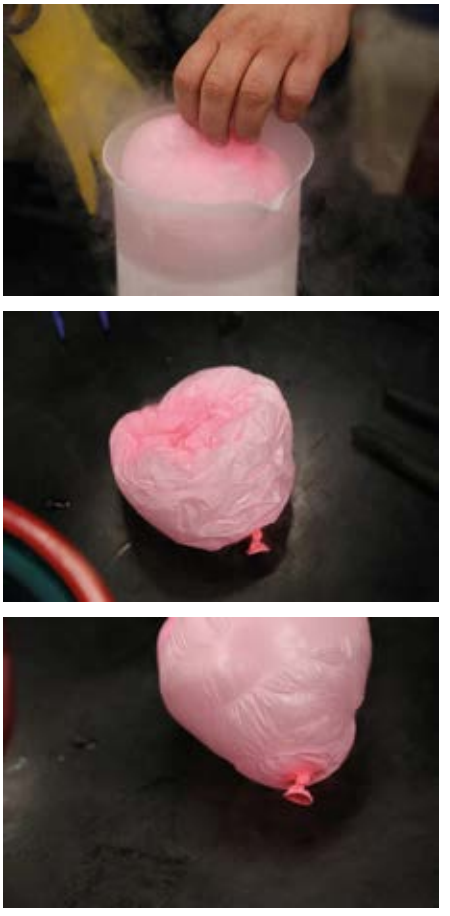

FIGURE 4 A balloon increases in volume when removed from the beaker and placed on a bench at room temperature.

Numerical representation of Charles' law

The relationship between temperature and volume is represented in Figure 5. A cylinder filled with a gas is kept at a constant pressure of 1 atm. The cylinder is heated, and the change in volume is recorded to determine the effect of temperature on the volume of a gaseous system. Table 2 summarises the data gathered in the experiment.

Figure 5 shows the graph extrapolating to a volume of 0 mL. The y-axis intercept here is –273.15°C. This value is exactly 0 Kelvin (K).

The graph in Figure 5 demonstrates that there is a linear relationship between temperature and volume. Therefore, volume is proportional to temperature ($V \propto T$). This relationship can therefore be represented mathematically as:

$$\frac{V}{T} = \text{a constant}$$

Any change in temperature or volume can be calculated from the equation:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

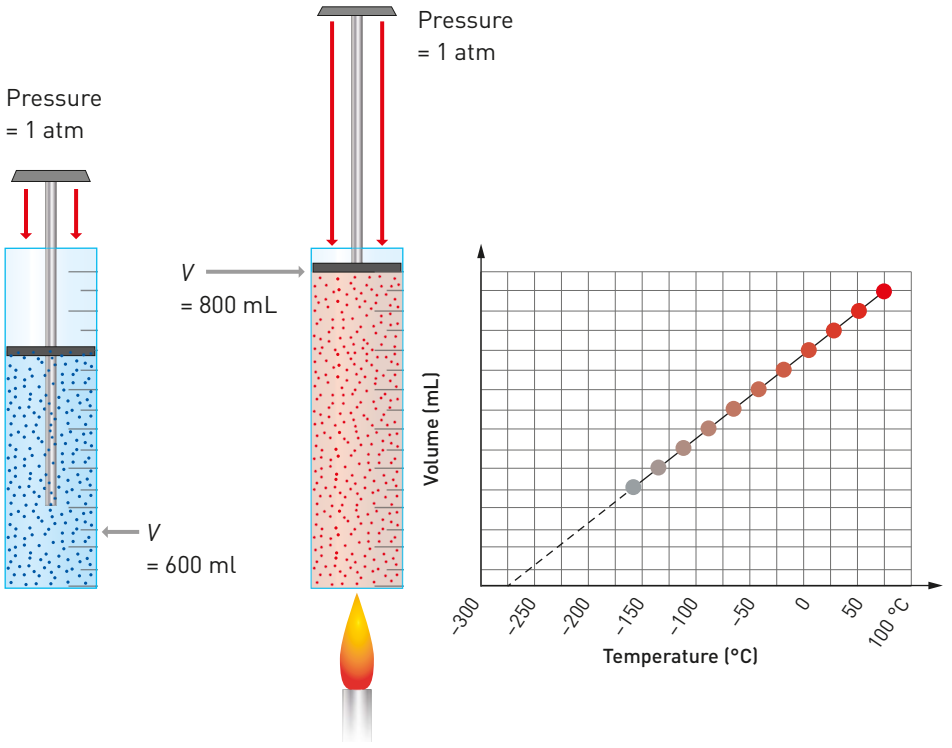

FIGURE 5 The pressure (1 atm) applied to the cylinder is identical throughout the experiment. When the temperature is increased, the volume that the gas occupies also increases.

WORKED EXAMPLE 14.2B

At 293.15 K the volume in a cylinder is 200 mL. If the temperature is raised to 335.15 K determine the volume of the cylinder.

SOLUTION

1 Substitute values into the equation

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{0.200}{293.15\ \text{K}} = \frac{V_2}{335.15\ \text{K}}$$

2 Rearrange to solve for V_2.

$$V_2 = \frac{0.200}{293.15} \times 335.15$$
$$= 0.229\ \text{L}$$
$$= 229\ \text{mL}$$

Pressure versus volume – Boyle's law

In a variable volume system, in order to increase the pressure, the volume must decrease. If the volume decreases, the molecules within the system have less space to move in. Therefore, the molecules hit the walls of the container more often, increasing the pressure. Boyle's law, which states that pressure is inversely proportional to volume in fixed temperature systems.

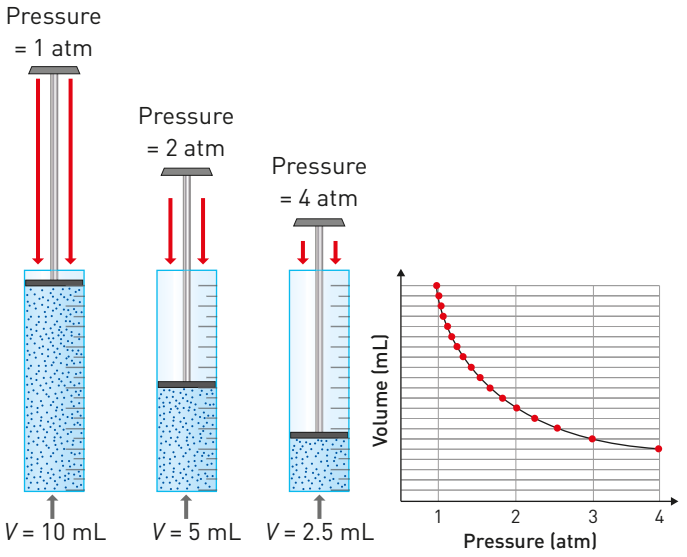

FIGURE 6 The temperature of the cylinder stays the same, but when the plunger is depressed, pressure increases and volume decreases.

Numerical representation of Boyle's law

The relationship between pressure and volume is represented in Figure 6. A syringe filled with a gas is kept at a constant temperature. The plunger on the syringe is pushed down. As the plunger is pushed down further, more pressure is required to keep pushing it. The volume within the syringe decreases the further down the plunger is pushed and the pressure increases. Therefore, as pressure increases, volume decreases. Table 2 summarises the data gathered in the experiment.

TABLE 2 Data obtained from the experiment in Figure 7

Volume (L)	Pressure (atm)	$P \times V$
10	1	1 × 10 = 10
5	2	2 × 5 = 10
2.5	4	4 × 2.5 = 10

molar volume
the volume occupied by 1 mole of a gas at a specified temperature and pressure

standard temperature and pressure (STP)
standard conditions for temperature and pressure, when gases are measured at 1 atm and 0°C or 273.15 K

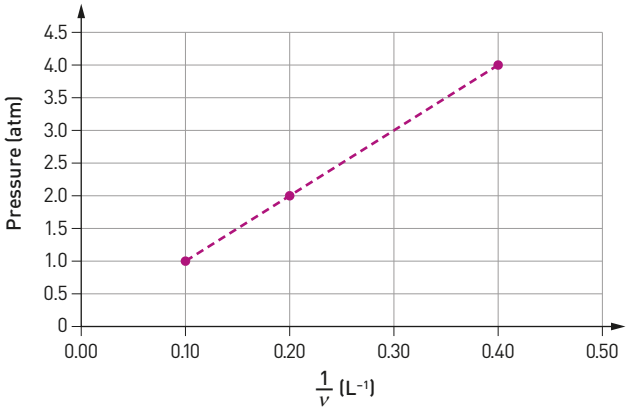

FIGURE 7 Graph of the relationship between pressure and $\frac{1}{V}$.

When the pressure is graphed against $\frac{1}{V}$, a linear relationship is seen (Figure 7). The graph demonstrates that there is a linear relationship between pressure and inverse volume. Therefore, pressure is proportional to $\frac{1}{V}$ ($P \propto \frac{1}{V}$). This relationship can be represented mathematically as:

$$P \times V = \text{a constant}$$

As with Gay-Lussac's and Charles' law, any change in pressure or volume at a constant temperature can be calculated using the following equation:

$$P_1 V_1 = P_2 V_2$$

WORKED EXAMPLE 14.2C

Table 3 demonstrates that at 1 atm of pressure, a sample of argon gas occupies 10 L of a cylinder. If the pressure is increased to 1.7 atm, what volume of the cylinder will the gas occupy?

SOLUTION

1 Substitute values into the equation.

$$P_1 V_1 = P_2 V_2$$
$$1 \text{ atm} \times 10 \text{ L} = 1.7 \text{ atm} \times V_2$$

2 Rearrange to solve for V_2.

$$V_2 = \frac{1 \text{ atm} \times 10 \text{ L}}{1.7 \text{ atm}}$$
$$= 5.9 \text{ L}$$

3 Does this answer make sense? The pressure has increased, so volume should decrease.

standard laboratory conditions (SLC) standard conditions for temperature and pressure under laboratory conditions, when gases are measured at 1 atm and 25°C or 298.15 K

Molar volume – Avogadro's hypothesis

The relationship between pressure, volume and temperature can be extended to the number of atoms or molecules in a gaseous system. Avogadro hypothesised that under identical conditions of temperature and pressure, equal volumes of gases contain an equal number of gas particles. Therefore, one mole of any gas (6.02×10^{23} molecules), at the same temperature and pressure, will occupy a specific volume (Figure 8). This is called the **molar volume**.

Standard conditions – analysis of gases at the same temperature and pressure

To compare the gases in two different systems, scientists keep the systems at the same standard conditions, which are sets of conditions that are kept constant when measuring gases. By keeping temperature and pressure constant, the only variables should be volume and number of moles.

- At **standard temperature and pressure (STP)**, 0°C/273.15 K and 1 atm, 1 mole of any gas occupies 22.4 dm^3, which is the equivalent of 22.4 L. In 1982, IUPAC changed their definition of STP to be 273.15 K and 100 kPa and this is the definition used in the data booklet. Under these conditions, one mole of a gas occupies 22.7 dm^3 or 22.7 L.
- At **standard laboratory conditions (SLC)**, 25°C/298.15 K and 1 atm, 1 mole of any gas occupies 24.5 dm^3, which is the equivalent of 24.5 L.

Numerical representation of Avogadro's law at STP and SLC

The relationship between moles and volume is represented by the equation:

$$n = \frac{V}{V_m}$$

where n is the number of moles, V is the volume (L) and V_m is the molar volume at STP (22.4 L).

At STP this equation becomes:

$$n_{STP} = \frac{V}{22.4}$$

At SLC this equation becomes:

$$n_{SLC} = \frac{V}{24.5}$$

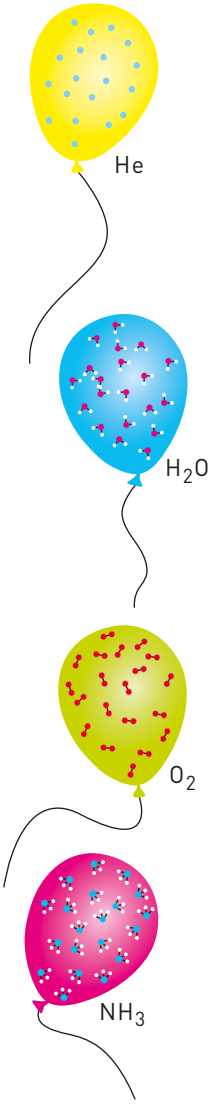

FIGURE 8 Four balloons of the same volume, at the same temperature and pressure, have the same number of gas particles.

WORKED EXAMPLE 14.2D

How many moles of nitrogen gas are there in a 40.0 L cylinder at STP?

SOLUTION

Substitute values into the equation.

$$n = \frac{V}{V_m}$$

$$n(N_2) = \frac{40\ L}{22.4\ L\ mol^{-1}}$$

$$= 1.79\ mol$$

WORKED EXAMPLE 14.2E

What is the mass of helium in a 20.0 L cylinder at STP?

SOLUTION

1 Determine the number of moles from the molar volume equation.

$$n = \frac{V}{V_m}$$

$$n(He) = \frac{20\ L}{22.4\ L\ mol^{-1}}$$

$$= 0.89\ mol$$

2 Determine the mass by using the mole equation.

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

$$= 0.89\ mol \times 4.0\ mol^{-1}$$

$$= 3.57\ g$$

WORKED EXAMPLE 14.2F

2.3656 g of an unknown monatomic gas fills a balloon to 632.34 mL at STP. What is the identity of the gas?

SOLUTION

1 Determine the number of moles from the molar volume equation.

$$n = \frac{V}{V_m}$$

$$n(unknown) = \frac{0.63234\ L}{22.4\ L\ mol^{-1}}$$

$$= 2.82 \times 10^{-2}\ mol$$

2 Determine the molar mass by using the mole equation.

$$M(unknown) = \frac{m}{n}$$

$$= \frac{2.3656\ g}{2.82 \times 10^{-2}\ mol}$$

$$= 83.9\ g\ mol^{-1}$$

The gas with a molar mass of 83.9 g mol^{-1} is krypton. This can be found by looking at a periodic table (Appendix 1).

CASE STUDY 14.2

High-altitude sports

As you travel higher into the atmosphere, the pressure exerted on the body decreases. This is because gas molecules in the atmosphere are attracted to Earth by gravitational forces. As there are more gas particles at the Earth's surface, there is more pressure. At higher altitudes, there are fewer gas molecules.

Mount Everest, the tallest mountain in the world at 8848 m high, has an air pressure of 33.7 kPa at its peak, compared to 101.3 kPa at sea level. The air has far fewer gas particles, a lower pressure and is referred to as being 'thinner'. At this high altitude, a breath of air contains approximately 33% of the oxygen that a typical breath at sea level contains. People who climb Mount Everest must become acclimatised to lower concentrations of oxygen in the air and often find themselves puffing after only a few steps.

Athletes sometimes train at higher altitudes before a competition to become acclimatised to higher altitudes and lower levels of oxygen. In response to lower concentrations of oxygen, the body produces more red blood cells, which deliver more oxygen to the muscles for energy. When the athletes return to sea level, their bodies retain the red blood cells for 10–20 days. The red blood cells can deliver more oxygen than normal to the muscles, resulting in an increase in energy and physical activity.

FIGURE 9 These climbers are at 7000 m on Mount Everest. They would have acclimatised to lower concentrations of oxygen in the air.

CHALLENGE 14.2B

Density calculations

Calculate the density (g mL^{-1}) of argon gas at standard temperature and pressure (STP).

CHECK YOUR LEARNING 14.2

Describe and explain

1 **Describe** the relationship between each of the following variables in one sentence.
   - **a** Pressure and volume
   - **b** Pressure and temperature
   - **c** Volume and temperature

2 **Explain** why heating the air inside a hot air balloon causes the balloon to rise into the atmosphere.

3 **Calculate** the volume of:
   - **a** 3.2 mol of water vapour at SLC
   - **b** 2.9 g of oxygen gas at STP
   - **c** 9.37×10^{21} nitrogen molecules at SLC.

4 **Calculate** the:
   - **a** mass of 120 mL of CO_2 at SLC
   - **b** number of H atoms in 525 mL of CH_4 at STP.

5 **Calculate** temperatures in Kelvin for the following values:
   - **a** –25°C
   - **b** 13°C
   - **c** 273°C

Apply, analyse and interpret

6 At a constant temperature, and at a pressure of 106.5 kPa, a gas sample occupies 120 mL of a container. **Deduce** what pressure would be required to increase the volume by 680 mL.

7 The graph in Figure 10 represents the data obtained when a volume of gas is heated at a constant pressure.

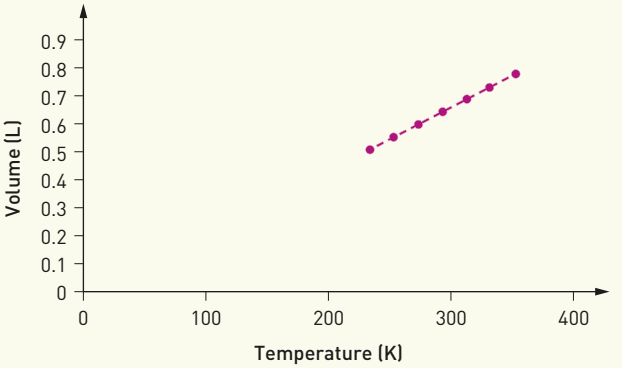

FIGURE 10 The effect on volume when temperature is raised in a variable volume system.

   - **a** **Determine** the constant that represents the relationship between volume and temperature.
   - **b** The temperature of the gas continues to increase until it reaches 473.15 K. **Determine** what volume the gas occupies at 473.15 K.

Investigate, evaluate and communicate

8 Theoretically, the volume of a gas could be 0 L.
   - **a** **Comment** on what would need to happen for this to occur.
   - **b** **Discuss** why is this theoretical.

9 An experiment found the density of nitrogen gas to be 1.2504 g L^{-1} at STP. **Predict** what would happen to the pressure of the system if the density of this gas was increased in a:
   - **a** fixed volume container?
   - **b** variable volume container?

obook assess

Check your obook assess for these additional resources and more:

- » Student book questions
  Check your learning 14.2
- » Challenge
  14.2A Gay-Lussac's law using Kelvin
- » Challenge
  14.2B Density calculations
- » Suggested practical
  14.2 Boyle's law

14.3 Ideal gases

KEY IDEAS

In this section, you will learn about:

- combining the gas laws into the ideal gas equation
- the universal gas constant *R*.

The gas laws covered in section 14.2 can be summarised as:

- $\frac{P}{T}$ = a constant
- $\frac{V}{T}$ = a constant
- $P \times V$ = a constant
- $V_m = \frac{V}{n}$ where molar volume is a constant at a specific temperature and pressure.

These laws can be combined to generate the **ideal gas** equation:

$$R = \frac{P \times V}{n \times T}$$

where R is the **universal gas constant** and temperature (T) is expressed in Kelvin.

ideal gas
a gas whose particles do not interact with one another and occupy no volume

universal gas constant (R)
the proportionality constant, defining gas behaviour under ideal conditions

The ideal gas equation, also called the universal gas equation, provides chemists with a simpler method of determining experimental values in gas experiments without having to complete multiple calculations.

Under standard conditions, where pressure is 1 atm, 1 mole is contained within a volume of 22.4 L and temperature is 273.15 K, the universal gas constant, R, can be calculated:

$$R = \frac{1 \text{ atm} \times 22.4 \text{ L}}{1 \text{ mol} \times 273.15 \text{ K}}$$
$$= 0.082\,05 \text{ L atm K}^{-1} \text{ mol}^{-1}$$

Study tip

These values are in the databook:
$R = 0.082$ L atm K^{-1} mol^{-1}
$= 8.314$ J K^{-1} mol^{-1}

To convert the universal gas constant to SI units, use $P = 101\,325$ N m^{-2} (where 1 N m^{-2} = 1 Pa) and $V = 0.0224$ m^3. This gives:

$$R = \frac{101\,325 \text{ N m}^{-2} \times 0.0224 \text{ m}^3}{1 \text{ mol} \times 273.15 \text{ K}}$$
$$= 8.314 \text{ N m K}^{-1} \text{mol}^{-1}$$

As 1 N m is the equivalent of 1 J, the universal gas constant is most commonly expressed as $R = 8.314$ J K^{-1} mol^{-1}.

Ideal gas equation

The ideal gas equation can be summarised as:

$$PV = nRT$$

where P is pressure (kPa), V is volume (L), n is the number of moles, T is temperature (K) and $R = 8.314$ J mol^{-1} K^{-1}.

WORKED EXAMPLE 14.3A

What volume does 0.94 mol of nitrogen (N_2) gas occupy at STP?

SOLUTION

1 Rearrange the ideal gas equation.

$$V = \frac{nRT}{P}$$

2 Solve for volume.

$$V(N_2) = \frac{0.94 \text{ mol} \times 8.314 \text{ J K}^{-1} \text{ mol}^{-1} \times 273.15 \text{ K}}{101.3 \text{ kPa}}$$

$$= 21.07 \text{ L}$$

As moles can be expressed as:

$$n = \frac{m}{M}$$

The ideal gas equation can be expanded to:

$$PV = \frac{mRT}{M}$$

WORKED EXAMPLE 14.3B

What volume does 6.00 g of oxygen (O_2) gas occupy at SLC?

SOLUTION

1 Rearrange the equation.

$$V = \frac{mRT}{MP}$$

2 Solve for volume.

$$V(O_2) = \frac{6.00\text{g} \times 8.314 \text{ J K}^{-1} \text{ mol}^{-1} \times 298.15 \text{ K}}{32 \text{ g} \times 101.3 \text{ k Pa}}$$

$$= 4.59 \text{ L}$$

3 Does this answer make sense? We know that 1 mol of O_2 has a mass of 32 g (2 × 16 g) and would occupy 24.5 L. Six grams is roughly 20% of 1 mol, so should occupy 20% of 24.5 L, about 4.6 L.

Study tip

There can be many ways of answering these types of questions. As long as your method is correct, the precise method is not important.

WORKED EXAMPLE 14.3C

What mass of carbon dioxide gas occupies a 2.00 L cylinder at STP?

SOLUTION

1 Rearrange the equation.

$$m = \frac{PVM}{RT}$$

2 Solve for mass.

$$m(CO_2) = \frac{101.3 \text{ kPa} \times 2.00 \text{ L} \times 44 \text{ g mol}}{8.314 \text{ J K}^{-1} \text{ mol}^{-1} \times 273.15 \text{ K}}$$

$$= 3.93 \text{ g}$$

FIGURE 1 A volcanic gas plume.

CHECK YOUR LEARNING 14.3

Describe and explain

1 **Explain** why there are two values for the universal gas constant, *R*.

2 **Explain** the two requirements of an 'ideal' gas.

Apply, analyse and interpret

3 **a** **Determine** the pressure 2 moles of H_2 gas would exert on a 250 mL cylinder at 10°C.

b A 0.1 g sample of He is used to fill a balloon. The sample has a pressure of 1.3 atm at 19°C. **Determine** what volume it occupies in the balloon.

c At STP, **determine** the mass of xenon gas that can be used to fill a 1.4 L cylinder.

4 **Compare** the following samples and **determine** which has the greater mass: 1.3 L of N_2 at 0.6 atm and 15°C or 0.5 L of water vapour at 123.8 kPa and 303.6 K.

Investigate, evaluate and communicate

5 **Prove** that the gas density equation is:

$$\rho = \frac{PM}{RT}$$

where ρ is density in g L^{-1} and *M* is molar mass.

6 The density of a sample of propane gas is 1.82 g L^{-1} at 1 atm and 21.1°C. **Prove** that the molecular formula of propane is C_3H_8.

Check your obook assess for these additional resources and more:

- » Student book questions
  Check your learning 14.3
- » Increase your knowledge
  14.3A Using the ideal gas equation
- » Increase your knowledge
  14.3B Ideal gases
- » Increase your knowledge
  Case study on ideal gases

SCIENCE AS A HUMAN ENDEAVOUR

14.4 The behaviour of gases and their effect on the body while scuba diving

KEY IDEAS

In this section, you will learn about:

- applying gas laws in scuba diving
- health risks related to scuba diving.

Scuba (self-contained underwater breathing apparatus) allows humans to explore and play under water. Scuba gear provides the oxygen required for divers to breathe under water. It is essential to have a thorough understanding of gas laws and their effect on the human body, so as not to cause harm or death.

Atmospheric pressure at sea level is 1 atm. Pressure increases by 1 atm every 10 m in depth. This compresses the body and has adverse effects if the diver does not know how to manage the extra pressure.

FIGURE 1 A scuba diver explores the Great Barrier Reef, near Port Douglas.

Effects of pressure on the body

When pressure on the body increases, the body is compressed and the volume of gas within the body decreases. Conversely, when pressure decreases, the volume of gas within the body increases. This phenomenon follows Boyle's law.

The bends – a problem while ascending

As pressure on the body decreases, nitrogen is forced from the bloodstream into the body's tissues. It is a similar effect to when opening a bottle of fizzy cordial. When the bottle is under pressure and the lid is then removed, large gas bubbles form rapidly, causing the bottle to overflow.

Nitrogen gas must be removed and can only be done so naturally and gradually, by the body. A slow ascension rate is the best method to remove the excess nitrogen. Taking breaks ('decompression stops') is also recommended as a diver ascends – the diver remains at a certain depth until it is safe to continue ascending. The nitrogen slowly moves from the tissues back into the bloodstream from where it is exhaled. If the nitrogen within the tissues is not removed slowly, if forms bubbles that can harm the body in several ways.

Furthermore, as the body ascends, any gaseous spaces will expand as the volume of air within them also expands. This can damage the fragile tissue surrounding the air sacs in the lungs, for example.

Decompression/hyperbaric chambers

A hyperbaric or decompression chamber (Figure 2) is used to treat divers with the bends by simulating the increased pressure experienced in deep dives. The chamber reduces this pressure gradually, allowing the body to reacclimatise to sea level pressure and remove nitrogen

bubbles from body tissues. If the treatment is given early enough, there is a significantly greater chance of a full recovery.

Decompression chambers are also used in saturation diving, which is when divers work as deep as 180 m to fix pipes on oil rigs. Without decompression chambers it would take about 6 days to get the divers safely to this depth, without causing illness. Instead, the divers live in a decompression chamber for 28 days at a time. After slowly acclimatising to pressures 19 times that of sea level, the divers are lowered to the sea floor. At the end of every day, the pressurised chamber is brought back to the surface and placed on a ship. The divers are kept at pressure over night and the chamber is lowered back into the water and down to depth. This process means the divers do not need to decompress at the end of each work day and reduces stress on the body.

FIGURE 2 A decompression chamber.

Barotrauma – a problem while descending

While descending, extra pressure is placed on the body, which damages tissues that contain or are surrounded by air spaces. As the body descends, the pressure it experiences increases. According to Boyle's law, this causes any gas spaces within the body to compress and reduce in volume. This can rupture the eardrums and cause sinus pain. To avoid these issues, divers are taught how to clear their ears and equalise the pressure in their body.

CHALLENGE 14.4

Charles' law and scuba diving

Does Charles' law apply to scuba diving (the diver or the oxygen tank)? Justify your answer, using your knowledge of the law.

CHECK YOUR LEARNING 14.4

Describe and explain

1 **Explain** why pressure varies when diving and what effect this has on gaseous volume within the body.
2 **Explain** how a decompression chamber works.

Investigate, evaluate and communicate

3 Pressure decreases and air becomes thinner with increasing altitude. Apply your knowledge of gas laws to **predict** what may happen to the human body at higher altitudes.

Check your obook assess for these additional resources and more:

- » Student book questions
  Check your learning 14.4
- » Challenge
  14.4 Charles' law and scuba diving
- » Weblink
  Hyperbaric chambers
- » Weblink
  Scuba divers descending and ascending

CHAPTER 14

Review

Chapter summary

14.1

- Gas properties can be explained by kinetic molecular theory. Pressure is measured as the force exerted, per unit of area, on the walls of a container.

14.2

- The relationship between the variables of pressure, volume and temperature are represented mathematically as:
- Gay-Lussac's law:

  $\frac{P}{T}$ = a constant

  $\frac{P_1}{T_1} = \frac{P_2}{T_2}$
- Charles' law:

  $\frac{V}{T}$ = a constant

  $\frac{V_1}{T_1} = \frac{V_2}{T_2}$
- Boyle's law:

  $P \times V$ = a constant

  $P_1 V_1 = P_2 V_2$
- Molar volume is the volume that one mole of a gas occupies at the same pressure and temperature. It can be measured at:
  - STP: standard temperature and pressure (1 atm (101.3 kPa) and 0°C (273.15 K)).
  - SLC: standard laboratory conditions (1 atm (101.3 kPa) and 25°C (298.15 K)).

  It is calculated as:

  $n = \frac{V}{V_m}$

14.3

- The ideal gas equation is $PV = nRT$ or $PV = \frac{mRT}{M}$. It is used to calculate the pressure, volume or temperature of a gas using the universal gas constant, R, which is equal to 8.314 J mol^{-1} K^{-1}.

14.4

- Water pressure increases with the depth of the water. This causes an increase in the volume of gases within the body. The decrease in pressure when ascending causes an increase in volume of gases within the body. If not managed appropriately, these changes can damage the tissues and organs.

Key terms

- gas laws
- ideal gas
- kinetic energy
- kinetic molecular theory
- molar volume
- pressure
- standard laboratory conditions (SLC)
- standard temperature and pressure (STP)
- temperature
- volume
- universal gas constant (R)

Key formulas

Pressure	$P = \frac{F}{A}$
Kinetic energy	$\text{KE} = \frac{1}{2} m v^2$
Relationship between temperature and pressure	$\frac{P}{T}$ = a constant
	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$
Relationship between temperature and volume	$\frac{V}{T}$ = a constant
	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Relationship between pressure and volume	$P \times V$ = a constant
	$P_1 V_1 = P_2 V_2$
Relationship between number of moles and volume	$n = \frac{V}{V_m}$
	$n_{\text{STP}} = \frac{V}{22.4}$
	$n_{\text{SLC}} = \frac{V}{24.5}$
Ideal gas equation	$PV = nRT$
Ideal gas equation	$PV = \frac{mRT}{M}$

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; *** = high.

Multiple choice

1 Kinetic molecular theory does not state that:
- **A** particles move in straight lines
- **B** particle movement is random
- **C** particle movement is circular
- **D** particle collisions are elastic and conserve energy.

2 According to Boyle's law, as the volume of a container decreases:
- **A** pressure decreases
- **B** temperature increases
- **C** pressure increases
- **D** temperature decreases.

3 The pressure of a container of gas at 50°C and 5 atm is increased by 97 kPa. What is the temperature increase of the container?
- **A** 62°C
- **B** 59.57°C
- **C** 50.02°C
- **D** 9.57°C

4 Which of the following has the most molecules?
- **A** 1.00 L of N_2 at STP
- **B** 2.5 L of O_2 at 2°C and 0.5 atm
- **C** 0.9 L of H_2O at SLC
- **D** 0.5 L of CO_2 at 50°C and 1.5 atm

5 What are the standard temperature and pressure conditions and the standard laboratory conditions?
- **A** STP: 0°C and 100 kPa. SLC: 0°C and 100 kPa
- **B** STP: 25°C and 1 atm. SLC: 0°C and 1 atm
- **C** STP: 0°C and 101.3 kPa. SLC: 25°C and 101.3 kPa
- **D** STP: 25°C and 101.3 kPa. SLC: 0°C and 101.3 kPa

Short answer

Describe and explain

★ **6** A pressure cooker is used to increase the pressure in a pot. **Explain** how it is able to cook food.

FIGURE 1 A pressure cooker

★ **7** A variable volume container of gas is kept at a constant temperature. **Explain** the impact on volume, temperature and pressure when the:
- **a** volume is increased
- **b** temperature is decreased.

Apply, analyse and interpret

★★ **8** A scientist wishes to inflate four balloons with the same number of molecules of four different gases. **Determine** the volume needed to occupy each balloon if they are filled with hydrogen (H_2), oxygen (O_2), helium (He) and argon (Ar).

★★★ **9** **a** A 50 mL vessel has a pressure of 104.5 kPa. If the volume is increased to 420 mL, **determine** the pressure of the system in atmospheres.
- **b** **Determine** the change in temperature needed to increase the pressure of a 117.3 kPa container at 20°C to 134.2 kPa.
- **c** The temperature of water at sea level is 25°C and the pressure is 101.3 kPa. If pressure increases by 1 atm at a 10 m depth, **determine** the temperature of the water at this depth.

★★★ **10** **a** A 40 L cylinder of O_2 at 13°C has a pressure of 98 kPa. **Determine** the mass of the gas.
- **b** **Determine** the volume occupied by 12 g of water vapour at a pressure of 67 kPa and a temperature of 123 K.
- **c** **Determine** the pressure of a 4 L gas container containing 67 mol of He at 5°C.

Investigate, evaluate and communicate

★★ **11** A student takes a plastic pump water bottle to school on a snowy winter's day. The student travels to school by car and the car heater is on. It takes almost an hour to get to school. On arriving at school, the student has a mouthful of water. A small amount of gas exits the bottle when it is opened. At lunchtime, the student takes the water bottle outside and accidentally leaves it there. The sides of the bottle become crumpled in.

Discuss what has happened to the water bottle to cause both effects.

★★ **12** Ping pong balls can become dented throughout the course of a game. **Propose** how could you remove the dent by using your knowledge of gas laws.

FIGURE 2 A dented ping pong ball

★★ **13** A turkey timer is used to indicate when a turkey is cooked. **Investigate** the turkey timer and use gas laws to explain how it works.

FIGURE 3 Turkey timer

★★★ **14** Tyres are a key safety feature of a car, responsible for the friction between the car and road, which allows for acceleration and deceleration, or breaking.

FIGURE 4 Tyres

a **Discuss** why tyre pressure should be checked on a cool day.

b **Predict** what would happen if you over-filled a car tyre (without it exploding). **Justify** your predictions.

c **Predict** what would happen if you under-filled a car tyre. **Justify** your predictions.

15 Based on what you know about pressure, **hypothesise** what might happen to a coke can if it sinks to the bottom of the Mariana trench? **Conduct** some research to see if you are correct

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CHAPTER 15

Aqueous solutions and molarity

Water is one of the most abundant and important molecules on Earth. Approximately 71% of Earth's surface is covered in water, with more than 96% of that water held in the oceans and seas. This is salt water so it is not suitable for human consumption unless the salt is removed. Incredibly, the remaining 4% (or just under) is all of the fresh water used to sustain the world's 7.6 billion life forms on our planet. Of this 4%, approximately one-third is trapped in the polar ice caps.

As water is one of the most essential molecules on Earth, it is important to recognise its structure and properties.

In 1766, British chemist and physicist Henry Cavendish noted that water formed when various gases in the atmosphere reacted with one another. Before this, water was believed to be an element, but Cavendish demonstrated that the reaction involved two different elements and therefore water must be a compound. Cavendish did not know at that time that the elements involved were hydrogen and oxygen.

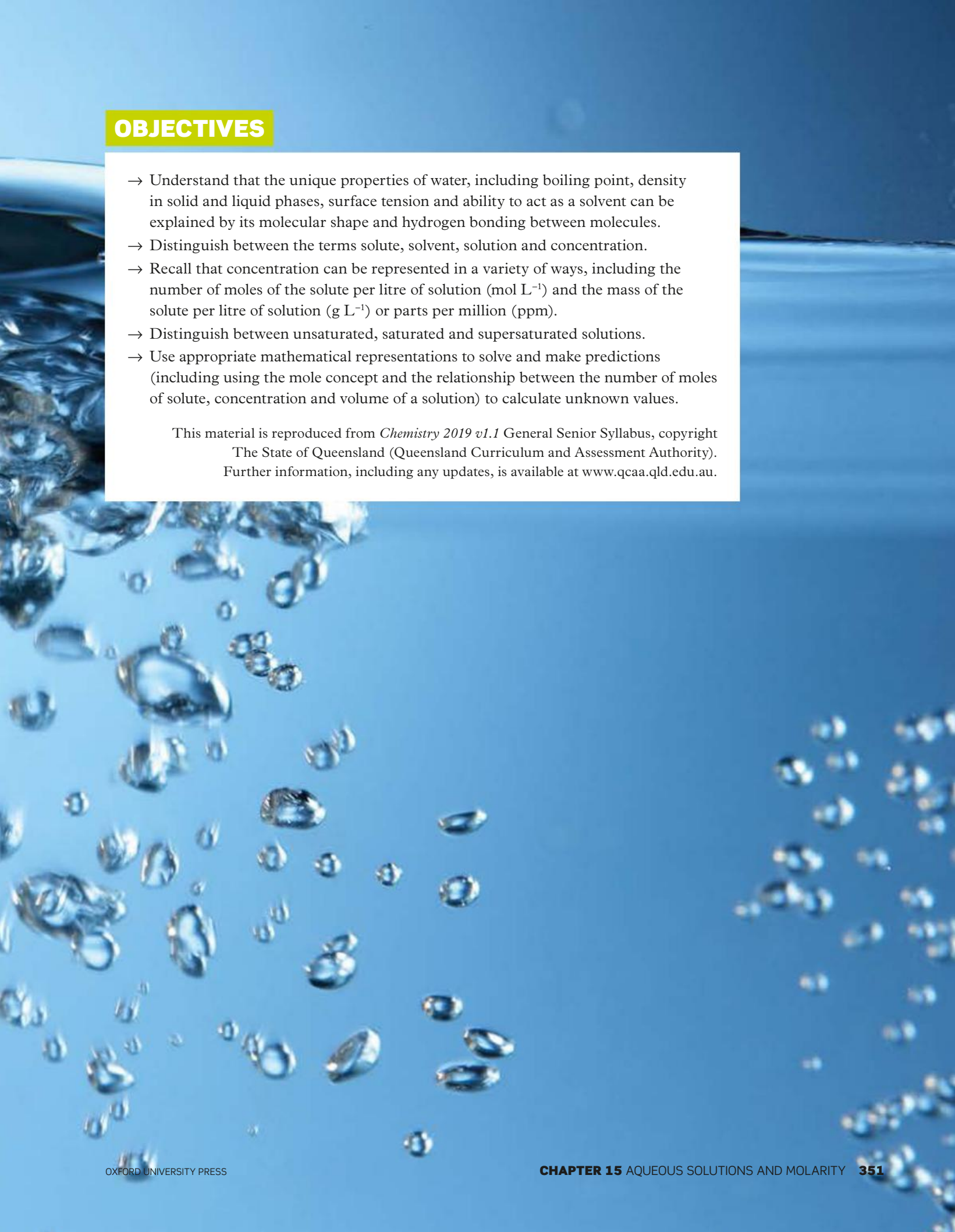
French chemist Antoine Lavoisier, the 'father of modern chemistry', built on Cavendish's work to name hydrogen, which derived from the Greek *hydro*, meaning 'water', and *gen*, meaning 'something produced'. Therefore, hydrogen was named for its ability to form water.

Not long after this, Cavendish hypothesised that water must be two parts hydrogen and one part oxygen. This was experimentally confirmed several years later when German chemist Johann Wilhelm Ritter passed an electric current through water, in a process called electrolysis. He was able to generate hydrogen gas at one electrode and oxygen gas at the other. Because there was twice as much hydrogen as oxygen, he concluded that water must be made of two parts hydrogen and one part oxygen.

Since then, water's behaviour has been researched extensively and scientists have discovered much about its properties and its involvement in reactions.

This chapter explores some of water's properties and its role in the chemistry of aqueous solutions.

FIGURE 1 Aqueous solutions consist of a substance dissolved in water.



OBJECTIVES

- → Understand that the unique properties of water, including boiling point, density in solid and liquid phases, surface tension and ability to act as a solvent can be explained by its molecular shape and hydrogen bonding between molecules.
- → Distinguish between the terms solute, solvent, solution and concentration.
- → Recall that concentration can be represented in a variety of ways, including the number of moles of the solute per litre of solution (mol L^{-1}) and the mass of the solute per litre of solution (g L^{-1}) or parts per million (ppm).
- → Distinguish between unsaturated, saturated and supersaturated solutions.
- → Use appropriate mathematical representations to solve and make predictions (including using the mole concept and the relationship between the number of moles of solute, concentration and volume of a solution) to calculate unknown values.

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15.1 Properties of water

KEY IDEAS

In this section, you will learn about:

- the shape and polarity of water
- the properties of water based on its shape and polarity.

Water is a unique and incredibly useful compound. You should recall from Chapter 12 that molecules have shapes based on the electrons contained within their valence shells. Oxygen has two lone pairs of electrons when it bonds covalently to two hydrogen atoms. This means that the water molecule adopts a bent structure due to the lone pairs of electrons on the oxygen, as seen in Figure 1.

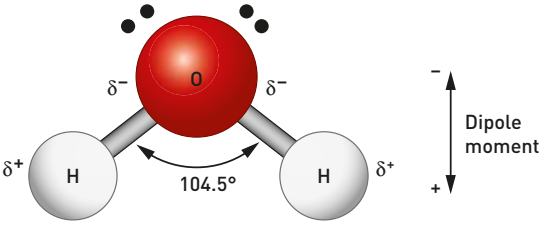

FIGURE 1 The bent shape of a water molecule (H_2O) and resulting polarity due to the electronegativity of oxygen.

Polarity

electronegativity
the attraction of a positively charged nucleus for the negatively charged electrons of a neighbouring atom

Oxygen is **electronegative**, so it pulls on the electrons in the covalent bonds, meaning that they spend more time closer to the oxygen nucleus. This means that in a O–H bond, the negative charge is closer to the oxygen atom than the hydrogen atom. Because water is bent, a **dipole** is formed in which oxygen has a partial negative charge and each hydrogen has a partial positive charge. The presence of partial charges in a molecule makes it a polar molecule.

dipole
a bond or a molecule that has a partial positive charge at one end and a partial negative charge at the opposite end

Intermolecular forces

As seen in Chapter 12, hydrogen bonds form by the electrostatic attraction between partially positive hydrogen and a partially negative oxygen, nitrogen or fluorine atom. In water, the partial positive charge on the hydrogen of one molecule is attracted to the partial negative charge on the oxygen of a different molecule (Figure 2), forming a hydrogen bond.

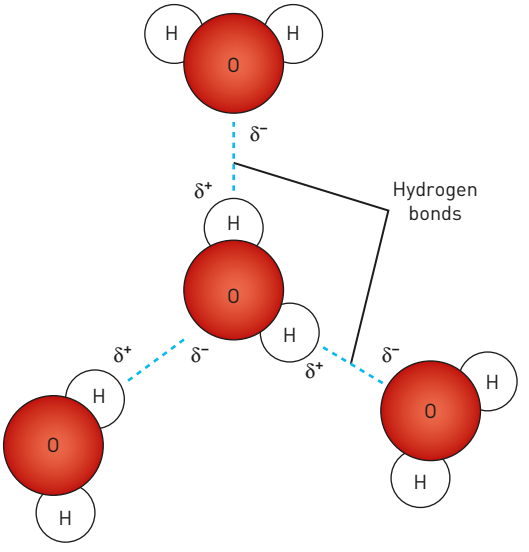

FIGURE 2 Hydrogen bonding between water molecules.

Properties of water

polarity
the electronegativity of atoms and the shape of a molecule gives the molecule an overall positive and negative side, creating a partial charge

The shape, **polarity** and intermolecular forces can be used to explain the properties of water, such as the:

- boiling point of water
- density of water in the solid and liquid states
- surface tension of water.

Boiling point

boiling point
the temperature at which a liquid becomes a gas

The **boiling point** of water is the temperature at normal atmospheric pressure at which it changes state from a liquid to a gas. In different states, the particles exhibit different behaviours.

- In the solid state, particles experience so little movement that they only vibrate. They pack tightly together because they are strongly attracted to one another and the solid maintains its shape as a result of the strength of this attraction (Figure 3). The force or energy of their attraction to one another is greater than the energy of their movement.
- In the liquid state, particles pack closely together. However, unlike in the solid state, they move readily over one another. This is why a liquid fills the space at the bottom of a container. In other words, the particles have enough energy to move, but not enough energy to separate from one another.
- In the gaseous state, particles have more energy than in the liquid phase. The attractions that held them together as a liquid are overcome by the energy involved with their movement. They travel in straight lines until they reach an object that they bounce off and change direction. In other words, they have too much energy to stay together.

Therefore, for water to change from a liquid to a gas, the particles must increase their movement to the point where the kinetic energy involved with their movement is greater than the energy that attracts, through dispersion forces and hydrogen bonding, one water molecule to another.

As temperature is a measure of the average kinetic energy of particles, boiling point signifies the temperature at which particles have enough energy to overcome intermolecular forces and transition between a liquid and a gas. For water, this temperature is 100°C or 373.15 K.

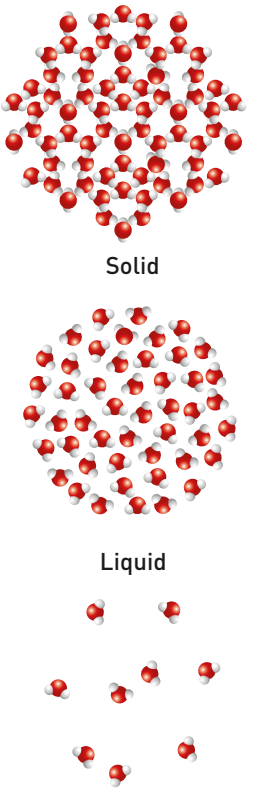

FIGURE 3 The interaction of water particles in solid, liquid and gaseous states.

Density

density
the mass of a substance per unit of volume

Density is a measure of the mass of a substance per unit of volume and can be demonstrated by the equation:

$$\rho = \frac{m}{V}$$

where ρ is density (g mL^{-1}), m is mass (g) and V is volume (mL).

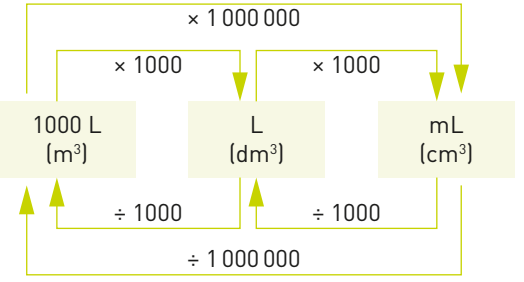

FIGURE 4 Unit conversions of volume.

The density of water is 1.00 g mL^{-1} at room temperature, meaning that 1.00 g of water occupies 1 mL of volume. Volume can be measured in millilitres (mL), centimetres cubed (cm^3), litres (L), decimeters cubed (dm^3) or metres cubed (m^3). Figure 4 demonstrates the conversion between units of volume.

The greater the mass in a fixed volume, the greater the density. When water freezes to become solid ice, the water particles align to leave spaces within its structure (Figure 3). This means that 1 L of ice has a lower mass than 1 L of liquid water and liquid water has a greater density than solid water. This is why ice floats on water (Figure 5). Water is exceptional because most compounds are more dense as solids than as liquids.

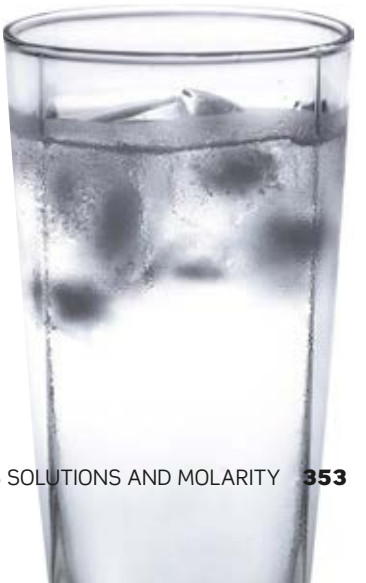

FIGURE 5 Ice floats on water because ice is less dense than liquid water.

Surface tension

surface tension
the ability of the surface of a liquid to resist an external force

hydrogen bonding
the strongest form of dipole–dipole interaction in which a partially negative atom (typically F, O or N) has an electrostatic interaction with a partially positive hydrogen atom

Surface tension is the ability of the surface of a liquid to resist an external force. As water molecules experience **hydrogen bonding**, the strongest of all intermolecular forces, the molecules are held together strongly.

Surface tension has various effects.

- **Beading**: Figure 5a demonstrates water beading on a leaf's surface. The outer surface of a leaf consists of a waxy material, made of lipids (fats). Because water is polar and the surface of the leaf is non-polar, the water molecules do not interact with it strongly; rather, they tend to interact with each other. Instead, the water rolls off the leaf in a droplet or bead.
- **Flotation**: Water molecules are strongly attracted to one another because of the strength of intermolecular hydrogen bonding. Objects that have a greater density than water can float on its surface (Figure 5b). This only occurs if the material that the object is made from does not interact with water. The object must also be able to distribute its surface area across the surface of the water.
- **Dripping**: Water forms droplets. For example, water adheres to the surface of a tap. As the mass of the water increases, it overcomes the strength of the hydrogen bonding that is holding the drop to the tap. Therefore, the mass of the water drop under the force of gravity becomes greater than the forces holding the water together and to the tap, so it falls (Figure 6c).

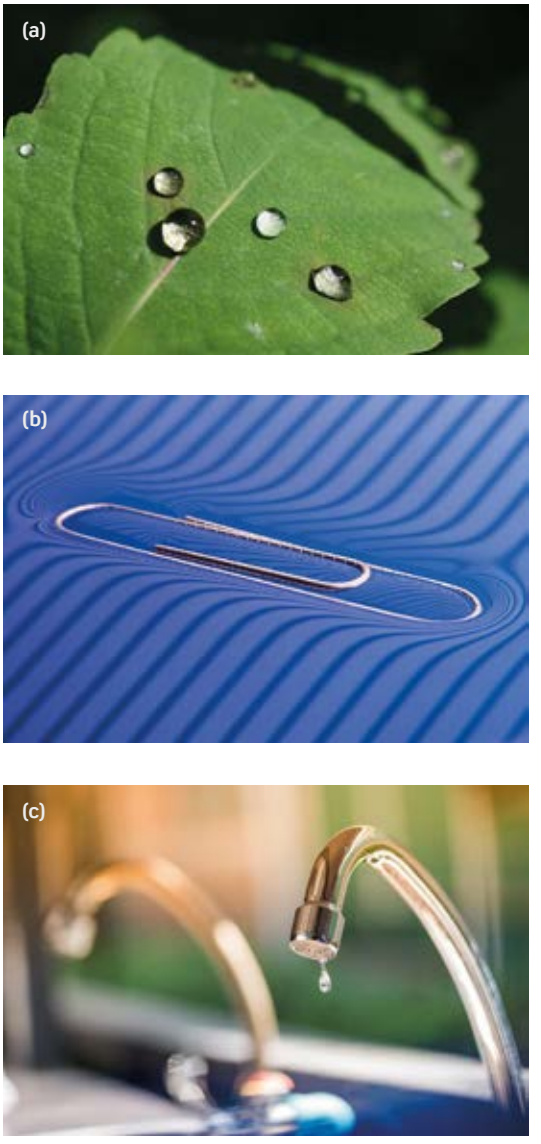

FIGURE 6 The surface tension of water produces various effects. (a) Water on a leaf demonstrates beading. (b) A paperclip on water demonstrates flotation. (c) Droplets falling from a tap demonstrate dripping.

Solutions

solution
a mixture of a solute dissolved in a solvent

solute
the minor component of a solution; is dissolved in the solvent

solvent
the major component of a solution; what the solute is dissolved in

A **solution** is the name given to the mixture resulting when a **solute** dissolves in a **solvent**. In order to prepare a solution, a sample of solute is measured (either by weighing on a balance or measuring a volume), and then added to a volume of solvent. The solution can be stirred or heated to enable the solute to dissolve faster.

- If all the solute dissolves and it is possible to dissolve more, the solution is **unsaturated**.
- If no more solute will dissolve, the solution is **saturated**.
- If a saturated solution is heated, more solute often dissolves. If the solute remains in solution as it cools, this is a **supersaturated** solution and is often used to produce crystals of the solvent.

A solution in which water acts as the solvent is called an **aqueous solution** (from the word *aqua*, meaning 'water'). It is given its own state in chemical equations, represented as (aq).

unsaturated
a solution that can dissolve more solute

saturated
a solution in which no more solute can dissolve

supersaturated
a solution that has been heated to dissolve more solute than can be dissolved under equilibrium conductions; on cooling the solute remains in solution in a metastable state, until a disturbance (seed crystal, speck of dust, tapping the container) starts the crystallisation process

aqueous solution
a mixture of a solute dissolved in water

FIGURE 7 The Cave of the Crystals, at Naica, Mexico has giant crystals grown over thousands of years from saturated solutions percolating underground.

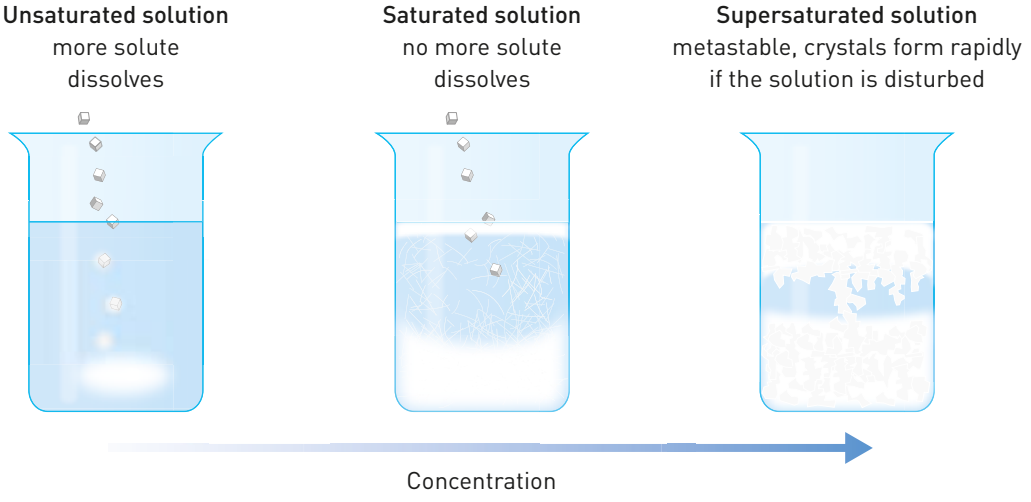

FIGURE 8 The degree of saturation increases as concentration increases.

Study tip

Remember that the properties of molecules depend on the atoms they are made of, the shape of the molecule and its intermolecular forces.

WORKED EXAMPLE 15.1

In seawater, what is the:

a solute?

b solvent?

c solution?

SOLUTION

a The solute (what has been dissolved) is the salt.

b The solvent (what the solute is dissolved in) is the water.

c The solution (combination of solute and solvent) is the seawater.

CHALLENGE 15.1

Gas density

Sulfur hexafluoride (SF_6) is a gas at room temperature. Scientists fill a glass fish tank with this gas. They construct a boat out of aluminium foil and sit the boat on the top of the gas. The boat floats.

Explain, using your knowledge of intermolecular forces, why the boat floats.

CASE STUDY 15.1

Formation of snowflakes in the atmosphere

Snowflakes are incredibly beautiful and unique. No two snowflakes are identical, and the various formations look amazing under a microscope.

All air contains water vapour. This is gaseous water that has evaporated from the land and into the atmosphere. Air can hold a finite amount of water vapour; however, hotter air can hold more water vapour. As air cools, the water vapour condenses into droplets and precipitates as rain.

FIGURE 9 Snowflake

Humidity is a measure of the concentration of water vapour in the atmosphere and depends on temperature. At 100% humidity, air is saturated. However, air humidity can exceed 100%, in which case it is said to be supersaturated. Supersaturation only happens when the temperature of the atmosphere drops and air still contains the water that was vaporised at the higher temperature.

In order to form snowflakes, the atmosphere must be supersaturated with water and super cool (0°C or below). The air must also contain pollen or dust. The crystals begin to form on the pollen or dust particles, which starts the snowflake forming in a process called 'seeding'.

FIGURE 10 Snowflakes during a snowfall, under natural conditions at low temperature.

CHECK YOUR LEARNING 15.1

Describe and explain

1 **Describe** the shape of the water molecule.
2 **Explain** why a water molecule is polar and forms hydrogen bonds.
3 Use your understanding of the intermolecular bonding of water to **explain** why water expands when it freezes.

Apply, analyse and interpret

4 A small pellet (a sphere of radius 1.4 cm) of osmium has the same mass as a beryllium sphere the size of a tennis ball. **Consider** how this is possible.

Investigate, evaluate and communicate

5 A chemist wishes to grow a copper sulfate crystal. Use your knowledge of saturation to research and **develop** a method for this. **Justify** your steps based on your knowledge of saturation.

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- » Challenge
  15.1 Gas density
- » Weblink
  Crystal formation
- » Weblink
  Saturation

15.2 Concentration

KEY IDEAS

In this section, you will learn about:

- concentration as a measure of the amount or mass of substance in a solution
- units used to measure concentration
- calculating concentration
- converting between units of concentration
- dilution calculations.

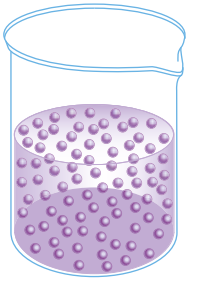
concentration
a measure of the number of particles in a specific volume of solution; units of concentration include mol L^{-1} (M), g L^{-1} and mg L^{-1} (ppm)

concentrated solution
a solution containing many particles per unit volume of solution

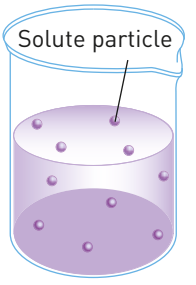
dilute solution
a solution containing few particles per unit of volume

molarity
molar concentration, units are moles per litre, mol L^{-1} or M

Concentration is a measure of the amount of substance in an accurately known volume of solution. A **concentrated solution** contains a higher amount of substance, whereas a **dilute solution** contains a lower amount of substance. However, these terms are used differently depending on the context of the situation. Concentration is measured in many different ways, depending on the nature of the solute and the solvent.



Concentrated solution



Dilute solution

FIGURE 1 A concentrated solution contains a higher amount of solute than a dilute solution.

Units of concentration

There are a number of units used to measure concentration, including molarity (mol L^{-1} or M), grams per litre (g L^{-1}) and parts per million (1 ppm = 1 mg L^{-1}).

Molarity (mol L^{-1} or M)

The most common unit of concentration used by chemists is the number of moles of a substance per litre of solvent. This is called **molarity**, or molar concentration, and is represented in units of moles per litre, mol L^{-1} or M. Either of these units can be used to calculate concentration.

Molar concentration, or molarity, is calculated as:

$$c = \frac{n}{V}$$

or

$$n = c \times V$$

where c is concentration (mol L^{-1} or M), n is number of moles (mol) and V is volume (L). If volume is given in mL, then it must be converted to L first.

The concentration of a substance can also be represented by placing its formula in square brackets. A solution of 0.8 M KI would be represented as [KI] = 0.8 M.

Study tip

Do not confuse M (mol L^{-1}) (concentration) with *M* (molar mass). You can use the symbol M_r for molar mass if this helps to distinguish the terms.

WORKED EXAMPLE 15.2A

A 200 mL solution contains 1.2 mol of calcium chloride ($CaCl_2$). Calculate its concentration.

SOLUTION

1 Convert the volume to litres.

$$200 \text{ mL} \times \frac{1 \text{ L}}{100 \text{ mL}} = 0.200 \text{ L}$$

2 Calculate the unknown value (concentration).

$$c = \frac{n}{V}$$

$$[CaCl_2] = \frac{1.2 \text{ mol}}{0.200 \text{ L}}$$

$6 \text{ mol L}^{-1} = 6 \text{ M}$

However, in a laboratory situation it is more realistic to weigh a mass of a chemical and dissolve it in water, then calculate its concentration. This calculation is a multistep process, where moles are calculated first, then concentration.

WORKED EXAMPLE 15.2B

2.5 g of sodium chloride is completely dissolved in 56 mL of water. Calculate the concentration of the sodium chloride solution.

SOLUTION

1 Convert the volume to litres.

$$56 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.056 \text{ L}$$

2 Calculate the number of moles.

$$n = \frac{m}{M}$$

$$n(\text{NaCl}) = \frac{2.5 \text{ g}}{58.4 \text{ g mol}^{-1}}$$

$$= 4.28 \times 10^{-2} \text{ mol}$$

3 Calculate the unknown value (concentration).

$$c = \frac{n}{V}$$

$$[\text{NaCl}] = \frac{4.28 \times 10^{-2} \text{ mol}}{0.056 \text{ l}}$$

$0.76 \text{ mol}^{-1} = 0.76 \text{ M}$

WORKED EXAMPLE 15.2C

A student is required to make a solution of 0.0500 M sodium hydroxide. They find 0.920 g in the chemical store. What volume of water would the student need to dissolve the solute in?

SOLUTION

1 Calculate the number of moles.

$$n = \frac{m}{M}$$

$$n(\text{NaOH}) = \frac{0.92 \text{ g}}{40 \text{ g mol}^{-1}}$$

$$= 2.3 \times 10^{-2} \text{ mol}$$

2 Calculate the unknown value (volume).

$$V = \frac{n}{C}$$

$$V(\text{NaOH}) = \frac{2.3 \times 10^{-2} \text{ mol}}{0.05 \text{ mol L}^{-1}}$$

$$= 0.460 \text{ L}$$

$$= 460 \text{ mL}$$

Study tip

Multistep calculations in chemistry can get quite messy, so it is essential that you develop good techniques early. Highlight the important information (numbers, units, chemicals or key words), and then add the letters that represent them (e.g. put the letter *n* above mole values or *m* above mass values). Last, put a question mark next to the value you are asked to calculate.

Then look at all of the formulas and chemical equations, and find a pathway that will allow you to use the values given to calculate the answer.

Often, scientists want to determine the concentration of ions in a solution. This is essential when calculating the pH of a solution (covered in Chapter 17 and later in Unit 3) or the concentration of an intravenous saline solution. To determine the concentration of ions, an extra step must be added in which the number of moles of the ion is determined from the number of moles of the solute. This requires a mole ratio statement, which is demonstrated in step 3 of Worked example 15.2D.

WORKED EXAMPLE 15.2D

1.32 g of sodium phosphate (Na_3PO_4) is completely dissolved in 137 mL of water. Calculate the molar concentration of the sodium ions (Na^+) in the solution.

SOLUTION

1 Convert volume to litres

$$37 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.137 \text{ L}$$

2 Calculate the number of moles

$$n = \frac{m}{M}$$

$$n(Na_3PO_4) = \frac{1.32 \text{ g}}{163.94 \text{ g mol}^{-1}}$$

$$= 8.0517 \times 10^{-3} \text{ mol}$$

Note: A greater number of significant figures has been retained than necessary. Remember to only round in the final step of calculations to avoid rounding errors.

3 Determine the number of moles of the ion using a mole ratio. Make sure that you keep the previous answer in the calculator; otherwise, rounding may result in an incorrect answer.

$$n(Na^+) = n(Na_3PO_4) \times 3$$

$$= 8.0517 \times 10^{-3} \times 3 \text{ mol } Na^+ / \text{ mol } Na_3PO_4$$

$$= 2.4155 \times 10^{-2} \text{ mol } Na^+ \text{ ions}$$

As one Na_3PO_4 formula unit contains 3 Na^+ ions, 10 Na_3PO_4 formula units contain 30 Na ions and so on. The pattern is that the number (or mole) of Na_3PO_4 formula units must be multiplied by 3.

4 Determine the concentration of the ion.

$$c = \frac{n}{V}$$

$$[Na] = \frac{2.4155 \times 10^{-2} \text{ mol}}{0.137 \text{ L}}$$

$$= 0.176 \text{ M}$$

Study tip

Always convert volume into litres when completing concentration calculations. Although included in the worked examples, showing working for this step is not required.

Grams per litre (g L^{-1})

Another common unit of concentration is grams of substance per litre. There are several reasons for using g L^{-1} instead of mol L^{-1}. The molar mass of a compound may not be known if it is a new compound and still being investigated or if the substance is impure.

CASE STUDY 15.2A

Nutrition information panels

One of the best-known uses of concentration as mass per unit of volume is on the nutritional panels on the back of food packaging. This nutrition information is often expressed in grams per 100 grams, for solid foods, or grams per 100 mL, for liquid foods (Figure 2). There is not enough space on the label to list all the individual substances in foods so generally fats, proteins and carbohydrates are grouped and the information refers to the overall mass of these within the product.

nutrition information

NUTRITION INFORMATION
SERVINGS PER PACKAGE: About 9
SERVING SIZE: 25g (6 Biscuits)

	QUANTITY PER SERVING	%DAILY INTAKE* (PER SERVING)	QUANTITY PER 100g
ENERGY	493 kJ	5.6%	1,970 kJ
PROTEIN	2.0 g	3.9%	7.9 g
FAT, TOTAL	5.0 g	7.2%	20.0 g
- SATURATED	1.0 g	4.1%	3.9 g
CARBOHYDRATE	15.6 g	5.0%	62.6 g
- SUGARS	0.3 g	0.4%	1.3 g
SODIUM	212 mg	9.2%	848 mg

*BASED ON AN AVERAGE ADULT DIET OF 8700 KJ. ALL VALUES CONSIDERED AVERAGES UNLESS OTHERWISE INDICATED

nutrition information

NUTRITION INFORMATION (Average)
SERVINGS PER PACK: 10 SERVING SIZE: 200 mL

	QUANTITY PER SERVING	% DAILY INTAKE* (PER SERVING)	QUANTITY PER 100 G
ENERGY	306 kJ	4%	153 kJ
PROTEIN	1.2 g	2%	0.6 g
FAT, TOTAL	<1 g	<1%	<1 g
- SATURATED	0 g	0%	0 g
CARBOHYDRATE	18.0 g	6%	9.0 g
- SUGARS	16.6 g	18%	8.3 g
DIETARY FIBRE	0.4 g	1%	0.2 g
SODIUM	16 mg	0.7%	8 mg
VITAMIN C	40 mg	100% RDI#	20 mg

*BASED ON AN AVERAGE ADULT DIET OF 8700 KJ. ALL VALUES CONSIDERED AVERAGES UNLESS OTHERWISE INDICATED

FIGURE 2 The nutritional information of a solid food product (left) and a liquid food product (right).

Concentration is calculated as:

$$c = \frac{m}{V} \text{ or } m = c \times V$$

where c is concentration (g L^{-1}), m is mass (g) and V is volume (L).

This is the same calculation as previously seen, but moles is replaced with mass (in grams).

WORKED EXAMPLE 15.2E

Calculate the concentration of a solution of 0.042 g protein dissolved in 241 mL of water.

SOLUTION

1 Convert mL to L.

$$241 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.241 \text{ L}$$

2 Calculate the unknown (concentration).

$$c = \frac{m}{V}$$

$$c(\text{protein}) = \frac{0.042 \text{ g}}{0.241 \text{ L}}$$

$$= 0.17 \text{ g L}^{-1}$$

WORKED EXAMPLE 15.2F

A chemist wishes to make 52 mL of a solution of sugar of concentration 0.94 g L^{-1}. What mass of sugar is required?

SOLUTION

1 Convert mL to L.

$$52 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.052 \text{ L}$$

2 Calculate the unknown (mass).

$$m = c \times V$$

$$m(\text{sugar}) = 0.94 \text{ mol L}^{-1} \times 0.052 \text{ L}$$

$$= 0.049 \text{ g}$$

$$= 0.049 \text{ g} \times \frac{100 \text{ mg}}{1 \text{g}}$$

$$= 49 \text{ mg}$$

Parts per million (ppm or mg L^{-1})

The unit parts per million (ppm) is used to measure very low concentrations.

To fully understand ppm, consider the density of water, which is 1.00 g mL^{-1} at room. This means that a millilitre of water has a mass of 1.00 g. As there are 1000 mL of water in a litre, 1 L of water has a mass of 1.00 kg.

One-millionth of 1 kg is 1 mg. Therefore, one-millionth of a litre is also 1 mg. So in water, 1 ppm is equivalent to 1 mg per litre of water (mg L^{-1}).

Concentration is calculated as:

$$c = \frac{m}{V} \text{ or } m = c \times V$$

where c is concentration (mg L^{-1} or ppm), m is mass (mg) and V is volume (L).

WORKED EXAMPLE 15.2G

125 mL of a sodium chloride solution was made from 0.0024 g of the salt. Calculate its concentration in ppm.

SOLUTION

1 Convert g to mg.

$$0.0024 \text{ g} \times \frac{1000 \text{ mg}}{1 \text{ g}} = 2.4 \text{ mg}$$

2 Determine the unknown (concentration).

$$c = \frac{m}{V}$$

$$c(\text{NaCl}) = \frac{2.4 \text{ mg}}{0.125 \text{ L}}$$

$$= 19.2 \text{ ppm}$$

WORKED EXAMPLE 15.2H

What mass of potassium hydroxide (KOH) is present in 250 L of a 3.26 ppm solution?

SOLUTION

$$m = c \times V$$

$$m(\text{KOH}) = 3.26 \text{ mg L}^{-1} \times 250 \text{ L}$$

$$= 815 \text{ mg}$$

$$= 0.815 \text{ g}$$

Converting between concentration units

When converting between units of concentration, consider the numerator and denominator of the equations. All of the concentration equations discussed so far express volume in litres (L). Therefore, when converting between units, volume can be disregarded.

Figure 3 shows a shortcut to convert between units. To calculate mass from a mole amount, multiply the mole amount by molar mass. This is demonstrated through the mole equation. Therefore, to convert from mol L^{-1} to g L^{-1}, multiply the concentration in mol L^{-1} by molar mass. To convert g L^{-1} to mol L^{-1}, divide the concentration in g L^{-1} by molar mass.

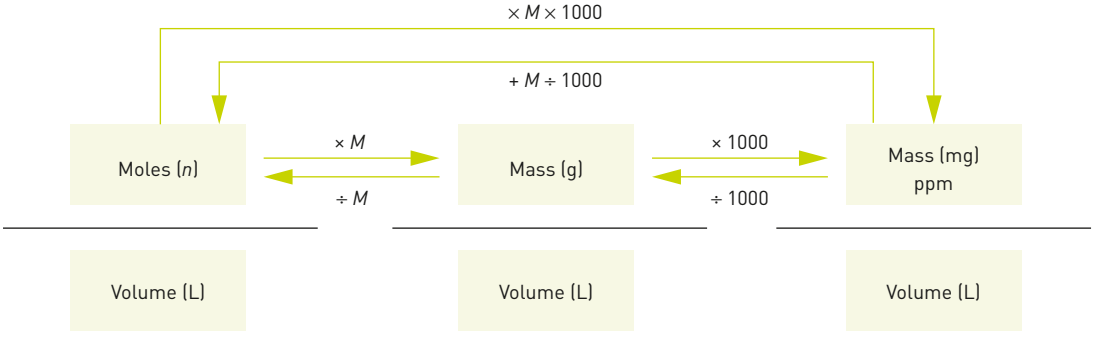


FIGURE 3 Conversion tool for concentration units mol L^{-1}, g L^{-1} and ppm, where M is molar mass.

To convert from g L^{-1} to mg L^{-1} (ppm), multiply g L^{-1} by 1000. To convert from mg L^{-1} (ppm) to g L^{-1}, divide by 1000.

WORKED EXAMPLE 15.2I

Convert the following concentration values.

a 2.6 ppm LiOH to M

b 5.7 g L^{-1} to ppm

c 0.76 g L^{-1} of HCl to M

d 2 M HNO_3 to ppm

SOLUTION

a 2.6 ppm ÷ 1000 = 0.0026 g L^{-1}

$$\frac{0.0026 \text{ g L}^{-1}}{23.9 \text{ g mol}^{-1}} = 1.1 \times 10^{-4} \text{ mol L}^{-1}$$

$$= 1.1 \times 10^{-4} \text{ M}$$

b 5.7 g L^{-1} × 1000 mg g^{-1} = 5700 mg L^{-1}

$$= 5.7 \times 10^{3} \text{ ppm}$$

c 0.76 g L^{-1} ÷ 36.5 g mol^{-1} = 0.021 M

d 2 × 63 g L^{-1} = 126 g L^{-1}

126 g L^{-1} × 1000 mg g^{-1} = 126 000 ppm

$$= 1.26 \times 10^{5} \text{ ppm}$$

Sometimes it is easier to calculate a concentration in a different unit and convert between units to determine an answer. This is seen in Worked example 15.2J. Either method can be used and both result in a correct answer.

WORKED EXAMPLE 15.2J

5.2 mg of NaCl is dissolved in 3.24 L of water. What is its concentration in mol L^{-1}?

SOLUTION

Using conversion

1 Calculate the concentration in ppm.

$$c = \frac{m}{V}$$

$$c(\text{NaCl}) = \frac{5.2 \text{ mg}}{3.24 \text{ L}}$$

$$= 1.6 \text{ ppm}$$

2 Convert to mol L^{-1}.

$$[\text{NaCl}] = \frac{1.6 \text{ ppm}}{1000 \text{ mg g}^{-1}}$$

$$= 1.6 \times 10^{-3} \text{ g L}^{-1}$$

$$\frac{1.6 \times 10^{-3} \text{ g L}^{-1}}{58.4 \text{ g mol}^{-1}} = 2.7 \times 10^{-5} \text{ mol L}^{-1}$$

$$= 2.7 \times 10^{-5} \text{ M}$$

Using mole calculations

1 Calculate the number of moles.

$$n = \frac{m}{M}$$

$$n(\text{NaCl}) = \frac{0.0052 \text{ g}}{58.4 \text{ g mol}^{-1}}$$

$$= 8.89 \times 10^{-5} \text{ mol}$$

2 Calculate the unknown value (concentration).

$$c = \frac{n}{V}$$

$$c(\text{NaCl}) = \frac{8.89 \times 10^{-5} \text{ mol}}{3.24 \text{ L}}$$

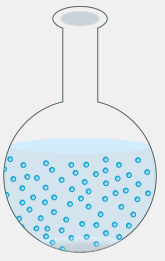
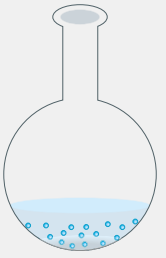
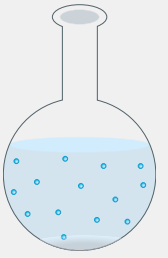
$$= 2.7 \times 10^{-5} \text{ M}$$

Dilution

dilution
a method for lowering the concentration of a solution

Often, solutions must be diluted before a chemist can analyse them. To perform a **dilution**, a small volume is transferred to a new flask and water or another solvent is added (Table 1).

TABLE 1 Changes in concentration and number of particles during the process of dilution

	 Original solution	Sample of original solution	 New solution: sample	 New solution: sample made up to 300 mL with water
Number of particles	100	15	15	15
Volume	250 mL	37.5 mL	37.5 mL	300 mL

In Table 1, the original solution contains 100 molecules in 250 mL of water. From this solution, a 37.5 mL sample, containing 15 molecules, is measured into a measuring cylinder. The sample is then poured into a new beaker, where it still contains 15 molecules in 37.5 mL. The sample is then made up to 300 mL with water to dilute it. The new solution now contains 15 molecules in 300 mL of solution and is therefore more dilute.

Assuming that the number of particles is the equivalent to the number of moles, several assumptions can be made:

- The concentration of the original solution is the same as the concentration of the sample taken from it.
- The sample and the new diluted solution contain the same number of moles; they both contain 15 molecules.

Recall that to calculate the number of moles from a solution, the following formula is used:

$$n = c \times V$$

As the sample and the diluted solution have the same number of moles, it can be deduced that the concentration multiplied by the volume of the sample will result in the same value as the concentration multiplied by the volume of the dilute solution. Therefore, the dilution equation can be used:

$$c_1 V_1 = c_2 V_2$$

c_1 and V_1 are the concentration and volume of the sample. c_2 and V_2 are the concentration and volume of the diluted solution.

WORKED EXAMPLE 15.2K

A 5 mL sample of a 50 mL solution of 0.2 M HCl is diluted to 250 mL. What is the concentration of the new solution?

SOLUTION

$$c_1 V_1 = c_2 V_2$$

$$0.2 \text{ mol}^{-1} \times 0.005 \text{ L} = c_2 \times 0.250 \text{ L}$$

$$c_2 = \frac{0.2 \text{ mol L}^{-1} \times 0.005 \text{ L}}{0.250 \text{ L}}$$

$$= 0.004 \text{ mol L}^{-1}$$

$$= 0.004 \text{ M}$$

Note: There is no need to convert volumes to litres as long as units remain constant for all volumes.

CHALLENGE 15.2

Lead paint

Lead paint has been banned since 1978 because lead is a toxic substance. However, some older homes still contain lead-based paint on walls or woodwork. A chemist performed an analysis to determine whether the concentration of lead in a sample of wall paint is safe. The recommended maximum concentration of Pb^{2+} ions is 0.05 M.

A 20mL sample of paint was reacted with dilute nitric acid to convert the insoluble $PbCO_3$ in the paint into soluble $Pb(NO_3)_2$ and this was then diluted with water to a total volume of 500 mL. The diluted lead(II) nitrate solution was then reacted with potassium iodide to form a precipitate of lead(II) iodide. The mass of the precipitate was 0.56 g.

1 What is the concentration of lead(II) ions in the 500 mL of dilute solution?
2 Is the lead(II) in the paint at a safe concentration?

CASE STUDY 15.2B

Atmospheric chemistry and climate change

Earth's atmosphere is a complex, interacting system. An understanding of the concentrations of the several gases that make up the atmosphere, including synthetic chemicals such as chlorofluorocarbons (CFCs), is essential to understanding atmospheric chemistry and phenomena such as climate change and pollution. Sensitive instruments are used to measure the concentration of gas in air parts per million (ppm); and parts per billion (ppb).

Data has been recorded from before 1970 to 2016. A summary of this data from the United States EPA (Environment Protection Authority) website states that, in the atmosphere, the global average concentration of:

- carbon dioxide (CO_2) has increased from 280 to 399.5 ppm
- methane (CH_4) has increased from 722 to 1834 ppb
- ozone (O_3) has increased from 237 to 337 ppb
- nitrous oxide (N_2O) has increased from 270 to 328 ppb.

Concentrations of greenhouse gases have increased significantly and are now higher than ever recorded. For example, the concentration of CO_2 is approximately 100 ppm higher than at any time in the last 800 000 years. This data has been collected from the drilling of ice cores in the Arctic and Antarctic to determine atmospheric conditions when the ice was deposited hundreds of thousands of years ago.

Water vapour (H_2O) makes up 95% of all greenhouse gases. It has a concentration of approximately 4% in the atmosphere. As temperatures increase, as a result of human activity, the amount of water vapour in the atmosphere also increases. Water vapour amplifies the warming effect of CO_2, with recent studies suggesting that it as much as doubles the effects of CO_2. Essentially, the presence of extra water vapour in the atmosphere makes temperature changes caused by CO_2 more substantial.

CHECK YOUR LEARNING 15.2

Describe and explain

1 **Describe** why the concentration units g L^{-1} and ppm might be used instead of mol L^{-1}?

2 **Convert** between the following units.
- **a** 0.056 M of $K_2Cr_2O_7$ to ppm
- **b** 25 ppm KOH to g L^{-1}
- **c** 82 g L^{-1} of H_3PO_4 to M

3 **Calculate** the concentration of the following solutions in mol L^{-1}.
- **a** 0.30 mol of NaCl in 50.0 mL of water
- **b** Iodide ions when 0.9 g of KI is dissolved in 20 mL of water
- **c** 4.93×10^{21} HCl molecules in 125 mL of water

4 **Calculate** the concentration of the following solutions in ppm.
- **a** 0.0032 mol of HNO_3 in 345 mL of water
- **b** 0.093 g of NaOH in 100 mL of water
- **c** 4.32×10^{-15} molecules of H_2SO_4 in 150 mL of water

Apply, analyse and interpret

5 A student analyses a solution by chromatography and generates the calibration curve in Figure 4. The blue dots represent a set of standards that have been used to determine the concentration of the sample (blue dot).
- **a** The graph is measuring 40 ppm for the sample. **Critique** the validity of the data. What two things could be done to make the data more valid?
- **b** **Demonstrate**, showing calculations, how to dilute the sample so that it is within the range of the standards.

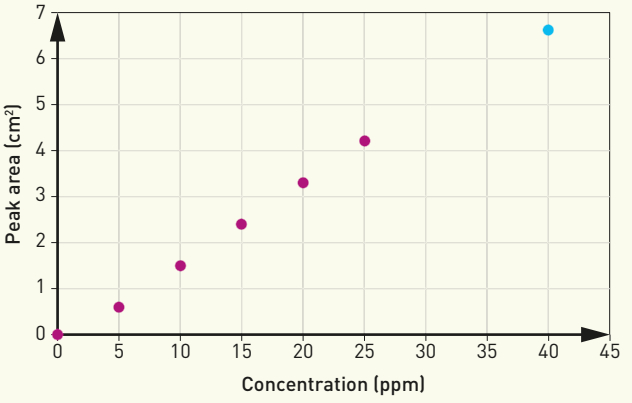


FIGURE 4 A calibration curve for a set of standards (purple) and a sample (blue).

Investigate, evaluate and communicate

6 A chemist has accidentally spilt 10 g of sodium acetate (CH_3COONa) into 125 mL of water. **Propose** how they could make a 10 ppm solution from this solution.

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15.3 Strength versus concentration of acids and bases

KEY IDEAS

In this section, you will learn about:

- the nature of acids and bases
- the difference between strong and concentrated acids and bases.

A very common misconception in chemistry is that the concentration of a solution in some way correlates to its strength. Specifically, strength refers to acid and base chemistry. Acids and bases will be covered in more detail in Chapters 17 and 18. However, in the context of aqueous solutions and concentration, it is important to understand the distinction between strength and concentration of an acidic or basic solution.

Acid and base strength

Acids and bases react differently depending on their strength.

Strong acids and bases

The strength of an acid refers to its ability to dissociate, or break down into its cation (positively charged ion) and anion (negatively charged ion), in water. This process is called ionisation or dissocation. Hydrochloric acid (HCl) is a **strong acid**. In water, HCl dissociates according to the chemical equation:

strong acid
an acid that completely dissociates in water

$$HCl(g) + H_2O(l) \rightarrow Cl^-(aq) + H_3O^+(aq)$$

As HCl dissociates completely into ions in water, no HCl remains. This is what makes HCl a strong acid.

The same principle applies for bases. Sodium hydroxide (NaOH) is a **strong base** because it dissociates completely in water according to the chemical equation:

strong base
a base that completely dissociates in water

$$NaOH(g) + H_2O(l) \rightarrow Na^+(aq) + OH^-(aq) + H_2O(l)$$

NaOH is a strong base because it dissociates completely into Na+ and OH- ions when dissolved.

Weak acids and bases

Ethanoic acid (CH_3COOH), commonly known as acetic acid, is a **weak acid**. It is safe for human consumption as a dilute solution and is the acid in vinegar. It partially dissociates in water according to the chemical equilibrium equation:

weak acid
an acid that only partially dissociates in water

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

The main difference between the chemical equations used to represent complete and partial dissociation is the arrow used. In partial dissociation, there are both forward and back reactions occurring simultaneously. So, CH_3COOH begins to dissociate, driving the reaction forward and producing CH_3COO^- and H_3O^+ ions. In the reverse or back reaction, CH_3COO^- and H_3O^+ react together to form CH_3COOH. This is represented using the equilibrium arrow $\rightleftharpoons$.

As the concentration of H_3O^+ ions formed in this reaction is very small, CH_3COOH is a weak acid. In other words, the reaction goes in the forward direction to only a very small extent.

weak base
a base that only partially dissociates in water

Ammonia (NH_3) is a **weak base**. It partially ionises in water according to the chemical equation:

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

As the concentration of OH^- ions formed is very low, the basicity of the solution is much less than if an equivalent amount of the strong base sodium hydroxide has been added. Therefore, ammonia is considered a weak base.

TABLE 1 Some common weak and strong acid and bases

Strong acids	Weak acids	Strong bases	Weak bases
Hydrochloric acid (HCl)	Ethanoic acid (CH_3COOH)	Sodium hydroxide (NaOH)	Ammonia (NH_3)
Nitric acid (HNO_3)	Carbonic acid (H_2CO_3)	Potassium hydroxide (KOH)	Methylamine (CH_3NH_2)
Sulfuric acid (H_2SO_4)	Hydrofluoric acid (HF)	Lithium hydroxide (LiOH)	Sodium carbonate (Na_2CO_3)

Strength vs concentration

The confusion between 'strong' and 'concentrated' can be easily understood. When a drink, such as cordial, is concentrated, it has a stronger taste to it and this may be the basis for the misconception. In chemistry, 'strength' and 'concentration' are very separate concepts with different meanings and cannot be treated in the same way.

A strong acid can be concentrated. This means that there are many acid molecules present and that all of these molecules are ionised in water (Figure 1).

A weak acid can also be concentrated. This means that there are many acid molecules present but only a few of these molecules are ionised in water.

A strong acid can be dilute. This means that there are only a few acid molecules present and that all of these molecules are ionised in water.

A weak acid can also be dilute. This means that there are only a few acid molecules present and only a few of these are ionised in water.

This can also be applied to bases.

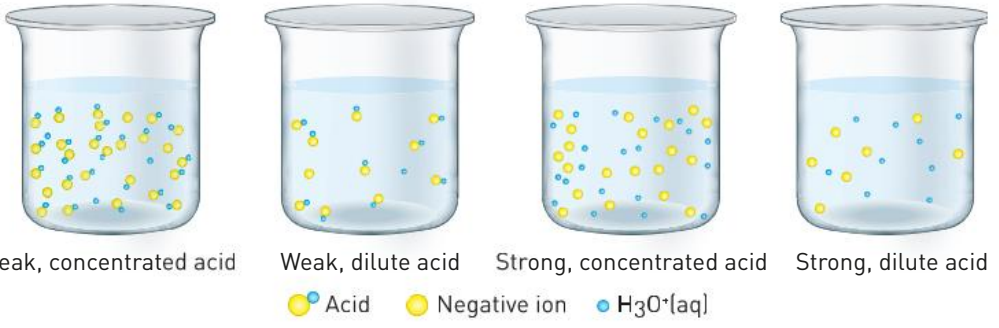

FIGURE 1 Acids can be strong or weak, and concentrated or dilute.

Methods for determining strength and concentration

A simple method for determining the strength of an acid is to measure its conductivity. A strong acid completely dissociates, so it contains a large number of charged ions. More charged ions increase the conductivity of the solution. When connected to an electric circuit, this results in a higher voltage and a brighter light bulb due to the higher number of changed ions (Figure 2). A lower conductivity implies that there are very few ions in solution, indicating a weak acid.

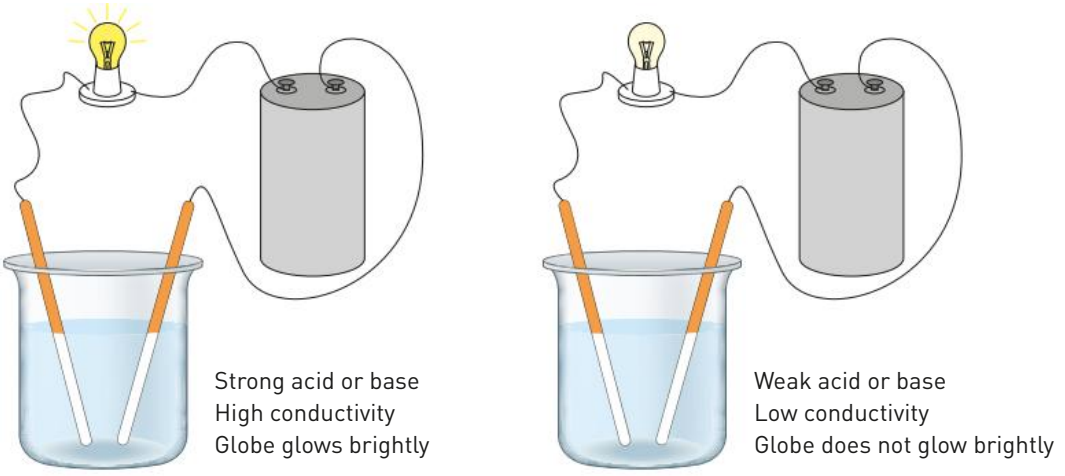

FIGURE 2 The strength of an acid or base can be determined by measuring its electrical conductivity.

The strength of an acid or base can also be determined by measuring its pH with universal indicator (see Chapter 17). Universal indicator changes colour in a solution depending on whether an acid or base is present and how strong it is (Figure 3). A strong acid has a pH of 0–2 and a weak acid has a pH of 3–6. Neutral solutions, such as water, have a pH of 7. Strong bases have a pH of 13–14 and weak bases have a pH of 8–12.

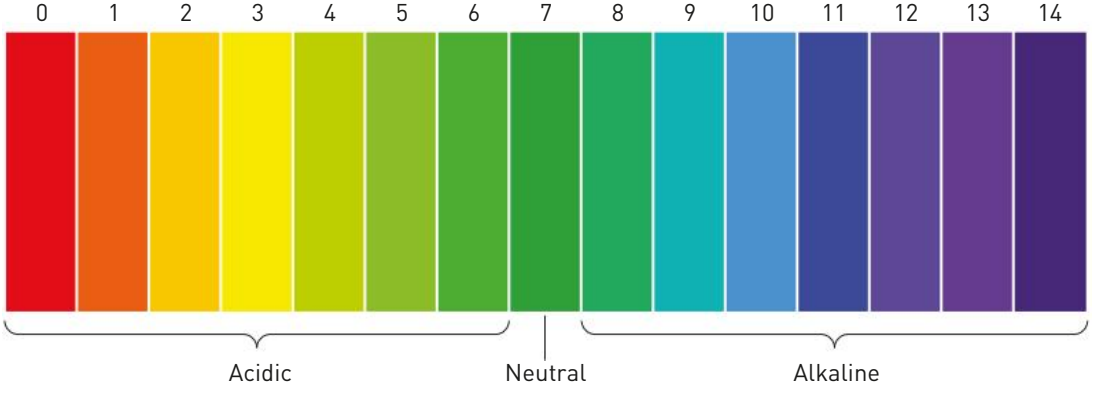

FIGURE 3 pH is indicated by the colours of universal indicator in acidic and basic solutions. The strongest acids are on the left at a pH of 0, neutral substances are neither acidic nor basic, and the strongest basic (alkaline) substances are on the right at a pH of 14.

Litmus paper can be used to determine whether a solution is either acidic or basic. Red litmus paper turns blue in basic solution. Similarly, blue litmus paper turns red in acidic solution (Figure 4). Because litmus paper does not differentiate the strength of an acid, it is not often used in higher levels of chemistry.

The most accurate way to measure pH is to use a pH probe (Figure 5). This is an electronic device that, when calibrated, can determine the pH of a solution to decimal place readings.

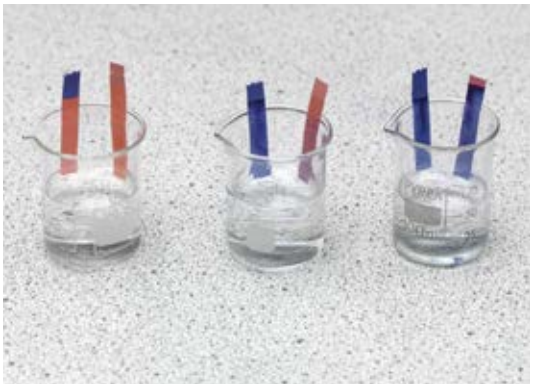

FIGURE 4 Litmus paper is used to determine whether a solution is acidic or basic.

FIGURE 5 A pH probe can measure the strength of a solution very accurately.

CHALLENGE 15.3

Ionisation of triprotic acid

Phosphoric acid (H_3PO_4) is triprotic because it contains three hydrogen ions (H^+). This means that it undergoes ionisation three times to lose each of these ions to eventually form the phosphate ion (PO_4^{3-}).

Write the three balanced chemical equilibirum equations for the ionisation of phosphoric acid in water.

CASE STUDY 15.3

Acids and bases in digestion

Food is broken down and absorbed into the bloodstream throughout the digestive tract. The stomach is an essential organ in the digestive system because it is the only place where food is broken down in an acidic environment.

The stomach contains hydrochloric acid (HCl), a strong acid, in a concentration of approximately 0.1 M. The stomach is lined with acid-resistant mucus, which stops the acid from eating through the lining of the stomach. This acidic environment is essential for the digestion of proteins and absorption of nutrients. It also ensures that any bacteria and microorganisms are killed before they can enter the small intestine, where they can cause potential harm.

However, the acid in the stomach cannot leave it. The oesophagus, which connects the mouth to the stomach, and the small intestine, in which the partly digested food passes from the stomach, cannot withstand acidic environments. As the food leaves the stomach, it is neutralised by bile, which is produced in the liver and stored in the gall bladder. This makes the digested food neutral and able to move through the small intestine without damaging it.

If the acid-rich contents of the stomach are regurgitated, they burn the throat and mouth, causing pain and irritation. This is the main cause of heartburn, which is treated by ingesting antacids – chemicals that help to neutralise the acid. Sodium bicarbonate ($NaHCO_3$) is an excellent alternative.

As a weak base, when $NaHCO_3$ reacts with HCl, it only partially ionises to form hydrogen carbonate (H_2CO_3) according to the chemical equation:

$$HCl(aq) + NaHCO_3(aq) \rightleftharpoons NaCl(aq) + H_2CO_3(aq)$$

This means that not all of the HCl reacts, which is very important. If a strong base was used, all of the HCl would react with the base, causing the stomach contents to become neutral. In a neutral environment, proteins are not digested, so nutrients are not absorbed into the blood and bacteria and microorganisms will not be killed.

Once formed, H_2CO_3 is highly unstable and decomposes into water and carbon dioxide according to the chemical equation:

$$H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$$

The decomposition of H_2CO_3 is one reason for the formation of carbon dioxide gas in the stomach and subsequent burping.

FIGURE 6 Antiacids neutralise stomach acid.

CHECK YOUR LEARNING 15.3

Describe and explain

1 **Define** 'ionisation'.
2 **Identify** what all strong acids have in common.
3 **Identify** what all strong bases have in common.

Apply, analyse and interpret

4 **Sketch** a diagram to illustrate the difference between the strong acid HCl and the weak acid HNO_2.
5 **Compare** a concentrated acid and a dilute acid.
6 When acids and bases react together, it is called a neutralisation reaction. **Consider** why this term is used. Use a balanced chemical equation to support your answer.

Investigate, evaluate and communicate

7 **Predict** if more, less or the same amount of base would be required to neutralise 20 mL of 0.1 M strong acid than would be required to neutralise 20 mL of 0.1 M weak acid. **Justify** your reponse.
8 **Investigate** why people who suffer from bulimia also experience tooth decay and mouth ulcers.

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CHAPTER 15

Review

Chapter summary

15.1
- Water has a bent shape, making it a polar molecule with a dipole or a dipole moment. This allows it to form hydrogen bonds with other water molecules. Water's boiling point (100°C at atmospheric pressure), density (1.00 g mL^{-1}) and surface tension can be explained by its structure and intermolecular bonding.
- Solutions are formed when a solute is dissolved in a solvent. If more solute can dissolve, the solution is unsaturated. If no more can dissolve, it is saturated. If a saturated solution is heated to dissolve more solute and then cooled, it can be saturated.

15.2
- Concentration is a measure of the number of particles in a specific volume of solution. This is measured in moles per litre (mol L^{-1} or M), grams per litre (g L^{-1}) or ppm (equivalent to mg L^{-1}).
- The equation to calculate concentration in mol L^{-1} is $c = \frac{n}{V}$, where volume is expressed in litres.
- The equation to calculate concentration in g L^{-1} is $c = \frac{m}{V}$, where mass is expressed in grams and volume is expressed in litres.
- The equation to calculate concentration in ppm is $c = \frac{m}{V}$ where mass is expressed in milligrams and volume is expressed in litres.
- Solutions can be diluted, and the concentration of the dilute solution can be determined from the equation $C_1 V_1 = C_2 V_2$.

15.3
- The strength of an acid or a base is determined by its ability to ionise in water. Acids and bases that ionise readily are strong, whereas those that undergo partial ionisation are weak. This contrasts with their concentration, which is a measure of the number of molecules in a volume of solvent. Concentrated solutions contain more molecules in a given volume than dilute ones.

Key terms

- aqueous solution
- boiling point
- concentrated solution
- concentration
- density
- dilute solution
- dilution
- dipole
- electronegativity
- hydrogen bonding
- molarity
- polarity
- saturated
- solute
- solution
- solvent
- strong acid
- strong base
- supersaturated
- surface tension
- unsaturated
- weak acid
- weak base

Key formulas

Molar concentration	$c = \frac{n}{V}$
Concentration	$c = \frac{m}{V}$
Dilution equation	$c_1 V_1 = c_2 V_2$
Density	$\rho = \frac{m}{V}$

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; *** = high.

Multiple choice

1 The substance in which a solute dissolves to form a solution, is called a:
- **A** solvent
- **B** saturate
- **C** solute
- **D** substance.

2 A solution in which no more solute can dissolve is referred to as:
- **A** a solution
- **B** supersaturated
- **C** saturated
- **D** unsaturated.

3 Which of the following is not a unit of concentration?
- **A** M
- **B** mol L^{-1}
- **C** ppm
- **D** kJ L^{-1}

4 What volume of water must be added to 40 mL of 2 M HCl to make a 0.4 M solution?
- **A** 200 mL
- **B** 2 L
- **C** 160 mL
- **D** 240 mL

5 0.056 g of LiOH is dissolved to make 9.6 L of solution. Which answer is not a correct concentration of this solution?
- **A** 2.4×10^{-4} M
- **B** 5.8×10^{-3} g L^{-1}
- **C** 5.8 ppm
- **D** 5.8×10^{-3} M

6 Which is the correct method of converting mL to L?
- **A** × 1000
- **B** ÷ 1000
- **C** × 100
- **D** ÷ 100

7 A solution that forms crystals is referred to as:
- **A** supersaturated
- **B** saturated
- **C** unsaturated
- **D** none of the above.

FIGURE 1 Crystals forming in a solution.

8 Which is the correct formula for the dilution equation?
- **A** $cV_1 = V_2$
- **B** $c_1V_2 = c_2V_1$
- **C** $c_1c_2 = V_1V_2$
- **D** $c_1V_1 = c_2V_2$

9 A strong base has a pH of:
- **A** 1–2
- **B** 4–5
- **C** 7
- **D** 13–14

10 Which of the following is a weak acid?
- **A** diluted HCl
- **B** ammonia
- **C** carbonic acid
- **D** potassium hydroxide

Short answer

Describe and explain

★ **11** **Explain** why density is often referred to as a unit of concentration.

★ **12** Water running from a tap curves and runs sideways when a charged rod is placed near it. **Explain**, using the properties of water, why this phenomenon occurs.

★**13** Water is referred to as a 'universal solvent'. **Explain** why.

FIGURE 2 Water is a universal solvent

Apply, analyse and interpret

★★★**14** **Apply** your knowledge to complete the following molarity calculations

- **a** What volume of a solution is required to make a 0.2 M KNO_3 solution from 0.5 g of KNO_3?
- **b** What is the molarity of a solution made from 0.2 g of KOH in 45 mL of water?
- **c** What mass of NaCl must be added to 450 mL of water to generate a 0.5 M solution?

FIGURE 3 Salt added to water

★★★**15** **a** **Determine** the concentration of a solution prepared from 4.22×10^{-21} formula units of $CaCl_2$ in 50 mL of water. Express your answer in g L^{-1}.

- **b** **Determine** the concentration of dichromate ions in a solution containing 2.6 mg $K_2Cr_2O_7$ in 105 mL of water. Express your answer in g L^{-1}.
- **c** **Determine** the mass of potassium permanganate, in grams, in a 246 mL solution of 0.6 g L^{-1} $KMnO_4$.

★★★**16** **a** 0.045 g of $FeCl_3$ is used to make a 2.70 L solution. **Determine** the concentration in ppm.

- **b** **Determine** what mass of NaOH is required to make a 2.7 ppm solution in 80 mL of water.
- **c** **Determine** what volume of water should be added to 0.01 g of $NaNO_3$ to make a 5 ppm solution.

★★★**17** **a** **Determine** how many millilitres, of a 2.2 M $CaCO_3$ solution are needed to prepare 120 mL of a 0.25 M solution.

- **b** 50.0 mL of a $BaCl_2$ solution is diluted to 425.0 mL has a new concentration of 2.0 M. **Determine** its original concentration.
- **c** 25 mL of a 1.0 M solution is used to make a diluted 0.25 M solution. **Determine** what volume of water must be added.

Investigate, evaluate and communicate

★★**18** A scientist performs an experiment and determines that a saturated sodium chloride solution had a boiling point of 108.7°C, whereas water has a boiling point of 100°C. **Evaluate** the scientist's findings and explain why the salt water has a higher boiling point.

FIGURE 4 Measuring the boiling point of water

★★**19** **Predict** and **justify** whether pure water or salt water has a higher surface tension.

★★**20** A student finds two solutions on a laboratory bench. The solutions are marked A and B. The student knows, from earlier in the lesson, that the solutions are salt water or sugar water, but does not remember which is which. To **identify** them, they pour a small amount of A into a beaker and then slowly pour in the same volume of solution B. The student notices that the solutions form two layers, where solution A sits on top of solution B. The salt water was made by mixing 15.25 g of NaCl in 50 mL of water. The sugar solution was made by dissolving 50 g of sugar in 62.5 mL of water. **Identify** solutions A and B. Use calculations to **justify** your answer.

FIGURE 5 Two unknown clear solutions

★★★**21** **Investigate** why birds are one of the animals most affected by an oil spill. Use your knowledge from this chapter to communicate your findings.

FIGURE 6 Pelicans affected by oil spills

Check your obook assess for these additional resources and more:

- **Student book questions**
  Chapter 15 Review
- **Revision notes**
  Chapter 15
- **assess quiz**
  Auto-correcting multiple choice quiz
- **Flashcard glossary**
  Chapter 15

CHAPTER 16

Solubility and identifying ions in solution

The identification of ions in a solution is an essential part of analytical, medical, food and environmental chemistry.

Environmental scientists need to determine the identity and concentration of various pollutants in order to avoid environmental and health-related disasters. Similarly, food chemists must be able to determine the identity and quantities of nutrients in a sample.

When donating blood, the blood is separated and analysed to determine the blood type as well as the various levels of nutrients and biochemicals present in the blood.

This chapter focuses on the application of the concept of solubility to identify various ions in unknown solutions and relates this knowledge to real-world contexts.

FIGURE 1 Silver nitrate reacts with potassium chromate to produce silver chromate, which precipitates out of solution as a red solid.

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OBJECTIVES

- → Apply solubility rules to determine the products of reactions and to predict if a precipitate will form.
- → Determine the presence of specific ions in solutions based on evidence derived from chemical reactions, including precipitation and acid–carbonate reactions.
- → Construct and use appropriate representations, including ionic formulas, chemical formulas, chemical equations and phase descriptions for chemical species, to communicate conceptual understanding, solve problems and make predictions.
- → Explain the relationship between the solubility of substances (ionic and molecular) in water and the intermolecular forces between species in the substances and water molecules.
- → Recognise that changes in temperature can affect solubility and recall that most gases become less soluble as solvent temperature increases, while most solutes become more soluble as the solvent temperature increases.
- → Interpret, analyse and evaluate data and solubility curves to communicate conceptual understanding, solve problems and make predictions.

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PRACTICALS

MANDATORY PRACTICAL	**16.1** Precipitation reactions Can be found on your obook assess
SUGGESTED PRACTICAL	**16.2** Investigating the effect of temperature on solubility

16.1 Identifying ions in solution

KEY IDEAS

In this section, you will learn about:

- precipitation reactions and the use of solubility rules
- identifying ions in a solution.

ion
an atom or molecule that has gained or lost one or more electrons and is charged

ionic substance
a substance containing a cation and an anion

cation
a metal atom that has lost one or more electrons and is positively charged

anion
a non-metal atom that has gained one or more electrons and is negatively charged

ionic bond/ionic interaction
the electrostatic interaction between a positive cation and negative anion

precipitation reaction
a reaction between two soluble ionic substances that forms a solid product

precipitate
the solid product formed as a result of a precipitation reaction

aqueous
the state given to a substance that is dissolved in water

solution
a mixture of a solute dissolved in a solvent

An **ion** is an atom or molecule that has gained (one or more) electrons to become negatively charged or lost (one or more) electrons to become positively charged (see Figure 1).

Ionic substances are substances consisting of ions – **cations** and **anions** – held together by electrostatic forces. These opposite charges are the reason why **ionic bonds** (or **ionic interactions**) are so strong. An ionic bond is an electrostatic attraction between a cation and an anion.

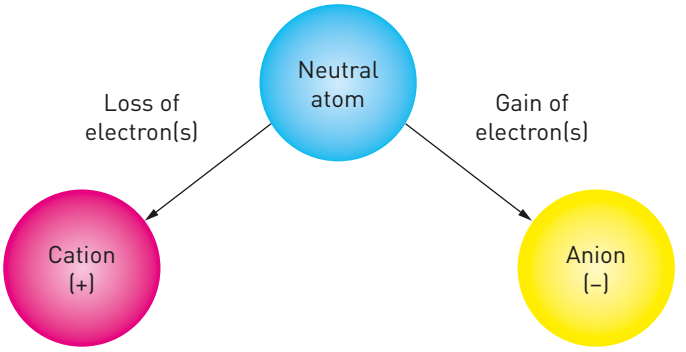

FIGURE 1 A neutral atom loses one or more electrons to become a cation or gains one or more electrons to become an anion.

Precipitation reactions

In a **precipitation reaction**, a **precipitate** (solid) is formed when two **aqueous** ionic **solutions** are mixed together. Because an ionic substance consists of cations and ions, the cation and anion swap partners during a reaction to become different substances.

In the reaction between lead(II) nitrate and sodium chloride, lead(II) and sodium ions are cations. Cations are always written first in the name of an ionic substance. Nitrate and chloride are the anions, and they are written last. When these two substances react, they swap partners to form lead(II) chloride and sodium nitrate, according to the equation:

$$Pb(NO_3)_2(aq) + 2NaCl(aq) \rightarrow PbCl_2(s) + 2NaNO_3(aq)$$

However, many mixtures of ionic reactants do not result in the formation of a precipitate. To determine whether a reaction occurs to produce a precipitate, you can use a set of **solubility** rules.

Solubility rules

The **solubility** rules provide a guideline for the behaviour of ionic substances at 25°C and 1 atm. These rules are provided in the QCAA data book and in Table 1.

solubility
the maximum amount of solute that will dissolve in a known quantity of solvent at a given temperature

TABLE 1 Simple rules for the solubility of salts in water

1	All nitrate (NO_3^-) salts are soluble.
2	Most salts of Na^+, K^+ and NH_4^+ are soluble.
3	Most chloride (Cl^-) salts are soluble; notable exceptions are AgCl and Hg_2Cl_2.
4	Most sulfate (SO_4^{2-}) salts are soluble; notable exceptions are $BaSO_4$, $PbSO_4$ and $CaSO_4$.
5	Most hydroxides (OH^-) are only slightly soluble; the important soluble hydroxides are NaOH, KOH and $Ca(OH)_2$.
6	Most sulfide (S^{2-}), carbonate (CO_3^{2-}) and phosphate (PO_4^{3-}) salts are only slightly soluble.

Consider the following reaction:

$$Pb(NO_3)_2 + 2NaOH \rightarrow Pb(OH)_2 + 2NaNO_3$$

To determine the solubility of the substances in the equation, the identity of the anions and cations are used with the solubility rules.

- Lead(II) nitrate: The anion is nitrate (NO_3^-). Rule 1 states that all nitrates are **soluble**. A soluble substance is one that **dissolves** in water. Therefore, the **state** of $Pb(NO_3)_2$ is aqueous or (aq).
- Sodium hydroxide: The anion is hydroxide (OH^-). Rule 5 states that most hydroxides are only slightly soluble, except NaOH, KOH and $Ca(OH)_2$. Also, rule 2 states that most salts of Na^+ are soluble. Therefore, NaOH is soluble and is assigned the aqueous (aq) state.
- Lead(II) hydroxide: The anion is hydroxide (OH^-). Rule 5 states that most hydroxides are only slightly soluble, except NaOH, KOH and $Ca(OH)_2$. Therefore, $Pb(OH)_2$ is **insoluble**. An insoluble substance forms a solid precipitate and is assigned the solid (s) state.
- Sodium nitrate: The anion is nitrate (NO_3^-). Rule 1 states that all nitrates are soluble. Therefore, $NaNO_3$ is soluble and is assigned the (aq) state.

soluble
dissolves in a solvent

dissolve
to dissociate in a solvent

state
the state of matter assigned to each chemical reactant or product – solid (s), liquid (l) gas (g), or aqueous (aq)

insoluble
not able to dissolve in a solvent

The chemical equation can now be rewritten to include the states:

$$Pb(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Pb(OH)_2(s) + 2NaNO_3(aq)$$

If this reaction does not result in the formation of a solid, then it is not a precipitation reaction.

WORKED EXAMPLE 16.1A

Determine the products of a reaction between silver nitrate and potassium chloride.

SOLUTION

silver nitrate + potassium chloride → silver chloride + potassium nitrate

WORKED EXAMPLE 16.1B

Determine whether these substances are soluble in water. Give reasons to support your answers.

a KOH
b NH_4Cl
c Na_2SO_4
d $Mg(OH)_2$

SOLUTION

a Rule 2 states that most K^+ ions are soluble. Furthermore, rule 5 states that KOH is soluble. Therefore, KOH is soluble.
KOH(aq)

b Rule 2 states that most NH_4^+ ions are soluble. In rule 3, the list of exceptions for soluble chlorides does not include NH_4^+. Therefore, NH_4Cl is soluble.
$NH_4Cl(aq)$

c Rule 2 states that most Na^+ ions are soluble. Rule 4 states that most sulfate salts are soluble, and sodium is not an exception. Therefore, Na_2SO_4 is soluble.
$Na_2SO_4(aq)$

d Rule 5 states that most hydroxides are only slightly soluble, and magnesium is not an exception. Magnesium is also not included in the list of cations that are usually soluble (rule 2). Therefore, $Mg(OH)_2$ is insoluble.
$Mg(OH)_2(s)$

Slightly soluble substances

slightly soluble
only partially dissolves in a solvent, forming a solution of low concentration

solvent
the major component of a solution – what the solute is dissolved in

dissociate
to break apart

spectator ion
an ion that does not participate in a chemical reaction

Some ionic compounds are classified as **slightly soluble**. This means that some of the chemical dissolves in the **solvent**, but most remains as a solid. This occurs because the equilibrium between the ions in solution and the ionic solid strongly favours the direction of formation of the solid. The ions are so strongly attracted to one another (i.e. attractions between the anions and cations are very large) that they are less likely to **dissociate** or dissolve.

Ionic equations

When representing precipitation reactions, the **spectator ions** are often excluded from the chemical equation. A spectator ion does not participate in a chemical reaction and its state does not change throughout the reaction. In the precipitation reaction between sodium chloride and silver nitrate, the following balanced chemical equation is generated:

$$NaCl(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)$$

In this reaction, sodium (Na^+) and nitrate (NO_3^-) ions do not participate in the reaction and do not change state – they are spectator ions. The chemical equation can be rewritten to demonstrate the ions in solution before and after reacting:

$$Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \rightarrow Na^+(aq) + NO_3^-(aq) + AgCl(s)$$

Because $Na^+(aq)$ and nitrate $NO_3^-(aq)$ are unchanged as reactants and products, they are cancelled out of the equation:

$$\cancel{Na^+(aq)} + Cl^-(aq) + Ag^+(aq) + \cancel{NO_3^-(aq)} \rightarrow \cancel{Na^+(aq)} + \cancel{NO_3^-(aq)} + AgCl(s)$$

The equation can be rewritten as an **ionic equation**, which excludes the spectator ions and only represents the reacting ions. Therefore, the ionic equation is:

$$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$$

Study tip

Reactants and products can only be cancelled out of an equation if they have an equal number on each side.

ionic equation
a chemical equation that includes only the ions that change state during a chemical reaction

Experimental evidence for ions in solution

Chemists are often required to determine the identity of ions in an 'unknown' solution. Precipitation and acid–carbonate reactions can be used for this purpose. The solubility rules can be used to identify various anions and cations in a solution. Alternatively, acid–carbonate reactions produce carbon dioxide gas, which can be identified in various ways.

Identifying carbonates with acids and lime water

Most carbonates (CO_3^{2-}) are only slightly soluble. However, when the carbonate anion is paired with a sodium (Na^+), potassium (K^+) or ammonium (NH_4^+) cation, they are soluble. Because other salts are slightly soluble, it can be difficult to determine whether a precipitate has formed.

Carbonate ions can be identified in a solution by first reacting them with an acid to form a salt, carbon dioxide and water. For example, sodium carbonate (Na_2CO_3) reacts with hydrochloric acid (HCl) to form sodium chloride salt (NaCl), carbon dioxide gas (CO_2) and water (H_2O) as a liquid (l):

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

The sodium and chloride ions are aqueous as both reactants and products. Therefore, the ionic equation for this process is:

$$CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$$

The formation of bubbles is a key indicator that a gas has been formed in a reaction. Because carbon dioxide is a clear odourless gas, it is impossible to determine the identity of the gas without further testing.

Carbon dioxide extinguishes fire. Therefore, a basic test is to place a lit match in the gas formed by the reaction and observe the behaviour of the flame. If the flame is extinguished, there may be carbon dioxide, but some other gases also extinguish flames. To confirm the presence of carbon dioxide, the gas produced in the reaction can be collected and used in a second reaction.

Lime water ($Ca(OH)_2(aq)$) is a useful substance for determining whether a gas is carbon dioxide. When carbon dioxide is bubbled through lime water (a clear solution), a calcium carbonate precipitate is formed (Figure 2), according to the equation:

$$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$$

FIGURE 2 Bubbling carbon dioxide through lime water causes a precipitate of calcium carbonate to form, seen as formation of a cloudy suspension.

CHALLENGE 16.1A

Limestone

Geologists use diluted HCl to test whether a rock is a limestone (calcium carbonate). If the rock bubbles when doused with HCl it is identified as a limestone. Based on what you have learned, predict the type of gas the limestone produces when dissolved with HCl. Write an equation to reflect this reaction.

FIGURE 3 Geologist at a rock outcrop

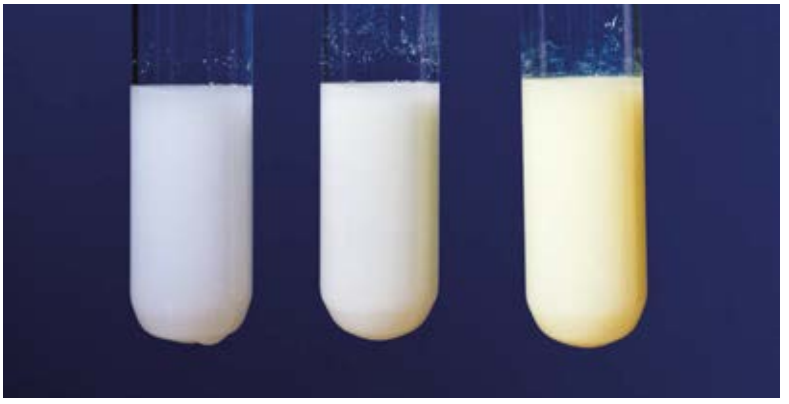

FIGURE 4 Silver chloride (AgCl) precipitate is white (left), silver bromide (AgBr) is cream (centre), and silver iodide (AgI) is yellow (right).

Identifying halogens with silver

Chloride (Cl^-), bromide (Br^-) and iodide (I^-) ions all form solid ionic substances with a silver cation (Ag^+). Solubility rule 3 verifies that chloride ions precipitate when reacted with silver ions. Although rule 3 does not mention bromide and iodide, it can be extended to group 7 elements, the halogens. Therefore, to determine the presence of halogen ions in solution, the solution is reacted with silver nitrate ($AgNO_3$). For example, a solution of silver nitrate reacts with sodium chloride to form a white precipitate of silver chloride:

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

According to the solubility rules, the sodium cation and the nitrate anion will not precipitate and therefore remain in an aqueous state as reactants and products. Any solid formed is a result of a reaction between silver and chloride:

$$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s) \text{ (white precipitate)}$$

Silver nitrate forms a precipitate with all halogens, which can be identified by their different colours. Silver bromide is a cream solid, and silver iodide is yellow (Figure 3).

$$Br^-(aq) + Ag^+(aq) \rightarrow AgBr(s) \text{ (cream precipitate)}$$

$$I^-(aq) + Ag^+(aq) \rightarrow AgI(s) \text{ (yellow precipitate)}$$

Identifying metal cations with sodium hydroxide and ammonia

Sodium hydroxide is a useful ionic compound for identifying metal cations. Most sodium compounds are soluble, and most hydroxides are slightly soluble. This means that any precipitate formed must be a result of the hydroxide reacting with a different metal cation.

If a metal is not a transition metal, it will precipitate as a white solid. This makes it very difficult to identify the metal and further testing is required to distinguish one metal from another.

Calcium, magnesium and aluminium form white precipitates when reacted with sodium hydroxide. To identify the metal, excess NaOH is added. If the precipitate re-dissolves, then the metal is aluminium (Figure 5).

Calcium and magnesium do not re-dissolve on addition of excess NaOH. A further test must be performed to distinguish between them. Ammonia (NH_3) is added to the original solution, and if a white precipitate forms, magnesium is present. Calcium will also produce a precipitate when ammonia is added, but to a far lesser extent than magnesium (Figure 5).

Transition metals

The solubility rules do not explain the solubilities of all metal ions. The ions of transition metals form coloured solutions, and precipitate with unique colours on reaction with sodium hydroxide (Table 2).

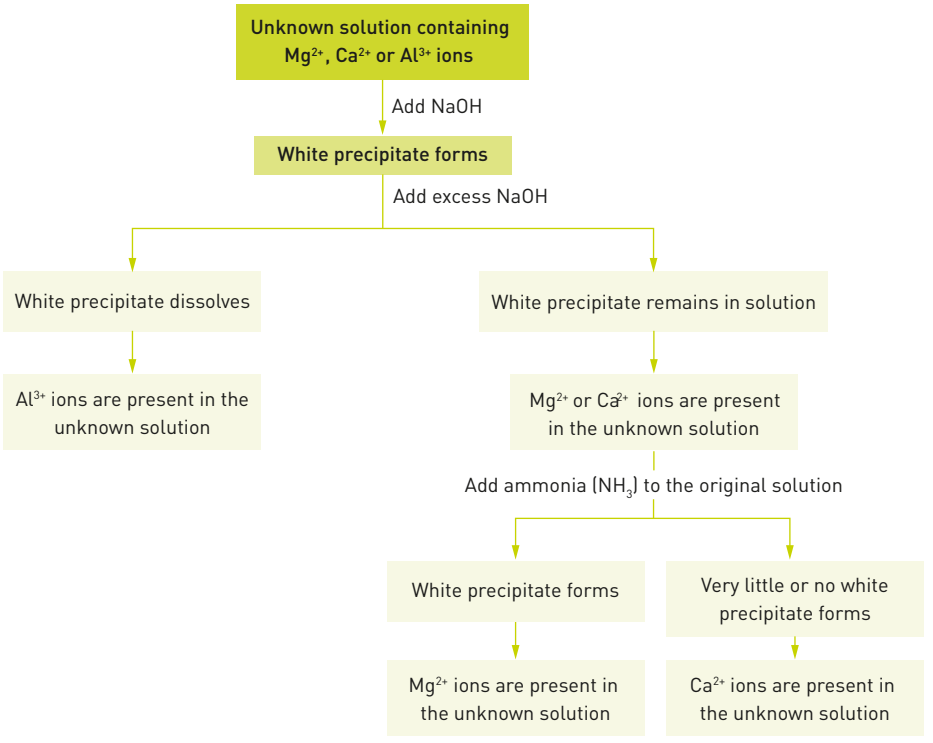

FIGURE 5 Flow chart for the identification of calcium, magnesium and aluminium ions using sodium hydroxide and ammonia.

The hydroxide reacts with the metal ions to form an insoluble coloured precipitate (Figure 5). If sodium hydroxide is added in excess, there will be no metal ions left in the solution, and it will result in a clear solvent.

Study tip

In ionic substances, always remember to write the cation first and then the anion.

TABLE 2 Colours of hydroxide precipitates of some common transition metals

Name	Formula	Colour
Iron(II)	Fe^{2+}	Green-brown
Iron(III)	Fe^{3+}	Brown-dark red
Copper(II)	Cu^{2+}	Blue
Nickel(II)	Ni^{2+}	Green

CHALLENGE 16.1B

Identifying ionic solutions

A chemist has four unknown solutions, labelled 1–4, and knows that they contain either iron(II), copper, aluminium or magnesium. Develop an overall method to determine which solution would contain each metal cation.

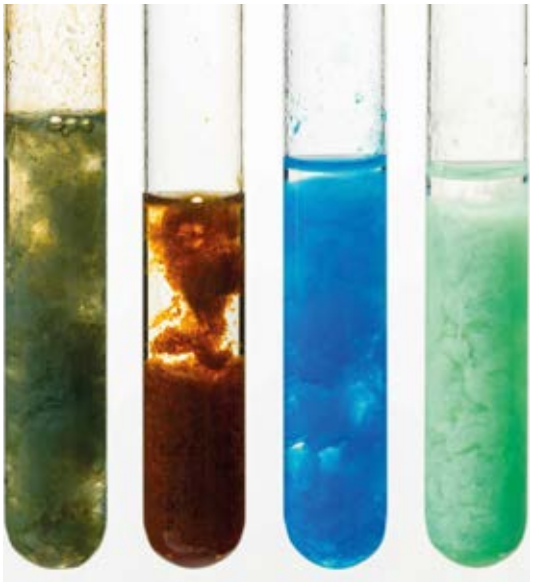

FIGURE 6 Left to right: precipitates of iron(II), iron(III), copper(II) and nickel(II) when reacted with sodium hydroxide.

CASE STUDY 16.1

Identifying mercury pollutants

Since the Industrial Revolution, more pollutants have entered the atmosphere, causing problems such as acid rain, respiratory issues, poisoning of waterways and the bioaccumulation of toxic chemicals in marine life.

One such pollutant is mercury(II) ($Hg^{2+}(aq)$), which has many negative health effects, even at low concentrations. When bonded to any soluble anion ($anion^-$), $Hg^{2+}(aq)$ can be identified by adding sodium iodide to the water sample. The reaction forms mercury(II) iodide ($HgI_2(s)$), a red-orange precipitate:

$$Hg^{2+}(aq) + anion^-(aq) + Na^+(aq) + 2I^-(aq) \rightarrow HgI_2(s) + Na^+(aq) + anion^-(aq)$$

The overall ionic equation for this process is:

$$Hg^{2+}(aq) + 2I^-(aq) \rightarrow HgI_2(s)$$

This test was especially important for the inhabitants of Minamata Bay, on the west coast of Japan, during the 1900s. In 1908, a factory was built close to the bay to produce chemicals required for fertilisers. Around 1932, the factory expanded to synthesise other, more dangerous chemicals, such as ethanal (CH_3CHO) and polyvinyl chloride (the plastic known as PVC).

FIGURE 7 A memorial to the tragedy at Minamata Bay.

The production of ethanal required a catalyst, which was mercury sulfate ($HgSO_4$) at the time. The workers did not realise that a side reaction was taking place, in which mercury sulfate was converted into methyl mercury chloride (CH_3HgCl). This highly toxic compound was dumped into the bay between 1951 and 1968.

In 1956, a young girl was taken to hospital with symptoms of convulsions and difficulty both in speaking and walking. When her home and the neighbouring area were inspected, more people were found to be suffering the same symptoms. Although doctors could not determine the cause, they knew that this epidemic disease was attacking the central nervous system of patients and named it Minamata disease.

Fishing families were found to be the main sufferers because the bay had a large fishing industry, mainly of tuna. By 4 November 1956, the cause was determined to be heavy metal poisoning. At the Minamata Disease Municipal Museum, a memorial stands as a tribute to the tragedy (Figure 7).

CHECK YOUR LEARNING 16.1

Describe and explain

1 **Explain** the behaviour of substances that are slightly soluble.

2 **Explain** why a cation must always occur with an anion.

Apply, analyse and interpret

3 **Determine** whether these substances are soluble in water. Give reasons to support your answers.

- **a** $NaNO_3$
- **b** $Pb(OH)_2$
- **c** Hg_2Cl_2
- **d** KCl
- **e** $BaSO_4$

4 **Determine** whether the following reactions are precipitation reactions. **Explain** why.

- **a** $Pb(NO_3)_2 + 2NaCl \rightarrow PbCl_2 + 2NaNO_3$
- **b** $NaCl + KNO_3 \rightarrow NaNO_3 + KCl$
- **c** $FeSO_4 + Mg(NO_3)_2 \rightarrow Fe(NO_3)_2 + MgSO_4$

Investigate, evaluate and communicate

5 A student wants to perform qualitative and quantitative tests on a precipitation reaction.

- **a** **Discuss** why a precipitation reaction is considered qualitative.
- **b** **Devise** a way to make a precipitation reaction quantitative.
- **c** **Propose** a method to determine the amount of chloride ions in a sample of sea water (containing sodium chloride), using a precipitation reaction as the first step.

6 A student in your class claims that all ions participate in a chemical reaction. **Evaluate** this statement and communicate a response, either affirming or negating their statement.

You can find the following resources for this section on your obook assess:

- » Student book questions
  Check your learning 16.1
- » Mandatory practical
  16.1 Precipitation reactions
- » Challenge
  16.1A Limestone
- » Challenge
  16.1B Identifying ionic solutions

16.2 Solubility

KEY IDEAS

In this section, you will learn about:

- the solubility of ionic and molecular substances in water
- the effect of temperature on the solubility of substances in water.

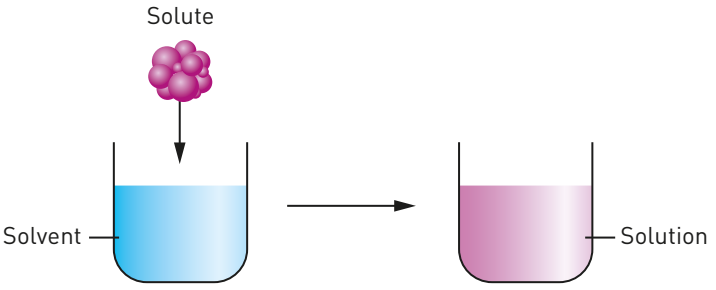

FIGURE 1 A solute dissolving in a solvent to form a solution.

As discussed throughout Units 1 & 2 Chemistry, solubility is the ability of a substance (whether it be a solid, liquid or gas) to dissolve in a solvent (Figure 1). If it dissolves easily, the substance is soluble. If it does not dissolve, the substance is insoluble.

Solubility depends on the strength of the intermolecular bonding that holds the molecules or ions within the substance together, the way that the solvent molecules can interact with the solute, and the temperature of the solvent. By altering the intermolecular forces, the properties and temperature of the solvent will change the solubility of a substance.

Solubility of substances in water

The ability of substance to dissolve in water depends on the:

- attraction between the substance and the polar water molecules (intermolecular bonding)
- attractive forces within the substance that are holding one particle to another.

Solubility of ionic substances

dipole–dipole interaction
two polar bonds or molecules interacting with each other by the partial charges

A solution is formed when a **solute** dissolves in a solvent. Salts are ionic compounds that contain a positive cation (metal) and a negative anion (non-metal). Because salts consist of ions, they contain whole positive and negative charges, which are electrostatically attracted to one another. This is a strong type of bonding (stronger than **dipole–dipole interactions**) because it involves whole charges (ionic substances), as opposed to partial charges (polar covalent molecules). When a salt (the solute) is dissolved in water (the solvent) to form a salt solution, the ionic substance with whole charges (e.g. NaCl) will interact with a polar substance (e.g. H_2O), which has partial positive and negative charges.

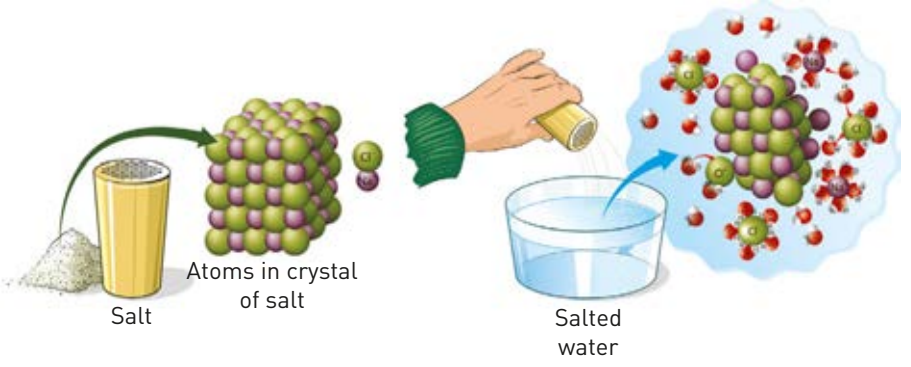

FIGURE 2 Water molecules dissociate anions from cations when they interact.

In water, the salt dissociates (breaks apart) as the sodium cation (Na^+) is attracted to the partially negative part of water (the oxygen atom), and the chloride anion (Cl^-) is attracted to the partially positive part of water (the hydrogen atoms).

It is a misconception that because ionic substances have whole positive and negative charges, which are stronger than partial charges, the salt will not dissociate. This is true if there is only one water molecule interacting with one salt molecule, but in reality, this is not the case. Many water molecules will interact with the cations and anions in the salt. The sum of all of the partial charges attached to many water molecules provides them with a very strong attractive force to pull the ionic substance apart (Figure 2). This process is called dissolving.

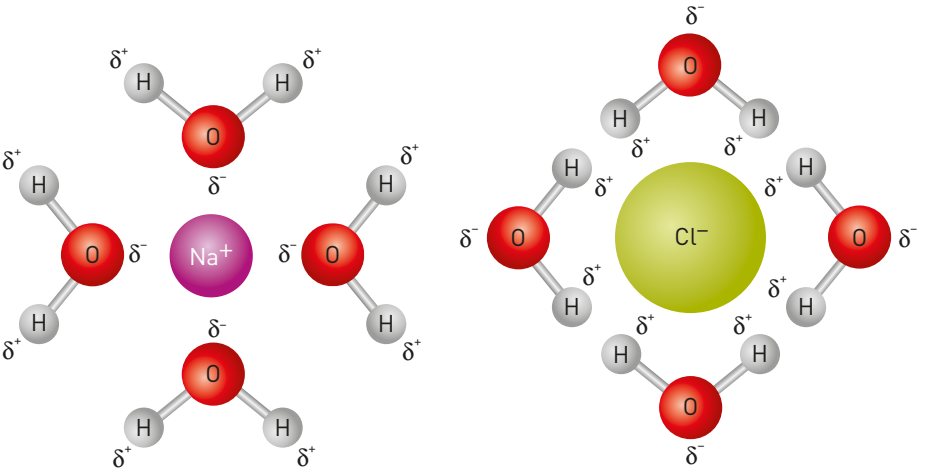

FIGURE 3 A hydrated Na^+ cation and Cl^- anion and their attraction to water molecules.

Once the salt has dissolved in the water, the ions are hydrated and are surrounded by water molecules (Figure 2). Figure 3 demonstrates the cation interacting with the partially negative part of water (oxygen) and the anion interacting with the partially positive part of water (hydrogen).

For an ionic substance to dissolve (or be soluble) in water, it must form intermolecular bonds with the water, which replace any existing bonds between cations and anions. In sodium chloride, the existing bonds within the sodium chloride solute involve the ionic interaction between Na^+ and Cl^-. The ionic interaction is replaced with an **ion–dipole interaction**, in which the ions are attracted to the **dipole** of water (Figure 3).

ion–dipole interaction
an ion with a whole charge interacting with a partial charge caused by a polar bond or molecule

dipole
a bond or a molecule that has a partial positive charge at one end and a partial negative charge at the opposite end

Solubility of molecular (covalent) substances in water

Molecular substances can dissolve in water if they are polar. If a molecular substance has the same properties as water (i.e. they are both polar), the permanent dipoles within the molecules attract each other, enabling the molecular substance (solute) to better dissolve in water (solvent).

Often molecular substances are non-polar and contain only dispersion forces, which allows them to interact with neighbouring molecules. Because dispersion forces create a temporary dipole (or a dipole moment), these molecules are still able to interact with the solvent water molecules, but not to the same extent as a polar molecule . Therefore, non-polar molecular substances are much less soluble in water.

If a molecular substance is a non-polar liquid, and it is poured into water, it will form a layer underneath or on top of the water, depending on its density. A substance with a higher density (i.e. heavier) will form the bottom layer in a flask, whereas the less dense substance will form the top layer.

The interaction between molecular substances and water is weaker than the interaction between ionic substances and water. This is because ionic substances have whole charges on their ions, whereas molecular substances are only partially positive and partially negative. Therefore, ionic substances generally have a greater solubility than molecular substances.

molecular substance
a substance made up of covalently bonded molecules in which non-metal atoms are bonded together by the sharing of electron(s) between two atoms

Study tip

Molecular substances are covalent molecules that contain only non-metals. Ionic substances are not bonded but are electrostatically attracted by the positive and negative charges of their species.

CASE STUDY 16.2

Mining for gold in the oceans

One of the richest sources of gold on Earth is the oceans. Approximately 100 million trillion dollars' worth of gold is in the oceans, but scientists have not yet been able to extract it for any profit.

Earth's mountains are composed of many different minerals, including limestone (calcium carbonate ($CaCO_3$)). There are many other minerals within these rock formations, including silver and gold minerals. As water (containing ionic substances) runs across the surface and through the cracks within the limestone, silver and gold become soluble and dissolve in the water. The minerals then travel with the water down the mountain, into rivers and finally out to sea.

The concentration of dissolved gold within the oceans is very low, in the order of parts per trillion (ppt). This is equivalent to one drop in a lake of water. More precisely, the concentration of gold in Earth's oceans is 0.000 004 mg L^{-1} (ppm) or 0.004 μg L^{-1} (ppb). This is the equivalent of 1 g of gold per 100 million tonnes of water.

There are undissolved gold deposits on the sea floor, and these are often encased in other rock deposits. The cost of travelling to the ocean floor far outweighs the profit involved. Once there, drilling and rock-breaking equipment would be required to extract the gold. In addition, rocks would need to be lifted to the surface and transported back to land for further treatment and purification before they could be sold as gold. The cost of this process far outweighs the price of the gold, and the process would not be profitable.

A methodology for the precipitation and collection of gold from oceans is still being researched. Although no method has yet been developed to reduce the solubility of the gold and precipitate it from the oceans, such a method would need to be worth the cost of the research involved, given its low concentrations.

FIGURE 4 Mining for gold on the sea floor.

Temperature and solubility

Solubility is defined as the maximum amount of solute that dissolves in a known quantity of solvent at a given temperature. Once the maximum amount of solute has been reached and no more can dissolve, the solution has become saturated.

If the temperature of a solution is increased, the solubility of a substance changes. An increase in temperature causes an increase in kinetic energy. This energy causes substances to become more or less soluble in water, depending on whether they are:

- polar or non-polar
- solid, liquid or gas.

SUGGESTED PRACTICAL

16.2 Investigating the effect of temperature on solubility
Go to your obook assess »

Solubility curves

The relationship between temperature and the solubility of substances can be plotted on a solubility curve (Figure 5). A solubility curve is a graphical representation of the mass of solute that can dissolve in 100 g of water (y-axis) at a certain temperature (x-axis).

Several key pieces of information can be determined from a solubility curve.

- The line represents the saturation point of a **solution** at a given temperature.
- A **solution** is unsaturated if it sits below the line.
- A **solution** is supersaturated if it sits above the line.
- A positive trend indicates that solubility increases as temperature increases.
- A negative trend indicates that solubility decreases as temperature increases.

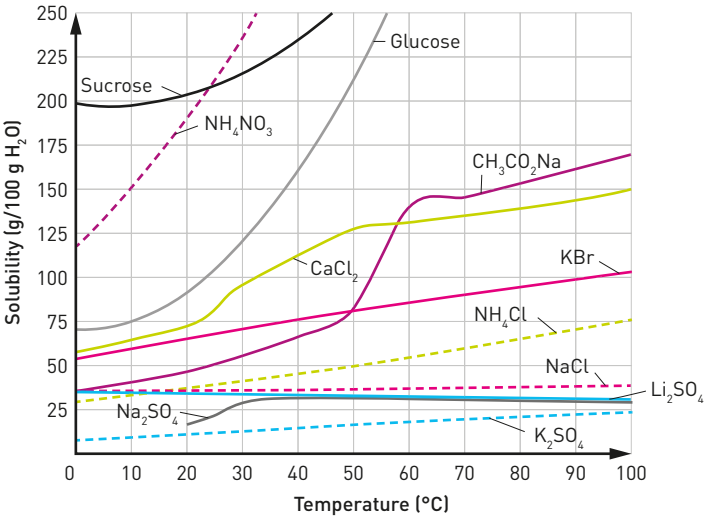

FIGURE 5 Solubility curves of some substances in water.

WORKED EXAMPLE 16.2A

At 50°C, 140 g of solid $CaCl_2$ is added to 100 mL of water. Determine whether all of the solute will dissolve.

SOLUTION

The green $CaCl_2$ line on the solubility curve (Figure 5) shows that at 50°C, a bit more than 125 g of the solute dissolves in 100 g (or 100 mL) of the solvent (water). Therefore, not all of the solute will dissolve.

Study tip

When a graph has a lot of data use different colours and dashed lines to clearly represent all data.

WORKED EXAMPLE 16.2B

At 90°C, 19.5 g of solid KBr is added to 20 mL of water. Determine whether all of the solute will dissolve.

SOLUTION

The purple KBr line on the solubility curve (Figure 5) shows that at 90°C, 100 g of the solute dissolves in 100 g (or 100 mL) of the solvent (water).

$\frac{100\text{ mL}}{20\text{ mL}} \times 19.5\text{ g of solute in 100 g}$

97.5 g is less than 100 g. Therefore, all of the solute will dissolve.

WORKED EXAMPLE 16.2C

At 21°C, 4.2 g of solute dissolves in 25 mL of water. Determine its solubility in g/100 g of water.

SOLUTION

1. 1 mL of water is 1 g. Therefore, 25 mL is 25 g.
2. In 100 g of water, there are:
   $\frac{100\text{ g}}{25\text{ g}} \times 4.2\text{ g} = 16.8\text{ g of solute}$
3. Therefore, there are 16.8 g of solute in 100 g of water and its solubility is: 16.8 g/100 g at 21°C

WORKED EXAMPLE 16.2D

3.6 mg of solute dissolves in 5.2 mL of water at 15°C. Determine its solubility in g/100 g of water.

SOLUTION

1. 1 mL of water is 1 g. Therefore, 5.2 mL is 5.2 g.
2. 3.6 mg = 0.0036 g of solute
3. In 100 g of water, there are:
   $\frac{100\text{ g}}{5.2\text{ g}} \times 0.0036\text{ g} = 0.07\text{ g}$ of solute
4. Therefore, there is 0.07 g of solute in 100 g of water and its solubility is: 0.07 g/100 g at 15°C

CHALLENGE 16.2

The solubility of lithium sulfate

Why does the solubility of lithium sulfate (Li_2SO_4) decrease as temperature increases?

Effect of temperature on the solubility of gases

Atmospheric gases such as nitrogen (N_2), oxygen (O_2) and carbon dioxide (CO_2) have very low solubilities in water. As temperature increases and kinetic energy increases, intermolecular bonds between the dissolved gas particles and the solvent water molecules are broken. Because there are very weak interactions between the water molecules and the gas molecules, these forces break more easily than the interactions between water molecules, and therefore dissolved gases evaporate more easily than water.

The best example of this phenomenon is water boiling. As the temperature of water increases, the solubility of various gases in the water decreases (Figure 6), and these gases (in addition to water vapour) form gas bubbles.

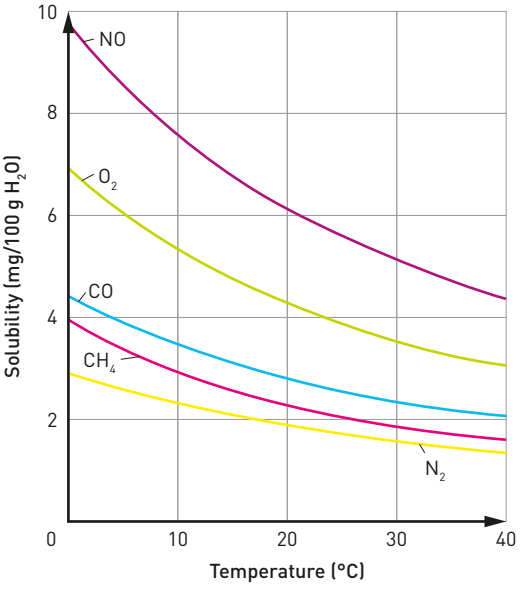

FIGURE 6 The solubility of gases in water decreases as temperature increases.

Gas particles are dissolved in solvents under high pressures. A soda stream pumps carbon dioxide gas into a bottle at high pressures, which forces the gas to dissolve in the water solvent to form fizzy water. When the bottle has been stored for a period of time, the pressure inside the bottle increases as the gas particles come out of solution. When the lid of the bottle is unscrewed, the pressure above the drink is released. Bubbles of gas form and rise to the top of the drink, reducing the amount of gas dissolved in the drink.

CHECK YOUR LEARNING 16.2

Describe and explain

1 **Explain** why most substances are more soluble as temperature increases.

2 **Identify** two reasons why a substance may be insoluble in water.

Apply, analyse and interpret

3 **Organise** the following molecules by their solubility, from strongest to weakest: non-polar molecular substance, ionic substance and polar molecular substance.

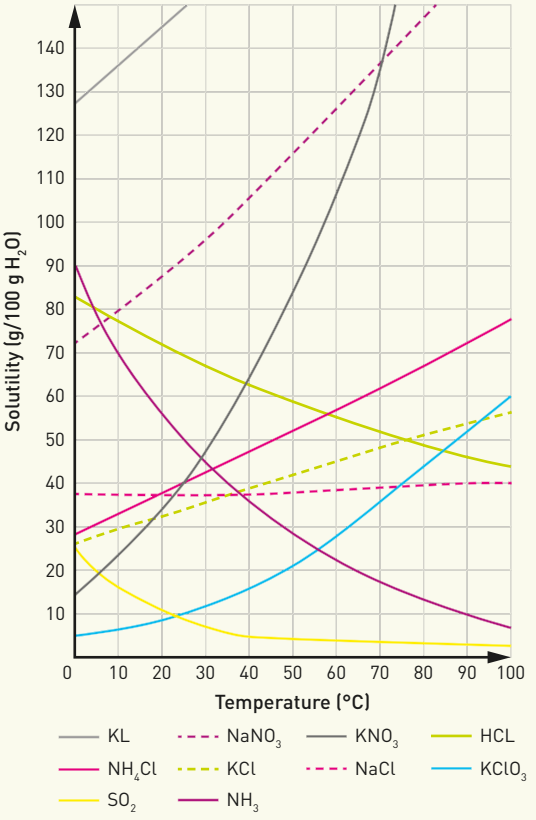

FIGURE 7 The solubilities of some ionic and molecular (covalent) substances in water.

4 a **Analyse** the solubility curves (Figure 5) and determine the:
   - i solubility of NH_4Cl at 70°C
   - ii solubility of $NaNO_3$ at 50°C
   - iii temperature required to produce a saturated solution containing 30 g of $KClO_3$
   - iv amount of $NaNO_3$ dissolved in 30 mL of water at 40°C
   - v amount of KI dissolved in 2.5 mL of water at 10°C.

   b **Consider** how a student could make a supersaturated solution of KCl.

5 **Consider** what the line on solubility diagrams represents. Determine what a line in the upward direction indicates.

Investigate, evaluate and communicate

6 Ammonia (NH_3), sulfur dioxide (SO_2) and hydrogen sulfide (H_2S) are polar molecules that are gases at room temperature. **Predict** whether their solubility curves would differ from that of a non-polar molecule. **Justify** your response.

7 The solubility of KCl at 76°C (50 g/100 g of H_2O) is the same as the solubility of KNO_3 at 32°C. **Evaluate** whether this observation is supported by the diagram or not and **hypothesise** why this may be.

8 **Research** the properties of the three materials honey, baby oil and milk. Using your knowledge of non-polar liquid substances, indicate what order the layers will form when added to water.

You can find the following resources for this section on your obook assess:

- » Student book questions
  Check your learning 16.2
- » Suggested practical
  16.2 Investigating the effect of temperature of solubility
- » Challenge
  16.2 The solubility of lithium sulfate
- » Video weblink
  Density layers

SCIENCE AS A HUMAN ENDEAVOUR

16.3 Blood chemistry

KEY IDEAS

In this section, you will learn about:

- the composition of blood
- soluble substances in blood plasma.

centrifuge
an instrument used to spin samples and separate them on the basis of their densities

centrifugal force
the apparent outwards force on a mass being rotated

blood plasma
the component of blood that contains dissolved substances

suspension
a heterogeneous mixture containing solute particles that settle to the bottom and out of the solvent

nutrient
a molecule that the body cannot synthesise on its own and must be obtained through the diet

Blood is a complex mixture of cells and dissolved substances. It can be easily separated to analyse its components and their composition. This information is useful for analysing various bodily processes to determine whether the body is functioning normally or whether medical intervention is required.

A **centrifuge** is an instrument that can be used for separating blood into its components. The blood is placed in a sealed vial and spun within the instrument for a period of time. The **centrifugal force** causes heavier components to settle to the bottom of the vial, whereas the lighter components remain on top.

Blood separates into three layers (Figure 1).

1. The bottom layer, consisting of red blood cells, is the densest component of blood and makes up 45% of its total volume. Red blood cells contain the protein haemoglobin, which carries oxygen around the body and removes carbon dioxide. These cells are made in the bone marrow and have a life span of 120 days.
2. The middle layer, consisting of white blood cells and platelets, occupies less than 1% of the volume. White blood cells are important for the correct functioning of the immune system because they fight infection, and attack viruses and bacteria that make their way into the body.
3. The top layer is plasma, which is the least dense component, occupying 55% of the blood's total volume.

Plasma

Plasma is an aqueous solution, making it the most diverse component of blood because it contains many solutes that are carried around the body for various functions. **Blood plasma** is 92% water. Plasma is referred to as a saline solution because it contains ionic substances (salts) to better enable substances to dissolve.

Plasma is the **suspension** that holds the red and white blood cells, carrying them around the body. It aids in the exchange of **nutrients** and wastes into and out of cells as well as various bodily organs.

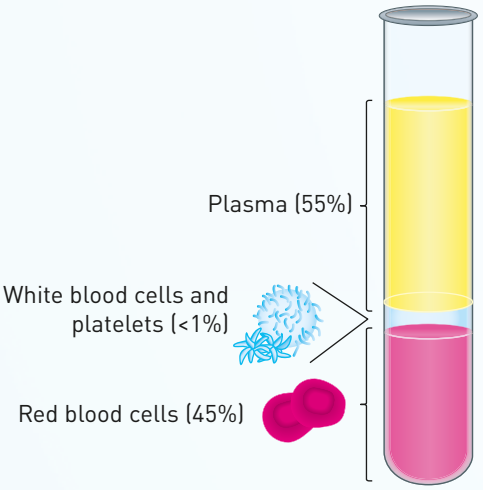

FIGURE 1 A vial of blood after being spun by a centrifuge to separate its components.

Solubility of substances in plasma

Any substance that is carried through the body in the blood's plasma must be highly soluble in water at body temperature (37°C). Transport must happen quickly to provide the body with the nutrition that it requires, keeping it functioning and healthy.

Proteins

Proteins make up 8% of blood plasma, which acts as a protein reservoir to the body. Plasma proteins have a variety of functions, including the transport of molecules and the function of the immune system.

The three primary proteins found in blood plasma are albumin 55%, globulins (37%) and fibrinogen (7%). The remaining 1% is made up of enzymes and hormones, such as insulin. It is essential that blood is kept at a nearly neutral pH of 7.4 because these proteins are only soluble in blood at this pH.

Gases

Small amounts of gases dissolve in the blood plasma and are transferred around the body, from red blood cells to the cells and organs that require them, and back again. Carbon dioxide is more soluble than oxygen.

Electrolytes/minerals

Electrolytes are substances that form ions in solution and are therefore able to conduct electricity. Because they are ionic salts, they can relatively easily dissolve in water. These ions include sodium, which is the most abundant ion in the body and has many actions, including in the correct functioning of nerve and muscle cells. Low sodium has been blamed for muscle cramping, and a common treatment is to eat a pinch of table salt (NaCl). Other common electrolytes are K^+, Ca^{2+}, Mg^{2+} and Cl^-. After physical activity, electrolyte levels can be low due to sweating. Sports drinks can replace lost electrolytes to keep the body functioning effectively.

electrolyte
a substance that forms ions in solution and is therefore able to conduct electricity

Nutrients

Nutrients are molecules that the body cannot synthesise on its own. Instead, they must be obtained from food. Some nutrients readily dissolve in blood plasma after consumption, resulting in a higher concentration, and are able to travel throughout the body to be used quickly. Glucose is one example of this type of molecule.

Glucose is absorbed quickly. This is why sugary foods give a sudden burst of energy. However, when the sugar supply is not sustained, the body crashes from its sugar high. Complex carbohydrates, consisting of many sugar molecules bonded together, provide longer lasting energy and are a healthier option.

CHALLENGE 16.3

Buffers in blood

Blood contains a carbonic acid/bicarbonate buffer, which ensures that it remains at a constant pH of 7.4. It works using the following two simultaneous equations:

$$H^+(aq) + HCO_3^-(aq) \rightleftharpoons H_2CO_3(aq) \rightleftharpoons H_2O(l) + CO_2(g)$$

What is a buffer? How does it work?

CHECK YOUR LEARNING 16.3

Describe and explain

1 Plasma consists of 90% water. **Identify** two advantages of this composition.
2 **Explain** the function of red and white blood cells in the body.
3 **Identify** the composition of sweat? Where do its components come from?
4 What components of plasma can influence its ability to act as a solvent? **Explain** why.

Investigate, evaluate and communicate

5 **Investigate** why iron is required for both the formation of red blood cells and their functioning.
6 **Discuss** why gases dissolve more readily in blood than in water.

You can find the following resources for this section on your obook assess:

» Student book questions
Check your learning 16.3

» Challenge
16.3 Buffers in blood

» Video weblink
Centrifuge

» Weblink
Electrolytes

CHAPTER 16

Review

Chapter summary

16.1
- When two aqueous ionic substances react, they can switch partners to form ionic products, where one product precipitates out as a solid:

$$AB(aq) + CD(aq) \rightarrow AD(s) + CB(aq)$$

- The six solubility rules are used to determine whether an ionic substance is soluble (aqueous) or insoluble (solid).
- Ionic equations are a simplified version of a precipitation reaction equation, where only the ions involved in the reaction are represented. Spectator ions (which do not change state) are not involved in the reaction and so are cancelled out of in the equation:

$$A^+(aq) + D^-(aq) \rightarrow AD(s)$$

16.2
- Precipitation reactions can be used in the identification of ions in an unknown solution.
- Soluble substances dissolve in a solvent. Solutes dissolve in solvents that have similar properties. For example, water, which is a polar molecule, dissolves polar solutes. Ionic substances have whole charges, resulting in a significant interaction between the solute molecules and the water molecules. Polar molecular substances also experience a large interaction with water because both contain partial positive and negative charges. Non-polar substances only interact through dispersion forces, the weakest form of intermolecular bonding, and therefore have the lowest solubility in water.
- Temperature has a direct relationship with solubility. Solubility is defined as the amount of solute that can dissolve in 100 g of water solvent at a certain temperature. In general, an increase in temperature (and an increase in kinetic energy) results in an increase in solubility. However, gases become less soluble with increasing temperature.

16.3
- Blood can be separated to measure and analyse its components; 55% is plasma, and 92% of plasma is water. Therefore, substances that dissolve in plasma must be soluble in water. Other components of plasma affect its ability to act as a solvent. These include the presence of ionic substances and a pH that is maintained at 7.4.

Key terms

- anion
- aqueous
- blood plasma
- cation
- centrifugal force
- centrifuge
- dipole–dipole interaction
- dissociate
- dissolve
- electrolyte
- insoluble
- ion
- ion–dipole interaction
- ionic bond/ionic interaction
- ionic equation
- ionic substance
- molecular substance
- nutrient
- precipitate
- precipitation reaction
- slightly soluble
- solubility
- soluble
- solute
- solution
- solvent
- spectator ion
- state
- suspension

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; *** = high.

Multiple choice

1 Identify which one of the following ionic substances is insoluble in water.
- **A** NaOH
- **B** KCl
- **C** AgCl
- **D** $NaNO_3$

2 Identify which one of the following ionic substances is soluble in water.
- **A** Hg_2Cl_2
- **B** $BaSO_4$
- **C** $PbCl_2$
- **D** K_2SO_4

3 Ionic substances have the highest solubility in water because they:
- **A** are polar molecules containing partial charges
- **B** are non-polar molecules
- **C** are polar molecules containing whole charges
- **D** contain whole charges on ions.

4 Non-polar molecules are the least soluble in water because they contain:
- **A** dispersion forces, creating a temporary dipole and therefore weaker attraction between molecules
- **B** a permanent dipole but a large amount of kinetic energy and therefore weaker attraction between molecules
- **C** ionic interactions, but these become weaker as they dissolve because cations and anions move away from one another in the water
- **D** dispersion forces, creating a permanent dipole and therefore weaker attraction between molecules.

5 The solubility of some substances in water:
- **A** increases as temperature increases
- **B** increases as temperature decreases
- **C** decreases as temperature increases
- **D** all of the above.

6 What colour precipitate is silver iodide?
- **A** yellow
- **B** white
- **C** brown
- **D** cream

7 The colour of rust indicates which transition metal is present?
- **A** Iron (II)
- **B** Iron (III)
- **C** Manganese (IV)
- **D** Copper (II)

FIGURE 1 Rust forms when a particular metal is exposed to oxygen.

8 Which of the following is not a molecular compound?
- **A** H_2O
- **B** NF_3
- **C** H_2S
- **D** Fe_2O_3

9 What colour will a metal precipitate be if a non-transition metal reacts with sodium hydroxide?
- **A** it will not form a precipitate
- **B** colourless
- **C** white
- **D** green

10 Which of the following is not a rule for the solubility of salts in water?

- **A** most chloride (Cl^-) salts are soluble
- **B** all nitrate (NO_3) salts not soluble
- **C** most salts of Na^+ and NH^{4+} are soluble
- **D** most sulphate (SO_4^{2-}) salts are soluble.

Short answer

Describe and explain

★**11 Explain** how a centrifuge is used to separate mud into water and soil.

FIGURE 2 Balanced centrifuge with samples loaded.

★**12 Explain** why some molecular substances are soluble in water, while others form layers with water.

★**13 Explain** why a precipitation reaction only forms a solid. Use an example to demonstrate.

★**14 Explain** why the solubility of all substances is converted to or measured in g/100 g of water.

Apply, analyse and interpret

★★**15** The structures of glucose and sucrose are shown in Figure 3. Using the solubility curve (Figure 4) and the structures of these molecules, **interpret** the difference in their solubility data.

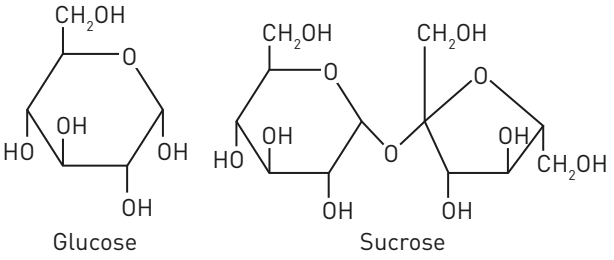

FIGURE 3 Glucose and sucrose

★★**16** KBr and K_2SO_4 both contain potassium cations but have very different solubility data. **Consider** what differences between these substances is causing this variation. **Deduce** why KBr has a higher solubility than that of K_2SO_4.

★★**17** Write the products of the following precipitation reactions and **determine** whether they are soluble or insoluble.

- **a** sodium sulfate + barium nitrate →
- **b** lead(II) nitrate + sodium hydroxide →
- **c** aluminium chloride + potassium hydroxide →

★★★**18 a** Analyse the solubility curves in Figure 2 to **determine** the:

- **i** solubility of NH_4NO_3 at 20°C
- **ii** solubility of K_2SO_4 at 80°C

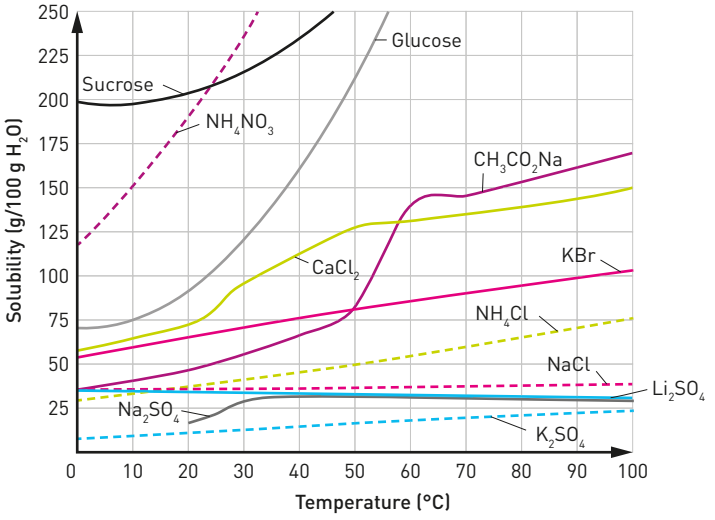

FIGURE 4 Solubility curves of several compounds.

iii temperature required to produce a saturated solution of glucose if 25 g of glucose is dissolved in 20 mL of water

iv solubility of sucrose in 40 mL of water at 30°C

v solubility of $CaCl_2$ in 9 mL of water at 40°C.

b **Explain** how a student could make a supersaturated solution of NH_4NO_3 at 20°C.

Investigate, evaluate and communicate

★★ **19** A student claims that a supersaturated solution cannot be made in a laboratory without a hot plate. **Evaluate** this claim and respond using the solubility rules.

★★★ **20** **Devise** a method to determine whether carbonate, bromide, aluminium and calcium ions are present in an unknown sample. Present your method as numbered steps, detailing the equipment that could be used in a laboratory.

★★★ **21** A scientist has reached the final step in the determination of unknown ions in a sample of waste water. The scientist cannot distinguish between chloride, iodide and bromide ions because they are all present in the same solution and their distinctive colouring is unclear.

Argue whether the scientist could use colour to draw any conclusions. **Investigate** other techniques for distinguishing between these ions.

FIGURE 5 A scientist is investigating an unknown solution.

★★★ **22** **Investigate** which sports drinks contain which electrolytes. Use this information to determine the best sports drink for replacing potassium in the human body. Are there other alternatives for humans to replace potassium in their bodies?

FIGURE 6 Different types of sports drinks

Check your obookassess for these additional resources and more:

- **Student book questions**
  Chapter 16 Revision questions
- **Revision notes**
  Chapter 16
- **assess quiz**
  Auto-correcting multiple choice quiz
- **Flashcard glossary**
  Chapter 16

CHAPTER 17

The pH scale

The word 'acid' comes from the Latin *acidus*, meaning sour. Acids may be concentrated or dilute, strong or weak. Weak acids such as the acetic acid in vinegar can be dangerous if they are highly concentrated. A base is often thought as the opposite of an acid, and many people underestimate the toxicity of a base. The pH scale is used as a measure of the strength of an acid or base.

OBJECTIVES

→ Use the Arrhenius model to explain the behaviour of strong and weak acids and bases in aqueous solutions.
→ Recall that pH is dependent on the concentration of hydrogen ions in solution.
→ Use the pH scale to compare the levels of acidity or alkalinity of aqueous solutions.

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PRACTICALS

MANDATORY PRACTICAL **17.1** Investigating the properties of strong and weak acids

FIGURE 1 Lemons have a sour taste due to the citric acid (pH 2–3) they contain. This may seem like a strong acid; however, it is usually very dilute in the fruit and therefore it is not considered harmful.

17.1 The Arrhenius model of acids and bases

KEY IDEAS

In this section, you will learn about:

- how Arrhenius acids increase the concentration of hydrogen ions in water
- how Arrhenius bases increase the concentration of hydroxide ions in water
- alkalis, which are bases that are soluble in water
- strong acids, which can easily release a hydrogen ion to form a hydronium ion in water
- weak acids, which cannot easily release a hydrogen ion to form a hydronium ion in water
- how Arrhenius described a strong base as a substance that easily dissociates to form hydroxide ions
- how a concentrated acid or base has a large number of acid or base molecules (respectively) in an aqueous solution.

acid
a chemical compound that donates a hydrogen ion

Lemons, vinegar and soft drinks all contain **acids**. Acids (and their 'partner' bases) were first described by Robert Boyle in 1680 – he described how some substances changed the colour of some natural dyes. In the 1700s, acids were described as having a sour taste and the ability to react with limestone to produce a gas. Today, acids such as citric acid and acetic (ethanoic) acid are often used in lollies to intentionally produce the popular sour flavour (Figure 1).

FIGURE 1 Sour lollies can contain a variety of different acids.

base
a chemical compound that accepts a hydrogen ion

alkali
a base that is soluble in water

Often, **bases** are considered the matching partner to acids. Bases, such as sodium bicarbonate, are bitter to taste and feel slippery. Like acids, bases are corrosive and can produce chemical burns on skin. Many detergents (including toothpaste) are water-soluble bases, known as **alkalis**.

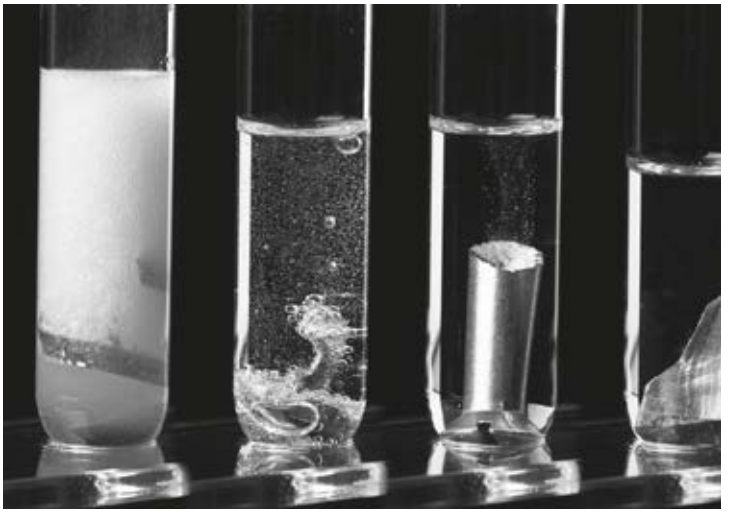

FIGURE 2 Acids react with metals to form a metal salt and hydrogen gas.

Arrhenius model of acids and bases

In chemistry, we need better methods to test acids and bases because it is not feasible to taste every chemical to determine whether they are an acid or a base.

In 1884, a scientist named Svante Arrhenius examined a group of materials with the properties of acids and found that when they were dissolved in water, the solutions had very high concentrations of hydrogen ions. As a result, Arrehnius defined acids as the group of chemicals that can produce a high concentration of hydrogen ions in an aqueous solution.

For example, a hydrochloric acid aqueous solution is composed of hydronium ions and chloride ions dissolved in water. When HCl gas is dissolved in water, the HCl molecules combine with water molecules to form H_3O^+ (hydronium ions) and chloride ions:

$$HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

The hydrogen ion does not stay with a single water molecule. Instead, it constantly shuffles between different water molecules. For this reason, when a chemist refers to the concentration of hydrogen ions in an acid, they are also including the hydronium ions in their calculations.

FIGURE 3 Most detergents and toothpaste are bases that dissolve in water (alkali).

Once Arrhenius had defined an acid as a material that was able to produce a high concentration of hydrogen ions when mixed with water, he started working on the group of chemicals known as bases. He found that when a base was dissolved in water to form an alkali, it resulted in a high concentration of hydroxide ions (OH^-).

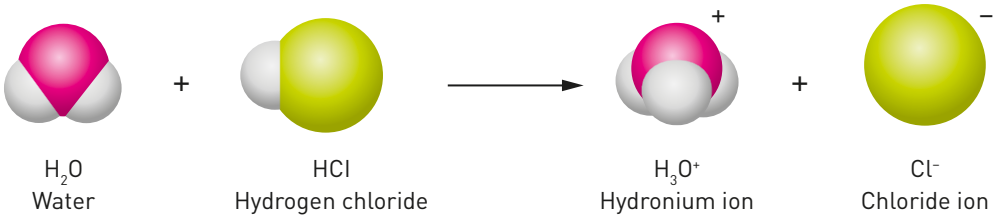

FIGURE 4 An Arrhenius acid dissociates in water to form hydronium ions (H_3O^+). In this example, water reacts with the HCl molecules to form hydronium ions (H_3O^+) and Cl^- ions, by formation of new O–H bonds and breaking of H–Cl bonds.

Study tip

A hydrogen atom consists of a single positive proton that has a single negative electron moving in the space around it. If the electron is removed (when a hydrogen ion is formed), only the proton remains. For this reason, a hydrogen ion is also known as a proton.

One of the most common examples of this is when sodium hydroxide (NaOH) is mixed with water. Sodium hydroxide is an ionic compound, made up of sodium ions (Na^+) and hydroxide ions (OH^-) that dissociates (separate) in water. Arrhenius tested many different bases that behaved in the same way and concluded that all bases increase the concentration of hydroxide ions in a water solution:

$$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$$

Study tip

Disassociation to form ions is not the same as ionisation. Dissociation occurs when an ionic compound separates into its separate ions. Ionisation occurs when the molecule reacts with water to produce ions.

Concentrated and dilute

As seen in Chapter 15, concentrated solutions contain more solute than the same volume of dilute solutions. Therefore, 1 L a concentrated salt solution contains more salt (NaCl) molecules than 1 L of a dilute salt solution. The same holds true for a concentrated acid or base solution. A concentrated solution of carbonic acid has a high number of carbonic acid (H_2CO_3) molecules. A dilute solution of sodium hydroxide contains a low number of sodium and hydroxide ions. This does not change how readily the acid or base molecules release hydrogen ions or hydroxide ions. Instead, it is a measure of the number of acid or base molecules there are in a set volume of solution.

Weak and strong acids

Acids are often described as weak or strong. These terms refer to the extent to which the acid reacts with water to form hydronium ions. Acids such as nitric acid, sulfuric acid and hydrochloric acid will completely react to form hydronium ions. These acids are often described as strong acids because they almost completely ionise (convert into ions) in water. This means a strong acid (completely forms hydronium ions) can be either dilute or concentrated, depending on the concentration of the acid in the solution. A dilute hydrochloric acid solution only has a low concentration of hydronium ions and chloride ions. A concentrated solution of ethanoic acid (a weak acid) has many CH_3COOH molecules in the solution, but only a few release their hydrogen ions to form hydronium (H_3O^+), ions and ethanoate (CH_3COO^-) ions.

In the Arrhenius model, strong bases easily release hydroxide ions (OH^-) into the solution. The existence of weak bases can be explained in terms of a more comprehensive theory of acids and bases called Bronsted–Lowry, which is covered in Unit 3.

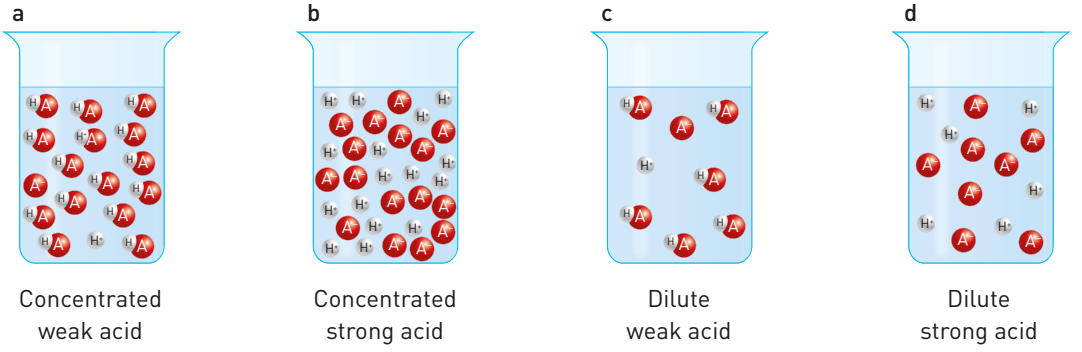

FIGURE 5 (a) A concentrated weak acid has many acid molecules but few free hydrogen ions. (b) A concentrated strong acid has many acid molecules that have released their hydrogen ions into the solution. (c) A dilute weak acid solution has few acid molecules that do not easily release their hydrogen ions. (d) A dilute strong acid solution has few acid molecules, but most have released their hydrogen ions into the solution.

CASE STUDY 17.1

Not all hydrogen leads to acid

A molecule with a large number of hydrogen atoms, is not necessarily an acid. Methane (CH_4) has four hydrogen atoms covalently attached to the carbon atom. The covalent bond is very strong, meaning the hydrogen atoms are not easily released when mixed with water. Similarly, glycerol (Figure 6) has three OH groups in its molecule, but the OH groups do not form hydroxide ions. This means glycerol does not act as a base.

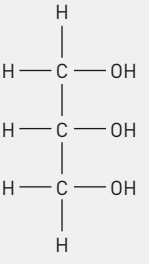

FIGURE 6 Glycerol does not act as a base, even though it has three OH groups.

MANDATORY PRACTICAL
17.1 Investigating the properties of strong and weak acids
Go to page 455 »

Limitations of the Arrhenius model

Despite the many advantages of the Arrhenius model of acids and bases, some substances do not fit this model. The Arrhenius model requires all acids and bases to be able to produce hydrogen ions or hydroxide ions respectively. However, some substances (e.g. NH_3), can still act as a base despite not containing hydroxide ions. Other compounds (e.g. SO_3) do not contain a hydrogen but react with water to release a hydrogen ion.

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(l)$$

$$H_2SO_4(l) + H_2O(l) \rightarrow H_3O^+(aq) + HSO_4^-(aq)$$

As a result, two scientists Brønsted and Lowry independently suggested a modification of the Arrhenius model. They suggested that acids release hydrogen ions (protons, H^+) and bases bond with hydrogen ions (protons). The Brønsted–Lowry model is covered in Unit 3 Chemistry.

CHECK YOUR LEARNING 17.1

Describe and explain

1 **Identify** the properties of:
   - **a** an acid
   - **b** a base.

2 **Define** 'alkali'.

3 **Explain** how Arrhenius defined:
   - **a** an acid
   - **b** a base.

Apply, analyse and interpret

4 Below is a list of properties of acids and bases. Using a Venn diagram, **classify** each property as belonging to an acid, a base, or both.
   - sour taste
   - bitter taste
   - corrosive
   - feels slippery
   - produces a chemical burn on skin
   - produced in the stomach
   - ionises in water
   - disassociates in water

5 Using the Arrhenius model, **differentiate** between a dilute strong acid and a concentrated weak acid.

Investigate, evaluate and communicate

6 Ammonia (NH_3) is often described as a base even though it does not contain hydroxide ions. **Discuss** how the Arrhenius model has been modified to account for this observation.

Check your obook assess for these additional resources and more:

- » Student book questions
  Check your learning 17.1
- » Mandatory practical
  17.1 Investigating the properties of strong and weak acids
- » Weblink
  Arrhenius model
- » Weblink
  Acids and food

17.2 The pH scale

KEY IDEAS

In this section you will learn about:

- how aqueous solutions with a pH below 7 are acids
- how aqueous solutions with a pH of 7 are neutral
- how aqueous solutions with a pH above 7 are basic
- how acidic solutions with low pH values have high concentrations of hydrogen ions
- indicators, which can be used to identify the strength of an acid.

pH
a numerical scale that is used to determine the level of acid or base in an aqueous solution

neutral
an aqueous solution that has a pH of 7

The **pH** scale is a numerical system for determining whether an aqueous solution is acidic or basic. Most commonly, the pH scale is measured between 0 and 14. Substances such as pure water are considered **neutral** (neither acidic nor basic) and are assigned a pH of 7. Acidic solutions have a pH of less than 7. The lower the pH value, more acidic the solution and the more concentrated the hydronium ions in that solution. Most detergents, such as ammonia and caustic soda, are basic and have a pH of greater than 7. Bleach is a strong base with a pH of 12.

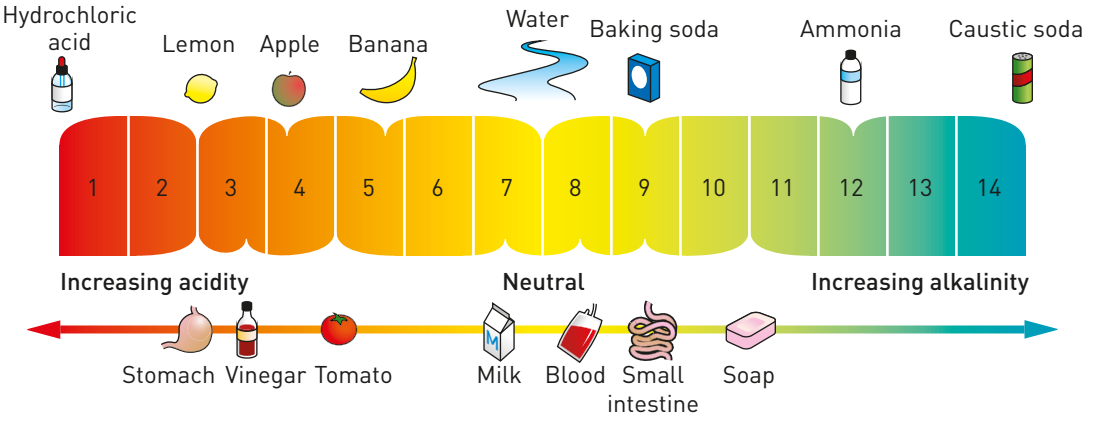

FIGURE 1 The pH scale is used to determine the relative strength of an acid or base. Acids have a pH below 7, bases have a pH above 7 and neutral substances have a pH of 7.

Study tip

A change in pH from 0 to 14 covers a change in hydronium ion concentration from 1 M to 1 × 10^{-14} M, i.e. 14 orders of magnitude. The pH scale is a useful way to compress large changes in concentration, so they are easier to think about.

There are many solutions that have a pH outside the 0–14 range; for example, any hydrochloric acid solution of concentration > 1 M. The reason for this is explained in the next section.

Study tip

Be careful not to confuse alkalinity with basicity. Alkalinity is the term used to describe water's ability to resist becoming acidic, and basicity is a measurement on the pH scale.

pH and hydrogen ion concentrations

According to the Arrhenius model, acids are substances that increase the concentration of hydronium ions in aqueous solutions. Strong acids react completely to form hydronium, while weak acids only partially react. In the Arrhenius model of acids and bases, bases have high concentrations of hydroxide ions (OH^-). These hydroxide ions can bind to free hydrogen ions (H^+) to form water (H_2O). As a result, the concentration of hydrogen ions decreases. The concentration of free hydrogen ions can therefore be used to indicate the 'power' of the acids and bases. pH is a calculation based on the concentration of hydrogen/hydronium ions, using the formula:

$$pH = -\log_{10} [H^+]$$

where $[H^+]$ is the concentration of H^+ (M).

Many sources quote pH as meaning 'power of hydrogen' or *puissance d'hydrogen* (French for 'strength of hydrogen'). Alternatively, the 'p' is thought to be a mathematical symbol that refers to the negative logarithmic scale, and the H refers to the concentration of hydrogen. Therefore, pH is the negative logarithmic scale of hydrogen ions.

TABLE 1 The pH of a solution is an indication of the hydrogen concentration

	Hydrogen concentration (M)	pH
Highly acidic solution	$0.1 = 10^{-1}$	1
Weakly acidic solution	$0.000\ 001 = 10^{-6}$	6
Neutral	$0.000\ 0001 = 10^{-7}$	7
Weakly basic solution	$0.000\ 000\ 001 = 10^{-9}$	9
Highly basic solution	$0.000\ 000\ 000\ 000\ 01 = 10^{-14}$	14

Using a log 10 scale to convert hydrogen concentration to a pH scale means that each pH value has a hydrogen concentration 10 times larger or smaller than the pH value one pH unit away. The equation for pH can be rearranged to determine the concentration of hydrogen ions in a solution:

$$[H^+] = 10^{-pH}\ (M)$$

An aqueous solution with a pH of 2 has ten times more hydronium ions in the same volume than a solution with a pH of 3. The lower the pH value, the higher the hydronium ion concentration.

WORKED EXAMPLE 17.2

If H^+ concentration is 0.001 M, what is the pH?

SOLUTION

0.001 M can be written as 1.0×10^{-3} M.

$$\begin{aligned} pH &= -\log_{10}[H^+] \\ &= -\log_{10}[1\,0^{-3}] \\ &= -(-3) \\ pH &= 3 \end{aligned}$$

Hint: Use a calculator to enter –log (0.001) to get pH = 3.

Indicators

Indicators are dyes that have different coloured forms when they are protonated and unprotonated. The indicator may be in solution or be infused into a paper. Indicators are weak acids. Different indicators change colour in different pH ranges (Figure 2).

indicator
a substance that changes colour in the presence of high and low hydrogen ion concentrations

$$H\text{–indicator} + H_2O \rightleftharpoons H_3O^+ + \text{indicator}$$
$$\text{Colour 1} \rightleftharpoons \text{Colour 2}$$

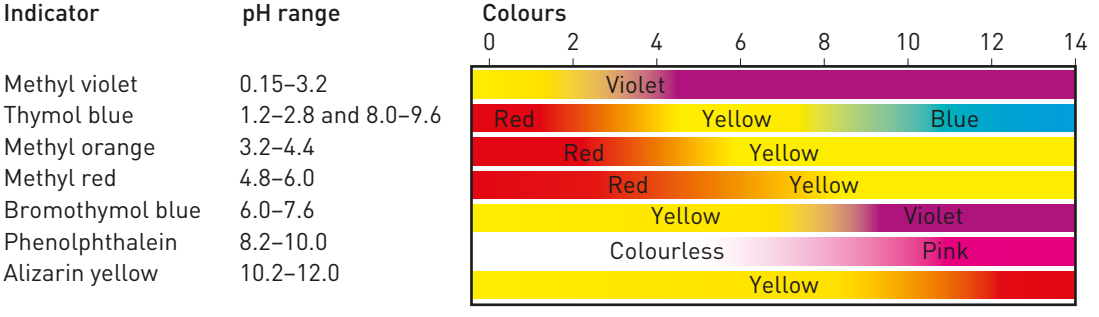

FIGURE 2 Different indicators change colour at different pH values.

These colour changes can be used to indicate the pH of the solution. Many common indicators were originally purified from plants. For example, the juice of a red cabbage, for example, is green in a basic solution, blue in a neutral solution and red in an acidic solution.

One of the most common indicators is universal indicator. Universal indicator is a mixture of many different indicators, including methyl red, bromothymol blue and phenolphthalein. This means it changes colours several times over a wide pH range. It is red in highly acidic solutions (pH 0–3), yellow in weakly acidic solutions (3–6), green at neutral pH (7), blue in a weakly basic solutions (8–11) and purple in a strongly basic solution (11–14).

Although indicators are a useful visual tool to determine the pH of a substance, most chemists use pH meters to more accurately measure the pH of substances. A pH meter measures the voltage of a solution that varies with different hydrogen ion concentrations. pH meters are more accurate than indicators because they can measure values to 0.01 of a pH value.

CHALLENGE 17.2

pH and flower colour

Two colours of hydrangea flowers act as an indicator for the pH of the soil. Hydrangeas have different coloured flowers depending on the presence of aluminium ions and the acidity of the soil.

In a neutral or basic soil (pH ≥ 7), aluminium ions in the soil form aluminium hydroxide. Aluminium hydroxide cannot move into the roots of the plant. Hydrangeas in this soil type will have blue flowers. In acidic soil (pH < 7), there are few hydroxide ions. This leaves the aluminium ions free to form other complex molecules that can move into the plant, reacting with pigments in the petals to produce pink–red flowers.

What colour will the flowers of a hydrangea be if there is no aluminum in acidic soil? Provide the reasoning that will support your answer.

FIGURE 3 The colour of hydrangea flowers depends on soil acidity and the presence of aluminium ions.

CHECK YOUR LEARNING 17.2

Describe and explain

1 **Identify** each of the following pH values as that of an acidic, neutral or basic solution.
   - **a** pH 3.1
   - **b** pH 7
   - **c** pH 11.3
   - **d** pH 1.3

2 **Explain** the three different meanings of the term 'pH'.

3 **Describe** how an indicator can be used to determine if a substance is an acid or a base.

Apply, analyse and interpret

4 Use Table 2 to **determine** the colour of each indicator at:
   - **a** pH = 2
   - **b** pH = 10.

TABLE 2 pH indicators for acids and bases

Indicator	pH at which colour changes	Strong acid pH	Strong base pH
Bromophenol blue	4.0	Yellow	Blue
Methyl red	5.1	Red	Yellow
Phenolphthalein	9.3	Colourless	Red

5 Interpret Table 3 to **determine** which plants will grow well with low levels of hydronium ions in the soil.

FIGURE 4 Vegetables grow well in different soils, depending on pH.

TABLE 3 Vegetable optimum pH levels

Vegetable	Optimal pH
Avocado	6.5–7.5
Broccoli	6.0–7.0
Cabbage	5.6–6.6
Capsicum	6.0–8.0
Cucumber	5.0–6.0
Garlic	5.0–6.0
Mushroom	6.2–6.8
Onions	7.0–8.0
Spinach	5.0–7.0
Tomato	5.5–7.0

Investigate, evaluate and communicate

6 If the concentration of hydrogen ions in a solution increases, **predict** what changes there would be in the pH value.

7 The optimum pH for water in a swimming pool is considered to be 7.4. A student claimed that this was a neutral pH. **Evaluate** this claim, commenting on whether or not you agree or disagree.

FIGURE 5 Measuring the pH of a swimming pool

Check your obook assess for these additional resources and more:

- **» Student book questions**
  Check your learning 17.2
- **» Challenge**
  17.2 pH and flower colour
- **» Weblink**
  pH indicators
- **» Weblink**
  pH in pools

CHAPTER 17

Review

Chapter summary

17.1
- Arrhenius acids increase the concentration of hydrogen ions in water.
- Arrhenius bases decrease the concentration of hydroxide ions in water.
- Bases that are soluble in water are called alkalis.
- A strong acid reacts completely with water to form hydronium ions.
- A weak acid does not react completely with water to form hydronium ions.
- A concentrated acid or base has a large number of acid or base molecules (respectively) in an aqueous solution.

17.2
- Aqueous solutions with a pH below 7 are acidic.
- Aqueous solutions with a pH of 7 are neutral.
- Aqueous solutions with a pH above 7 are basic.
- Acidic solutions with low pH values have high concentrations of hydronium ions.
- Indicators can be used to identify the hydronium ion concentration of a solution or its pH.
- Alkalinity describes the ability of water to resist becoming an acid.
- Acidity is a measure of how easily an acid is able to produce hydronium ions.

Key terms

- acid
- alkali
- base
- indicator
- neutral
- pH

Key formulas

pH calculation	$pH = -\log_{10}[H^+]$ $[H_3O^+] = 10^{-pH}\,(M)$

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; ***=high.

Multiple choice

1 Identify which of the following is the best definition of an acid.

- **A** An acid is a substance that tastes sour.
- **B** An acid is a substance that produces hydroxide ions.
- **C** An acid is a substance that produces hydronium ions in water.
- **D** An acid is a substance that reacts with water to produce hydrogen.

2 Equal mole amounts of substance X and substance Y were used to make aqueous solutions of the same volume. The solution made from substance X had a high concentration of hydronium ions, while the solution containing substance Y had a low concentration of hydronium ions. Use this information to determine the correct statement.

- **A** X has a higher concentration of acid than Y.
- **B** Y has a higher concentration of acid than X.
- **C** Y is a stronger acid than X.
- **D** Y is an alkali.

3 Identify the best description of an indicator.

- **A** a dye that changes colour
- **B** a dye that changes colour in a basic solution
- **C** a dye that changes colour in an acidic solution
- **D** a dye that changes colour in different hydronium ion concentrations

4 A student placed a drop of a liquid on blue litmus paper (which becomes red in acidic solutions) and observed that there was no colour change. This meant the student could assume that the substance was:

- **A** a base
- **B** an acid
- **C** neutral
- **D** not an acid.

5 An aqueous solution becomes yellow when universal indicator is added. What does this indicate about the pH of the solution?

- **A** pH = 1
- **B** pH = 5
- **C** pH = 7
- **D** pH = 14

Short answer

Describe and explain

★ **6** **Describe** the Arrhenius model of acids and bases.

★ **7** **Explain** how an aqueous solution of H_2SO_4 can be either concentrated or dilute, but can only be strong (and not weak).

★ **8** **Explain** why an aqueous solution of HCl is able to conduct electricity.

Apply, analyse and interpret

★★ **9** **Determine** what pH you would expect for the following solutions.

- **a** $[H_3O^+] = 10^{-4}$
- **b** $[H_3O^+] = 10^{-7}$
- **c** $[H_3O^+] = 10^{-11}$

★★ **10** A chemist describes a base as being able to completely disassociate (separate into ions) in an aqueous solution. **Consider** what this suggests about the strength of the base.

★★ **11** **Reflect on** why more than one indicator should be used when trying to determine if a substance is an acid, or does not alter the pH of the solution, or is a base.

★★ **12** **Determine** what type of indicator should be used to determine the pH of a cloudy (opaque) substance.

★★ **13** **Determine** whether the addition of water changes the strength or the concentration of an acid.

★★ **14** **Apply** the Arrhenius definition of a base to explain how $Ca(OH)_2$ qualifies as a base even though it is relatively insoluble in water.

Investigate, evaluate and communicate

★★ **15** **Discuss** how a difference of 1 pH unit (from pH 3 to pH 4) changes the concentration of hydrogen ions.

★★ **16** **Develop** a diagram to represent a:

- **a** dilute aqueous solution of the weak acid, ethanoic acid (CH_3COOH)
- **b** concentrated aqueous solution of the strong base potassium hydroxide (KOH).

Check your obook assess for these additional resources and more:

- » Student book questions: Chapter 17 Revision questions
- » Revision notes: Chapter 17
- » assess quiz: Auto-correcting multiple-choice quiz
- » Flashcard glossary: Chapter 17

CHAPTER 18

Reactions of acids

Knowing how an acid reacts with other substances allows scientists to predict the products of the reaction. Acid spill kits use chemicals to react with the spilt acid, producing molecules that are not as toxic or dangerous to humans or surrounding objects. Knowing how acids react in the environment around us provides an understanding of how the acids in polluted rain affect metals or marble in the environment. Understanding these chemical reactions, gives us the opportunity to control these reactions and prevent further damage to natural and manmade structures.

OBJECTIVES

→ Understand and apply the reactions of acids with bases, metals and carbonates to determine reactants and products.

→ Construct and use appropriate representations including ionic formulas, chemical formulas and chemical equations, to symbolise the reactions of acids and bases; and ionic equations to represent the reacting species and products in these reactions.

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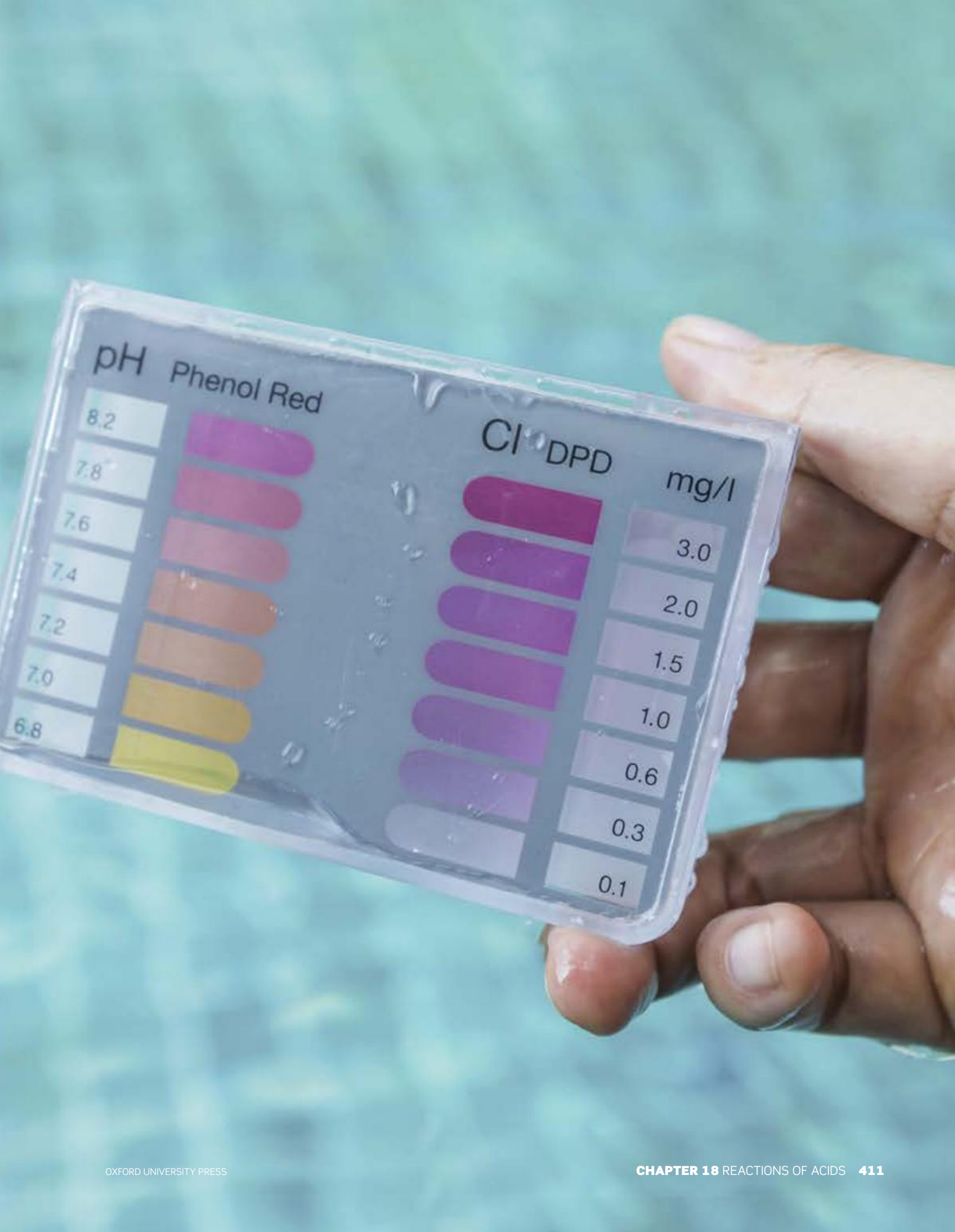
PRACTICALS

SUGGESTED PRACTICAL **18.2** Reactions of acids with bases, metals and carbonates

FIGURE 1 Swimming pool test kits can be used to determine the amount of acid or base needed to adjust a pool's pH to the desired level.


pH
Phenol Red
8.2
7.8
7.6
7.4
7.2
7.0
6.8
Cl
DPD
mg/l
3.0
2.0
1.5
1.0
0.6
0.3
0.1

18.1 Neutralisation reactions

KEY IDEAS

In this section you will learn about:

- neutralisation reactions, which involve mixing an acid and a base to produce water and a metal salt
- how a completed neutralisation reaction between a strong acid and a strong base will have a pH of 7
- how a strong acid may only be partially neutralised by a weak base
- titration, which involves dripping an acid or a base into a substance with an unknown pH until the known indicator changes colour.

In the Chapter 17 you learnt about the Arrhenius model, which described acids as substances that dissociate (separate) in water, increasing the concentration of hydrogen (or hydronium) ions in the solution. For example:

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

In contrast, bases were described as compounds that react with water to produce hydroxide ions. For example:

$$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$$

When an acid and a base are mixed together, the hydrogen ion and hydroxide ion react to produce water and a metal salt:

$$\text{acid} + \text{base} \rightarrow \text{water} + \text{metal salt}$$

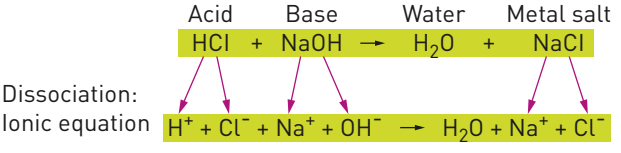


FIGURE 1 In a neutralisation reaction, the hydrogen and hydroxide ions bond to form water.

neutralisation the reaction between an acid and a base to produce water and a metal salt

ionic equation of neutralisation the equation obtained when spectator ions are removed from a neutralisation chemical equation, e.g. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

This process of an acid reacting with a base is called **neutralisation**. In the example in Figure 1, Na^+ and Cl^- are spectator ions that do not take part in the reaction. The spectator ions combined are called a salt (an ionic compound that separates completely in water). If the spectator ions are removed from the equation, neutralisation can be written as an **ionic equation of neutralisation**:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

or

$$H_3O^+(aq) + OH^-(aq) \rightarrow 2H_2O(l)$$

If a strong acid (that easily releases its hydrogen ions) reacts with a strong base (that produces a lot of hydroxide ions), then the availability of hydrogen ions and hydroxide ions decreases as they combine to form water molecules. As a result, the pH of the acid increases, and the pH of the base decreases, and the pH approaches a neutral 7 the indicator, bromothymol blue, which changes colour between pH 6.0 and 7.6 is a good choice here.

If a strong acid reacts with a weak base (with few free hydroxide ions), some free hydrogen ions remain in the solution, resulting in the pH still being less than 7, and the solution stays acidic. Likewise, if a strong base reacts with a weak acid, some free hydroxide ions will

remain, causing the pH to stay greater than 7. If bromothymol blue had been used as an indicator for this reaction, it would have changed colour too late when the acid was in excess. Methyl red would be a better indicator to use this time because it changes colour between pH 4.8 and 6.0.

FIGURE 2 Titrating (dripping) a strong base into an acid until it becomes neutralised (as shown by the indicator) is a way of determining the strength of the acid.

Titrations

One way to determine the concentration of an acid or a base is to use a neutralisation reaction. Titration is the process of slowly neutralising a solution of unknown concentration by dripping (or titrating) an acid or a base of known concentration into it. An indicator is added at the start of titration. The **equivalent point** is the point in the titration when the acid and base are present in equivalent amounts, and the reaction is complete. The point at which the indicator changes colours is the **end point**.

equivalence point the point in a titration at which a reaction is complete

end point the point in a titration at which the indicator changes colour

For an accurate analysis, an indicator should be chosen that changes colour very close to the equivalence point (Figure 3). For example, an unknown concentration of acid will become slowly neutralised by slowly adding a base. The concentration of the unknown acid can then be determined by calculating the number of moles of the known base it took for the solution to reach the pH determined by the indicator.

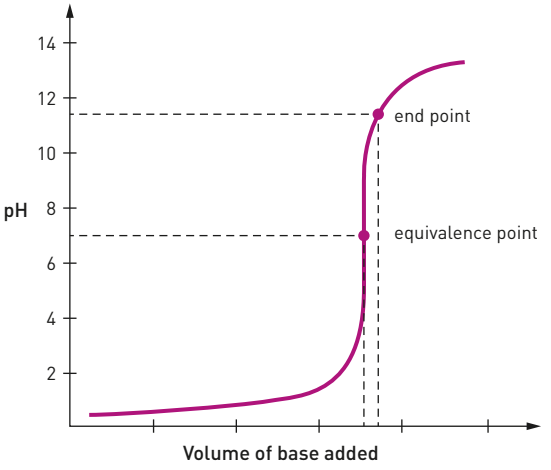


FIGURE 3 A titration curve for a strong acid and a strong base

CHALLENGE 18.1

Acid spills

A common compound used to neutralise strong acid spills in the laboratory is sodium carbonate anhydrous or soda ash (Na_2CO_3). When soda ash reacts with sulfuric acid (H_2SO_4) it produces a salt solution. Explain, using a chemical equation how this occurs.

Study tip

The ability to identify common acids and bases will help you to recognise a neutralisation reaction:
acid + base → water + metal salt

CASE STUDY 18.1

Indigestion and antacids

Neutralisation reactions occur all around us and even inside us. Hydrochloric acid is produced by cells in the stomach in order to kill bacteria that may be in our food. Occasionally, too much acid is produced, forcing its way up into the oesophagus. The oesophagus is not protected by a lining of mucus, and the acid can damage the cells lining the oesophagus. The most common treatment for this acid reflux or indigestion is to take antacid tablets. Most antacid tablets contain a weak base (aluminium hydroxide and sodium bicarbonate) that neutralises the stomach acid, preventing further damage to the oesophagus.

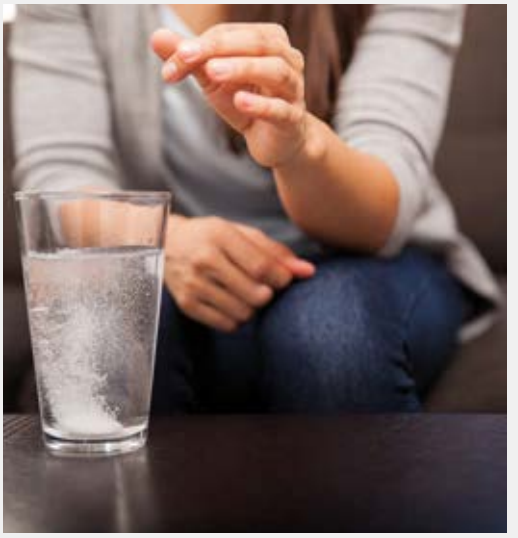


FIGURE 4 An antacid contains a weak base that treats indigestion or acid reflux by neutralising the stomach acid.

CHECK YOUR LEARNING 18.1

Describe and explain

1 **Define**:
   - **a** reactant
   - **b** product
   - **c** neutralisation reaction
   - **d** neutral pH.
2 **Construct** a word equation for a neutralisation reaction.
3 **Explain** how a titration is used to determine the concentration of an acid.
4 **Explain** what titrations are used for when investigating acids and bases.

Apply, analyse and interpret

5 **Construct** the ionic equation for the neutralisation of:
   - **a** hydrochloric acid and potassium hydroxide
   - **b** sulfuric acid and sodium hydroxide.
6 **Deduce** if a solution will become neutralised (pH 7), if equal volumes of a strong acid and a weak base are mixed, will the solution become neutralised (pH).

Investigate, evaluate and communicate

7 **Investigate** an example of neutralisation that occurs in your life. Write a word and ionic equation for this reaction.

Check your obook assess for these additional resources and more:

- » Student book questions
  Check your learning 18.1
- » Challenge
  18.1 Acid spills
- » Weblink
  Neutralisation and breathing
- » Weblink
  Titrations

18.2 Further acid reactions

KEY IDEAS

In this section you will learn about:

- the reaction of acids with most metals to produce a metal salt and hydrogen gas
- how different metals react with acid at different rates
- the pop test, which can be used to determine if the gas produced is hydrogen
- the reaction of acids with carbonates to produce water, metal salt and carbon dioxide
- how a lit splint can be used to test for the presences of carbon dioxide.

Acids and metals

Acids react with metals to produce a metal salt and hydrogen gas:

$$\text{acid} + \text{metal} \rightarrow \text{metal salt} + \text{hydrogen}$$

The type of metal salt produced depends on the metal and acid. For example, hydrochloric acid reacts with magnesium to produce magnesium chloride and hydrogen gas:

$$2HCl(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)$$

In this reaction, the chloride ion is a spectator ion. Therefore, the ionic equation is:

$$2H^+(aq) + Mg(s) \rightarrow Mg^{2+}(aq) + H_2(g)$$

This can be written as a general ionic equation for acids and metals:

$$2H^+(aq) + M(s) \rightarrow M^{2+}(aq) + H_2(g)$$

where M represents the metal.

Pop test

pop test
a test to determine the presence of hydrogen gas

The **pop test** is a test to determine if a gas produced is hydrogen (Figure 1). The hydrogen gas is usually collected in a test tube before being exposed to a lit splint or match. The heat from the flame is enough to initiate a combustion reaction between the collected hydrogen gas and oxygen in the surrounding air, to produce water. The reaction is accompanied by an audible 'pop' sound.

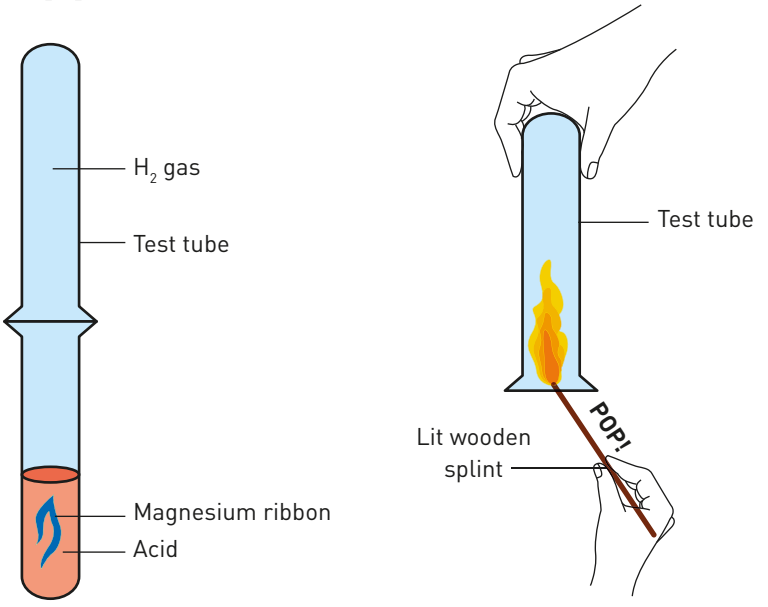


FIGURE 1 The pop test can be used to determine if a gas that is produced in a reaction is hydrogen.

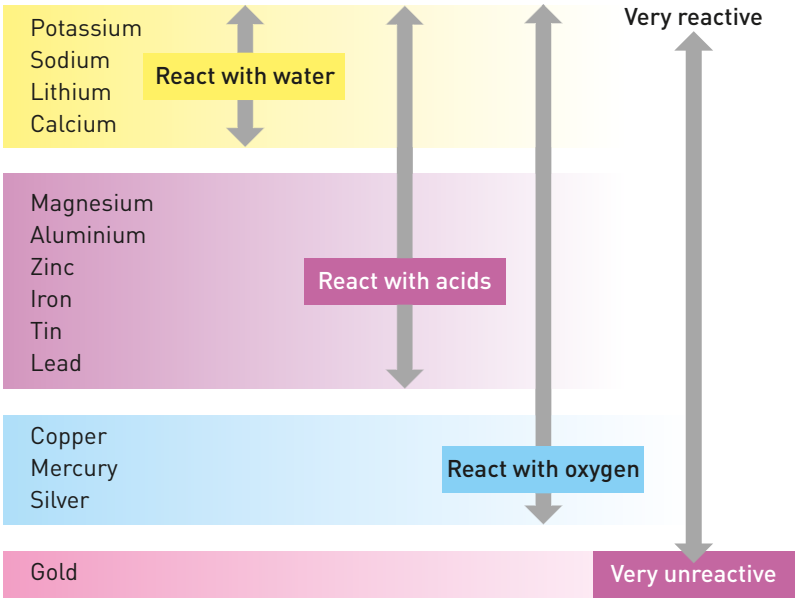


FIGURE 2 Different metals have different levels of reactivity. This is noticeable when the metal is placed in acid.

Reactivity of metals

Chapter 2 discussed the different properties of the elements (or trends) across the periodic table. This is relevant to the way different metals react with acids. Some metals (e.g. magnesium and aluminium) are very reactive with acid (Figure 2). This means the reaction occurs very quickly, with hydrogen gas often seen bubbling off the metal. Other metals (e.g. lead and tin) are much less reactive with acid. The reaction occurs much slower and can take many weeks to occur to any noticeable extent. There are even a few metals that do not react with acid; for example, copper, mercury and silver.

Acids and carbonates

Acids react with carbonates to produce water, a metal salt and carbon dioxide:

$$\text{acid} + \text{carbonate} \rightarrow \text{water} + \text{metal salt} + \text{carbon dioxide}$$

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

The chloride ions and sodium ions are spectator ions in this reaction. If they are removed from the equation, the ionic equation becomes:

$$2H^+ + CO_3^{2-} \rightarrow H_2O + CO_2$$

FIGURE 3 A lit splint requires oxygen to keep burning. Placing a lit splint into a test tube of carbon dioxide makes the flame to go out.

Lit splint test

The **lit splint test** is used to determine if a gas produced is carbon dioxide. The carbon dioxide gas is usually collected in a test tube before being exposed to a lit splint or match (Figure 3). As the flame cannot burn without oxygen being present, when the lit splint is lowered into the carbon dioxide filled test tube, the flame goes out.

lit splint test
a test to determine the presence of carbon dioxide gas

CASE STUDY 18.2

Acid rain

Humans have relied on burning coal and oil for energy for many centuries. Coal often contains sulfur, and combustion of coal can produce sulfur dioxide gas (SO_2). Over time, the gas breaks down in the atmosphere, as it does with the SO_2 gas produced in volcanic eruptions.

The number of combustion reactions has increased because of industrialisation. This means the amount of SO_2 in the atmosphere has increased, allowing the gas time to react with the water in the atmosphere:

$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(g)$$
$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(g)$$

As the gas reacts with the water in the atmosphere it starts to dissolve and the reaction falls to the ground as **acid rain**. Acid rain reacts with exposed metal or marble ($CaCO_3$), as seen in the following reactions:

$$H_2SO_4(l) + Fe(s) \rightarrow H_2(g) + FeSO_4(aq)$$
$$H_2SO_4(l) + CaCO_3(s) \rightarrow CO_2(g) + H_2O(l) + CaSO_4(aq)$$

The iron (II) sulfate and calcium sulfate are soluble in water so the metal and marble objects erode over time.

acid rain
rain that has been made acidic due to atmospheric pollution containing sulfur and nitrogen oxides from burning fossil fuels

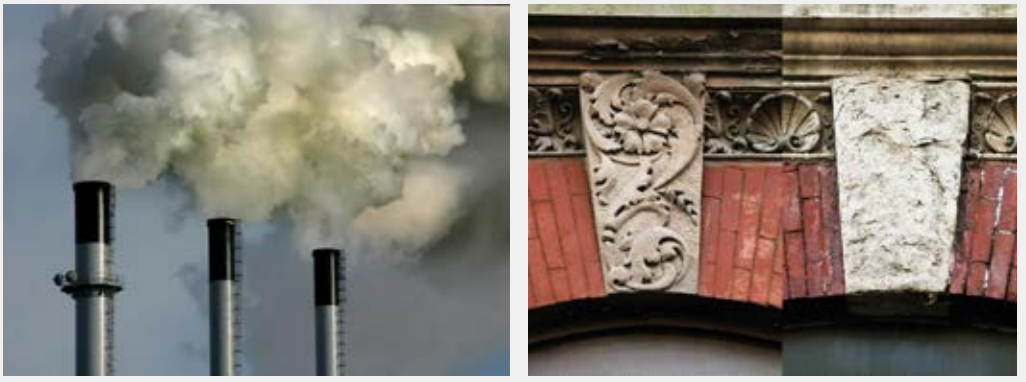


FIGURE 4 The reaction between industrially produced sulfur dioxide (SO_2) and water in the atmosphere produces acid rain. The acid reacts with the calcium carbonate in marble statues, causing them to break down.

CHECK YOUR LEARNING 18.2

Describe and explain

1 **Recall** the general equation (showing predicted products) for the reaction of:
   - **a** acid + base
   - **b** acid + metal
   - **c** acid + carbonate.

2 **Recall** the general ionic equation for each of the general equations shown in question 1.

Apply, analyse and interpret

3 **Apply** the general equations to predict the products of:
   - **a** $H_2SO_4(aq) + Mg(OH)_2(s)$
   - **b** $H_2SO_4(aq) + Mg(s)$
   - **c** $H_2SO_4(aq) + MgCO_3(s)$

4 **Apply** ionic equations for all of the reactions shown in the Case study 18.2.

5 **Contrast** the pop test and the lit splint test.

Investigate, evaluate and communicate

6 The shells of many sea organisms consist of calcium carbonate. **Investigate** the impact of the increasing acidity of the oceans on the survival of these animals. Provide chemical reactions to support your answer.

7 Consider Case study 18.2. **Investigate** acid rain, considering its causes, effects and possible solutions.

Check your obook assess for these additional resources and more:

- » Student book questions
  Check your learning 18.2
- » Suggested practical
  18.2 Reactions of acids with bases, metals and carbonates
- » Weblink
  Pop test
- » Weblink
  Karst cave

CHAPTER 18

Review

Chapter summary

18.1
- Neutralisation reactions involve mixing an acid and a base to produce water and a metal salt.
- A completed neutralisation reaction will have a pH of 7.
- A strong acid may only be partially neutralised by a weak base.
- Titration involves dripping an acid or a base into a substance with an unknown pH until the known indicator changes colour.

18.2
- Acids react with most metals to produce a metal salt and hydrogen gas.
- Different metals react with acid at different rates.
- A pop test can be used to determine if a gas is hydrogen.
- Acid reacts with carbonates to produce water, metal salt and carbon dioxide.
- A lit splint can be used to test for the presence of carbon dioxide.

Key terms

- acid rain
- end point
- equivalence point
- ionic equation of neutralisation
- lit splint test
- neutralisation
- pop test

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; ***=high.

Multiple choice

1 Consider the reaction:
$CH_3CH_2NH_3^+ + NaOH \rightarrow CH_3CH_2NH_2 + H_2O$
Identity which of the molecules is a base.
- **A** $CH_3CH_2NH_3^+$
- **B** NaOH
- **C** $CH_3CH_2NH_2$
- **D** H_2O

2 The end point in an acid–base titration occurs when:
- **A** the pH is 7
- **B** neutralisation is complete
- **C** the indicator changes colour
- **D** equal amounts of the acid and base are combined.

3 Antacids are used to neutralise the acid produced by cells in the stomach. Antacids contain:
- **A** acid
- **B** weak base
- **C** metal
- **D** water.

4 In the reaction between sodium carbonate and hydrochloric acid, what products will be produced?
- **A** Salt and water
- **B** Only water and carbon dioxide
- **C** Salt, water and hydrogen
- **D** Salt, water and carbon dioxide

5 In the equation:
$NaOH(aq) + HCl(aq) \rightarrow H_2O(l) + NaCl(aq)$,
the chloride ion is the:
- **A** acid
- **B** base
- **C** spectator ion
- **D** metal salt.

Short answer

Describe and explain

★ **6** **Define** a neutralisation reaction by providing the reactants and products.

★ **7** **Identify** the products from the reaction of an acid with a metal.

★ **8** **Identify** the products from the reaction of an acid with a carbonate.

★ **9** **Construct** the ionic equation for the reaction between sodium hydroxide and sulfuric acid.

★★ **10** Fertilisers often contain metal salts. Select the acid and base that could be used to produce:

a ammonium nitrate

b potassium phosphate.

Apply, analyse and interpret

★★ **11** **Complete** Table 1 by determining the products of the reaction between reactant 1 and reactant 2.

TABLE 1 Products and reactions

Reactant 1	Reactant 2		
	Hydrochloric acid	Sulfuric acid	Nitric acid
Sodium carbonate			
Iron (II) carbonate			
Copper (II) carbonate			

★★★ **12** **Determine** the products of the following reactions.

a Calcium hydroxide + sulfuric acid

b Lithium hydroxide + hydrochloric acid

c Calcium carbonate + nitric acid

d Potassium hydroxide + sulfuric acid

e Calcium carbonate + nitric acid

Investigate, evaluate and communicate

★★ **13** The pH of a swimming pool should be maintained at 7.5. The water in a pool is found to have a pH of 5, so the pH needs to be raised to an appropriate pH. **Propose** a safe solution to this problem.

★★★ **14** **a** Bee stings and ant bites often contain acidic molecules. **Determine** what you might use to neutralise these molecules remaining on the skin surface.

b **Determine** why the neutralizing chemical used in part (a) be used on the skin's surface rather than being injected close to the injury. Provide reasonsing to support your answer.

★★★ **15** A lit splint is used to determine if a gas is either hydrogen or carbon dioxide. **Justify** why you would you not use a lit splint to test if a gas is oxygen.

★★★ **16** Plants need to absorb a variety of plant nutrients to survive and grow.

a **Evaluate** the data in Figure 1 to determine what nutrients a plant would be able to absorb in a soil of pH 4 to 5.

b **Consider** how a plant's ability to absorb nutrients would be affected if lime (a base) was added to the soil in part (a) so that the pH increased to 8.

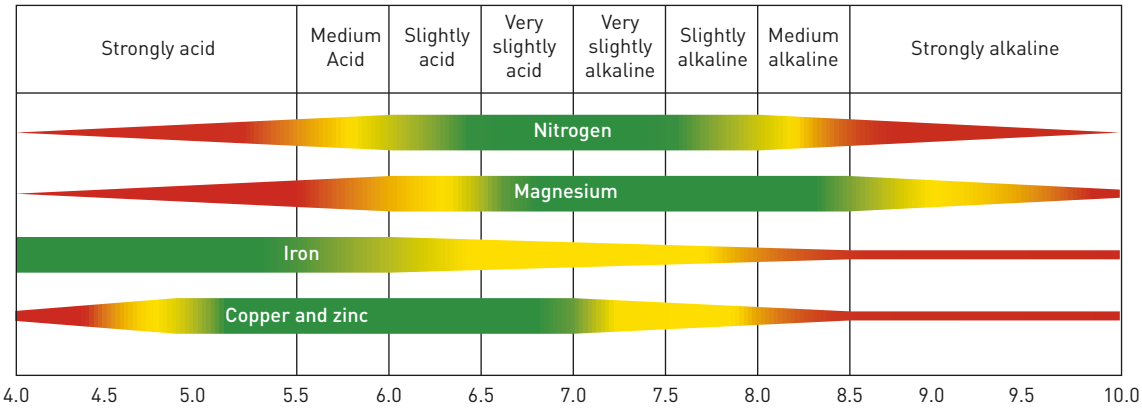


FIGURE 1 Soil pH affects the availability of soil nutrients.

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CHAPTER 19

Rates of chemical reactions

The efficiency of the rate of production of chemicals on an industrial scale is dependent on the rates of chemical reactions. It determines the speed with which chemical products can be formed and the amount of product that can be formed in a given time period. A fast reaction ensures that products are formed within a relatively short period of time.

The food industry also relies heavily on knowledge of rates of reaction. Fruit is stored for long periods of time (some apples are stored for up to two years) before it is sold in supermarkets and consumed. Fruit must be stored in a way that slows the breakdown of biomolecules that lead to rotting, bruising and decay.

Rates of reaction can be explained by kinetic molecular theory, which states that particles are in constant random motion. For a reaction to occur, particles must make contact with each other and have sufficient energy to react. This is the foundation that led to the development of collision theory. This theory was first proposed by German chemist Max Trautz in 1916. He also pioneered the understanding of concepts of activation energy and chemical energy.

British chemist William Lewis proposed a very similar theory two years later. Because of the breakdown in communications during World War I, both men were credited with this discovery.

This chapter investigates rates of reactions, the factors that affect reaction rates, the use of catalysts and the representation of chemical reactions in energy profile diagrams.

FIGURE 1 Chemical reactions occur at different rates depending on temperature, pressure, concentration, surface area and catalysts.

OBJECTIVES

- → Explain how varying the conditions of chemical reactions, including temperature, surface area, pressure (gaseous systems), concentration and the presence of a catalyst, can affect the rate of the reaction.
- → Use the collision theory to explain and predict the effect of concentration, temperature, pressure and surface area on the rate of chemical reactions by considering the structure of the reactants and the energy of particles.
- → Construct and explain Maxwell–Boltzmann distribution curves for reactions with and without catalysts.
- → Recognise that activation energy (E_a) is the minimum energy required for a chemical reaction to occur and is related to the strength and number of the existing chemical bonds; the magnitude of the activation energy influences the rate of a chemical reaction.
- → Sketch and use energy profile diagrams, including the transition state and catalysed and uncatalysed pathways, to represent the enthalpy changes and activation energy associated with a chemical reaction.
- → Explain how catalysts, including enzymes and metal nanoparticles, affect the rate of certain reactions by providing an alternative reaction pathway with a reduced activation energy, hence increasing the proportion of collisions that lead to a chemical change.
- → Use appropriate mathematical representations to calculate the rate of chemical reactions by measuring the rate of formation of products or the depletion of reactants.
- → Analyse experimental data, including constructing and using appropriate graphical representations of relative changes in the concentration, volume and mass against time.

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PRACTICALS

MANDATORY PRACTICAL

19.1 Investigating the rate of chemical reactions

19.1 Factors affecting the rate of a chemical reaction

KEY IDEAS

In this section, you will learn about:

- collision theory and the requirements for a successful reaction
- factors affecting the rate of a reaction
- how to manipulate a reaction to make it faster or slower.

collision theory the theory that states that reactants must collide with the correct orientation and sufficient energy for a reaction to occur

activation energy the minimum amount of energy required in a collision for a reaction to occur

Scientists manipulate chemical reactions by knowing how to speed up or slow down reactions, which is essential in many different areas. Industrial chemists sometimes need to speed up chemical reactions to produce their products to meet deadlines. Food chemists want to slow down the rate of decay of nutrients and molecules within products to extend shelf life within a supermarket. For this reason, an understanding of rates of chemical reactions is essential.

Collision theory

Collision theory states that reactants must collide in order to react. However, not all collisions form products. The success of a collision depends on its orientation and energy.

- Reactants must collide with the correct orientation, or in the correct position. Figure 1 demonstrates molecules of ethene (C_2H_4) and hydrogen chloride (HCl) reacting according to the chemical equation:

$$C_2H_4(g) + HCl(g) \rightarrow C_2H_5Cl(g)$$

  Unless the molecules collide in the correct orientation, the collision is unsuccessful.
- Reactants must collide with a minimum amount of energy (**activation energy**) to rearrange and form products. If C_2H_4 and HCl in Figure 1 do not collide with sufficient energy, even a correct orientation does not result in a successful collision.

Study tip

Collisions can occur without resulting in a chemical reaction, so it is essential to use the term 'successful collision' when explaining whether a chemical reaction results in the formation of a product.

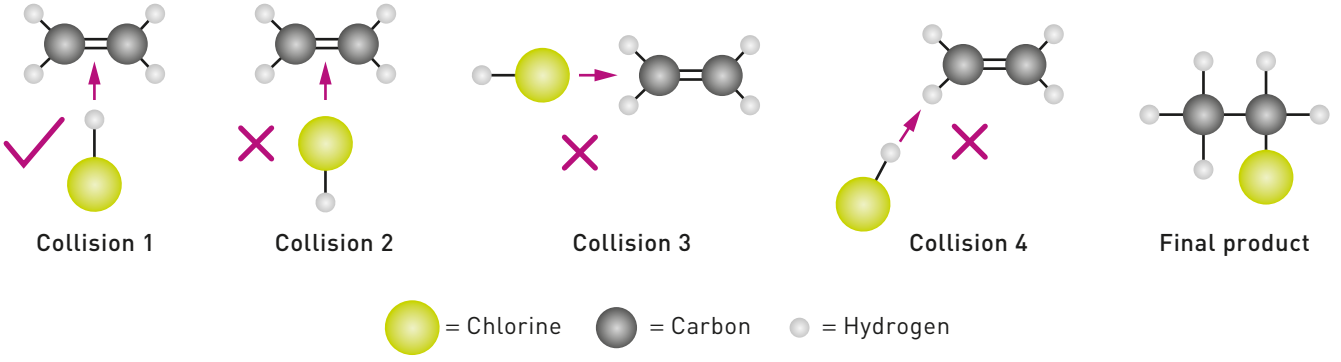


FIGURE 1 The correct and incorrect orientations of ethene (C_2H_4) and hydrogen chloride (HCl) molecules that result in successful and unsuccessful collisions. Collision 2 is not successful because of the repulsion between the slightly negatively charged chlorine atom and the electrons in ethane's double bond.

MANDATORY PRACTICAL
19.1 Investigating the rate of chemical reactions
Go to page 457 »

Rates of reaction

Reactions can take anywhere from a few hours to more than a few hundred thousand years, so it is important that scientists can increase the rate of a chemical reaction (to make it happen faster).

Reaction rates can be increased by:

- adding energy
- manipulating the activation energy
- increasing the number of successful collisions.

This can be accomplished by changing five main factors, depending on the state of the reactants and products. They are:

- **concentration**: liquid, aqueous and gaseous reactants
- **pressure**: gaseous reactants only
- **surface area**: solids reactants only
- **temperature**: solids, liquids, aqueous and gaseous reactants
- **catalysts**: solids, liquids, aqueous and gaseous reactants.

concentration
a measure of the number of particles in a specific volume of solution; units of concentration include mol L^{-1} (M), g L^{-1} and mg L^{-1} (ppm)

pressure
the force exerted, per unit area, by one substance upon another substance

surface area
a measure of the total surface area available to react in a chemical reaction

temperature
a measure of the heat energy of a system

catalyst
a substance that lowers the activation energy of a reaction and increases the rate of reaction without participating in the chemical reaction

Concentration

Concentration is the measure of the number of particles in a given volume. An increase in concentration means an increase in the number of particles in a given volume. If the number of particles increases, there is a higher chance of reactants colliding and therefore a higher chance of successful collisions. If there are more successful collisions, then there will be a faster chemical reaction. Therefore, using concentrated reactants increases the rate of a reaction. Using dilute solutions, which have fewer reactant particles in a given volume, results in fewer collisions (Figure 2) and a slower reaction.

Concentration can be measured as the number of sand molecules in a sample of sand, the number of salt molecules in a sample of saltwater or the number of oxygen molecules in the atmosphere. In a solid, concentration is the number of particles in a larger volume of a solid substance. In aqueous solutions, concentration refers to the amount of solute in a volume of solvent. In gases, concentration refers to the number of gas molecules in the volume of a container.

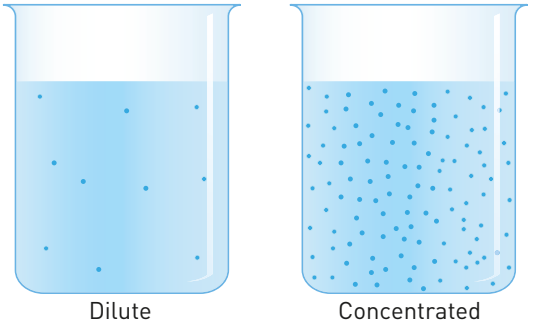


FIGURE 2 A dilute solution has fewer particles than a concentrated solution so results in fewer collisions.

Pressure

Pressure is the force exerted per unit area by one substance on another substance. For gases, this is the pressure exerted by gas particles when they collide with the walls of a container. An increase in the number of particles in a system results in an increase in pressure.

To increase the pressure of a gaseous system, you can:

1 Decrease the volume of the container. This results in more particles per unit of volume and therefore a greater concentration of particles (Figure 3). For example, 1000 oxygen molecules in a 1 L container is less concentrated than 1000 oxygen molecules in a 100 mL container.

Study tip

When explaining the effect of pressure, always make a link to concentration. For example, an increase in pressure is due to an increase in concentration of particles. This will result in more successful collisions and a faster rate of reaction.

2 Increase the number of particles within the container. In a fixed volume system, the volume cannot be decreased. Therefore, to increase the pressure, more gas must be added to the reaction container, increasing the number of particles and the concentration (Figure 3).

Increase number of particles ←

Decrease volume →

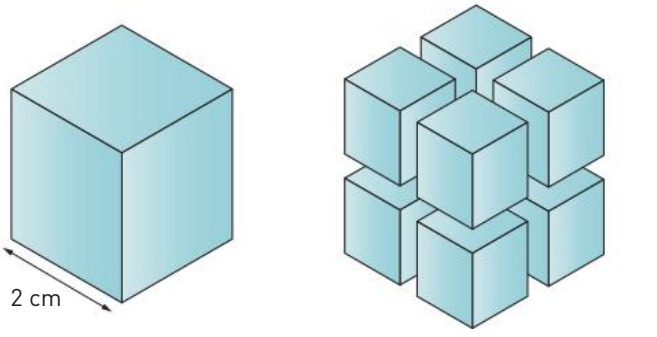
n	Original container with 30 gas molecules	Original container with 15 gas molecules	Half of the original container with 15 gas molecules
V	1 L	1 L	500 mL or 0.5 L
c	$c = \frac{n}{V}$ $= \frac{30}{1}$ = 30 particles per litre	$c = \frac{n}{V}$ $= \frac{15}{1}$ = 15 particles per litre	$c = \frac{n}{V}$ $= \frac{15}{0.5}$ = 30 particles per litre

FIGURE 3 The effect on concentration when pressure is increased by increasing the number of particles or decreasing the volume *n* represents the number of particles, *V* is volume and *C* is the concentration.

In both methods, an increase in pressure is due to an increase in the concentration of particles, increasing the number of successful collisions and resulting in a faster rate of reaction.

Surface area

Reactions occur at the surface of a solid reactant. Therefore, the greater the surface area, the greater the number of collisions. Figure 4 shows a 2 cm × 2 cm cube with a surface area of 24 cm^2. When the cube is cut into eight smaller 1 cm × 1 cm cubes, they have a total surface area of 48 cm^2. Therefore, smaller solids have larger surface areas. This increases the number of successful collisions and increases the rate of reaction.



Total surface area
= 2 cm × 2 cm × 6 sides
= 24 cm^2

Total surface area
= 1 cm × 1 cm × 6 sides × 8 cubes
= 48 cm^2

FIGURE 4 The surface area calculation for a bigger 2 cm × 2 cm cube and smaller 1 cm × 1 cm cubes.

Temperature

kinetic molecular theory
the theory that states that all particles are in constant random motion

kinetic energy
the energy associated with moving particles

The activation energy (E_a) of a reaction is the minimum amount of energy that a collision must have to start a reaction.

Kinetic molecular theory states that all particles are in constant, random motion (Chapter 14). The energy involved with this motion is called **kinetic energy**. Temperature is a measure of heat energy. Therefore, when a reaction is heated, energy is transferred to the reactants, making them move faster. An increase in temperature results in an increase in energy.

If the temperature of a reaction increases, more particles have energy greater than the activation energy ($E > E_a$). Therefore, there are more successful collisions and an increased rate of reaction.

Maxwell–Boltzmann representation of the temperature of a reaction

Some molecules in a system have more energy than others. The Maxwell–Boltzmann distribution (Figure 5), discussed in Chapter 14, represents the number of particles in a system and the energy that each has. The distribution is usually plotted with a line representing the activation energy (E_a) for the reaction.

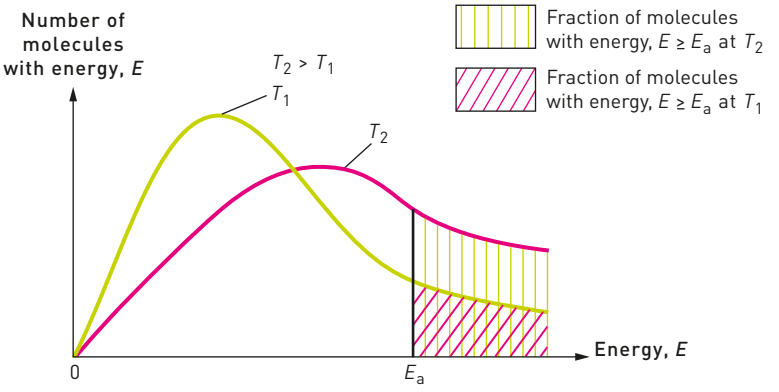


FIGURE 5 The Maxwell–Boltzmann distribution demonstrates the energy level of particles in a system. T_1 is a lower temperature than T_2.

Several key pieces of information can be gained from this distribution.

- Each particle in a reaction has a different amount of energy.
- The peak of the distribution represents the average kinetic energy; therefore, the higher the peak, the higher the temperature of the system.
- At higher temperatures, there are more particles with higher energy levels.
- At higher temperatures, there are more particles with energy greater than the activation energy ($E > E_a$).

In Figure 5, the distribution clearly demonstrates that at the higher temperature (T_2), there are more particles with energy greater than the activation energy ($E > E_a$) than there are at the lower temperature T_1. This results in more successful collisions and a faster reaction.

Measuring the rate of a chemical reaction

The rate of a chemical reaction is measured as a change over time. This change can be the loss in mass of a solid reactant, the volume of gas generated, the formation of a solid precipitate or a change in colour.

Mathematically, the rate of a chemical reaction can be measured in two ways. The first is to create a graph of the change in a reaction over time. Figure 6 represents the mass of a reactant lost over time. The graph demonstrates that the rate of the chemical reaction starts faster, and gradually slows over time.

The second measurement involves using a mathematical formula to calculate the rate of reaction at a certain time in the experiment or to calculate the overall (average) rate of reaction at the end:

$$\text{Rate} = \frac{\text{change in mass of reactant}}{\text{time}}$$

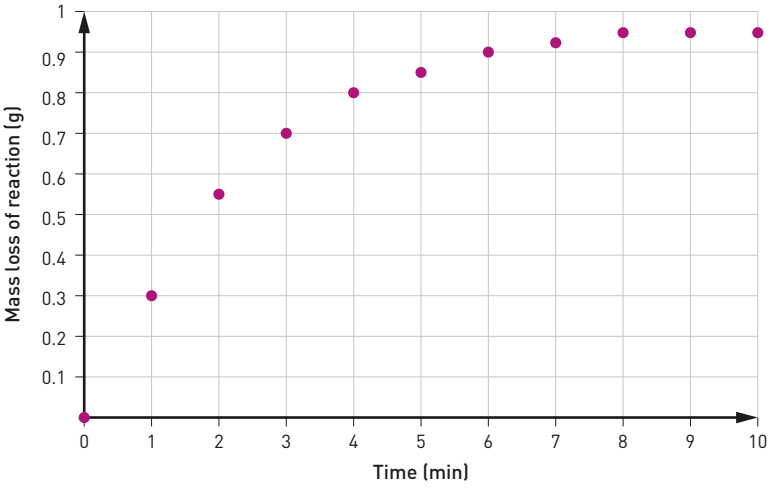


FIGURE 6 The mass of a reactant lost over time.

WORKED EXAMPLE 19.1A

Calculate the rate of reaction for the reaction shown in Figure 6:

a at 1 minute **b** at 4 minutes **c** for the overall reaction.

SOLUTION

a $\text{Rate} = \frac{\text{mass loss}}{\text{time}}$

$= \frac{0.3\text{ g}}{1\text{ min}}$

$= 0.3\text{ g min}^{-1}$

b $\text{Rate} = \frac{\text{mass loss}}{\text{time}}$

$= \frac{0.8\text{ g}}{4\text{ min}}$

$= 0.2\text{ g min}^{-1}$

c $\text{Rate} = \frac{\text{mass loss}}{\text{time}}$

$= \frac{0.95\text{ g}}{10\text{ min}}$

$= 0.095\text{ g min}^{-1}$

If the rate of the reaction was only measured as the final loss in mass over the total time, the rate of the reaction would look very different because it is an average rate calculation. By monitoring the mass loss at various time intervals throughout the experiment, a more detailed picture of the rate of reaction can be observed, particularly how the rate changes over the course of the reaction.

Mole calculations and rates of reaction

A theoretical end point for a reaction can be calculated using stoichiometry. For a reaction between solid magnesium metal and hydrochloric acid, magnesium chloride and hydrogen gas are produced according to the equation:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

If the amount (mole) of reactant is known, the amount of product can be calculated.

WORKED EXAMPLE 19.1B

A student reacts 2.00 g of magnesium ribbon with an excess of hydrochloric acid at SLC. Determine the theoretical volume of hydrogen gas that could be generated.

SOLUTION

1 Write a balanced chemical equation, including states.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

2 Calculate the moles of magnesium in 2.00 g.

$n(Mg) = \frac{m}{M}$

$= \frac{2.00\text{ g}}{24.3\text{ g mol}^{-1}}$

$= 0.0823\text{ mol}$

3 Use the mole ratio to calculate the moles of hydrogen gas.

$n(H_2) = n(Mg)$

$= 0.0823\text{ mol}$

4 Calculate the volume of hydrogen gas at SLC, where $V_m = 24.5\text{ L mol}^{-1}$ (see databook and Chapter 14).

$V(H_2) = n \times V_m$

$= 0.0823\text{ mol} \times 24.5\text{ L mol}^{-1}$

$= 2.02\text{ L}$

Theoretically, the reaction will generate 2.02 L of hydrogen gas before it stops.

A key part of preparation for an experiment is calculating the theoretical yield, so this can be compared with the experimental value.

Mass lost over time

Mass can be lost from a solid reactant in a chemical reaction. For example, in the reaction between solid zinc and hydrochloric acid, mass is lost when hydrogen gas is formed as the gas can escape from an open contained into the air:

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

First, the mass of hydrochloric acid and the mass of the zinc are measured on an electronic balance (Figure 7). Hydrochloric acid is then poured onto the zinc and the initial mass is recorded. The subsequent loss of mass, which occurs as hydrogen gas escapes, is measured in time intervals until the reaction reaches a constant mass.

A graph of mass (on the *y*-axis) against time (on the *x*-axis) is then plotted (Figure 8). The data in the graph should never reach zero because zinc chloride will remain in the flask when the reaction is complete.

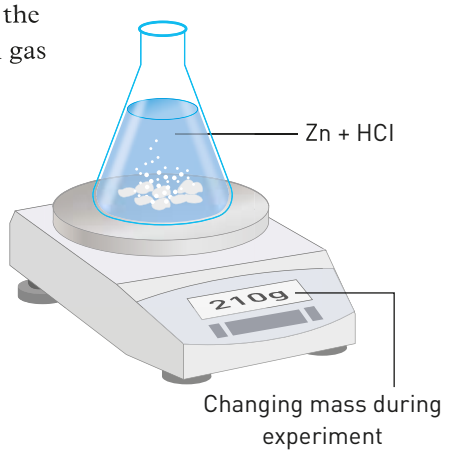


FIGURE 7 The process of measuring mass lost on an electronic balance.

Volume of gas formed over time

In this method, the gas generated is used to displace a known volume of water. Figure 9 demonstrates a gas being generated in a conical flask. The gas moves through the connecting tube into the trough of water and into a gas jar, which has graduated measurements marked on its side. The water is displaced as the gas enters the jar and the volume of gas generated can be measure.

A graph of volume (on the *y*-axis) against time (on the *x*-axis) is then plotted (Figure 10). The volume of gas is measured using the graduated markings on the beaker that collects the gas.

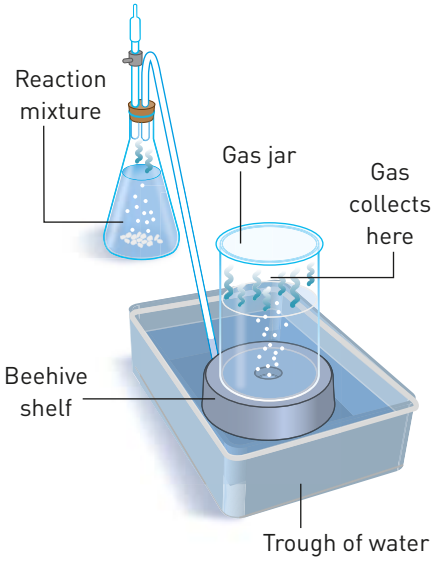


FIGURE 9 The process of measuring gas volume by water displacement.

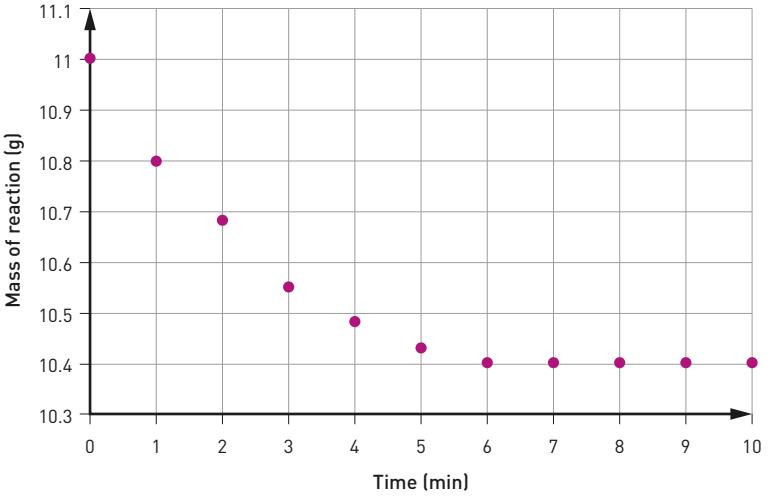


FIGURE 8 Graph of mass lost over time.

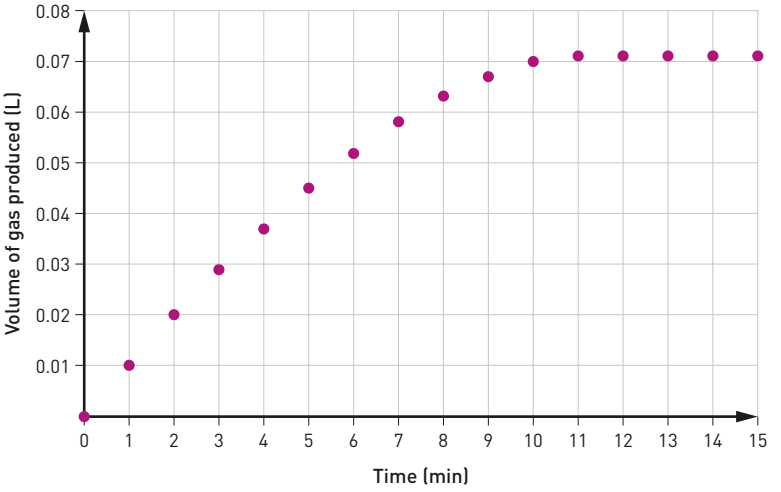


FIGURE 10 Graph of volume of gas produced over time.

CASE STUDY 19.1

The problem with dust – the sugar refinery explosion

At 7.00 pm on 7 February 2008, an explosion rang out across the town of Port Wentworth, Georgia, USA. The blast originated in the centre of the Georgia Sugar Refinery, located on the banks of the Savannah River.

The refinery was, built in 1916 and its construction features were outdated. Many staff members said the machinery was antiquated, some of it being close to 30 years old.

The explosion took place in the centre of the refinery, where the sugar was stored before being bagged and packaged. The storage of the sugar led to a build-up of sugar dust within the air. The building was equipped with ventilation fans that extracted dust particles from the air, but these were found to contain enough dust to cause further explosions.

Refined sugar is sucrose, its molecular formula is $C_{12}H_{22}O_{11}$. Sucrose undergoes combustion according to the chemical equation:

$$C_{12}H_{22}O_{11}(s) + 12O_2(g) \rightarrow 12CO_2(g) + 11H_2O(g) \qquad \Delta H = -5645 \text{ kJ mol}^{-1}$$

Sugar is naturally highly combustible; however, as a fine powder or dust it can spontaneously combust. The ignition source was probably static electricity or a spark; however, spontaneous combustion was not ruled out.

The fire took 7 days to extinguish because of the age of the building and the quantity of sugar that continued to act as a fuel source. The task of putting out the fire was made worse because the sugar that did not combust melted at temperatures above 186°C.

The explosion killed 14 people and severely injured 36.

FIGURE 11 The aftermath of the Georgia Sugar Refinery explosion.

CHECK YOUR LEARNING 19.1

Describe and explain

1 **Identify** and **describe** the two requirements of collision theory.

2 **Describe** the factors that increase the rate of a chemical reaction.

Apply, analyse and interpret

3 Some students are conducting an experiment by reacting 5.00 g of calcium carbonate ($CaCO_3$) chips with 50.00 mL of 2.00 M hydrochloric acid (HCl(aq)). The reaction produced carbon dioxide gas, water and calcium chloride solution according to the chemical equation:

$CaCO_3(s) + 2HCl(aq) \rightarrow CO_2(g) + H_2O(l) + CaCl_2(aq)$

a Assuming that the hydrochloric acid is in excess, **calculate** the volume of carbon dioxide gas produced at SLC.

b The volume of gas generated was measured and recorded in Table 1. **Sketch** a graph of the rate of reaction.

TABLE 1 Results data of volume of gas produced from a reaction between calcium carbonate and hydrochloric acid

Time (min)	Volume of gas (L) produced
0.00	0
0.30	0.2
1.00	0.35
1.30	0.45
2.00	0.5
2.30	0.525
3.00	0.53
3.30	0.53
4.00	0.54
4.30	0.54
5.00	0.54

c **Calculate** the overall rate of reaction.

d **Compare** the theoretical and experimental results.

The students wish to repeat the experiment to increase the rate of reaction.

e **Identify** three things that they could incorporate into their methodology and how each would increase the rate of the reaction.

f **Identify** factors that would have no affect on the rate of this reaction.

4 Iron rusts according to the equation:

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

This reaction occurs in the presence of water.

A car manufacturer has found that their car parts are rusting before the painting and assembly processes. **Determine** what the manufacturer should do to slow down the rust or avoid it entirely.

Investigate, evaluate and communicate

5 In a reaction between sodium thiosulfate ($Na_2S_2O_3(aq)$) solution and hydrochloric acid (HCl(aq)), a yellow precipitate forms according to the equation reaction:

$Na_2S_2O_3(aq) + 2\,HCl(aq) \rightarrow S(s) + SO_2(g) + 2NaCl + H_2O(l)$

Develop a method for measuring the rate of the reaction.

6 **Discuss** the benefits of using granulated sugar instead of sugar cubes to sweeten tea and coffee.

FIGURE 12 Sugar cubes

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  19.1 Investigating the rate of chemical reactions
- » Weblink
  Collision theory
- » Weblink
  Rates of reaction

19.2 Energy profile diagrams

KEY IDEAS

In this section, you will learn about:

- sketching energy profile diagrams, including labelling transition state, E_a and ΔH
- the relationship between activation energy, bond strength and bond number
- the relationship between activation energy and rates of chemical reaction.

The law of conservation of energy states that energy cannot be created or destroyed. Chemical energy is potential energy (Chapter 9). This energy can be most readily converted into heat, observed as an increase in temperature, or motion, as kinetic energy. Energy differences between reactants and products result in either an **exothermic** reaction (where the products have less enthalpy than the reactants) or an **endothermic** reaction (where the products have more enthalpy than the reactants).

exothermic describes a reaction in which the reactants have more energy than the products; excess energy is released to the environment as heat

endothermic describes a reaction in which the reactants have less energy than products; the required energy is absorbed from the environment as heat

Energy profile diagrams

Energy profile diagrams represent the energy changes that occur in chemical reactions. These diagrams show the difference in energy within the bonds of reactants and products. It is similar to an **enthalpy** level diagram, discussed in Chapter 8.

In an energy profile diagram (Figure 1), energy is plotted on the *y*-axis and time, or reaction progress, is plotted on the *x*-axis. If the reactants contain more energy within their bonds than the products, energy is released as heat and the reaction is exothermic. If the reactants contain less energy than the products, energy must be absorbed from the environment and the reaction is endothermic.

energy profile diagrams a theoretical representation of the energy pathway that reactants must follow in order to form products

enthalpy energy stored within the chemical bonds of molecules, chemical energy or heat content

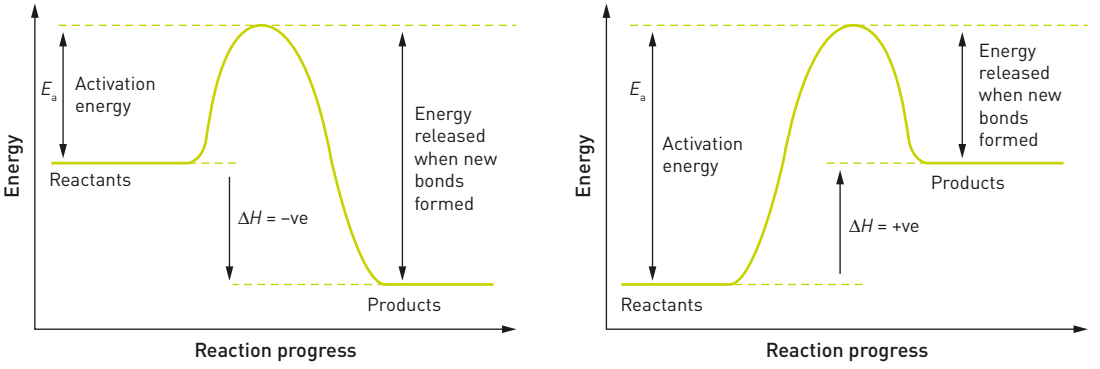

FIGURE 1 Energy profile diagrams for (a) exothermic and (b) endothermic reactions, including E_a (activation energy, energy required to break bonds) and ΔH (enthalpy).

Remember, from Chapter 8, the change in enthalpy is measured as:

$$\Delta H = H_{products} - H_{reactants}$$

An energy profile diagram includes a number of key features.

1 The left side of the curve or 'hill' is the energy that must be absorbed to break reactant bonds.
2 The peak of the curve is the **transition state** where energy is at its greatest and the arrangement of atoms is the most unstable. At this point, bonds are both breaking and forming into products.
3 The right side of the curve is the energy released when product bonds are formed.
4 If reactants have more energy than products, their energy is released into the environment as heat. Therefore, the change in enthalpy between reactants and products is negative ($-\Delta H$).
5 If products have more energy than reactants energy is absorbed from the environment as heat. Therefore, the change in enthalpy between reactants and products is positive ($+\Delta H$).

transition state the state that corresponds to the unstable arrangement of atoms, where bonds are breaking and re-forming, when they reach the highest energy level (activation energy) of a reaction pathway

Magnitude of activation energy

Reaction rates depend on the size of the activation energy. If the energy contained within the bonds of the reactants is relatively small (i.e. the bonds are weak), a smaller activation energy is required. Conversely, strong bonds contain larger amounts of energy, are harder to break apart and require a larger activation energy.

Bond energy, or bond strength, is the amount of energy required to separate atoms that are bonded together. The higher the bond energy, the more energy required to break the bond. Therefore, bond energy (strength) and the number of bonds also affect the activation energy and how readily a reaction occurs.

bond energy the amount of energy required to break a chemical bond in the gas phase, measure in kJ mol^{-1}

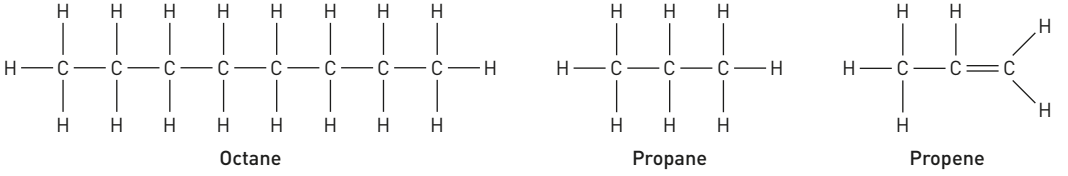

FIGURE 2 Structural formulas of octane, propane and propene.

Consider octane and propane (Figure 2). The C–C bonds in both molecules take around 348 kJ mol^{-1} to break (in the gas phase), and the C–H bonds are also of similar energy (413 kJ mol^{-1}). We would therefore expect octane and propane to react at similar rates for reactions that involved breaking C–C or C–H bonds.

Considering the difference between propane and propene, the double bond in propene has two components. For an addition reaction (where new bonds are formed and the double bond between C atoms is replaced with a single bond), only one of these components needs to be broken, requiring about 264 kJ mol^{-1} of energy, so we would predict that an addition reaction with propene would be faster than a reaction with propane that requires breaking a C–C single bond, which requires more energy.

CHALLENGE 19.2A

Inputs and outputs

Explain how energy inputs and outputs differ in endothermic and exothermic reactions. Specifically refer to activation energy, ΔH, energy required to break bonds, and energy evolved.

CASE STUDY 19.2

Cold and hot packs

It is possible to buy small plastic pouches that contain chemicals that can react and consume or produce heat. These heat and cold packs can be used for a variety of purposes such as warming hands and cooling injuries. The reactants are kept apart within the packaging and are reacted when needed.

In heat packs, two main chemical reactions can be used. These are the reaction between iron and oxygen and the precipitation reaction involving sodium acetate:

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

$$CH_3COONa(s) + 3H_2O(l) \rightarrow CH_3COONa.3H_2O(aq)$$

These exothermic reactions last for a short time before the energy is converted into another form.

Cold packs do not need to be kept in a freezer before use. Cold packs rely on endothermic reactions, which absorb heat from their surroundings and therefore feel cool to the touch. The main reaction used is the dissociation of ammonium nitrate in water:

$$NH_4NO_3(s) \rightarrow NH_4^+(aq) + NO_3^-(aq)$$

FIGURE 3 Heat packs work when two exothermic reactions produce heat.

CHALLENGE 19.2B

Complex energy profile diagrams

Explain why energy diagrams, such as the one shown below, could have a two-step profile?

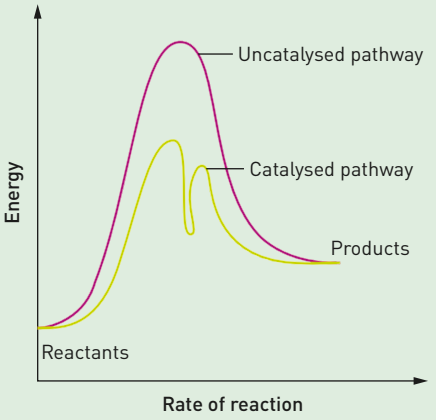

FIGURE 4 Complex energy diagram

CHECK YOUR LEARNING 19.2

Describe and explain

1 **Describe** activation energy. How does it affect the rate of a chemical reaction?

2 **Define** 'transition state'.

3 **Construct** an energy profile diagram for an endothermic reaction. Label both axes, reactants, products, E_a, ΔH, energy absorbed and energy released.

4 **Construct** an energy profile diagram for an exothermic reaction. Label both axes, reactants, products, E_a, ΔH, energy absorbed and energy released.

Apply, analyse and interpret

5 **Analyse** the incomplete energy profile diagram in Figure 5 for the reaction $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$. Methane and oxygen have an enthalpy of 877.1 kJ mol^{-1}. Carbon dioxide and water have an enthalpy of 74.8 kJ mol^{-1}. The activation energy for this reaction is 2632 kJ mol^{-1}. Use this information to complete the energy profile diagram for the reaction.

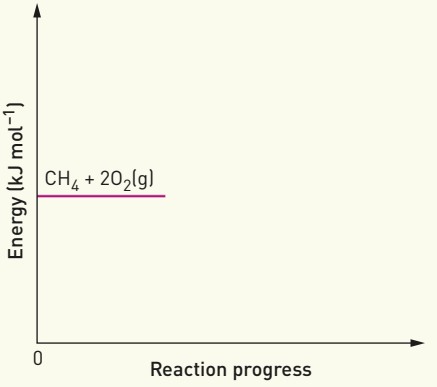

FIGURE 5 Energy profile diagram

The activation energy of the catalysed reaction is 500 kJ mol^{-1}. **Sketch** the catalytic pathway on the energy profile diagram.

Determine the E_a and ΔH for the reverse reaction? Is the reaction endothermic or exothermic?

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  19.2A Inputs and outputs
- » Challenge
  19.2B complex energy profile diagrams
- » Weblink
  Endothermic and exothermic reactions

19.3 Catalysts

KEY IDEAS

In this section, you will learn about:

- the effect of a catalyst on a reaction system
- representing a catalyst on a Maxwell–Boltzmann distribution
- representing a catalyst on an energy profile diagram
- enzymes and metal nanoparticles as catalysts.

catalyst
a substance that lowers the activation energy of a reaction and increases the rate of reaction without participating in the chemical reaction

homogeneous catalyst
a catalyst that is in the same state as the reactants

heterogeneous catalyst
a catalyst that is in a different state from the reactants

Catalysts increase the rate of a chemical reaction without participating in the reaction. They do this by lowering the activation energy of a reaction by reducing the energy required for a successful collision. Catalysts are essential in the human body, medicines, pharmaceuticals, the chemical industry and many other science-related fields. Catalysts include enzymes and metal nanoparticles.

There are two types of catalyst.

- **Homogeneous catalysts** are in the same state as the reactants. For example, in the dissociation of aqueous hydrogen peroxide ($H_2O_2(aq)$), an aqueous potassium iodide catalyst is used.
- **Heterogeneous catalysts** are in a different state from the reactants. For example, in the reaction between hydrogen and nitrogen gases to produce ammonia, a solid iron catalyst is used.

How catalysts work

Catalysts affect the rate of reactions by providing an alternative reaction pathway with reduced activation energy. The mechanisms by which catalysts lower activation energy are of great interest to scientists. For example, scientists know that heterogenous catalysts have high surface energy. When reactants strike the surface of the catalyst, they bond to it.

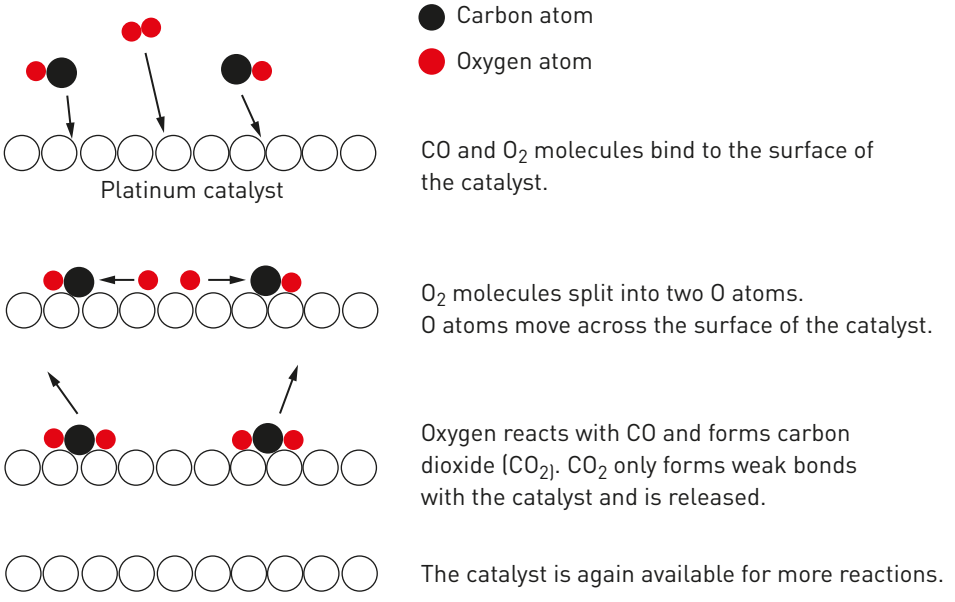

FIGURE 1 The action of a catalyst in lowering the energy required for reaction of carbon monoxide and oxygen. This is one of the reactions that take place inside catalytic converters in vehicles to reduce pollution.

The formation of these bonds weakens or breaks the bonds contained within the reactants. The atoms then rearrange on the catalyst's surface to form products (Figure 1).

Effect of catalyst on Maxwell–Boltzmann distribution

The catalysed reaction can be represented on a Maxwell-Boltzmann distribution as a line, similar to the one that represents activation energy. As a catalyst lowers the activation energy of a system, the number of particles with energy greater than the activation energy ($E > E_a$) increases (Figure 2).

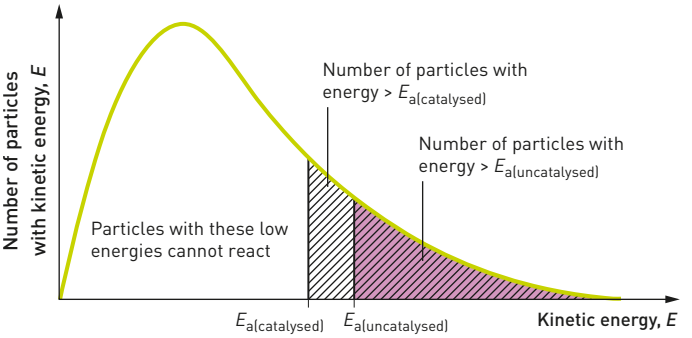

FIGURE 2 The Maxwell–Boltzmann distribution demonstrating activation energy (E_a) and the lowered E_a with a catalyst.

Effect of catalyst on energy diagram

Because a catalyst lowers the E_a of reaction, it can also be represented on an energy profile diagram. In Figure 3, the catalysed chemical pathway is represented as a broken line below the uncatalysed chemical pathway. The activation energy of the catalysed pathway (E_c) is lower than E_a, but the enthalpy of the reaction (ΔH) never changes.

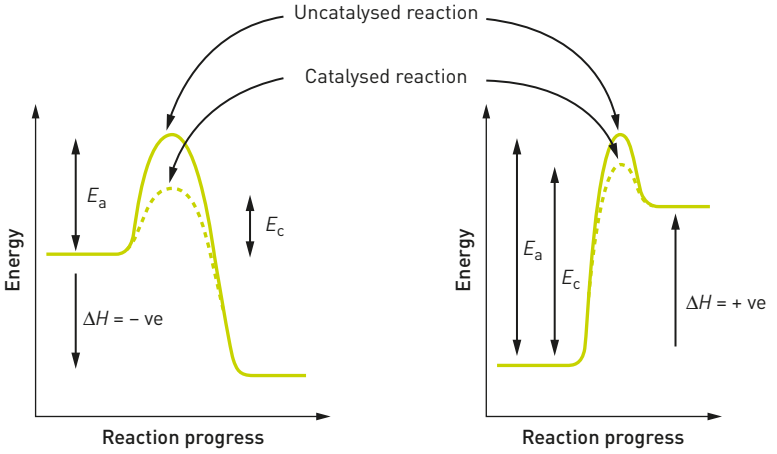

FIGURE 3 Energy profile diagrams of exothermic and endothermic reactions, demonstrating the regular reaction pathway and the catalysed reaction pathway with lowered activation energy. E_a = activation energy of uncatalysed reaction, E_c = activation energy of catalysed reaction.

Metal nanoparticles as catalysts

Metal nanoparticles act as heterogeneous catalysts. The catalyst is a solid, and the reactants are usually gases or liquids. The catalysts are easily recyclable as they do not need replacing unless impurities have reacted with the catalyst.

The metal nanoparticles are in the shape of very small beads or dispersed on a ceramic support. This maximises the surface area of the catalyst, increasing the number of reactant particles that can bind to it and form products and increasing the rate of chemical reaction.

The metal nanoparticles can also react with other chemicals to change the properties of the nanoparticles and therefore increase the number of types of chemical reactions.

These catalysts have a wide range of uses in the reaction of organic molecules to form complex molecules that are used as medicines.

Metal nanoparticles are also used as catalysts in the Fischer–Tropsch process in which carbon monoxide and hydrogen are converted to liquid hydrocarbons. Because carbon monoxide is detrimental to the environment as an atmospheric gas and hydrocarbons are used as fuels to provide energy to everyday processes, the presence of the catalyst quickly converts harmful gases into useful fuels.

Study tip

Remember that catalysts lower E_a and therefore provide an alternative reaction pathway. They do not lower the energy of reactants.

CASE STUDY 19.3

Fritz Haber and the Haber process

The Haber process was developed by German chemist Fritz Haber in the early 1900s. involved the reaction of hydrogen and nitrogen to form ammonia, according to the chemical equation:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H = -91 \text{ kJ mol}^{-1}$$

This process has been highly beneficial to society over the last century as the product, ammonia, is extensively used in fertilisers and has enabled the production of crops on a mass scale, and minimised the loss of nutrients from soil. It has been estimated that almost half the world's population depends on the nitrogen from artificial fertilisers for food.

Air, which is approximately 80% nitrogen, is the source of the nitrogen for this reaction. Hydrogen is sourced from the cheapest and most readily available source. Some methods for the formation of hydrogen are:

- reaction of hydrocarbons with water, in a process called steam reforming

$$CH_4(g) + H_2O(g) \rightarrow 3H_2(g) + CO(g)$$

- reaction of a metal with an acid

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

- electrolysis of water

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$

Steam reforming is the most common process for the production of hydrogen gas. This is because methane is sourced from natural gas and three molecules of hydrogen are formed for every one molecule of methane and water reacted, making it a highly effective and efficient process.

The reaction between hydrogen and nitrogen does not occur readily so a metal catalyst is used. Usually this is an iron-based catalyst, Fe_2O_3.

Haber also helped to develop chemical weapons, which were used in World War I. He led the development of chlorine gas and various other deadly gases for use in trench warfare.

Haber's wife, Clara Immerwahr, was a pacifist and a highly accomplished chemist. The first woman to achieve a PhD in chemistry in Germany, she assisted in the translation of his work and spent her life frustrated that a woman's role was limited to the home. The role of women in research was not, at the time, an accepted practice and therefore her ability to conduct research was limited. Not long after Haber was set to stage his first gas attack with the weaponry he had developed, Clara shot herself using his military pistol and died in her son's arms on 2 May 1915.

Haber was awarded the Nobel Prize in Chemistry for his work on ammonia synthesis in 1918. However, it was a controversial decision and resulted in many protests. The Nobel Committee considered that although Haber had worked on chemical weapons, his work on ammonia in agricultural and other processes had 'conferred the greatest benefit to mankind'.

CHALLENGE 19.3

The Haber-Bosch process

The Haber process is an industrial procedure used to form ammonia. It involves the reaction of gaseous hydrogen (H_2) and nitrogen (N_2) to form gaseous ammonia (NH_3). What could be done to increase the rate of this reaction? What safety precautions would need to be considered in this process?

FIGURE 4 The Haber-Bosch process was invented by Fritz Haber (pictured) and Carl Bosch.

CHECK YOUR LEARNING 19.3

Describe and explain

1 **Explain** the effect a catalyst has on a chemical reaction.

2 **Construct** a Maxwell–Boltzmann diagram representing particles in a reaction at 30°C and 60°C. Mark the activation energy and the activation energy of the catalysed reaction.

Apply, analyse and interpret

3 **Contrast** a homogeneous and a heterogeneous catalyst.

4 The dissociation of hydrogen peroxide (H_2O_2), using a potassium iodide catalyst, is an exothermic process that forms water and oxygen gas. **Apply** your understanding to **create** a fully labelled energy profile diagram of the dissociation of H_2O_2, including both the catalysed and uncatalysed pathways.

Investigate, evaluate and communicate

5 **Investigate** the production of ethanol by the fermentation of sugar and by the reaction of water with ethene.

- **a** Write a balanced chemical equation for each chemical reaction.
- **b** Identify the catalyst involved in each chemical reaction.
- **c** Comment on the environmental impact of each reaction.

6 **Investigate** metal nanoparticles as catalysts and provide a summary of two examples of their use.

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  19.3 The Haber process
- » Increase your knowledge
  Homogenous and heterogeneous catalysts
- » Weblink
  Biological catalysts

SCIENCE AS A HUMAN ENDEAVOUR

19.4 Enzymes as catalysts

KEY IDEAS

In this section, you will learn about:

- enzymes as biological catalysts
- the lock and key and induced fit models of enzyme activity.

enzyme
a type of protein that catalyses or accelerates biochemical reactions

Enzymes are biological molecules, mostly proteins, that are produced by organisms. Enzymes catalyse biochemical reactions by lowering the activation energy required for a chemical reaction to occur by stabilising the transition state.

Each enzyme generally catalyses one specific chemical reaction. For example, amylase is an enzyme found in saliva, which catalyses the breakdown of starch into glucose molecules before they are sent to the stomach for further breakdown.

Biological enzymes work best in a certain temperature and pH range. For example, human enzymes tend to work best at body temperature (37°C). Pepsin, an enzyme that catalyses the breakdown of proteins in the stomach, operates best at low pH such as the pH of 1.5 of stomach acid.

How enzymes work

Enzymes act by interacting with a reactant called a **substrate**. The substrate enters the active site of the enzyme and forms intermolecular bonds with it (Figure 1). The formation of these intermolecular interactions stabilises the transition state for the reaction, lowering the activation energy for the reaction andaccelerating the rate of reaction. Once the reaction is complete, the product is released from the enzyme to perform a biological function, leaving the enzyme unchanged.

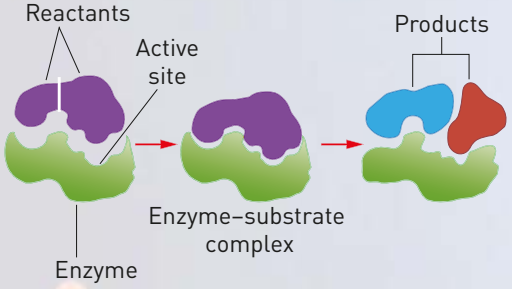


FIGURE 1 The enzyme-catalysed breakdown of a biological substrate occurs at the active site of the enzyme.

Many enzymes catalyse the joining together of two or more substrate molecules through the formation of bonds. The way that the substrates interact with the enzyme ensures that the substrate molecules are aligned with one another in the right orientation to react correctly (Figure 2). The activation energy of the reaction is lowered, allowing the substrate molecules to bond together. The new product molecule is released, leaving the enzyme unchanged.

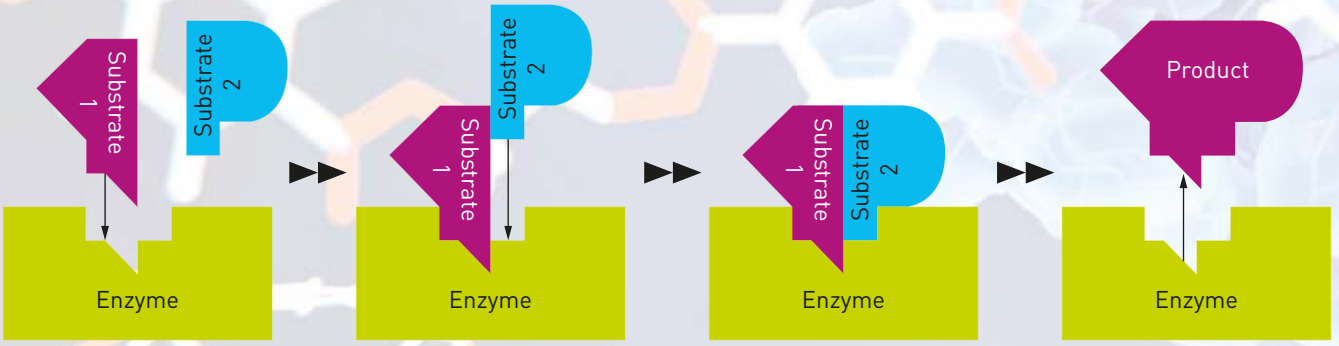


FIGURE 2 The enzyme-catalysed bonding of two biological substrates occurs at the active site of the enzyme.

Two models of enzyme action

The lock and key and induced-fit models explain how the active site of an enzyme interacts with a substrate molecule and is able to catalyse a reaction. It is difficult to say that one model occurs rather than the other, because there is no conclusive research either way. However, evidence seems to support the induced-fit model over the lock and key model.

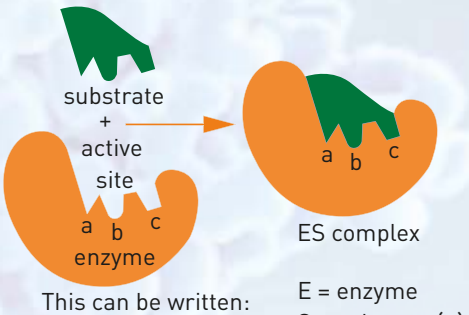
Lock and key model

The lock and key model suggests that the enzyme acts as a 'lock' and the substrate as a 'key'. One key must fit exactly into the lock before the substrate will break down (Figure 3) or react with another substrate molecule. If the key is too small, too large, has too many teeth or not enough teeth or is incorrectly shaped, it will not fit into the lock and therefore the enzyme will have no effect on the reaction of the substrate. In this model, the enzyme is rigid and unchangeable, and the key must be in the exact correct position to fit into the lock.

Induced-fit model

In the induced-fit model, the active site of the enzyme will detect the presence of the substrate when it aligns to it. The enzyme then curls around the substrate to better align it to the active site. This has been compared to a hand and a glove. The enzyme is not rigid, and its shape is not fixed. The enzyme is either entirely or partially flexible.

(a) The 'lock and key' model for enzyme action substrate



This can be written:
$E + S \rightleftharpoons ES \longrightarrow E + P$
complex

E = enzyme
S = substrate(s)
P = product(s)

(b) The 'induced fit' model for enzyme action; the shape of the enzyme's active site is modified to accommodate the substrate.

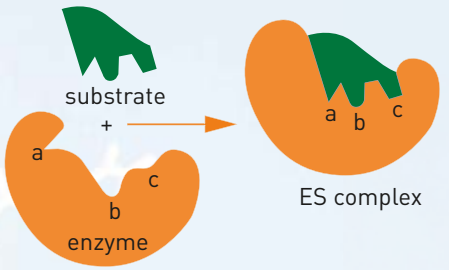


FIGURE 3 The lock and key model and induced fit model explain how an enzyme interacts with a substrate molecule.

CHECK YOUR LEARNING 19.4

Describe and explain

1 **Explain** the role of an enzyme in breaking down or synthesing a molecule.
2 **Summarise** the importance of an active site in an enzyme.

Apply, analyse and interpret

3 **Consider** the benefits of an enzyme being flexible in the 'induced-fit' model, over being rigid in the 'lock and key' model?

Investigate, evaluate and communicate

4 Based on your knowledge of the active site of an enzyme, **hypothesise** how an enzyme can become denatured and lose its ability to catalyse a reaction.
5 'Every enzyme generally only has one function'. **Evaluate** this statement and justify why it is correct.
6 **Investigate** what happens to the enzyme pepsin once it leaves the stomach. Comment on whether it still works and **justify** why or why not.

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Enzymes

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The 'lock and key' hypothesis

CHAPTER 19

Review

Chapter summary

19.1
- Collision theory states that reactants must collide in the correct orientation and with sufficient energy to result in a successful chemical reaction.
- To increase the rate of a reaction, the activation energy (minimum energy required for a chemical reaction to occur) or the number of collisions resulting in a successful collision must be increased.
- An increase in concentration, increase in pressure (and therefore decrease in volume), increase in surface area or increase in temperature all result in a faster chemical reaction.
- The Maxwell–Boltzmann distribution represents the number of particles in a reaction that have energy greater than the activation energy and successfully react.

19.2
- Energy profile diagrams are used to demonstrate the difference in chemical energy within the bonds of reactants and products. Reaction enthalpy is calculated as:

$$\Delta H = H_{products} - H_{reactants}$$

- If the reactants contain more energy than the products do, energy is released and the reaction is exothermic.
- If the products contain more energy than the reactants do, energy must be absorbed from the environment and the reaction is endothermic.
- The stronger the bonds broken into a reaction, the higher the activation energy and the slower the reaction. The weaker the bonds, the lower the activation energy and the faster the reaction.

19.3
- Catalysts provide an alternative pathway, which requires less activation energy, resulting in an increased rate of chemical reaction. Catalysts can be represented on the Maxwell–Boltzmann distribution as well as on an energy profile diagram.
- Enzymes and metal nanoparticles are examples of catalysts.

19.4
- Chemical reactions can be both fast and slow.
- Enzymes act as biological catalysts in a process that can be explained either by the lock model and key or induced-fit model.

Key terms

- activation energy
- bond energy
- catalyst
- collision theory
- concentration
- endothermic
- energy profile diagram
- enthalpy
- enzyme
- exothermic
- heterogeneous catalyst
- homogeneous catalyst
- kinetic energy
- kinetic molecular theory
- pressure
- surface area
- temperature
- transition state

Key formulas

Reaction enthalpy	$\Delta H = H_{products} - H_{reactants}$

Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; *** = high.

Multiple choice

1 Consider the chemical reaction:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Which of the following will not increase the rate of the reaction?

- **A** Crush the magnesium into a powder.
- **B** Increase the concentration of oxygen.
- **C** Add water to make the reactants aqueous.
- **D** Increase the temperature by heating the magnesium.

2 Consider the chemical reaction:

$$NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

Which of the following will not increase the rate of the reaction?

- **A** Increase the volume of the solution.
- **B** Increase the concentration of NaCl.
- **C** Increase the temperature.
- **D** Increase the concentration of $AgNO_3$.

3 Which of the following conclusions about the graph in Figure 1 is correct?

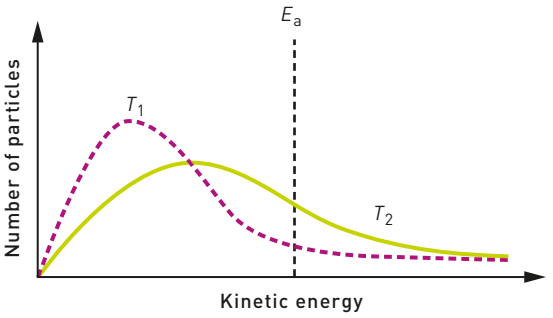

FIGURE 1 Kinetic energy against number of particles.

- **A** At T_2, fewer particles have energy greater than E_a; therefore, there will be a slower reaction at this temperature, compared to T_1.
- **B** At T_1, fewer particles have energy greater than E_a; therefore, there will be a faster reaction at this temperature, compared to T_2.
- **C** At T_1, more particles have energy greater than E_a; therefore, there will be a slower reaction at this temperature, compared to T_2.
- **D** At T_2, more particles have energy greater than E_a; therefore, there will be a faster reaction at this temperature, compared to T_1.

4 In an endothermic reaction:

- **A** the enthalpy value is positive
- **B** the reaction loses energy to the environment
- **C** the energy required to break bonds is smaller than the energy released when new bonds form
- **D** a catalyst will increase the activation energy.

5 In a catalysed chemical reaction:

- **A** the difference in enthalpy between products and reactants (ΔH) changes
- **B** an alternative reaction pathway with lower activation energy is provided
- **C** the kinetic energy of particles and therefore temperature increases
- **D** an alternative reaction pathway with higher activation energy is provided.

6 An enzyme acts as a catalyst for chemical reactions by:

- **A** lowering the activation energy by stabilising the transition state.
- **B** increasing the activation energy by stabilising the transition state.
- **C** lowering the activation energy by destabilising the transition state
- **D** increasing the activation energy by destabilising the transition state.

7 Metal nanoparticles are highly reactive as:

- **A** they have a large surface area.
- **B** they have a small surface area.
- **C** they can combine together to form large particles.
- **D** their metallic properties make them highly reactive.

8 Which of the following does not always affect the rate of a reaction?

- **A** surface area
- **B** temperature
- **C** pressure
- **D** oxidation state of reactants

9 What colour will a metal precipitate be if a non-transition metal reacts with sodium hydroxide?

A it will not form a precipitate

B colourless

C white

D green

10 Which of the following is not a rule for the solubility of salts in water?

A most chloride (Cl^-) salts are soluble

B all nitrate (NO_3) salts not soluble

C most salts of Na^+ and NH^{4+} are soluble

D most sulphate (SO_4^{2-}) salts are soluble

Short answer

Describe and explain

★ **11 Define** an enzyme.

★ **12 Describe** how temperature alters the rate of a reaction.

★ **13** A cold pack contains solid ammonium nitrate (NH_4NO_3) with water in separate pouches. When the pouches are broken, the water reacts to dissociate the ammonium nitrate in an endothermic process. Ammonium nitrate has an enthalpy value of 647 kJ mol^{-1}, while NH_4^+ and NO_3^- have a combined enthalpy value of 621 kJ mol^{-1}.

The ΔH value for the forward reaction is +26 kJ mol^{-1}, and the activation energy for the reverse reaction is 64 kJ mol^{-1}.

FIGURE 2 Cold packs are used to decrease swelling from injuries.

Sketch a fully labelled energy profile diagram for the forward reaction, including E_a and ΔH values.

★ **14 Describe** the difference between activation energy of an uncatalysed reaction and the activation energy of a catalysed reaction?

★★ **15** A beaker containing a clear solution of a slow reacting precipitation reaction is placed on a piece of paper which has a black cross written on it.

a **Describe** a method which could be used to measure the rate of reaction.

b **Explain** two ways that the rate of reaction be increased.

★★ **16** Hydrogen peroxide decomposes into oxygen gas and water, but the process is very slow.

a **Describe** a method that could be used to measure the rate of reaction.

b **Explain** three ways that the rate of reaction could be increased.

★★ **17 Determine** how the lock and key model differs from the induced fit model.

★★★ **18** A chemical reaction between 20.0 mL of 1 M sodium carbonate and 5.0 mL of 2.0 M hydrochloric acid produced the volume vs time graph in Figure 3.

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

a The temperature of the reaction is increased and the reaction is repeated. **Sketch** the resulting line on the graph.

b The volume of sodium carbonate is increased by 10 mL in a third version of this experiment. **Sketch** the resulting line on the graph.

c The temperature of the reaction flask is decreased in a fourth experiment, using the same amount of reactants as originally stated. **Sketch** the resulting line on the graph.

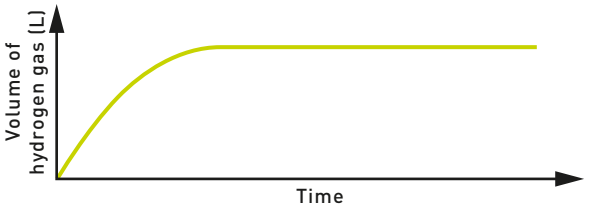

FIGURE 3 Hydrogen gas volume changes over time.

★★★ **19** Two experiments are conducted to investigate the factors affecting the rate of a chemical reaction between magnesium metal and hydrochloric acid:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

In the first experiment, a 2 cm strip of magnesium ribbon, with a mass of 0.9 g, is reacted with 3 mL of 1 M HCl. In the second experiment, a cubic piece of magnesium, measuring 0.5 × 0.5 × 0.5 cm, with a mass of 0.9 g, is placed in 3 mL of 2 M HCl.

a **Identify** how the rate of reaction could be measured in these experiments.

b **Identify** what factors are being investigated in this reaction.

c **Explain** what the major flaw is in this experiment.

d Is it possible to determine which reaction occurs faster? **Explain** why.

e When both reactions finish, and no more bubbles are generated, more magnesium is added to each flask with the purpose of making the reactions start again. **Describe** what happens after adding extra magnesium.

Apply, analyse and interpret

★★ **20** A chemistry student made the claim 'All collisions result in a chemical reaction. When the number of collisions is increased, the rate of reaction increases'. **Determine** why this student incorrect?

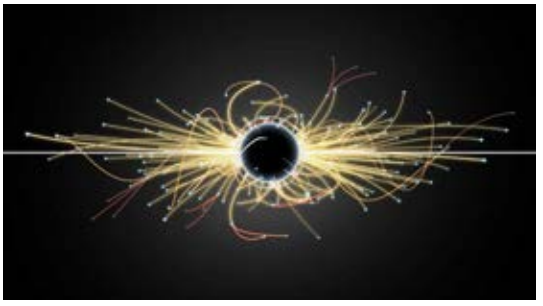

FIGURE 4 Particles colliding

★★ **21** **Determine** how you would measure the rate of reaction for combustion? **Sketch** a graph of the trend you would expect to observe. Is this qualitative or quantitative data?

★★ **22** **Determine** how you would measure the rate of reaction for precipitation? **Sketch** a graph of the trend you would expect to observe. Is this qualitative or quantitative data?

Investigate, evaluate and communicate

★★★ **23** **Consider** the diagram in Figure 5.

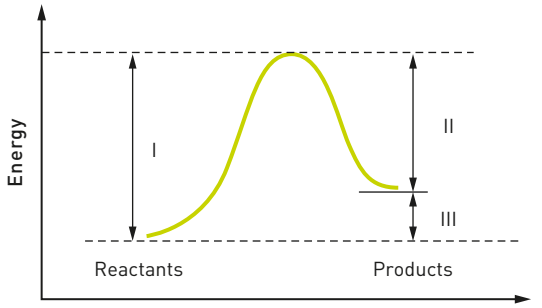

FIGURE 5 Energy diagram

a Is the reaction endothermic or exothermic? **Justify** your answer.

b **Determine** which of the three quantities (I, II or III) is affected by the addition of a catalyst?

c **Determine** which of the three quantities is affected when the reaction is switched between endothermic/exothermic and endothermic/exothermic?

d What would happen if the reaction were reversed? **Sketch** an energy profile diagram to **demonstrate** the effect.

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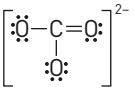
UNIT 2

Practice exam questions

Molecular interactions and reactions

Multiple choice

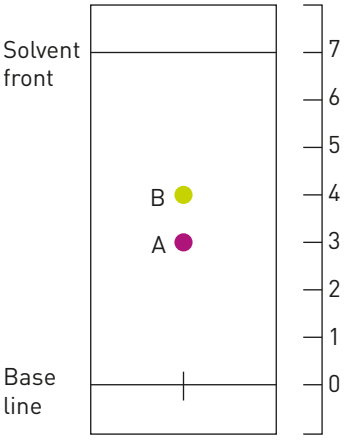
1 The shape of the carbonate ion below is described as:



A linear
B trigonal planar
C tetrahedral
D pyramidal.

2 Which of the following is an intermolecular force?
A Ionic bonding
B Covalent bonding
C Electrostatic bonding
D Hydrogen bonding

3 A forensic scientist tested the contents of a pill found at a crime scene. The diagram below shows the results of the chromatography test. The R_f of compound A was found to be:



A 0.43
B 0.57
C 0.75
D 0.92.

4 Which of the following statements is correct?
A A strong acid cannot be concentrated.
B A weak acid cannot be concentrated.
C A strong acid can be dilute.
D A strong acid cannot be dilute.

5 Which of the following substances are insoluble in water?
A $AgNO_3$
B NaOH
C AgCl
D $CuSO_4$

6 At 10°C, 175 g of substance X was dissolved in 100 mL of water. Use the solubility curve in Figure 5 in Section 16.2 to determine the possible identity of substance X.
A Sucrose
B Glucose
C $CaCl_2$
D KBr

7 An Arrhenius acid:
A increases the concentration of hydroxide ions in water
B increases the concentration of hydronium ions in water
C is a measure of how easily a solution is able to produce hydroxide ions
D easily disassociates to form hydroxide ions.

8 Which of the following is an ionic representation of a neutralisation reaction?
A $NaOH \rightarrow Na^+ + OH^-$
B $NH_3 + H^+ \rightarrow NH_4^+$
C $HCl + NaOH \rightarrow NaCl + H_2O$
D $H^+ + OH^- \rightarrow H_2O$

9 Collision theory states that the rate of a reaction can be increased by:

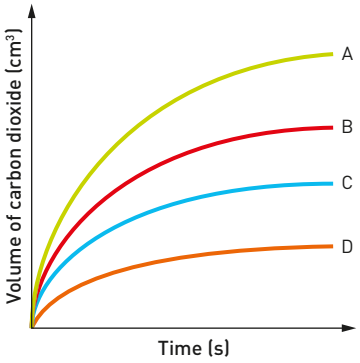
A decreasing the concentration of the reactants

B increasing the pressure of the solid reactants

C decreasing the surface area of the reactants

D increasing the temperature of the reactants.

10 Which of the graphs shows the reaction that has the most concentrated reactant?



A D

B B

C A

D C

Short answer

1 15.9 g of potassium dichromate ($Kr_2Cr_2O_7$) is dissolved in enough water to produce 250 mL of solution. **Determine** what the molarity of the solution is.

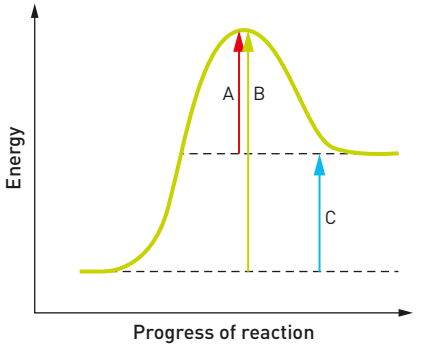
2 At 17°C, 26.5 mL of nitrogen gas at 1 Pa was collected in a flask. The next day the volume of the nitrogen was 27.1 mL with the pressure of 1 Pa. **Determine** the temperature on the second day.

3 Two iron nitrate solutions were poured into the same conical flask. The first solution contained 40.0 mL of 1.80 M $Fe(NO_3)_3$, and the second contained 21.5 mL of 0.808 M $Fe(NO_3)_3$. **Calculate** the final molar concentration of $Fe(NO_3)_3$ in the flask.

4 If the hydrogen ion concentration of a solution is 0.00001 M, **Calculate** the pH.

5 **Discuss** how nano-catalysts increase the rate of a reaction more than other catalysts.

6 The graph below represents the energy profile of a chemical reaction.



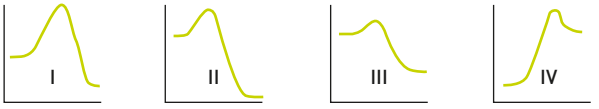
a **Identify** which arrow (A, B or C) represents the amount of activation energy in the reaction.

b **Identify** which arrow represents the change in energy between the reactants and products.

c **Determine** whether this reaction an exothermic reaction or an endothermic reaction. Provide reasoning to support your answer.

d How will the use of a catalyst affect the activation energy in this reaction? **Sketch** a graph to illustrate your answer.

e For the diagrams below **determine** the most easily ignitable fuel? Provide reasoning to support your answer.



7 Most of the chemical reactions that occur in living organisms involve enzymes such as catalase. Catalase aids in the breakdown of hydrogen peroxide (H_2O_2) into oxygen and water that occurs in the liver.

a **Define** the term 'enzyme'.

b **Predict** the effect of placing a piece of liver in a concentrated hydrogen peroxide solution.

c **Construct** a balanced chemical equation (including states) for this reaction.

d Hypothermia is a condition where the temperature of the body decreases below 35°C. **Explain** how this would affect the chemical reactions that occur in the human liver.

CHAPTER 20

Practical manual

The QCAA Chemistry General Senior Syllabus outlines a number of mandatory and suggested practicals for completion in Units 1 & 2. Suggestions for methodology and materials have been supplied in this chapter and are outlined in the table on the next page. However, the following is not prescriptive; schools may complete the mandatory or suggested practicals in any other form suited to their resources.

The experiments in this chapter have been trialled, and obvious cautions of hazards given; however, teachers are legally obliged to carry out their own risk assessment prior to undertaking any practical activity.

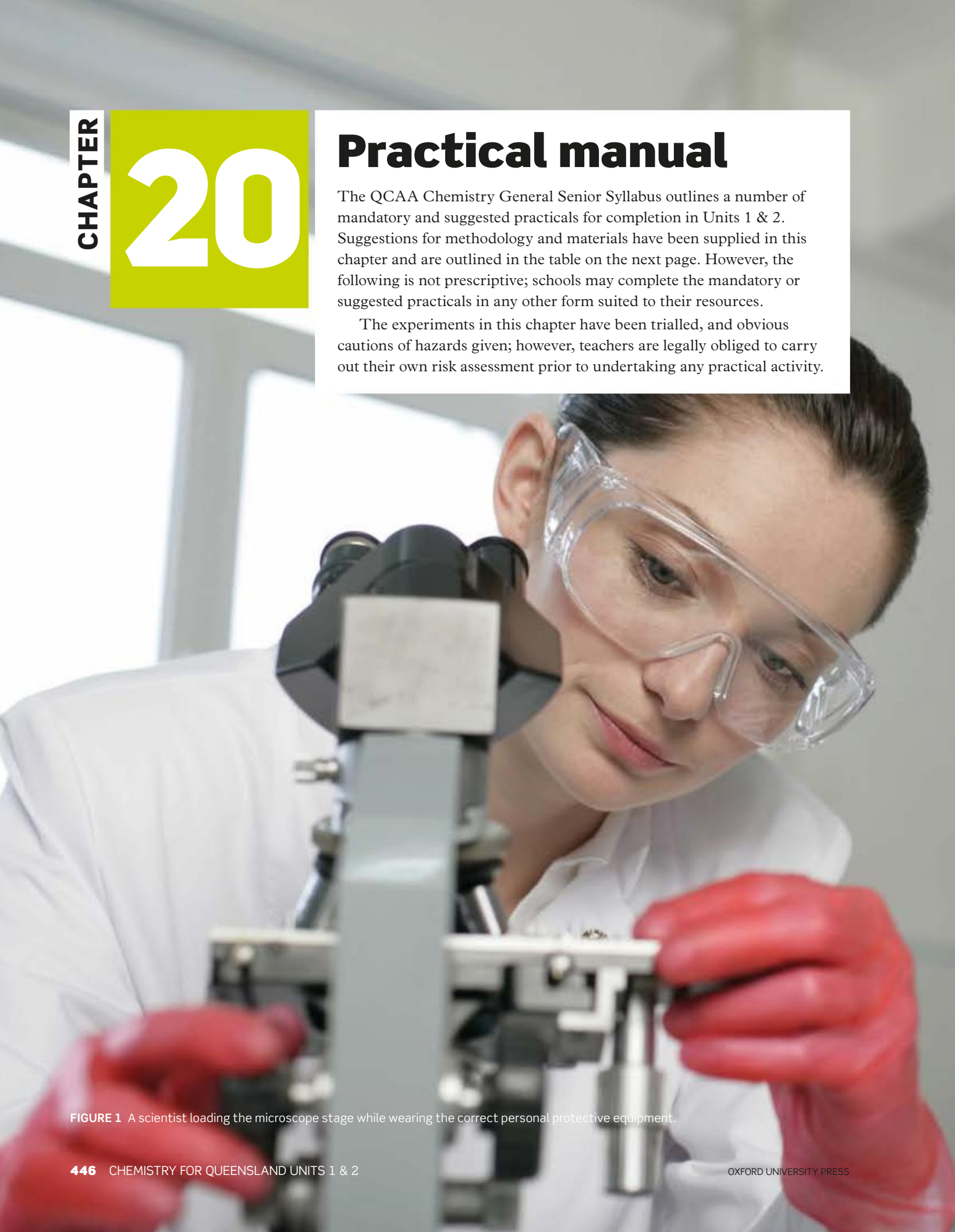


FIGURE 1 A scientist loading the microscope stage while wearing the correct personal protective equipment.

UNIT 1 PRACTICALS

MANDATORY	**8.5** Finding the percentage yield of magnesium oxide
MANDATORY	**9.3** Calorimetry: Measuring the enthalpy of a reaction
SUGGESTED	**2.4** Investigating periodic trends using a database
SUGGESTED	**2.5** Simulating Rutherford's gold foil experiment
SUGGESTED	**5.2** Conducting a flame test
SUGGESTED	**6.2** Separating mixtures
SUGGESTED	**7.4** Testing for saturation
SUGGESTED	**8.4** Investigating limiting reagents
SUGGESTED	**9.2A** Applying Hess's law
SUGGESTED	**9.2B** Measuring the molar heat of a chemical reaction

UNIT 2 PRACTICALS

MANDATORY	**12.2** Constructing 3D models of molecules
MANDATORY	**14.1** Investigating the properties of gases
MANDATORY	**16.1** Precipitation reactions Can be found on your obook assess
MANDATORY	**17.1** Investigating the properties of strong and weak acids
MANDATORY	**19.1** Investigating the rate of chemical reactions
SUGGESTED	**13.2** Separating the components of a mixture by paper chromatography
SUGGESTED	**14.2** Boyle's law
SUGGESTED	**16.2** Investigating the effect of temperature on solubility
SUGGESTED	**18.2** Reactions of acids with bases, metals and carbonates

8.5 MANDATORY PRACTICAL Finding the percentage yield of magnesium oxide

CAUTION: BUNSEN BURNERS MAY CAUSE BURNS; HANDLE THEM WITH CARE.

Derive the empirical formula of a compound from reactions involving mass changes.

Source: *QCAA Chemistry General Senior Syllabus 2019*

Context

Magnesium reacts with oxygen in the air to produce magnesium oxide. The reaction is known as combustion and is exothermic, meaning energy is produced. The mass of magnesium changes as it reacts with oxygen.

Aim

To determine the yield of magnesium oxide obtained from the combustion of magnesium ribbon in excess oxygen.

Materials

- Magnesium ribbon
- Balance (accurate to 0.0001 g)
- Bunsen burner
- Tripod
- Safety mat
- Clay triangle
- Crucible and lid
- Tongs

Safety

Ensure that you review how to light a Bunsen burner.

The crucible will become very hot. Always use the tongs to move the crucible or lid, and plan where you will place the hot item (i.e. on the safety mat).

Method

1. Set up the equipment as shown in Figure 1.
2. Heat the crucible and lid for approximately 3 minutes to remove any contaminants. Use the tongs to remove the crucible and lid from the flame and allow them to cool.
3. Use the balance to record the mass of the crucible and lid.
4. Use the balance to record the mass of approximately 3 cm of magnesium ribbon.
5. Place the magnesium ribbon in the crucible. Place the crucible lid over the crucible so that air can freely enter the container.
6. Place the crucible and lid on the clay triangle and tripod so that the magnesium ribbon will be over the heat of the Bunsen burner.
7. Light the Bunsen burner and ensure that the blue flame provides a maximum heat on the magnesium ribbon.

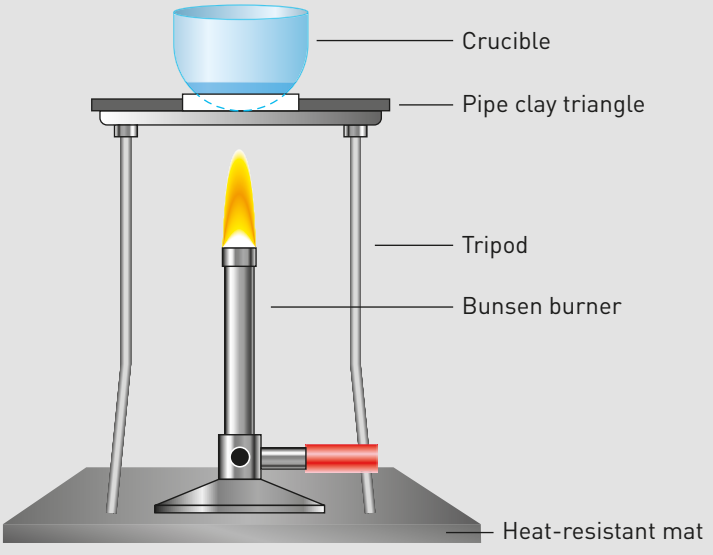


FIGURE 1 The equipment needed to determine the percentage yield of magnesium oxide.

8 Heat the crucible for approximately 10 minutes (until all the magnesium ribbon is a grey-white powder).

FIGURE 2 Magnesium after burning to form magnesium oxide in crucible.

9 Stop heating and allow the crucible, lid and contents to cool.

10 Record the mass of the crucible and its contents (Table 1).

11 Follow your teacher's instruction for disposal of the magnesium oxide.

Results

TABLE 1 Mass of the crucible, lid and contents

Item	Mass (g)
Crucible and lid	
Magnesium ribbon	
Crucible, lid and magnesium oxide	

Discussion

1 Use the mass of the magnesium ribbon to determine the number of moles of magnesium that was available for the reaction. Include the required number of significant figures and the uncertainties that may result from the measurements.

2 Subtract the mass of the magnesium from the mass of the magnesium oxide to determine the mass of the oxygen used in this reaction. Include the required number of significant figures and the uncertainties that may result from the measurements.

3 Determine the percentage composition of the magnesium and oxygen in the magnesium oxide.

4 Use the percentage composition to determine the empirical formula for magnesium oxide.

5 If the molar mass of magnesium oxide is 40.3044 g mol^{-1}, determine the molecular formula of magnesium oxide.

6 Write a balanced chemical equation for the reaction between the reactants, magnesium ribbon and oxygen, and the product, magnesium oxide.

7 Determine the molar ratio of the reactants to the product for this reaction.

8 Use the molar ratio to determine the theoretical yield of magnesium oxide from the reaction.

9 Use the experimental results to determine the percentage yield of magnesium oxide from the reaction.

10 Provide one systematic and one random error that could explain why the percentage yield may be less than 100%. Ensure your answer is specific to this reaction.

Conclusion

Evaluate the effectiveness of the production of magnesium oxide in this reaction.

9.3 MANDATORY PRACTICAL Calorimetry: Measuring the enthalpy of a reaction

CAUTION: HANDLE CHEMICALS WITH CARE AND CONSULT YOUR TEACHER AND RISK ASSESSMENT FOR EACH CHEMICAL USED IN THIS PRACTICAL. ALL CHEMICALS ARE IRRITANTS. HCl, NaOH and $CuSO_4$ ARE CORROSIVE. FOR ALL CHEMICALS, AVOID CONTACT WITH SKIN OR EYES BY WEARING THE CORRECT PERSONAL PROTECTIVE EQUIPMENT. WASH YOUR HANDS OR FLUSH YOUR EYES IF ANY CHEMICALS COME IN CONTACT WITH THEM. ZINC POWDER IS HIGHLY FLAMMABLE SO AVOID CONTACT WITH AN OPEN FLAME.

Conduct a calorimetry experiment to measure the enthalpy of a reaction.

Source: *QCAA Chemistry General Senior Syllabus 2019*

Context

Calorimeters provide an effective way of measuring the enthalpy changes in a chemical reaction because they measure the heat exchanged with the surroundings. In an aqueous solution, the specific heat capacity of water can be used to calculate the energy involved with the temperature change of the reaction. This can be calculated using the following equation:

$$q = mc\Delta T$$

where q is heat energy (J), m is the mass of water (g), c is the heat capacity of water (4.18 J g^{-1} $°C^{-1}$) and ΔT is the change in temperature (°C)

Energy can be converted to the amount of energy per mole of substance using the following equation:

$$\Delta H = \frac{E}{n}$$

Aim

To determine the enthalpy changes (ΔH) of chemical reactions by calorimetry.

Materials

- Calorimeter (if no calorimeter is available, an alternative is to place a Styrofoam cup with a lid in a beaker. The thermometer pierces the lid)

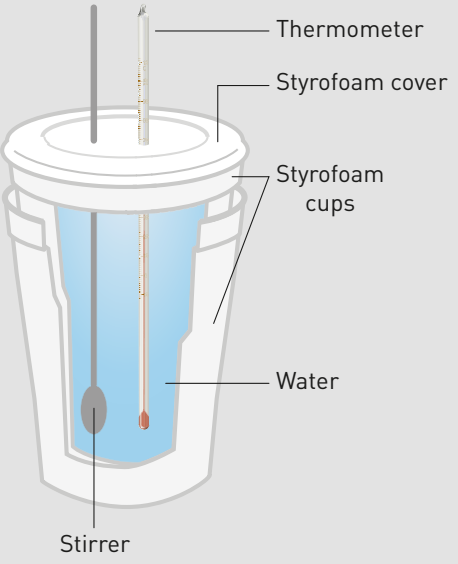


FIGURE 1 An alternative to a calorimeter

- 50 mL measuring cylinder (alternatively, a 50 mL glass pasteur pipette and pipette bulb provide more accurate measurements)
- Stopwatch
- Thermometer (digital is faster)
- Electronic balance and weigh boat
- 150 mL 2.0 M HCl
- 50 mL 2.0 M NaOH
- 0.9 g MgO
- 3.5 g $MgCO_3$
- 25 mL 1 M $CuSO_4$
- 6 g Zinc Powder

Safety

A laboratory coat, gloves and safety glasses should be worn in all experiments.

Latex gloves are recommended when handling 2.0 M HCl to avoid burning skin. If acid is spilt on skin or in the eyes, they must be immediately washed under the tap or using an eyewash station.

All solutions can be washed down the sink in dilute concentrations

Method

Part A: Measure the enthalpy of a neutralisation reaction

1 Add 50 mL of 2.0 M HCl to the calorimeter and record the temperature until it remains stable for 2 minutes. Record this as the initial temperature.

2 Add 50 mL of 2.0 M NaOH and stir constantly, record the maximum temperature that the solution reaches in the calorimeter in the data table.

Theoretical ΔH value is −58 kJ mol^{-1}

Part B: Measure the enthalpy of a chemical reaction: acid + metal oxide

1 Add 50 mL of 2.0 M HCl to the calorimeter and record the temperature until it remains stable for 2 minutes. Record this as the initial temperature.

2 Accurately weigh 0.9 g of MgO, using an electronic balance and weigh boat.

3 Add the MgO to the HCl in the calorimeter and stir constantly, recording the maximum temperature that the solution reaches in the calorimeter in the data table.

Theoretical ΔH value is −146 kJ mol^{-1}

Part C: Measure the enthalpy of a chemical reaction: acid + metal carbonate

1 Add 50 mL of 2.0 M HCl to the calorimeter and record the temperature until it remains stable for 2 minutes. Record this as the initial temperature.

2 Accurately weigh 3.5 g of $MgCO_3$, using an electronic balance and weigh boat.

3 Add the $MgCO_3$ to the HCl in the calorimeter and stir constantly, recording the maximum temperature that the solution reaches in the calorimeter in the data table.

Theoretical ΔH value is −90 kJ mol^{-1}

Part D: Measure the enthalpy of a displacement reaction

1 Add 25 mL of 1.0 M $CuSO_4$ to the calorimeter and record the temperature until it remains stable for 2 minutes. Record this as the initial temperature.

2 Accurately weigh 6.0 g of Zn powder, using an electronic balance and weigh boat.

3 Add the Zn powder to the $CuSO_4$ in the calorimeter and stir constantly, recording the maximum temperature that the solution reaches in the calorimeter.

Theoretical ΔH value is 217 kJ mol^{-1}

Results

Record the following in a table.

- Density of the solution (density of water in g mL^{-1})
- Volume of the solution
- Mass of the solution
- Specific heat capacity of the solution (SHC of water)
- Initial temperature (°C)
- Final temperature (°C)
- Change in temperature, ΔT (°C)

Discussion

1 Calculate q (in J) for the four reactions and convert it to kilojoules (kJ):

$$q = m \times c \times \Delta T$$

2 Calculate the number of moles for each solution by writing a balanced chemical equation for each reaction and determining the limiting reactant.

3 Calculate the change in enthalpy of each of the solutions, using the energy calculated in question 1 and number of moles calculated in question 2.

$$\Delta H = \frac{E}{n}$$

4 Calculate the percentage error for each of the reactions:

$$\%error = \frac{theoretical\ \Delta H}{\text{experimental } \Delta H} \times 100$$

5 What errors may be involved with the experiment and what is the specific impact do these errors have on the results (i.e. higher/lower)?

6 What assumption is being made by using the specific heat capacity of water in the calculations? Is this assumption justified?

Conclusion

Summarise the outcome of this practical.

12.2 MANDATORY PRACTICAL

Constructing 3D models of molecules

Construct 3D models (real or virtual) of linear, bent, trigonal planar, tetrahedral and pyramidal molecules.

Source: *QCAA Chemistry General Senior Syllabus 2019*

Context

Modelling is used in science to explain properties or processes that cannot be directly observed. Atoms were found to be arranged in three dimensions, which was particularly important when explaining their properties and the way they behaved in chemical reactions.

The simplest model of a molecule is the Lewis dot structure that shows the pairing of an atom's valence electrons. The valence shell electron pair repulsion (VSEPR) theory states that electron pairs repel each other in three dimensions, forming the shapes that influence the molecules properties.

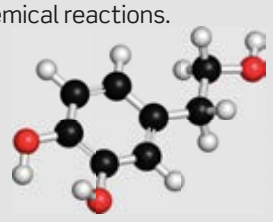


FIGURE 1 3D rendering of hydroxytyrosol olive oil antioxidant molecules.

FIGURE 2 The lewis structure of a krypton atom.

Aim

To construct 3D models of linear, bent, trigonal planar, tetrahedral and pyramidal molecules.

Materials

3D modelling chemical modelling kit

Procedure

1 Your chemical modelling kit will contain a number of different atoms that are usually different colours. Complete Table 1, which identifies the element, and the colour of that atom in your kit.

TABLE 1 The colour coding of elementary atoms in the chemical modelling kit

Elementary atom	Colour
Carbon	
Oxygen	
Hydrogen	
Nitrogen	

2 Use your chemical modelling kit to make the following molecules. Draw and label the shape of each molecule in Table 2.

TABLE 2 The 3D structure of molecules

Chemical formula	Lewis dot diagram	3D drawing of molecule	Name of molecular shape
O_2			
N_2			
H_2O			
CO_2			
CH_2O			
CH_4			
NH_3			
NO_2H			

Discussion

1 How many electrons do most stable atoms tend to have in their outside shell?

2 Why is hydrogen most stable with only two electrons in its outside shell?

3 Hydrogen and nitrogen are both diatomic molecules. How are covalent bonds formed between each molecule?

4 Without making a model, describe the molecular shape of carbon tetrabromide (CBr_4).

5 The model hydrogen atom in your kit is only able to form a single covalent bond with other atoms. Use the periodic table to determine what other elements this model atom could represent.

6 Ammonia (NH_3) and formaldehyde (H_2CO) each have three bonds around the central atom; however, their molecular shapes are different. Explain for this observation.

Conclusion

Summarise the outcomes of this practical.

14.1 MANDATORY PRACTICAL Investigating the properties of gases

CAUTION: HANDLE CHEMICALS WITH CARE AND CONSULT YOUR TEACHER AND RISK ASSESSMENT FOR EACH CHEMICAL USED IN THIS PRACTICAL.
HCl IS AN IRRITANT AND CORROSIVE. AVOID CONTACT WITH SKIN OR EYES BY WEARING THE CORRECT PERSONAL PROTECTIVE EQUIPMENT. WASH YOUR HANDS OR FLUSH YOUR EYES IF ANY CHEMICALS COME IN CONTACT WITH THEM.
HYDROGEN GAS IS HIGHLY FLAMMABLE. DO NOT USE ANY OPEN FLAMES AROUND THE EXPERIMENT.

Investigate the properties of gases to determine the molar volume of a gas.

Source: *QCAA Chemistry General Senior Syllabus 2019*

Context

The volume that a gas occupies is affected by temperature and pressure. At standard laboratory conditions (SLC), 25°C/298 K and 1 atm/101.325 kPa, 1 mole of any gas occupies 24.5 dm^3, which is the equivalent of 24.5 L. At SLC, the relationship between the moles of a substance and the volume it occupies can be summarised as:

$$V = n \times V_m$$

It is difficult to maintain the temperature and pressure of the labatory at 25°C/298 K and 1 atm/101.325 kPa, so the temperature and pressure of a laboratory can be measured and converted to SLC using the following conversion:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

By converting measured laboratory conditions to SLC, the molar volume of a gas can be measured at SLC.

Aim

To measure the molar volume of hydrogen gas produced in a reaction between magnesium and hydrochloric acid at SLC.

Materials

- 0.05 g Mg strip (approximately 3 cm) cleaned with emery paper
- 10 mL 2.0 M HCl
- 10 mL measuring cylinder (or 10 mL glass pipette and a pipette bulb)
- Electronic balance and weigh boat
- 100 mL conical flask
- Rubber stopper with tubing (to fit the conical flask)
- 100 mL gas syringe (to fit the tubing in the rubber stopper)
- Heatproof mat
- Retort stand, boss head and clamp
- Thermometer
- Plastic tweezers

Safety

A laboratory coat, gloves and safety glasses should be worn for all experiments. Use an acid spill kit for any acid split.

Method

1. Record room temperature and atmospheric pressure in a results table.
2. Clamp the gas syringe to the retort stand, ensuring it is clamped low enough to reach the conical flask. Ensure the plunger is all the way into the syringe.
3. Measure 10 mL of 2.0 M HCl into a measuring cylinder (alternatively use a pipette if available to increase accuracy) and transfer to the conical flask.
4. Place the conical flask on the heat proof mat and add the stopper. Connect the stopper to the gas syringe.
5. To ensure that there are no gas leaks between the syringe and the conical flask, pull the plunger out. If there are no leaks, it will return to its original position when it is released.
6. Using the balance and weigh boat, measure 0.05 g of magnesium ribbon. (NOTE: a lower mass

is acceptable, but higher masses may result in a volume of hydrogen gas that exceeds the volume of the gas syringe.)

7 Pull the cork out of the conical flask and quickly add the magnesium ribbon. Replace the cork as quickly as possible.

8 When the mixture has finished reacting, record the volume of hydrogen gas, which fills the syringe.

9 After the reaction is complete and no more gas is produced, there should be no more magnesium ribbon left in the conical flask. If there is, remove the stopper and use plastic tweezers to remove the remaining ribbon. Dry it and weigh it, then subtract this from the original mass.

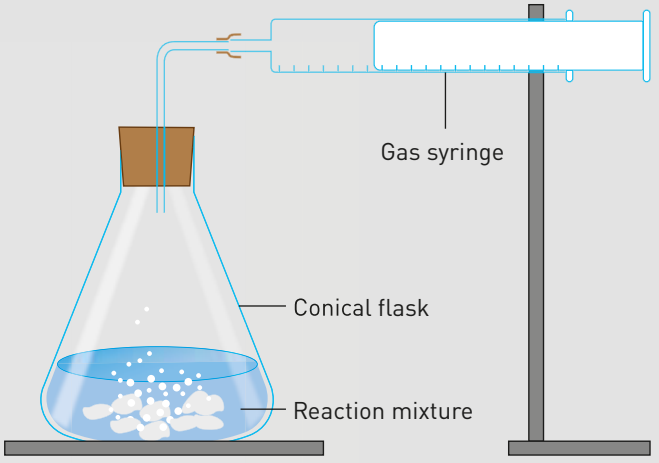


FIGURE 1 Experimental set-up for measuring hydrogen gas.

Results

Create a results table to measure the:

- temperature of the room
- atmospheric pressure
- mass of the magnesium strip
- volume of HCl used
- volume of gas produced in the reaction.

Discussion

1 Write a balanced chemical equation for the reaction between magnesium (Mg) and hydrochloric acid (HCl) to produce magnesium chloride ($MgCl_2$) and hydrogen gas (H_2).

2 Convert the volume of hydrogen collected at the temperature and pressure measured in the experiment to SLC, where V_2 is calculated, P_2 is 101.325 kPa and T_2 is 298 K.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

3 Calculate the number of moles of magnesium ribbon.

4 Calculate the number of moles of HCl.

5 Which reactant is in excess and which is limiting?

6 Calculate the number of moles of hydrogen gas produced.

7 Calculate the theoretical volume of hydrogen gas produced at SLC:

$$V = n \times V_m$$

8 What is the percentage yield of the gas produced?

$$\text{Percentage yield} = \frac{\text{experimental volume}}{\text{theoretical volume}} \times 100$$

9 Suggest errors to account for why the percentage yield may not be 100%. Note: errors must be specific to your yield (i.e. they should explain why the yield was above or below 100%.)

Conclusion

Summarise the outcomes of this practical.

17.1 MANDATORY PRACTICAL Investigating the properties of strong and weak acids

CAUTION: HANDLE CHEMICALS WITH CARE AND CONSULT YOUR TEACHER AND RISK ASSESSMENT FOR EACH CHEMICAL USED IN THIS PRACTICAL.
HCl IS AN IRRITANT AND CORROSIVE. ETHANOIC ACID IS CORROSIVE. FOR ALL CHEMICALS, AVOID CONTACT WITH SKIN OR EYES BY WEARING THE CORRECT PERSONAL PROTECTIVE EQUIPMENT. WASH YOUR HANDS OR FLUSH YOUR EYES IF ANY CHEMICALS COME IN CONTACT WITH THEM.
ELECTRICAL EQUIPMENT CAN CAUSE SHOCKS. CHECK THERE ARE NO DAMAGED WIRES OR POWER PACKS.

Investigate the properties of strong and weak acids.

Source: *QCAA Chemistry General Senior Syllabus 2019*

Aim

To investigate the conductivity and pH of a strong acid (hydrochloric acid) and a weak acid (ethanoic acid).

Materials

- Hydrochloric acid (0.1 M)
- Ethanoic acid (0.1 M)
- Power pack
- Ammeter
- Connecting wires
- Carbon electrodes
- Plastic well-plate
- Plastic pipettes
- Universal indicator
- pH meter

Safety

Wear safety glasses and lab coats. Use an acid spill kit for any acid that is spilt on the bench. If acid comes into contact with skin, immediately flush the skin with running water.

Method

Part A: Testing conductivity

1 Set up the apparatus as shown in Figure 1, ensuring that the volume of hydrochloric acid is enough to cover at least half of the electrodes.

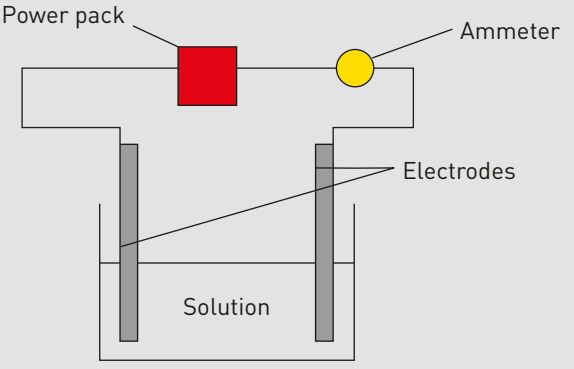


FIGURE 1 Testing the conductivity of an acid.

2 Record the volume of hydrochloric acid used.

3 Set the power pack to 4 V.

4 Turn on the power and record the current shown by the ammeter.

5 Repeat the experiment with the same volume of ethanoic acid.

Part B: Testing pH (universal indicator)

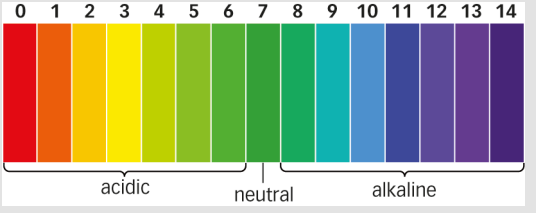


FIGURE 2 Universal pH indicator

1 Use a plastic pipette to place a few drops of hydrochloric acid into one of the wells in the well-plate.

2 Use a different plastic pipette to place a few drops of ethanoic acid into one of the wells in the well-plate.

3 Place one drop of universal indicator into each sample of acid.

4 Use a universal indicator pH colour chart to determine the pH of each acid.

Part C: Testing pH (pH meter)

1 Place the end of a pH meter into a sample of the hydrochloric acid. Record the pH.

2 Rinse the pH meter according to your teacher's instructions.

3 Place the end of a pH meter into a sample of the ethanoic acid. Record the pH.

Discussion

1 What ions are present in each of the acid solutions?

2 Compare the conductivity readings for each of the acids. If conductivity is directly proportional to the number of ions present in solution, which of the acids contained the most ions in solution?

3 Compare the pH levels for each of the acids. Which of the acids had the lowest pH?

4 Why do scientists prefer to use a pH meter instead of universal indicator to test pH levels?

5 Why was it important to use the same concentration for each acid?

6 From the data collected from this experiment, what can you infer about the relative ionisation of each of the acids used?

Conclusion

Evaluate the experimental data collected against the theory of strong and weak acids.

FIGURE 3 Scientist working with a pH meter.

19.1 MANDATORY PRACTICAL Investigating the rate of chemical reactions

CAUTION: HANDLE CHEMICALS WITH CARE AND CONSULT YOUR TEACHER AND RISK ASSESSMENT FOR EACH CHEMICAL USED IN THIS PRACTICAL. ALL CHEMICALS ARE IRRITANTS. HCl IS ALSO CORROSIVE. FOR ALL CHEMICALS, AVOID CONTACT WITH SKIN OR EYES BY WEARING THE CORRECT PERSONAL PROTECTIVE EQUIPMENT. WASH YOUR HANDS OR FLUSH YOUR EYES IF ANY CHEMICALS COME IN CONTACT WITH THEM. BUNSEN BURNERS MAY CAUSE BURNS, HANDLE THEM WITH CARE.

Investigating the factors which affect the rate of chemical reactions

Source: *QCAA Chemistry General Senior Syllabus 2019*

Context

The rate of a chemical reaction can be measured in order to determine the factors that affect it, such as surface area, temperature and concentration.

Calcium carbonate reacts with hydrochloric acid to form a solution of calcium chloride and carbon dioxide gas, according to the chemical equation:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

According to the law of conservation of mass, the mass of the reactants must be the same as the mass of the products.

FIGURE 1 Geologists use HCl to dissolve calcium carbonate (limestone).

Aim

To determine the effect that concentration, temperature and surface area have on the rate of a chemical reaction.

Materials

- 20 g large calcium carbonate chips
- 20 g small calcium carbonate chips (these can be washed and dried for Parts A and B. If they cannot be dried, 100 g is required for the whole experiment)
- 20 g powdered calcium carbonate
- 0.5 M, 1.0 M and 2.0 M hydrochloric acid
- Electronic balance and weigh boat
- Hot plate (alternatively a Bunsen burner with tripod, heatproof mat and gauze can be used)
- Thermometer
- 3 x 100 mL conical flasks
- 20 mL measuring cylinder
- 2 x 100 mL beakers
- Heatproof gloves
- Ice bath (iced water in a 250 mL beaker. Alternatively, 30 mL of 2.0 M HCl can be placed in a 100 mL beaker in the laboratory technician's fridge overnight and removed when needed for use in method 3).
- Stopwatch

Safety

A laboratory coat, gloves and safety glasses should be worn in all experiments.

If acid is spilt on the bench, use an acid spill kit.

The aqueous solution of calcium chloride formed should be decanted off the chips into a waste beaker and given to the laboratory technician for safe disposal. Very dilute solutions of 1 to 100 dilution can be disposed of down the sink.

Method

Part A: Surface area

1 Construct a results table of the mass lost by the chips over time. In the first 2 minutes, mass lost will be measured every 15 seconds. In the last 3 minutes, it will be measured every 30 seconds. Note, time zero is the initial mass of the calcium carbonate.

Note: If you are completing Part C, measure the temperature of the reaction and record it in the results table.

2 Place a 100 mL conical flask and a weigh boat on an electronic balance. Measure 20 mL of 2.0 M HCl in a measuring cylinder and add it to the conical flask on the balance. Tare the balance to zero.

3 Measure 20 g of large calcium carbonate chips into the weigh boat (which is already on the balance) and record the mass. This will be the initial mass of the chips.

4 Add the chips to the conical flask gently (so as not to cause the acid to splash) and start timing with the stopwatch. Important: Leave the weigh boat on the balance or the mass will be affected.

5 Record the mass lost by the chips in time intervals according to step 1. If the calcium carbonate is still losing mass after 5 minutes, continue recording the mass lost until it remains constant for 1 minute.

6 Decant the solution off the chips into a waste beaker. Rinse the chips with water three times, decanting off the water each time. Place the chips on some paper towel for drying so that they can be re-used.

7 Repeat steps 1–6 with 20 g of small calcium carbonate chips and 20 g of powdered calcium carbonate.

Part B: Concentration

Note: the small chips with 2.0 M HCl have already been measured in Part A.

1 Repeat Part A with small chips only and changing the concentration of HCl to 1.0 M and 0.5 M.

Part C: Temperature

Note: small chips with 2.0 M HCl at room temperature have been measured in Part A.

1 Repeat Part A with small chips only and changing the temperature of the 2.0 M HCl to 40°C and 60°C. Do this by adding the following step in between steps 1 and 2:

2 In a 100 mL beaker, add 30 mL of 2.0 M HCl. Place on a hotplate and heat to 40°C. Wear a heatproof glove while pouring 20 mL of the solution into a measuring cylinder and transfer this to the conical flask as instructed in Part A.

3 In a second 100 mL beaker, add 30 mL of 2.0 M HCl. Place the beaker in an ice bath and cool the acid to 10°C. Wear a heatproof glove while pouring 20 mL of the solution into a measuring cylinder and transfer to the conical flask as instructed in Part A.

Results

Record the results in a results table. Create a graph of mass lost over time.

Note: Plot the three sets of surface area data on the same graph. This should also be done for concentration and temperature.

Discussion

1 Calculate the number of moles of calcium carbonate.

2 Calculate the number of moles of 2 M, 1 M and 0.5 M HCl.

3 Which reactant is in excess in each reaction?

4 How does the trend represented on the graph represent the rate of a chemical reaction?

5 For each graph, explain the effect surface area, concentration and temperature have on the rate of a chemical reaction.

6 Why does the rate of the reaction eventually become zero? Use a mole calculation to justify your answer.

7 Why were small chips only used for concentration and surface area?

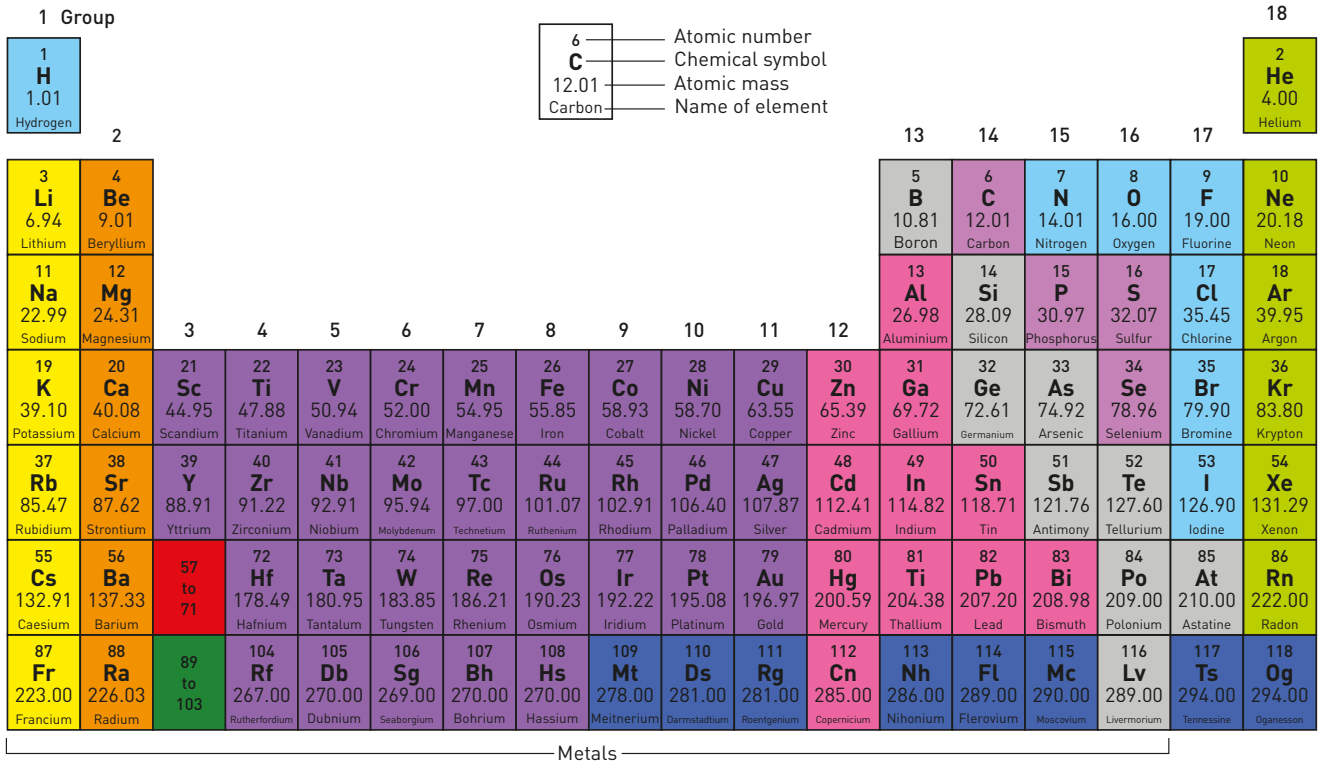
8 Can temperature be controlled in the experiment? What effect does this have on the results? How could it be minimised in future?

Conclusion

Summarise the outcome of the experiment.

APPENDIX

Periodic table



GLOSSARY

ΔH
the difference in enthalpy between reactants and products in a chemical equation

A

absorbs
the attraction of a substance within a sample to the stationary phase

absorption spectrum
the band of lines observed when an element absorbs light; viewed as a series of black lines on the visible light spectrum

abstract
a summary of the most important findings of a research article

accuracy
a measure of how close the measured value is to the true or accepted value

acid
a chemical compound that donates a hydrogen ion

acidic
has a pH of lower than 7

acid rain
rain that has been made acidic due to atmospheric pollution containing sulphur and nitrogen oxides from burning fossil fuels

activation energy
the minimum amount of energy required in a collision for a reaction to occur

addition reaction
a reaction in which the atoms in a reagent are added to an unsaturated compound

adsorb
the attraction of a substance within a sample to the stationary phase

affinity
the interaction of a substance within a sample within the mobile or stationary phases

alkali
a base that is soluble in water

alkane
a hydrocarbon that contains the maximum number of hydrogen atoms possible for the given number of carbon atoms; for linear or branched alkanes with n carbon atoms, the number of hydrogen atoms is $2n+2$; for a cycloalkane with one ring and n carbon atoms, the number of hydrogen atoms is $2n$

alkene
a hydrocarbon that contains one or more double bonds between carbon atoms

alkyne
a hydrocarbon that contains one or more triple bonds between carbon atoms

allotrope
one of the different forms in which a particular element may exist in a given physical state; these forms are due to differences in bonding between atoms of that element

alpha particle
a helium atom with removed electrons

amphoteric
can be basic or acidic

analytical chemistry
the branch of chemistry concerned with the detection and measurement of amount of substances

anion
a non-metal atom which has gained one or more electrons and has a negative charge

aqueous solution
a mixture of a solute dissolved in water

aqueous
a substance that is dissolved in water, or the state given to a substance that is dissolved in water

aromatic compound/arene
a cyclic unsaturated hydrocarbon compound that has delocalised clouds of electrons

atomic absorption spectroscopy (AAS)
the determination of the concentration of metal atoms in the gas phase by measuring the amount of light absorbed from a specific lamp

atomic radius
used to describe the size of an atom, which is measured by halving the distance between the nuclei of adjacent atoms

autotroph
an organism, such as a plant or alga, that produces its own food by transforming sunlight to chemical energy

Avogadro's constant
the number of atoms or molecules in one mole of a substance, equivalent to 6.02×10^{23} particles

B

balanced
having the same numbers of atoms on either side of a reaction arrow

base
a chemical compound that accepts a hydrogen ion

basic
has a pH of higher than 7

bent
the molecular shape in which the central atom has two covalent bonds and two lone pairs of electrons, forming a V-shape

benzene
the simplest aromatic compound with the molecular formula C_6H_6, where the carbon atoms are arranged in a flat (planar) six-membered ring, with one hydrogen atom attached to each carbon atom

biodiesel
a biofuel made from animal fats or vegetable oils

bioethanol
a biofuel containing ethanol made from sugar cane, wheat and corn

biofuel
a fuel that is produced from biomass (living matter)

blank solution
a solution that does not contain the compound being measured, but contains all other compounds that would be present in the sample being tested; this allows the response due to those other compounds to be subtracted from other measurements

blood plasma
the component of blood that contains dissolved substances

blue shift
a shift in the spectrum towards the blue region, indicating that a galaxy is moving away from the observer

boiling point
the temperature at which a liquid becomes a gas

boiling
vaporisation at a substance's boiling point, which is the temperature at which a liquid becomes a gas

bond angle
the angle between covalent bonds attached to the same central atom

bond energy
the amount of energy required to break a chemical bond in the gas phase, measure in kJ mol^{-1}

bond enthalpy
the amount of chemical energy stored within a single bond

bonding electrons
valence electrons that are bonded to another atom

C

calibration curve
the trend associated with plotting the peak area of a set of standards, from a chromatogram, against their concentration

calibration
the checking and adjustment of measured values of equipment so that they become more accurate

calorimeter
a device that measures the change in temperature of water when an endothermic or exothermic reaction takes place

catalyst
a substance that lowers the activation energy of a reaction and increases the rate of reaction without participating in the chemical reaction

cation
a metal atom that has lost one or more electrons and is positively charged

causation
when a change in a single variable effects a change in a second variable

cellular respiration
an exothermic reaction that converts glucose ($C_6H_{12}O_6$) and O_2 to CO_2 and H_2O; occurs in the mitochondria of heterotrophs and autotrophs

central tendency
the tendency for repeated measurements of the same value to be grouped around the mean, mode or median

centrifugal force
the apparent outward force of a mass being rotated

centrifugation
the technique for separating mixtures where tubes are placed into a centrifuge that spins in a circular motion very fast to push solids to the bottom

centrifuge
an instrument used to spin samples and separate them based on their densities

chemical energy
a form of potential (stored) energy generated by the attraction and repulsion of protons, neutrons and electrons, as well as the electron movement in a chemical substance

chemistry
an experimental science, concerned with the study of matter, and how substances can be combined or separated and how substances interact with energy

chloroplast
an organelle in plants cells that is the site of photosynthesis

chromatogram
the pattern of bands, spots or peaks formed on the chromatography paper to TLC plate demonstrating the separation of a mixture

closed system
a system in which gas particles cannot escape

colligative property
a property of a solution that depends on the ratio of the number of solute particles to the number of solvent molecules, but does not depend on the identity of the solute

collision theory
the theory that states that reactants must collide with the correct orientation and sufficient energy for a reaction to occur

colloid
a homogeneous mixture in which dispersed particles do not separate over time

column
the tube-like structure that contains the stationary phase ion column chromatography and through which the mobile phase and sample flow

column chromatography
a technique used to separate and purify individual components from mixtures of compounds

combination
a chemical reaction in which two atoms or elements react to form a molecule; also called a synthesis or addition reaction

combustion
a chemical reaction with oxygen to form a metal oxide, covalent compound or carbon dioxide and water

compound
a substance that is made up of atoms of different elements, joined together by chemical bonds

computational chemistry
a branch of chemistry in which the 3D structure of a molecule is generated by technology to assist in the determination of shape, intermolecular bonding and properties

concentrated solution
a solution containing many particles per unit volume of solution

concentrated
containing many particles per unit of volume

concentration
a measure of the number of particles in a specific volume of solution; units of concentration include mol L^{-1} (M), g L^{-1} and mg L^{-1} (ppm)

condensation
the process of turning a gas into a liquid by removing thermal energy; particles move slower and become more attracted to one another

condense
to change from a gas to a liquid

condensed electron configuration
an electron configuration that only states electrons in the highest energy levels, based on the nearest noble gas

confounding variable
an extraneous variable that changes systematically along with the variables being studied and offers an alternative explanation for the observations or measurements made

consistent
when measurements have overlapping uncertainty ranges

controlled variable
a condition that the experimenter tries to keep constant during an experiment, in order to minimise its effect on measurements or experiment outcomes

coordinate covalent bond
two electrons come from only one atom in a covalent bond

correlation
a link between a change in the independent variable and a change in the dependent variable; this does not mean that the changing independent variable caused the change in the dependent variable

covalent bonding
a sharing of electrons between two atoms

crystalline lattice
the ordered, repeating three-dimensional arrangement of atoms, molecules or ions in crystalline solids

crystallisation by evaporation
the process of evaporating a solvent to separate a solid substance that is dissolved in a liquid mixture

D

decantation
the technique by which liquid is carefully poured off the top of a liquid mixture to leave the undisturbed solid substance behind

decomposition
a chemical reaction in which a molecule breaks down into smaller molecules or even atoms on the application of heat or electricity

delocalised electrons
electron that is not retained in bonds between two atoms or associated with a single nucleus of an atom but is instead spread over several atoms or is highly mobile within a solid

density
how closely packed atoms are arranged; the mass of a substance per unit of volume

dependent variable (DV)
the variable that is measured or observed during an experiment, whose change is in response to changes made in the independent variable

deposition
the process of turning a gas into a solid by removing thermal energy; particles move slower and become more attracted to one another

desorption
the release of a substance within a sample from the stationary phase into the mobile phase

diatomic molecule
a molecule containing two atoms of the same element bonded together

dilute solution
a solution containing few particles per unit of volume

dilute
containing few particles per unit of volume

dilution
a method for lowering the concentration of a solution

dipole
a bond or a molecule that has a partial positive charge at one end and a partial negative charge at the opposite end

dipole–dipole interaction
two polar bonds or molecules interacting with each other by the partial charges

discrepant
when there is no overlap between the uncertainties of two measurements

discrete
individual or distinct data; usually used for qualitative data on column or bar graphs

dispersion force
the weak interaction between non-polar molecules

dissociate
to break apart

dissolve
when a substance dissociates in a solvent

distillation
the technique of separating a liquid from a liquid mixture by boiling the mixture into the liquid evaporates and is then collected by condensation

doppler effect
the change in a frequency of a wave as an object moves towards or away from an observer

double bond
two bonds between atoms where there are not enough electrons to complete the central atom

double displacement
a chemical reaction in which both the cation and anion of one reactant swaps with the cation and anion of the second reactant, forming new products

ductile
able to be drawn into a wire

E

electrolyte
a substance that forms ions in solution and is therefore able to conduct electricity

electromagnetic spectrum
the range of frequencies of electromagnetic radiation and their energies and wavelengths; electromagnetic radiation, or light, is comprised of photons

electron configuration
a representation of the electrons in an atom

electron dot
represent the number of valence electrons in a Lewis structure

electron gain enthalpy
the energy released or required to add an electron to a neutral atom or species in the gas phase (usually measured in kJ mol^{-1}), with formation of the –1 anion in the gas phase

electron
a negatively charged subatomic particle that exists outside of the nucleus

electronegativity
the attraction of a positively charged nucleus for the negatively charged electrons of a neighbouring atom

electrostatic force
an attraction between objects of opposing charges

electrostatic interaction
the attraction of the negative end of one molecule to the positive end of another molecule

element
a substance that is made up of one type of atom, defined by its atomic number

elute
come out of the bottom of a chromatography column

emission spectrum
the band of lines observed when an element emits light; viewed as a series of coloured lines that represent the wavelengths of electromagnetic radiation emitted by a source

emission
the process of emitting energy as light

empirical formula
the simplest whole-number ratio of the elements in a compound

endothermic
describes a reaction in which the reactants have less energy than products; the required energy is absorbed from the environment as heat

end point
the point in a titration at which the indicator changes colour

energy profile diagrams
a theoretical representation of the energy pathway that reactants must follow in order to form products

energy
the ability to do work

enthalpy
energy stored chemical substances, referred to as its chemical energy or heat content

enzyme
a type of protein that catalyses or accelerates biochemical reactions

equivalence point
the point in a titration at which a reaction is complete

error bar
a small line drawn on a plotted data point to show the uncertainty of the measured value

evaporation
the phase change from a liquid to a gas

excess reactant
the reactant that is not totally consumed by a chemical reaction

excited state
an energy state that is higher than the ground state; an electron is unstable in this state and emits energy as light to return to the ground state

exothermic
describes a reaction in which the reactants have more energy than the products; excess energy is released to the environment as heat

experimental yield
the actual amount of product that is produced in a chemical reaction

exponential
a graphed curved line where the dependent variable increases at a greater rate with every measurement

extraneous variable
a variable not being intentionally studied during an experiment that could have an unplanned or unintended effect on the results

F

film badge
a device worn by radiation-industry workers to detect and quantify the level of radiation exposure

filtration
the technique of pouring a liquid mixture through a filter so that the liquid passes through the filter, while the solid substance is caught in the filter

flame test
the analysis of metal ions by heating them in a Bunsen burner flame and measuring the light emitted when excited electrons return to their ground state

flotation
the technique for separating different substances in a powdered ore depending on whether they sink in, or float on, a given liquid

fossil fuel
a fuel formed from the remains of ancient organisms

free electron model of metallic bonding
a model of the bonding in metals that presents metals as ions surrounded by one or more valence electrons that are not associated with a particular atomic nucleus

freezing
the process of turning a liquid into a solid by removing thermal energy; particles move slower and become more attracted to one another

frequency
the number of waves (referring to wavelength) passing a point each second

fuel
a substance that reacts with oxygen to produce energy

full electron configuration
an electron configuration that includes every electron in an atom

G

gas chromatography
a technique used to separate and purify individual components from mixtures of compounds that can be vaporised

gas laws
the laws that explain the properties of gases and how they behave at different temperatures, volumes and pressures

Geiger counter
a device used to detect the presence of radiation (primarily beta radiation)

giant covalent network
an element or compound in which each atom is covalently bonded to other atoms to form an extended assembly of covalently bonded atoms

ground state
the lowest energy, and most stable, electron shell that an electron can occupy as it orbits the nucleus of an atom

group
a vertical column of the periodic table

H

half-life
the time it takes for half of the atoms of a radioisotope sample to decay

Hess's law
a law that states the enthalpy change for the overall reaction is equal to the sum of the enthalpy changes for the individual steps

heterogeneous catalyst
a catalyst that is in a different state from the reactants

heterogeneous mixture
a mixture with an inconsistent composition

heterotroph
an organism that must consume an autotroph or other organism to gain energy

high-performance liquid chromatography (HPLC)
a technique used to separate and purify individual components from mixtures of compounds by pumping a pressurised liquid containing the sample through a column of absorbent material

homogeneous catalyst
a catalyst that is in the same state as the reactants

homogeneous mixture
a mixture with a uniform composition

hydrocarbon
a molecular compound that only contains atoms of the carbon and hydrogen, for example, propane (C_3H_8)

hydrogen bonding
the strongest form of dipole–dipole interaction in which a partially negative atom (typically F, O or N) has an electrostatic interaction with a partially positive hydrogen atom

hypothesis
a proposed explanation used as a starting point for further investigation

I

ideal gas
a gas whose particles do not interact with one another and occupy no volume

independent variable (IV)
the variable that is altered systematically in a controlled way by the experimenter to test the effect on a related dependent variable

indicator
a substance that changes colour in the presence of different hydronium ion concentrations

insoluble
not able to dissolve in a solvent

intermolecular force
a force between molecules

intramolecular force
a force between the atoms in a molecule

ion
an atom or molecule that has gained or lost one or more electrons and is charged

ion–dipole interaction
an ion with a whole charge interacting with a partial charge caused by a polar bond or molecule

ionic bond/ionic interaction
the electrostatic interaction between a positive cation and negative anion

ionic bonding
the electrostatic (charge-based) attractive forces that holds ions together in ionic compounds

ionic compound
a compound consisting of positively charged cations and negative charged anions

ionic equation
a chemical equation that involves only the ions which change state during a chemical reaction

ionic equation of neutralisation
the equation obtained when spectator ions are removed from a neutralisation chemical equation e.g. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

ionic substance
a substance containing a cation and an anion

ionisation energy
the energy (in kJ mol^{-1}) required by a gaseous atom to remove an electron from that atom

ionisation
the formation of ions from atoms by the addition or removal of electrons

isotope
atoms of an element that have the same number of protons but differ in the number of neutrons

J

joule
a unit of energy

K

kinetic energy
the energy involved with the motion of particles, which can be increased by adding heat energy

kinetic molecular theory
the theory that states that all particles are in constant random motion

L

lattice energy
the energy released by formation of an ionic solid from ions in the gas phase that are infinitely separated; it is a measure of the stability of an ionic compound

law of conservation of energy
a law that states energy cannot be created or destroyed

Lewis structure
visual diagrams to show the number of electrons in the valence shell

limiting reactant
the reactant that is totally consumed when a chemical reaction is complete; this limits the amount of product that can be produced

line of best fit (trend line)
a line drawn on a graph joining as many points as possible and showing the general direction of the data; should be drawn with, approximately, an equal number of points above and below the line

linear (graphing)
a relationship where the *x*-values change proportionally with the *y*-values, leading to a straight line

linear (molecular shape)
the molecular shape in which the central atom has two covalent bonds and no lone pairs of electrons, forming a straight line

lit splint test
a test to determine the presence of carbon dioxide gas

lone pair of electrons
valence electrons that are not bonded to another atom

long-term renewable resource
a fuel that is being used faster than it can be produced, e.g. coal and oil

M

malleable
able to be pressed into other shapes

mass spectrometer
an instrument that ionises elements and molecules and produces spectra showing mass-to-charge ratio

mass spectrometry
an analytical technique in which elements are ionised and the ions are sorted according to their mass-to-charge ratio

mass spectrum
the spectrum obtained from a mass spectrometer; it displays mass-to-charge ratio on the *x*-axis and percentage abundance of each positively charged ion on the *y*-axis

matter
physical substances that have mass

Maxwell–Boltzmann distribution
a representation of the various kinetic energies of each individual particle in a system

measurement bias
a type of systematic error that occurs when equipment repeatedly makes the same miscalculation

melting point
the temperature at which a solid becomes a liquid

melting
the process of turning a solid into a liquid by applying thermal energy; particles move faster and start to separate

metallic bond
the attractive forces between metal ions and delocalised electrons that hold metals together

metalloid
an element that displays both metallic and non-metallic properties

mitochondria
a cell organelle that is the site of cellular respiration

mixture
a substance that is made up of two or more substances that are not chemically bonded

mobile phase
the phase that flows in chromatography, moving the components of a sample at different rates over the stationary phase

model
a simplified description of a process or phenomenon, often in a mathematical form

molar mass
the mass of one mole of a substance

molar ratio
the ratio of the amounts in moles of the reactants and products in a reaction

molar volume
the volume occupied by 1 mole of a gas at a specified temperature and pressure

molarity
molar concentration, units are moles per litre, mol L^{-1} or M

mole
the quantity of a substance containing the same number of particles as there are atoms in exactly 12.000 g of carbon-12, which is the same as Avogadro's number (6.02×10^{23})

molecular formula
the chemical formula of a compound that indicates the number of each elementary atoms present

molecular shape
the shape formed between a central atom and bonded atoms when covalent bonds form

molecular substance
a substance made up of covalently bonded molecules in which non-metal atoms are bonded together by the sharing of electron(s) between two atoms

molten
a melted substance

monatomic ion
an ion with only one atom

N

nanomaterial
a substance that contains particles with at least one size measurement ≤100 nm

nanotechnology
the study of how the control of individual atoms and molecules in the synthesis of nanomaterials can be applied across science and engineering fields

negative control
a test to make sure that an experiment is working as expected which should give no result

neutral
an aqueous solution that has a pH of 7

neutralisation
the reaction between an acid and a base to produce water and a metal salt

neutron
a neutral subatomic particle that exists inside the nucleus

non-bonding electrons
pairs of electron in an atom's outer shell that will not be shared because they have already paired up

nuclear reaction
a process used by radioisotopes to become more stable

nucleus
the dense structure within the atom

nutrient
a molecule that the body cannot synthesise on its own and must be obtained by eating foods

O

octet rule
atoms gain or lose electrons to reach eight electrons in their valence shells

origin
the line applied to a chromatogram to mark the point where the sample or standard was placed

outlier
a data point outside the rest of the data set

oxidation state/ionic charge
the difference between the positive charges from the protons and the negative charges from the electrons, represented as a superscript on the right-hand side of an elemental symbol

P

paper chromatography
an analytical technique for separating and identifying mixtures; the stationary phase is a thin strip of absorbent paper

parallax error
a random error that results from different values being read due to the angle of the observer

percentage abundance
the percentage of each isotope contained within a sample of element; it must always add up to 100%

percentage composition
the percentage that each element contributes to the mass of a compound

percentage error
the percentage difference between the measured experimental value and the known true value

percentage uncertainty
the uncertainty of measurements completed during an experiment expressed as a percentage of the final value

period
a horizontal row of the periodic table

periodic table
an organised presentation of elements by increasing atomic number

pH
a numerical scale that is used to determine the level of acid or base in an aqueous solution

photon
a particle and wave of light that carries energy and is absorbed or emitted from atoms

photosynthesis
an endothermic reaction that converts CO_2 and H_2O to glucose ($C_6H_{12}O_6$) and O_2 using sunlight; occurs in the chloroplasts of autotrophs

physical separation
the technique by which substances in a mixture can be easily removed by physically picking them out of the mixture

picometre
10^{-12} or 0.000 000 000 001 metres

polar
a polar bond or molecule is one which has a dipole

polarity
the electronegativity of atoms and the shape of a molecule gives the molecule an overall positive and negative end, creating a partial charge

polyatomic ion
an ion with two or more atoms

pop test
a test to determine the presence of hydrogen gas

positive control
a test to make sure that an experiment is working as expected which should give a positive result

ppb
parts per billion; a unit of concentration that measures the amount of substance in micrograms per litre

ppm
parts per million; a unit of concentration that measures the amount of substance in milligrams per litre

precipitate
the solid product formed as a result of a precipitation reaction

precipitation reaction
a reaction between two soluble ionic substances that results in the formation of a solid product

precision
the consistency a series of measurements

pressure
the force exerted, per unit area, by one substance upon another substance

producer
an autotroph; it produces its own food by transforming sunlight to chemical energy

proton
a positively charged subatomic particle that exists inside the nucleus

pure substance
a substance that is made up of a single type of element or compound

pyramidal
the molecular shape in which the central atom has three covalent bonds and one lone pair of electrons, forming a pyramid structure

Q

qualitative
data about types or categories, which may be represented by names, symbols, codes etc.

qualitative
relating to the non-numerical characteristics of a substance or object

quantitative
relating to numerical data about a substance, an object or a phenomenon

quantum theory
a theory that describes the behaviour of matter and energy at the atomic and subatomic level

R

radiation energy
energy that is a by-product of nuclear reactions

radioactive intermediate
a radioisotope that is very unstable and has an extremely short half-life

radioisotope
an unstable isotope that undergoes radioactive decay

random error
an unpredictable error in measurement or experimental procedure that has no detectable pattern

reactivity
how a substance reacts with other substances

red shift
a shift in the spectrum towards the red region, indicating that a galaxy is moving away from the observer

redox
a chemical reaction involving the transfer of electrons from one reactant to another

relative atomic mass (RAM)
the mass of an atom once the mass and percentage abundance of all isotopes have been considered, relative to the carbon-12 isotope

relative isotopic mass (RIM)
the mass of isotopes measured on a mass spectrum, relative to the carbon-12 isotope

renewable resource
a fuel that can be replaced at the same rate or faster than it is being used, e.g. wood or solar energy

replicate experiment
measurement that is repeated multiple times to give an indication of the precision of measurements; calibration to check that an instrument is giving accurate results and rectify if needed

retardation factor (R_f)
the ratio of the distances travelled by a component of a sample, from the origin, to the distance travelled by the mobile phase

risk assessment
an evaluation in a systematic way of the potential risks and their likelihood involved in doing an experiment or activity

retention time (R_t)
the time that each component is retained by the column

S

sample selection bias
a type of systematic error where a small sample is taken for testing from a larger amount of substance

saturated hydrocarbon
a compound of carbon and hydrogen that contains the maximum possible number of hydrogen atoms for a number of carbon atoms present

saturated
a solution in which no more solute can dissolve

scatterplot
a series of data points that are individually plotted on a horizontal and vertical axis

scientific method
systematic exploration of a phenomenon or topic using observation, measurement and experiment to support, disprove or modify hypotheses

scintillation counter
a device used to detect multiple forms of radiation energy

sedimentation
the process by which dense solid substances in a liquid mixture sink to the bottom of a container over time because of gravity

significant figures
the number of numerical figures that can be used to express a measured or calculated quantity pattern

single displacement
a chemical reaction in which a more reactive metal ion replaces a less reactive metal ion in a molecule

skeletal structure
a model in which the carbon chain is drawn as a zig-zag line with bends representing carbon atoms; carbon atoms are also at the end of the lines and each line represents a bond between carbons

slightly soluble
able to only partially dissolve in a solvent, forming a solution of low concentration

solubility
the ability of a substance to dissolve into another, or the maximum amount of solute which will dissolve in a known quantity of solvent at a given temperature

soluble
able to dissolve in a solvent

solute
the minor component of a solution – this substance is dissolved in the solvent

solution
a mixture of a solute dissolved in a solvent

solvent
the major component of a solution – this is what the solute is dissolved in

solvent front
the point on a chromatogram where the mobile phase reaches before the analysis is terminated

specific heat capacity
the amount of heat energy, in joules (J), that it takes to increase the temperature of a specific volume of water by 1°C (remembering that 1 g of water is 1 mL)

spectator ion
an ion that does not participate in a chemical reaction

stability
related to bond enthalpy

standard deviation
a statistical value that expresses how spread out a group of values are or by how much they differ from the mean value for the group

standard laboratory conditions (SLC)
standard conditions for temperature and pressure under laboratory conditions, when gases are measured at 1 atm and 25°C or 298.15 K

standard temperature and pressure (STP)
standard conditions for temperature and pressure, when gases are measured at 1 atm and 0°C or 273.15 K

standard
a sample of accurately known concentration of the chemical that is being tested for

state
the state of matter assigned to each chemical reactant or product – solid (s), liquid (l), gas (g) or aqueous (aq)

stationary phase
the phase to which the components of a chromatographic sample are absorbed

stoichiometry
the numerical relationship between the amount of reactants used and the products produced in a reaction

strong acid
an acid that completely dissociates in water

strong base
a base that completely dissociates in water

structural formula
a model that shows all atoms and all bonds but not necessarily in 3D, unless solid triangle and dotted lines are used

subatomic particle
a component that makes up, and exists within, the atom

sublimation
the process of turning a solid into a gas by applying thermal energy; particles move faster and separate from one another

substitution reaction
a reaction in which an atom or group of atoms in a molecule is replaced by another atom or group of atoms

supersaturated
a solution that has been heated to dissolve more solute than can be dissolved under equilibrium conditions; on cooling the solute remains in solution in a metastable state, until a disturbance (seed crystal, speck of dust, tapping the container) starts the crystallisation process

surface area
a measure of the total surface area available to react in a chemical reaction

surface tension
the ability of the surface of a liquid to resist an external force

suspension
a heterogeneous mixture containing solute particles which settle to the bottom and out of a solvent

synthetic chemistry
the branch of chemistry concerned with making substances in the laboratory and discovering new methods to transform substances into new substances

systematic error
a consistent repeatable error that occurs every time a piece of equipment is used

T

temperature
a measure of the average heat energy of the particles within a system

tetrahedral
the molecular shape in which the central atom has four covalent bonds and no lone pairs of electrons, forming a tetrahedral structure

theoretical yield
the amount of product predicted from stoichiometric calculations

thermochemical equation
a balanced chemical equation, including states, with a ΔH value that is positive or negative

thin-layer chromatography (TLC)
an analytical technique for separating and identifying mixtures; the stationary phase is typically a thin layer of silica gel, aluminium oxide or cellulose supported on a piece of glass or plastic

three-dimensional modelling
modelling in which atoms attached to the central atom are drawn in 3D to maximise the electron repulsion around the central atom

transition state
the state that corresponds to the unstable arrangement of atoms, where bonds are breaking and re-forming, when they reach the highest energy level (activation energy) of a reaction pathway

trigonal planar
the molecular shape in which the central atom has three covalent bonds and no lone pairs of electrons

trigonal pyramidal
the molecular shape in which the central atom has three covalent bonds and one lone pair of electrons, forming a pyramid structure

triple bond
a bond in which atom share three electrons

two-dimensional modelling
modelling in which central atom is drawn and all bonds are drawn evenly around it on a 2D plane

U

uncertainty
a numerical measurement that describes the observer's confidence in the accuracy and precision of a measurement

universal gas constant (*R*)
the proportionality constant, defining gas behaviour under ideal conditions

unsaturated hydrocarbon
a compound of carbon and hydrogen that contains double or triple bonds and therefore have fewer than the maximum number of hydrogen atoms for a particular number of carbon atoms

unsaturated
a solution that can dissolve more solute

V

valence
electrons located in the atom's outermost energy level only

valency
the combining power of an element as determined by the number of electrons in its outer shell (valence electrons)

vaporisation
the process of turning a liquid into a gas by applying thermal energy; particles move faster and separate from one another

vaporise
to change from a liquid to a gas

vapour pressure
the pressure of a vapour in contact with its solid or liquid form

vapour
a substance in the gaseous phase

variable
a condition or parameter that is changed or changes during an experiment

volatile substance
a substance that easily evaporates from a liquid to a gas

volume
a measure of the space occupied by a substance

VSEPR (valence shell electron pair repulsion) theory
a theory that states that electron pairs in bonds and lone electron pairs will repel each other around a central atom, forming a 3D shape

W

wavelength
the distance between two crests or troughs of a wave

weak acid
an acid that only partially dissociates in water

weak base
a base that only partially dissociates in water

weak solution
a solution containing few particles per unit volume of solution

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The front cover shows a mixture of household liquids including paint, milk, cooking oil and liquid soap. Each of these substances can be clearly seen because they have different densities and therefore do not mix.

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