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JACARANDA

CHEMISTRY 1

VCE UNITS 1 AND 2 | THIRD EDITION

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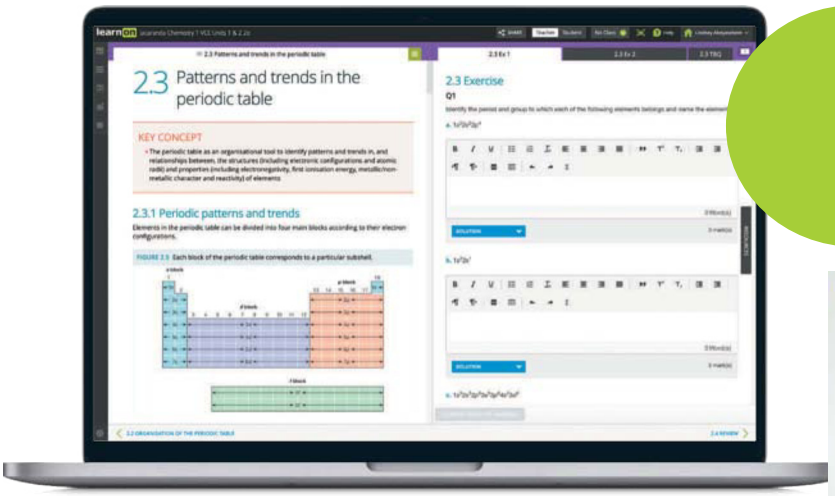
About this resource



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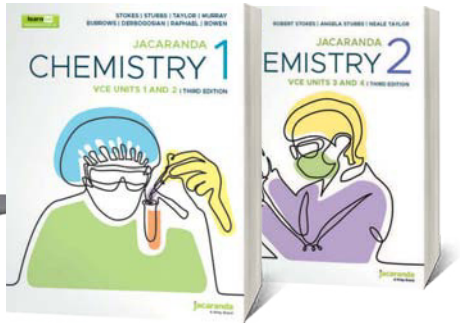

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CHEMISTRY 1 VCE UNITS 1 AND 2

THIRD EDITION

Developed by expert Victorian teachers for VCE students

Tried, tested and trusted. The NEW Jacaranda VCE Chemistry series continues to deliver curriculum-aligned material that caters to students of all abilities.

Completely aligned to the VCE Chemistry Study Design

Our expert author team of practising teachers and assessors ensures 100% coverage of the new VCE Chemistry Study Design (2023–2027).

Everything you need for your students to succeed, including:

- NEW! Access targeted questions sets including exam-style questions and all relevant past VCAA exam questions since 2013. Ensure assessment preparedness with practice SACs.
- NEW! Enhanced practical investigation support including practical investigation videos, and eLogbook with fully customisable practical investigations — including teacher advice and risk assessments.
- NEW! Teacher-led videos to unpack challenging concepts, VCAA exam questions, exam-style questions, investigations and sample problems to fill learning gaps after the COVID-19 disruptions.

Learn online with Australia's most

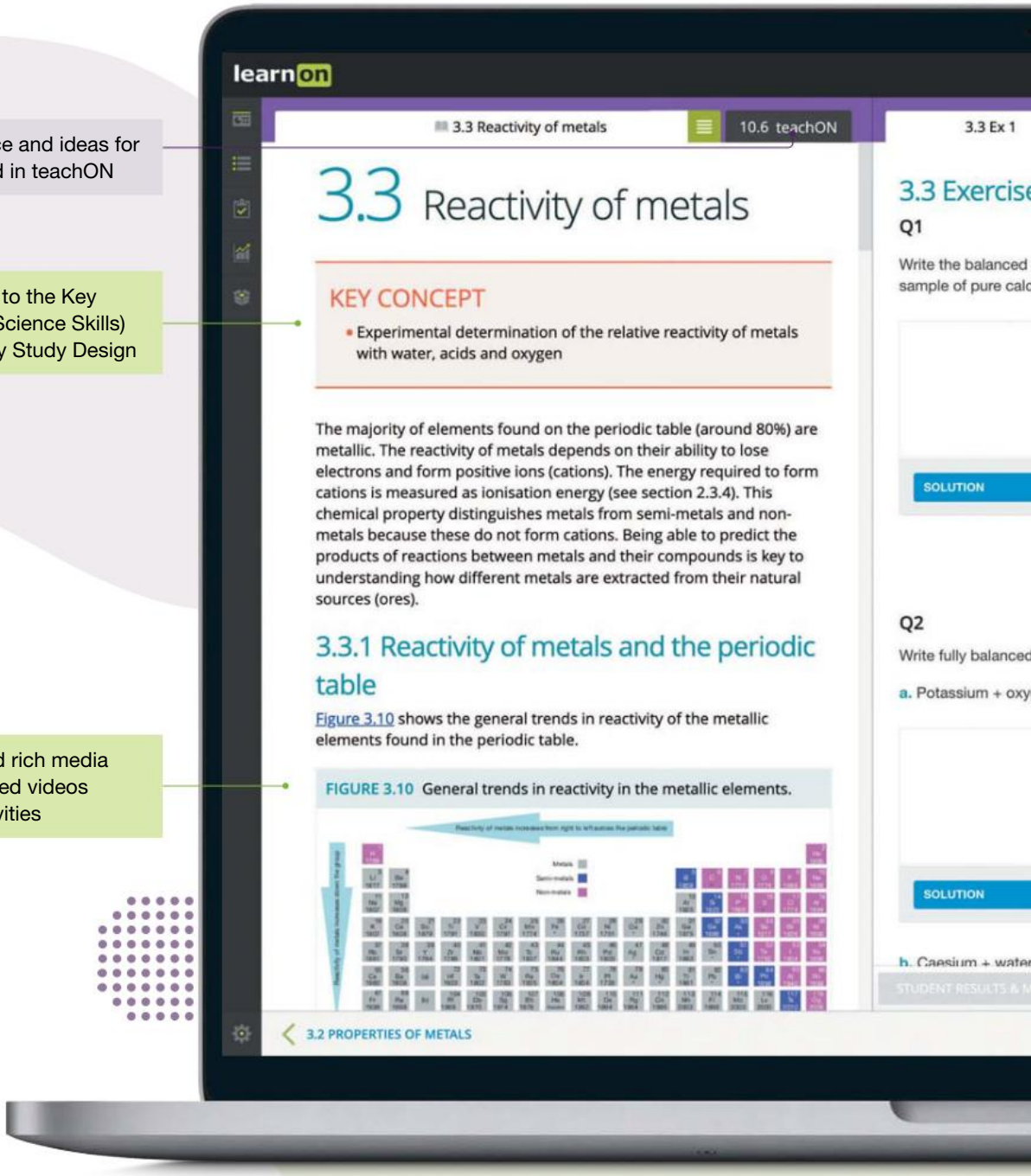
Everything you need for each of your lessons in one simple view

- **Trusted, curriculum-aligned theory**
- **Engaging, rich multimedia**
- **All the teacher support resources you need**
- **Deep insights into progress**
- **Immediate feedback for students**
- **Create custom assignments in just a few clicks.**

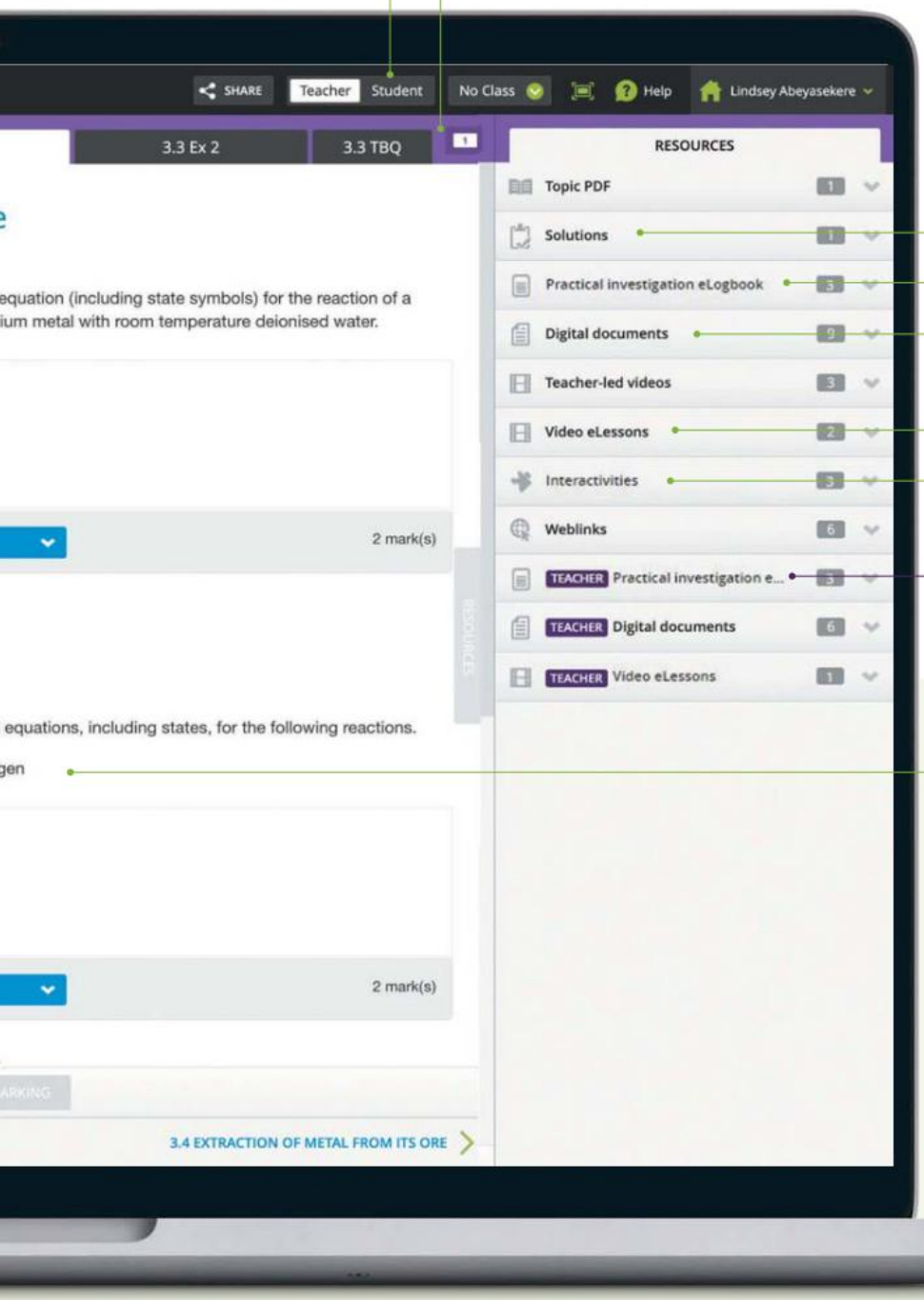
Practical teaching advice and ideas for each lesson provided in teachON

Each lesson linked to the Key Knowledge (and Key Science Skills) from the VCE Chemistry Study Design

Reading content and rich media including embedded videos and interactivities



powerful learning tool, learnON



Teacher and student views

Textbook questions

Fully worked solutions and sample responses

Practical investigation eLogbook

Digital documents

Video eLessons

Interactivities

Extra teaching support resources

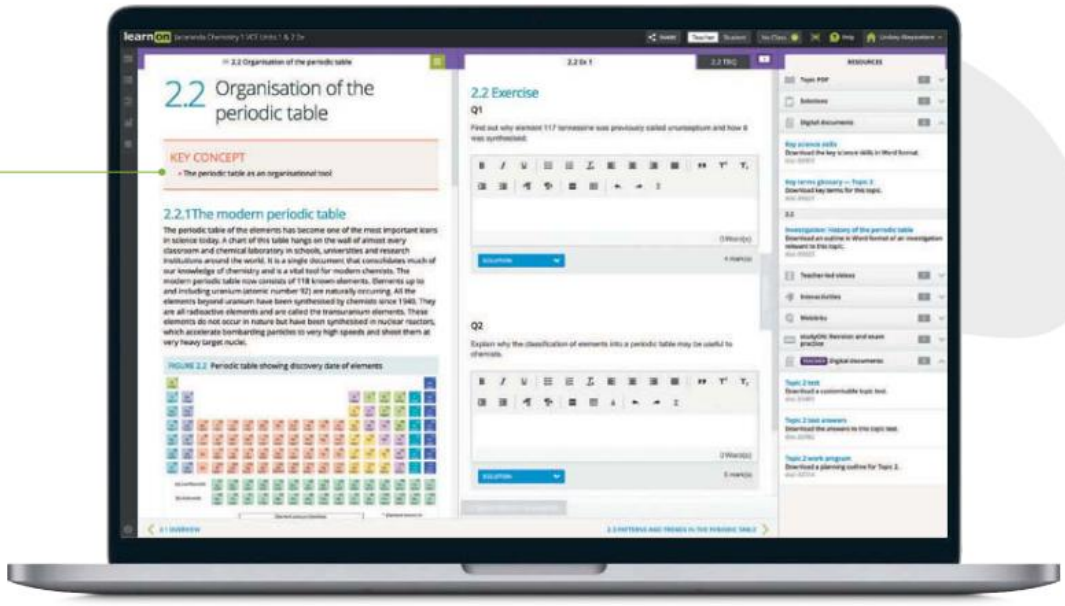
Interactive questions with immediate feedback

Get the most from your online resources

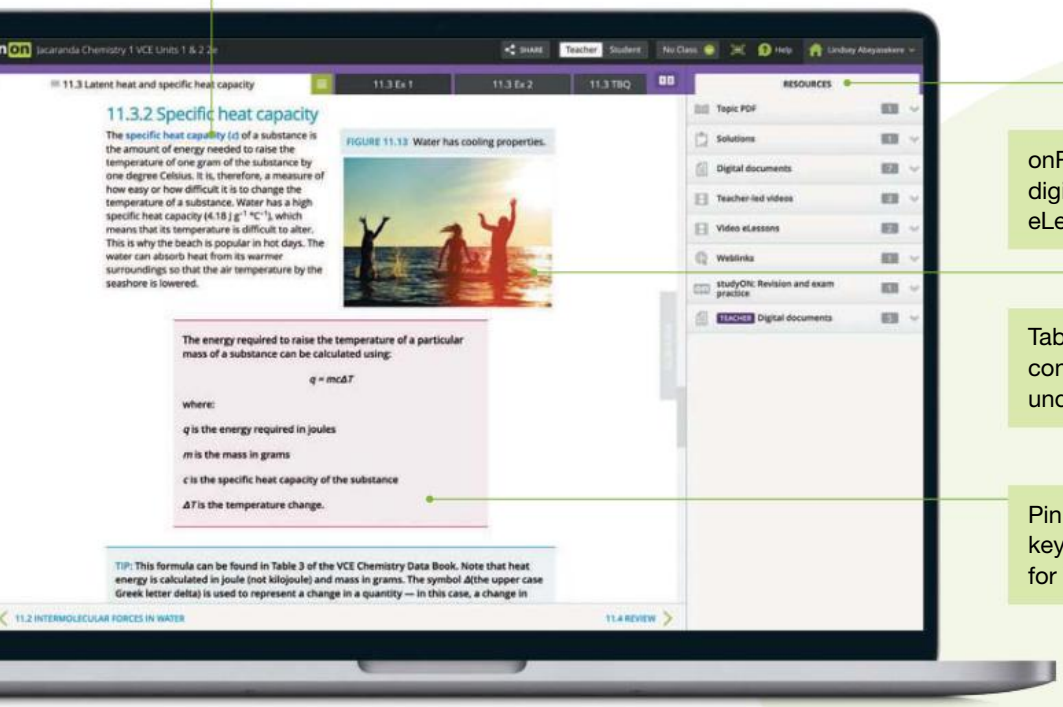
Online, these new editions are the complete package

Trusted Jacaranda theory, plus tools to support teaching and make learning more engaging, personalised and visible.

Each subtopic is linked to Key Knowledge (and Key Science Skills) from the VCE Chemistry Study Design.



Interactive glossary terms help develop and support scientific literacy.

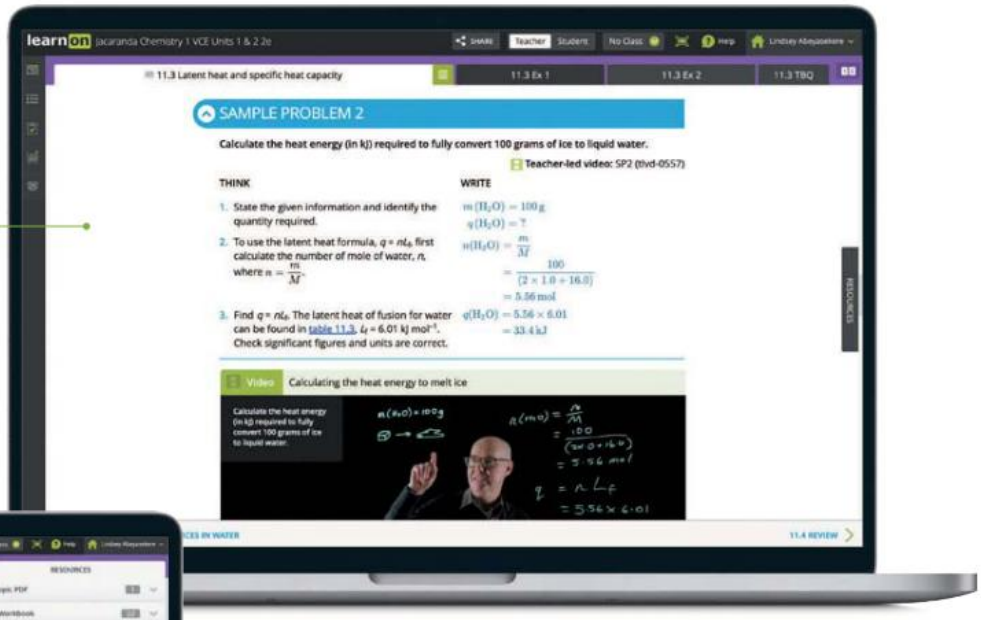


onResources link to targeted digital resources including video eLessons and weblinks.

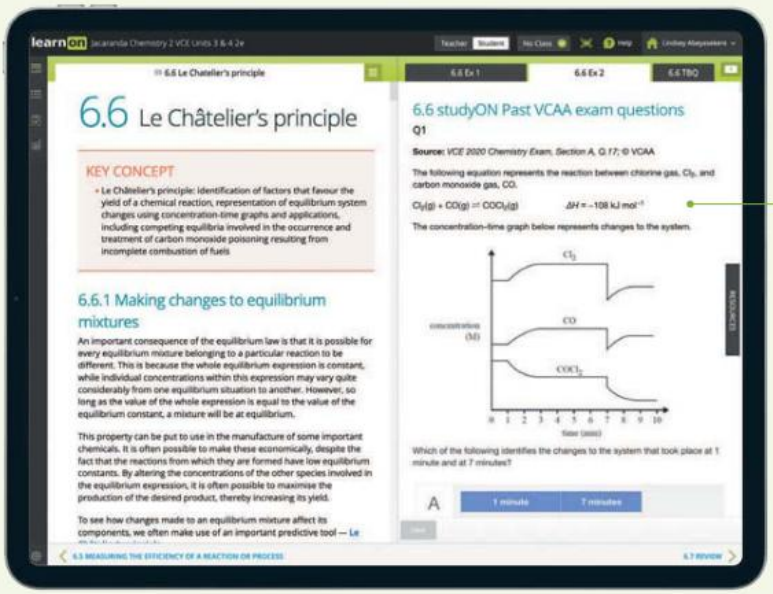
Tables and images break down content, allowing students to understand complex concepts.

Pink highlight boxes summarise key information and provide tips for VCE Chemistry success.

Sample problems break down the process of answering questions using a think/write format and a supporting teacher-led video.



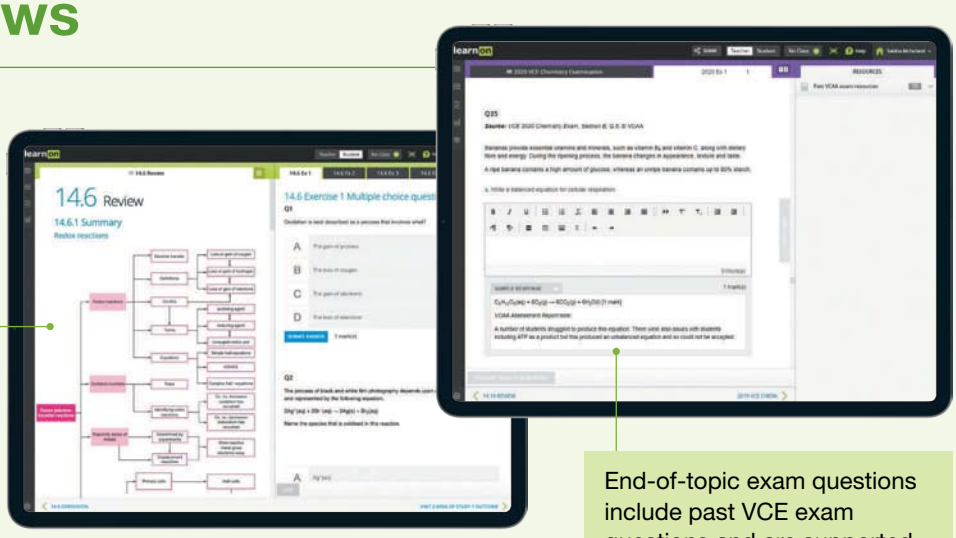
Practical investigations are highlighted throughout topics, and are supported by teacher-led videos and downloadable student and teacher version eLogbooks.



- Online and offline question sets contain practice questions and past VCAA exam questions with exemplary responses and marking guides.
- Every question has immediate, corrective feedback to help students to overcome misconceptions as they occur and to study independently — in class and at home.

Topic reviews

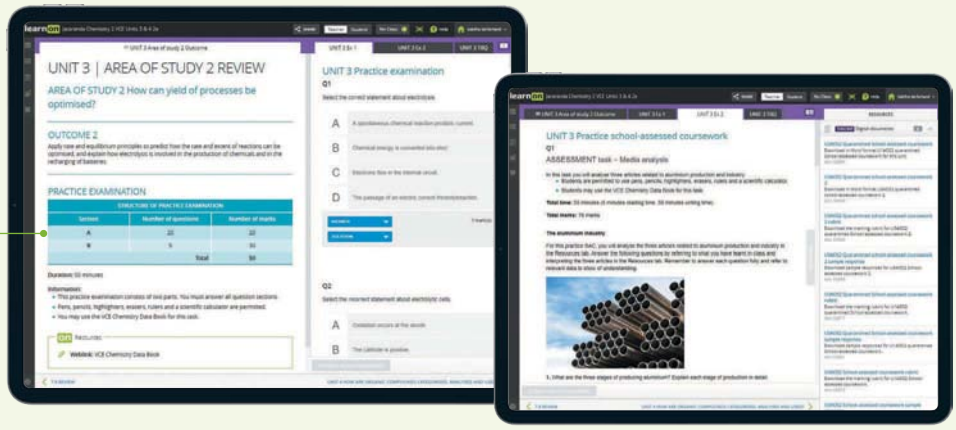
A summary flowchart shows the interrelationship between the main ideas of the topic. This includes links to both Key Knowledge and Key Science Skills.



End-of-topic exam questions include past VCE exam questions and are supported by teacher-led videos.

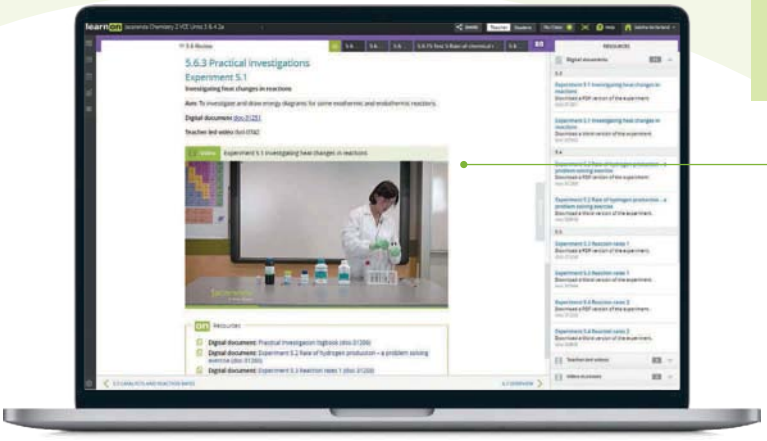
Area of Study reviews

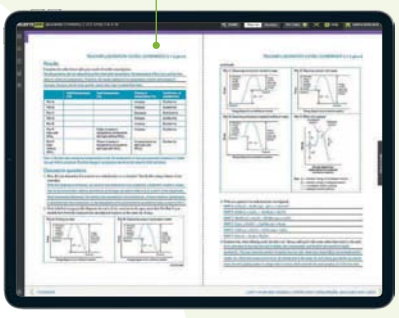
Areas of study reviews include practice examinations and practice SACs with worked solutions and sample responses. Teachers have access to customisable quarantined SACs with sample responses and marking rubrics.



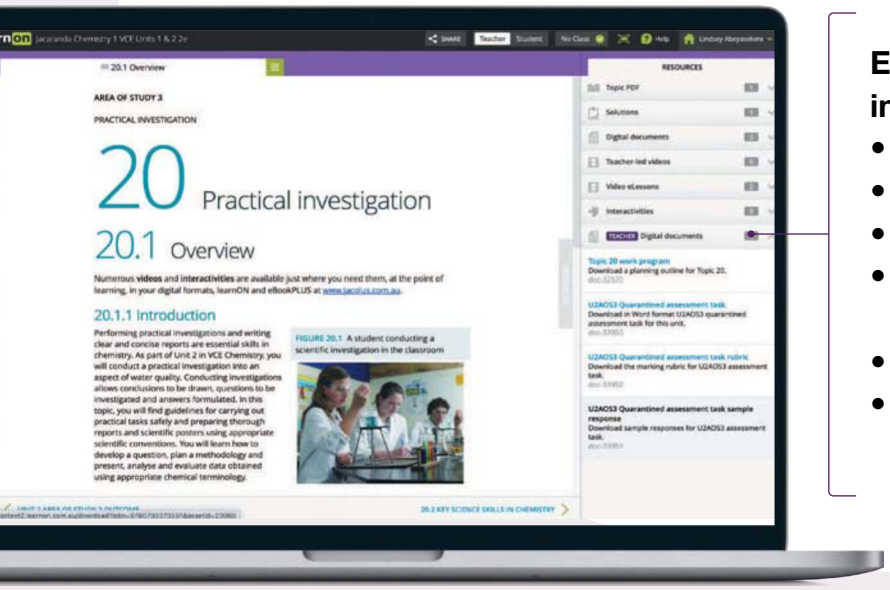
Practical investigation eLogbook

Enhanced practical investigation support includes practical investigation videos and an eLogbook with fully customisable practical investigations — including teacher advice and risk assessments.





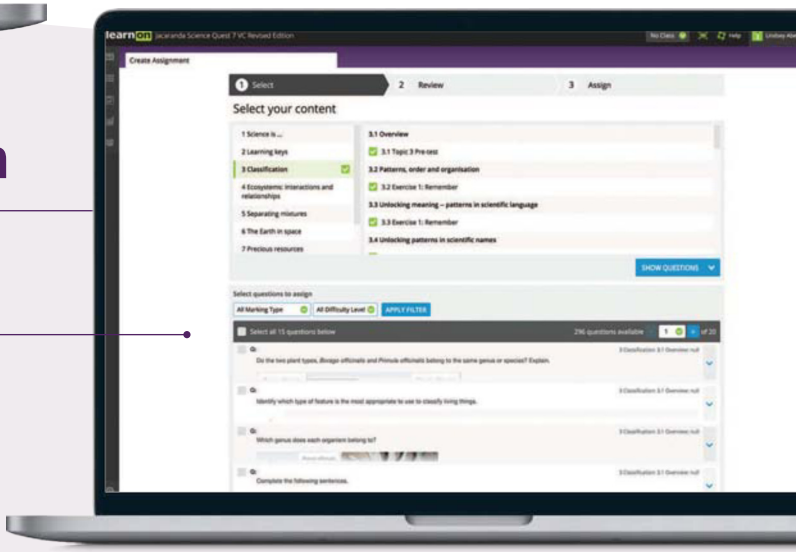
A wealth of teacher resources



Enhanced teacher support resources, including:

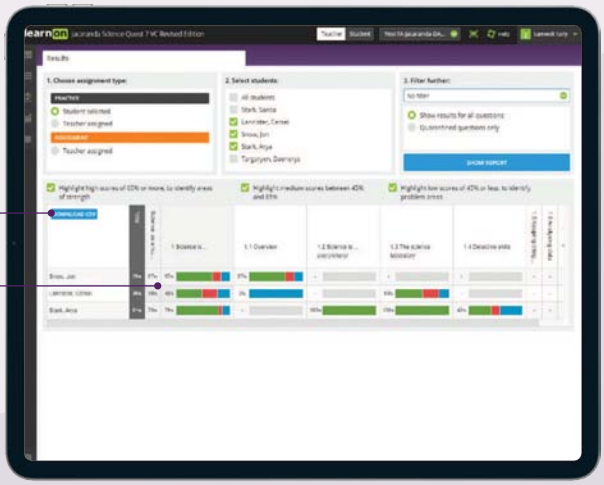
- work programs and curriculum grids
- teaching advice
- additional activities
- teacher laboratory eLogbook, complete with solution and risk assessments
- quarantined topic tests (with solutions)
- quarantined SACs (with worked solutions and marking rubrics).

Customise and assign



A testmaker enables you to create custom tests from the complete bank of thousands of questions (including past VCAA exam questions).

Reports and results



Data analytics and instant reports provide data-driven insights into progress and performance within each lesson and across the entire course.

Show students (and their parents or carers) their own assessment data in fine detail. You can filter their results to identify areas of strength and weakness.

Acknowledgements

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UNIT 1

How can the diversity of materials be explained?

AREA OF STUDY 1

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

OUTCOME 1

Explain how elements form carbon compounds, metallic lattices and ionic compounds, experimentally investigate and model the properties of different materials, and use chromatography to separate the components of mixtures.

1 Elements and the periodic table 3
2 Covalent substances 59
3 Reactions of metals 113
4 Reactions of ionic compounds 145
5 Separation and identification of the components of mixtures 187

AREA OF STUDY 2

How are materials quantified and classified?

OUTCOME 2

Calculate mole quantities, use systematic nomenclature to name organic compounds, explain how polymers can be designed for a purpose, and evaluate the consequences for human health and the environment of the production of organic materials and polymers.

6 Quantifying atoms and compounds 225
7 Families of organic compounds 259
8 Polymers and society 309

AREA OF STUDY 3

How can chemical principles be applied to create a more sustainable future?

OUTCOME 3

Investigate and explain how chemical knowledge is used to create a more sustainable future in relation to the production or use of a selected material.

9 Research investigations ..... online only

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1 Elements and the periodic table

KEY KNOWLEDGE

In this topic you will investigate:

Elements and the periodic table

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

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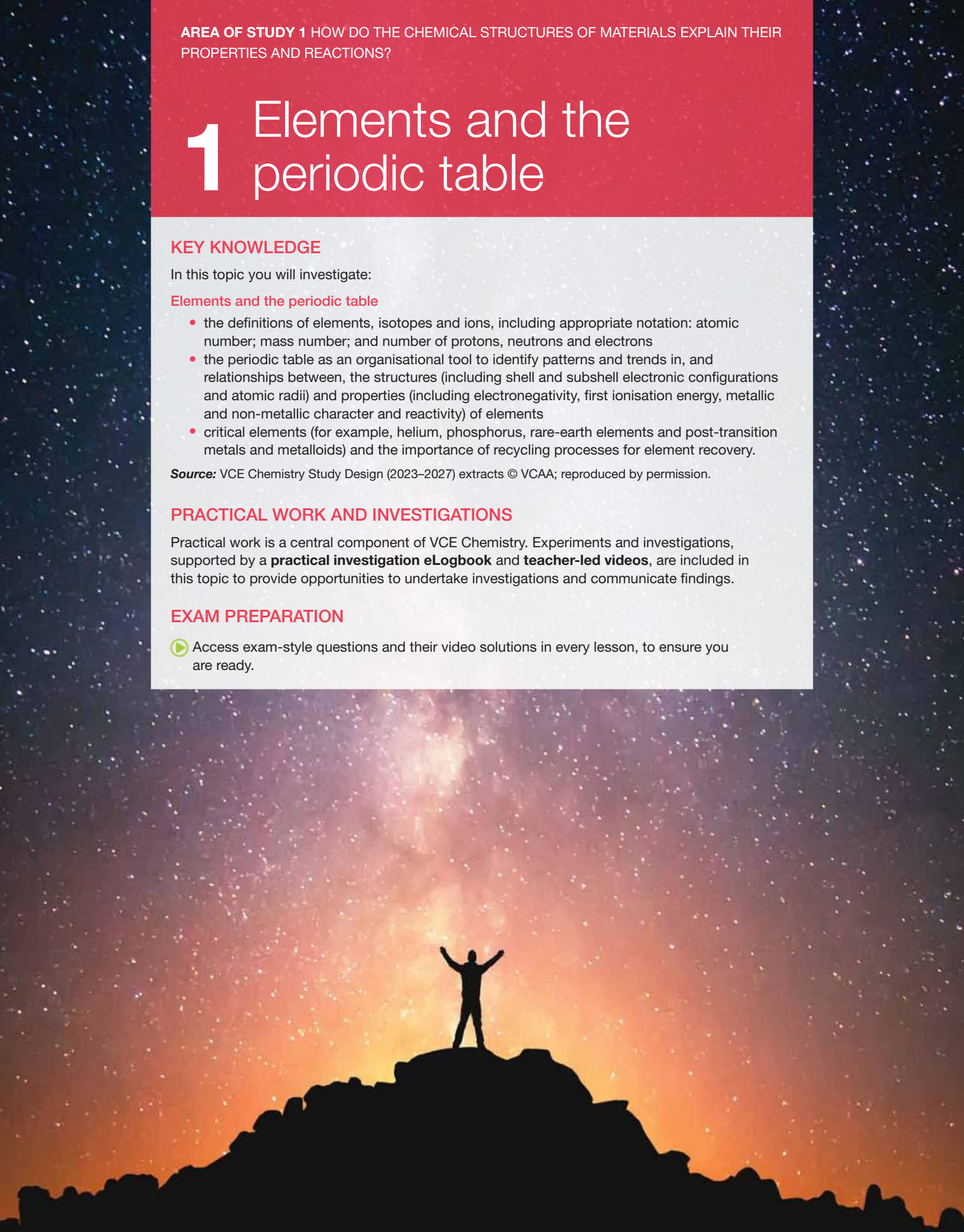
PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of VCE Chemistry. Experiments and investigations, supported by a **practical investigation eLogbook** and **teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

EXAM PREPARATION

Access exam-style questions and their video solutions in every lesson, to ensure you are ready.



1.1 Overview

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1.1.1 Introduction

Life is a mystery. Where did we come from? What are we made of? Scientists tell us that we are made of very small particles called atoms and that these atoms have their origin in stars. What a remarkable journey these atoms must have undertaken while being recycled over the billions of years since the origin of the universe. All matter is made up of atoms. Every material thing that you can see, smell and touch, that occupies space and has mass, is a form of matter.

FIGURE 1.1 Everything in the universe that has mass is composed of atoms.

Studying the structure and behaviour of matter — of which life, Earth and the universe are composed — has been ongoing. This topic introduces the fundamental structure and size of the building blocks of our universe, and how we have refined our theories to help us better understand our world. Atoms consist of even smaller subatomic particles and are amazingly 99.9 per cent empty space. There are 118 different atoms, known collectively as elements, which chemists organise into the periodic table. The periodic table is an indispensable tool, with remarkable patterns in its arrangement, that scientists use to predict the ways in which elements behave and react.

LEARNING SEQUENCE

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1.2 Elements 5
1.3 Electrons 12
1.4 The periodic table 25
1.5 Trends in the periodic table 32
1.6 Critical elements 43
1.7 Review 53

on Resources

Solutions	Solutions — Topic 1 (sol-0800)
Practical investigation eLogbook	Practical investigation logbook — Topic 1 (elog-1606)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 1 (doc-37067) Key idea summary — Topic 1 (doc-37068)
Exam question booklet	Exam question booklet — Topic 1 (eqb-0082)

1.2 Elements

KEY KNOWLEDGE

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

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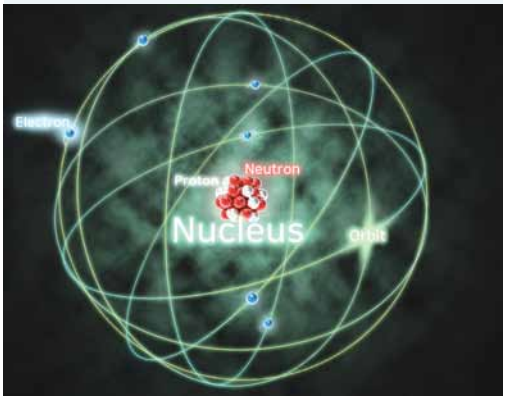
1.2.1 The structure of atoms

The atomic theory attempts to explain the structure of materials. According to this theory, all matter is made of **atoms**. Atoms are so small that it was not until 1981 that their images could finally be seen using the newly invented scanning tunnelling microscope. Due to their incredibly small size, models have developed to represent the internal structure of atoms.

One very useful model is the nuclear model of the atom proposed by Ernest Rutherford in 1911.

FIGURE 1.2 A modern take on New Zealand–born Ernest Rutherford's nuclear model of an atom.



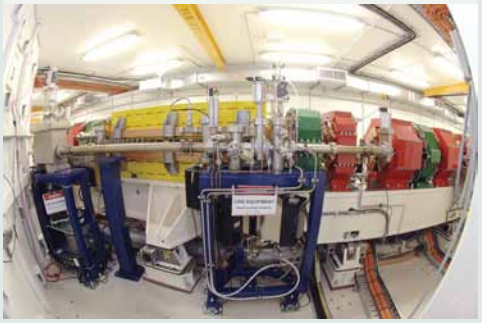
Rutherford's descriptions of an atom include:

- An atom is mostly empty space.
- An atom has a dense central structure called a nucleus.
- The nucleus, though its volume is very small relative to the atom as a whole, contains most of the mass of the atom.
- The nucleus is made up of positively charged particles called protons.
- The simplest nucleus is that of the hydrogen atom, which contains just one proton.
- All other atoms have nuclei that also contain neutrons.
- A neutron has no charge but has virtually the same mass as a proton.
- The empty space around the nucleus contains negatively charged particles called electrons.
- Electrons move very rapidly around the nucleus in orbits.
- Each electron has a definite energy and moves in a specific energy level. The mass of an electron is very much less than that of a proton or a neutron.

atom a neutral particle with a nucleus; the smallest constituent of an element

EXTENSION: Australia's particle accelerator

FIGURE 1.3 The Australian Synchrotron



The Australian Synchrotron is a particle accelerator in Victoria that is used by scientists to investigate the structure of matter. With over 4000 research visits per year, the fields of investigation are diverse and include agricultural science, environmental science, minerals analysis, medical investigations, materials science, cultural heritage, nanotechnology and forensics.

From the outside, the Australian Synchrotron resembles a football stadium. Inside, however, instead of footballs going in different directions, electrons are accelerated around a large loop (with a circumference of 216 metres) at almost the speed of light. The light is produced by high-energy electrons that are deflected into circular orbit by the 'synchronised' application of strong magnetic fields. The light produced is 1 million times brighter than the sun. The light, X-rays and infrared radiation produced is directed to a number of experimental workstations where many different experiments take place.

Protons, electrons and neutrons are called **subatomic particles**. Figure 1.4 and table 1.1 summarise the properties of these particles. Atoms that are neutrally charged have the same number of electrons and protons.

FIGURE 1.4 Subatomic particles in a nitrogen-14 atom

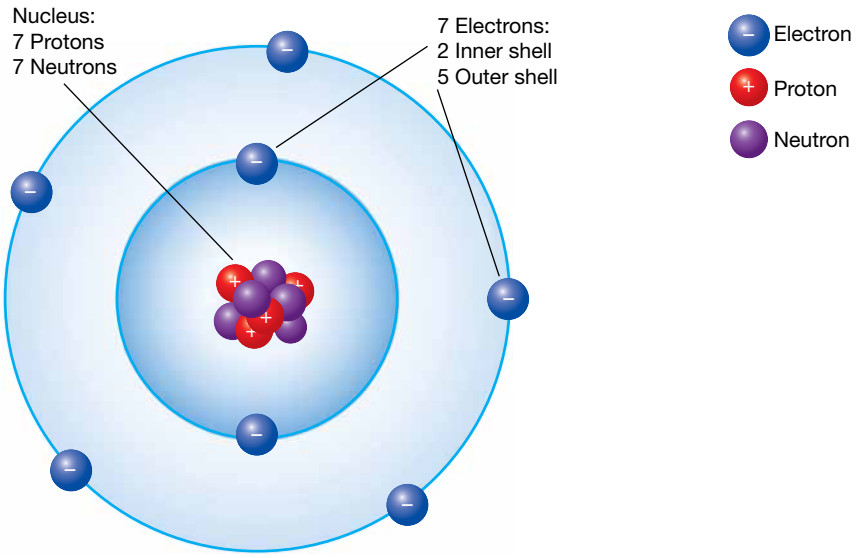


TABLE 1.1 Particles in an atom and their properties

Subatomic particle	Relative mass	Relative charge	Location
Electron	$\frac{1}{1837} = 0.0005$	–1	Outside nucleus
Proton	1	+1	In nucleus
Neutron	1	0	In nucleus

1.2.2 Elements

Atoms are not all the same. To date, chemists have identified 118 different types of atoms.

Elements are substances that contain only one type of atom. For example, pure oxygen contains only oxygen atoms and pure lead contains only lead atoms. Elements are defined by the number of protons in the nucleus.

The atoms of each element are classified based on the number of subatomic particles they have. Very few elements exist as individual atoms; examples are helium and neon.

subatomic particles particles in atoms: electrons, protons and neutrons

element a pure chemical species consisting of atoms of a single type

Elements

- Elements are substances that contain only one type of atom.
- An element is defined by the number of protons in the nucleus.
- Atoms contain the subatomic particles protons, neutrons and electrons.

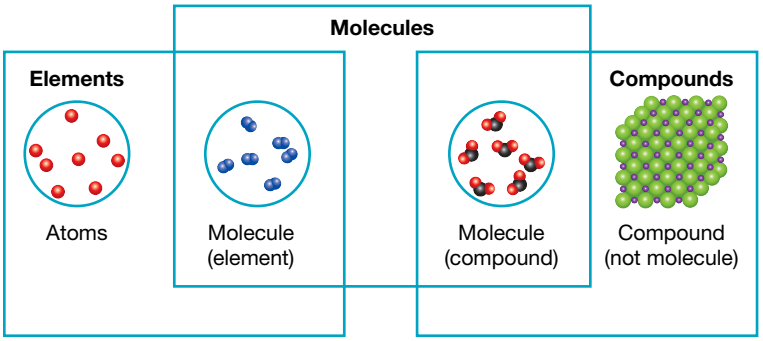
1.2.3 Atoms, elements, molecules and compounds

Molecules are substances that consist of two or more atoms that are chemically combined.

They contain either the same elements as in hydrogen gas, H_2, and oxygen gas, O_2, or different elements as in carbon dioxide, CO_2, and water, H_2O.

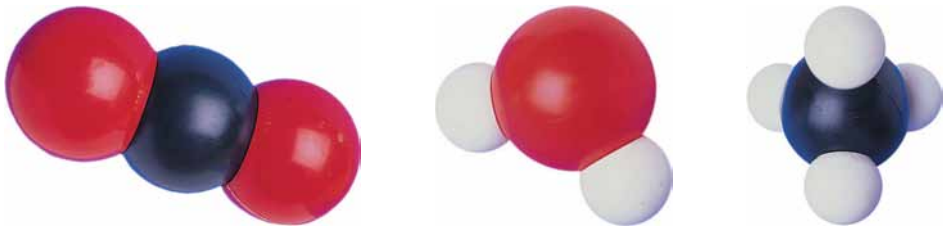
Compounds are substances that contain two or more elements but not all compounds are molecules. Some elements (such as carbon, forming diamond and graphite) and compounds exist as continuous lattice structures, and these are discussed in topic 2.

FIGURE 1.5 Elements, atoms, molecules and compounds



Molecules consist of two or more atoms chemically combined.

FIGURE 1.6 Space-filling models of molecules of (a) carbon dioxide, CO_2, (b) water, H_2O, and (c) methane, CH_4



Carbon dioxide, CO_2 (a)

Water, H_2O (b)

Methane, CH_4 (c)

1.2.4 Representing elements

Elements are represented by an element **symbol**, and with the atomic number and the mass number. Most symbols for elements come from the first letter or two letters of their names; for example, C for carbon and Cd for cadmium. Some atoms have symbols that have originated from a Greek or Latin name; for example, Au is the symbol for gold because gold was known in the past by its Latin name, aurum.

Atomic number (symbol *Z*)

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

molecule group of atoms bonded together covalently

compound substance consisting of two or more elements

symbol simplified representation of an element consisting of one or two letters

atomic number the number of protons in the nucleus of an atom of a particular element

Mass number (symbol *A*)

- The **mass number** (symbol *A*) is defined as the total number of protons and neutrons in an atom of an element.
- Protons have approximately the same mass as neutrons. The electron's mass is negligible compared with protons and neutrons. Therefore, the mass of an atom depends only on the number of particles in the nucleus.
- The elements are arranged in the periodic table in order of increasing atomic number. The relative atomic mass of each element is also shown on the table. Relative atomic mass is discussed in topic 7.

Isotopic symbols

An element is commonly represented as follows:

mass number $\rightarrow A$
atomic number $\rightarrow Z$ $\boldsymbol{E}$ $\leftarrow$ symbol for element

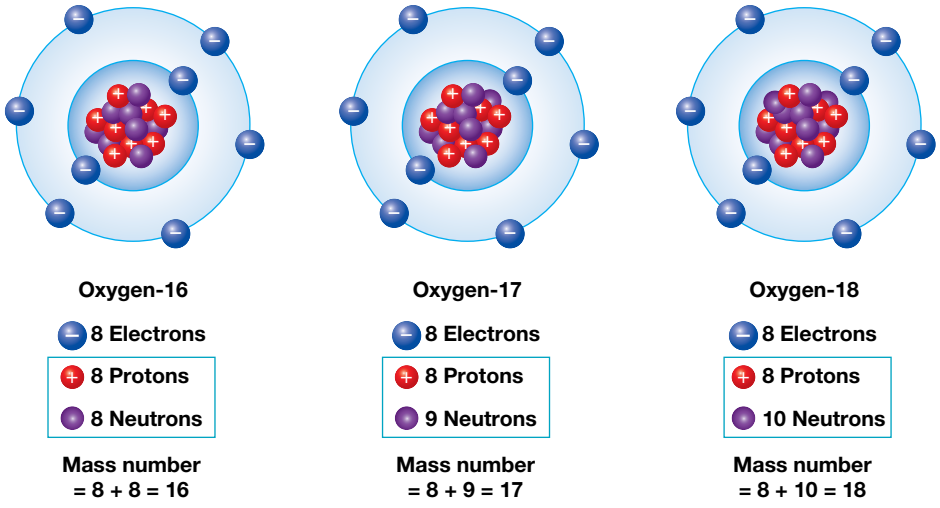
This is known as the **isotopic symbol** of an element. We can determine the number of neutrons in an atom by subtracting the atomic number, *Z*, from the mass number, *A*. For example, sodium, Na, has atomic number 11 and mass number 23, and can be represented as $^{23}_{11}\text{Na}$. An atom of sodium,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

1.2.5 Isotopes

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

Naturally occurring oxygen consists of three isotopes: $^{16}_{8}\text{O}$, $^{17}_{8}\text{O}$ and $^{18}_{8}\text{O}$. Isotopes are named by their element name followed by their mass number to distinguish them; for example, the isotopes of oxygen are oxygen-16, oxygen-17 and oxygen-18 (see figure 1.7). Aluminium has only one isotope, aluminium-27, $^{27}_{13}\text{Al}$.

FIGURE 1.7 Isotopes of oxygen



mass number the total number of protons and neutrons in the nucleus of a particular isotope of an element

isotopic symbol representation of an element as $^{A}_{Z}E$, where *E* is the symbol for the element, *A* is the mass number and *Z* is the atomic number

isotopes forms of an element with the same number of protons but different numbers of neutrons in the nucleus

Isotopes

Isotopes are atoms of the same element that have different numbers of neutrons; that is, they have the same atomic number but different mass numbers.

Isotopes can be used to date archaeological and geological features. Radiocarbon dating using the decay of carbon-14 isotopes is used to date organic material within (or nearby) Indigenous Australian rock art, for example.

FIGURE 1.8 The oldest reliably dated rock art in Australia is 28 000 years old.

tlvd-0509

SAMPLE PROBLEM 1 Representing isotopes using appropriate notation

a. Write the symbols for the atoms nitrogen-14 and nitrogen-15.
b. How many protons does each atom have?
c. How many neutrons does each atom have?
d. What are these atoms called?
e. Write the isotopic notation for each isotope of nitrogen.

THINK	WRITE
a. The symbol will not change, regardless of the atomic masses.	N
b. The number of protons of an element will not change, regardless of the atomic masses. The atomic number of N is 7, hence it has 7 protons.	7
c. The number of neutrons can be determined using the following: Number of neutrons = mass number (A) – atomic number (Z).	nitrogen-14: $14 - 7 = 7$ nitrogen-15: $15 - 7 = 8$
d. The name for atoms with the same number of protons but with different masses (different number of neutrons).	Isotopes
e. This must include the symbol, mass and atomic numbers in the correct format.	nitrogen-14: $^{14}_{7}\text{N}$ nitrogen-15: $^{15}_{7}\text{N}$

PRACTICE PROBLEM 1

a. Write the elemental symbol for the atoms of hydrogen-1, hydrogen-2 and hydrogen-3.
b. How many protons does each atom have?
c. How many neutrons does each atom have?
d. Write the isotopic notation for each species of hydrogen.

1.2 Activities

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1.2 Quick quiz on | 1.2 Exercise | 1.2 Exam questions

1.2 Exercise

1. MC An atomic particle has a net charge of zero and is found in the nucleus. Identify which type of particle it is.
   - A. Proton
   - B. Electron
   - C. Neutron
   - D. Positron
2. MC The particle that is represented by the symbol $^{121}_{51}Sb^{3+}$ has
   - A. 51 protons, 48 electrons and 121 neutrons.
   - B. 51 protons, 48 electrons and 70 neutrons.
   - C. 51 protons, 54 electrons and 121 neutrons.
   - D. 51 protons, 54 electrons and 70 neutrons.
3. Look up your periodic table to find the atomic number of each of the following elements.
   - a. H
   - b. Ne
   - c. Ag
   - d. Au
4. An atom has 13 protons and 14 neutrons. Identify the following.
   - a. Its atomic number
   - b. Its mass number
   - c. Its name
5. Find the symbols for elements with the following atomic numbers.
   - a. 5
   - b. 12
   - c. 18
   - d. 20
6. In the element argon, $Z = 18$ and $A = 40$. For argon, state the following.
   - a. The number of neutrons
   - b. The number of electrons
   - c. The isotopic symbol for this element
7. Determine the number of protons, neutrons and electrons in $^{79}_{35}Br^-$.
8. An atomic nucleus consists of one proton and one neutron. What is its isotopic symbol?

9. a. Complete the following table.

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

b. Identify any isotopes in the table.
c. Explain the difference between the isotopes.

10. Why do we identify an element by its atomic number rather than its mass number?

1.2 Exam questions

Question 1 (1 mark)

MC Which of the following species has a different number of neutrons from the rest?

A. $^{64}_{30}Zn$
B. $^{62}_{28}Ni$
C. $^{63}_{29}Cu$
D. $^{69}_{31}Ga$

Question 2 (36 marks)

Complete the following table.

Atom	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons	Name of element
$^{23}_{11}Na$						
$^{19}_{9}F$						
$^{28}_{14}Si$						
$^{56}_{26}Fe$						
$^{197}_{79}Au$						
$^{235}_{92}U$						

Question 3 (2 marks)

What is the general name for the group of atoms that includes carbon-12, carbon-13 and carbon-14? Identify the similarities and differences between these atoms.

Question 4 (3 marks)

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

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

Question 5 (1 mark)

Tellurium is element 52 and iodine is element 53. Explain why iodine atoms have less mass than less than tellurium atoms.

More exam questions are available in your learnON title.

1.3 Electrons

KEY KNOWLEDGE

- The structures (including shell and subshell electronic configurations and atomic radii) of elements
- The definitions of ions

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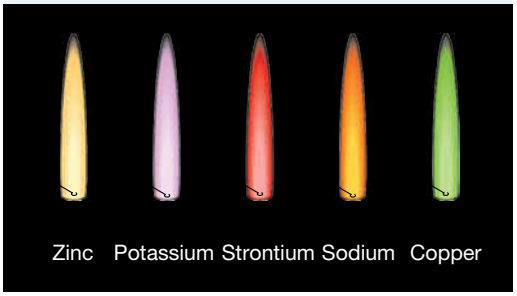
1.3.1 Exciting electrons

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

When white light is separated by a prism, a *continuous spectrum* of colour is observed. Every element emits light if it is heated by passing an electric discharge through its gas or vapour. This happens because the atoms of the element absorb energy and then lose it, emitting it as light. Passing the light emitted by an element through a prism produces an **atomic emission spectrum** for that element.

FIGURE 1.9 When wires with small amounts of different metal salts are placed in a flame, the electrons are excited and emit characteristic coloured light.



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

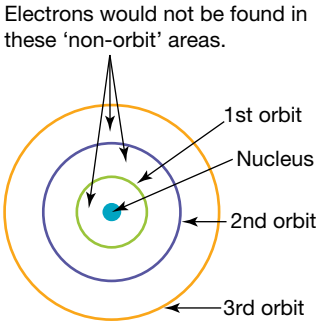
1.3.2 Bohr's energy levels

In 1913, Niels Bohr suggested an explanation for the emission spectrum by proposing a model for the hydrogen atom. His model proposed the following:

- Electrons of specific energy move around the central nucleus in circular orbits or energy levels. Electrons cannot exist between these orbits.
- Although an electron cannot lose energy while orbiting a nucleus, it can be given excess energy (by a flame or electric current) and then move to a higher orbit. If this happens, the electron has moved from the **ground state** (lowest energy level) to an **excited state**.
- When an electron drops back down to a lower, more stable orbit, the excess energy is given out as a **photon** or quantum of light. This is seen as a line of a particular colour on the visible spectrum.
- The energy given out is the difference in energy between the two energy levels. Since only certain allowed energy levels are possible, the energy released has specific allowed values, each corresponding to a line in the emission spectrum. This spectrum is different for each element, so it is often called the 'fingerprint' by which an element may be identified.

Some metallic elements can be identified simply by their characteristic flame colours when heated in a Bunsen burner flame. Copper burns with a blue–green flame, for example, and sodium burns with a yellow–orange flame.

FIGURE 1.10 The Bohr model of an atom



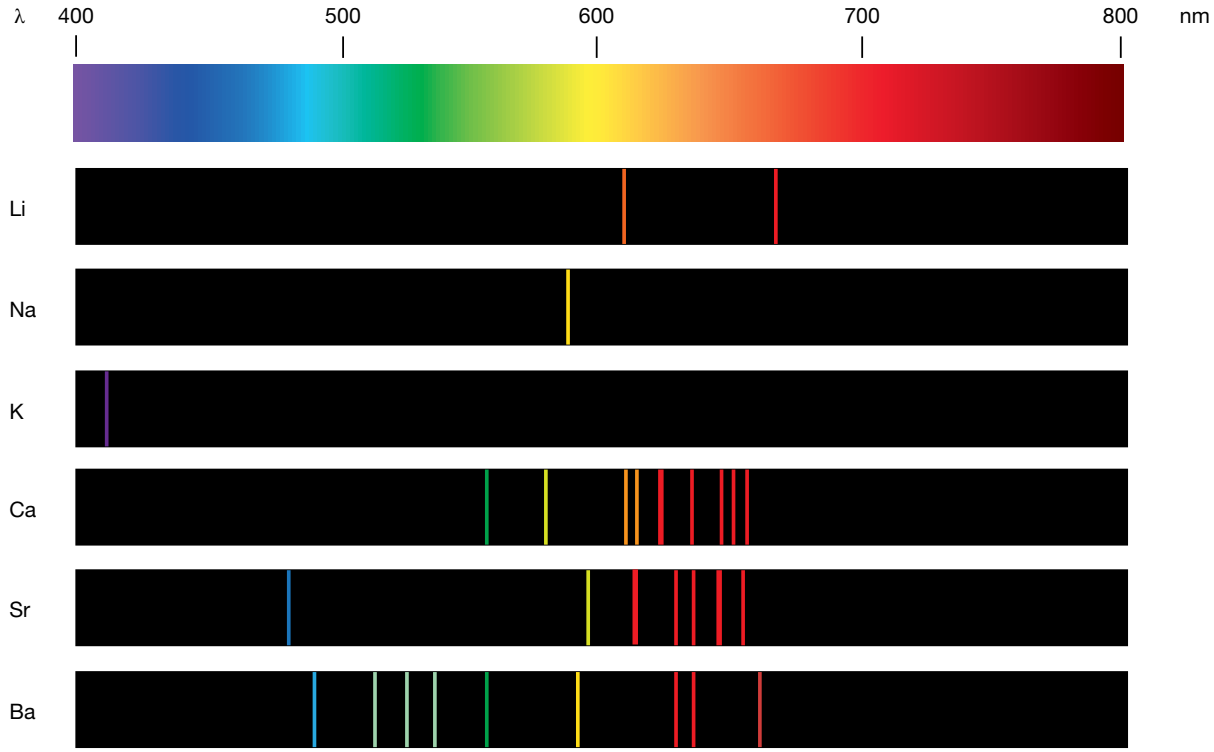
atomic emission spectrum a spectrum emitted as distinct bands of light of diagnostic frequencies by elements or compounds

ground state the least excited state of an atom, where the electrons are occupying the lowest possible energy levels

excited state raised to a higher than ground-state energy level

photon particle of light

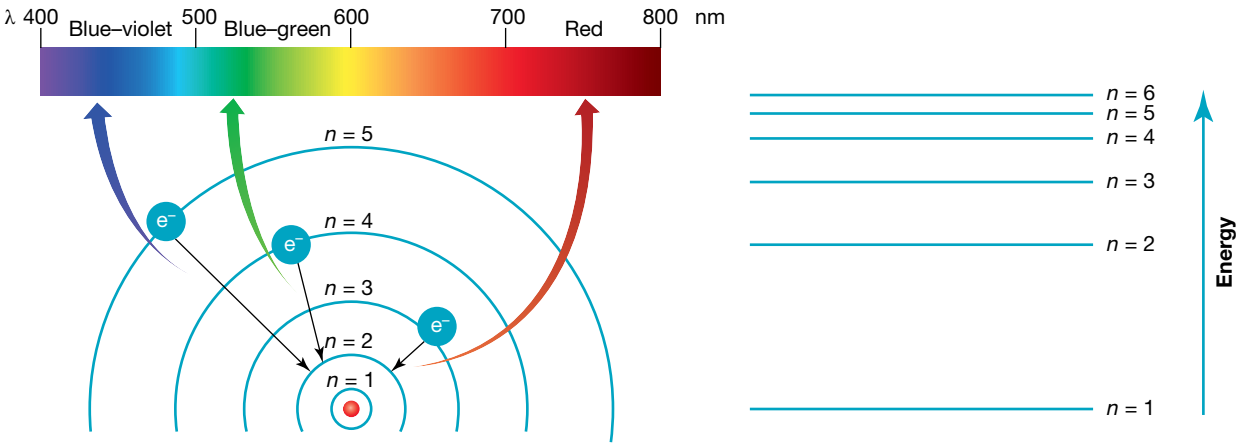
FIGURE 1.11 White light is a continuous spectrum (top). The emission spectra of various atomic elements consist of distinct lines that correspond to differences in energy levels.



Emission spectra and electron energy levels

- When a particular amount of energy is supplied to an atom, an electron can move from a lower energy level to a higher energy level. When the electron returns to a lower state, it emits a photon of energy equal to the difference of energy between the two levels.
- Atomic emission spectra provide evidence that electrons exist in specific energy levels.

FIGURE 1.12 Emission spectrum and energy levels; each electron transition produces a line of a different colour, with blue being the highest energy and shortest wavelength, λ.



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EXPERIMENT 1.1 online only

Flame tests of metal cations

Aim

To observe the characteristic flame colours of the metal ions K^+, Na^+, Li^+, Sr^{2+}, Cu^{2+}, Ca^{2+}, Ba^{2+}, and to identify an unknown metal ion

on Resources

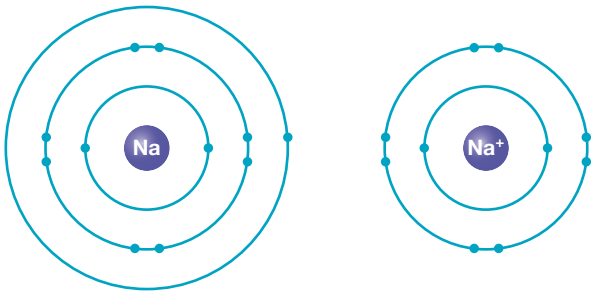
Weblink Bohr model

Electron shells

Electrons may be visualised as moving within a region of space surrounding the nucleus. The regions are called electron shells and are numbered 1, 2, 3 and 4. A definite energy level is associated with each shell; the innermost shell ($n = 1$) has the lowest energy level.

- To move further away from the nucleus, an electron must gain energy.
- If it gains enough energy to completely leave the atom, the particle that is left is no longer neutral and is called a positive ion. Sodium, Na, has 11 protons and 11 electrons. If it loses an outer shell electron, it becomes the positive ion Na^+ because it now has 11 protons and only 10 electrons.

FIGURE 1.13 Sodium easily loses an outer shell electron to become a sodium ion.



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

1.3.3 Electron configuration

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

Keep in mind the following when determining the electron configuration:

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

electron configuration the arrangement of electrons in the shells of an atom

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

Valence electrons

Valence electrons are the electrons in the outer shell of atoms.

Shell model diagrams

The electron configuration of an atom can be represented using shell model diagrams, such as those in figure 1.14. These show the electron shells as concentric rings around the nucleus, with the electrons marked on each ring, and help us to visualise the structure and behaviour of atoms.

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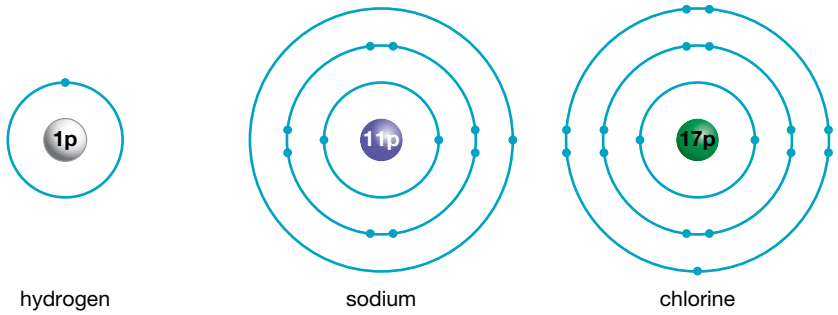
FIGURE 1.14 Shell model diagrams of hydrogen, sodium and chlorine; dots represent the electrons in the shells.



Limitations of the shell model

The shell model represents only part of the story of the atom. More discoveries are always being made that cause scientists to reconsider their models and their understanding of the atom.

The limitations of the shell model are as follows:

- The shell model doesn't really explain the various differences in energies between the electron shells. It seems to imply that all the electrons orbit the nucleus in exactly circular paths, like planets around a sun. We know from looking at molecules with electron tunnelling microscopes that they come in many different shapes and sizes, so this model does not fully explain every aspect of every atom.
- The order of the electrons filling the electron shells is not really explained by this model either. For example, compare calcium (2, 8, 8, 2) with scandium (2, 8, 9, 2) — why isn't the electron configuration of scandium 2, 8, 8, 3? Other models have been developed that are more complex and explain more of the data scientists have gathered.

1.3.4 From atoms to ions

Atoms gain or lose electrons to achieve more stable outer shell configurations; they are then called **ions**. When an atom becomes an ion, it is no longer neutrally charged, since the number of electrons is not equal to the number of protons (see table 1.2). Note that the number of protons remains the same. An atom that has lost electrons becomes *positively* charged and is called a **cation** (e.g. Na^+). An atom that has gained electrons becomes negatively charged and is called an **anion** (e.g. Cl^-).

valence electrons electrons in the outermost shell of an atom; largely determine chemical properties of an element and contribute to chemical bond formation

ion an atom that has lost or gained electrons and so has a charge

cation a positively charged ion

anion a negatively charged ion

FIGURE 1.15 All elements in the periodic table are neutral but some can become charged when they gain or lose electrons.

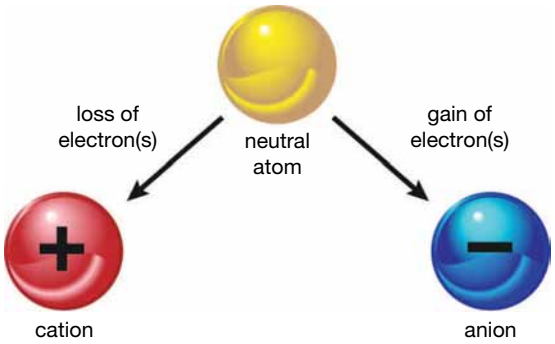


TABLE 1.2 Common atoms and their ions

Atom/ion	Symbol	Number of protons	Number of electrons
Sodium atom	Na	11	11
Sodium cation	Na^+	11	10
Chlorine atom	Cl	17	17
Chloride anion	Cl^-	17	18

Ions

Ions are formed when an atom gains or loses electrons; in other words, the atom becomes charged.

Metallic ion formation

The metallic elements are those on the left side of the mauve staircase in figure 1.16 (except hydrogen). Metals tend to lose electrons to *achieve a noble gas configuration in their outer shells*.

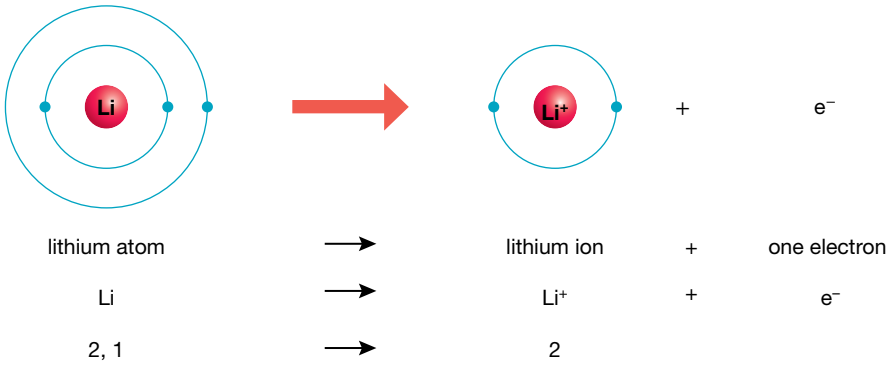
FIGURE 1.16 Periodic table (up to element 89) showing the division between metals and non-metals

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

Note: Numbers are used to identify periods and groups. Roman numerals are no longer used.

For example, lithium is a very reactive group 1 metal with one outer shell electron and the electron configuration 2, 1. In order to obtain the stable configuration of a full outer shell, the lone electron is lost (see figure 1.17). The electron configuration, 2, of the nearest noble gas, helium, results. Since the lithium cation has three protons but only two electrons, it has a net charge of +1. Charges are written as superscripts above and to the right of the element symbol; thus the lithium atom is now written as Li^+. This process can be represented by electron shell diagrams or in the simple equation form in figure 1.17.

FIGURE 1.17 The lithium atom has one valence electron, which it loses to form the lithium cation, which has a charge of +1.



Note: When an atom's net charge is +1 or −1, it is not necessary to include the numeral 1 in the superscript notation.

Consider the following groups and periods:

- The group 2 and group 13 metals contain two and three valence electrons respectively. They lose their outer shell electrons to form ions with charges of +2 and +3 respectively.
- The periods 2, 3 and 4 form simple ions with electron configurations identical to those of the closest noble gases. Each occupied energy shell contains the maximum number of electrons. Examples of their electron configurations are as follows:

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

When we name a metallic ion, we use the full name of the metal followed by the word 'ion' to distinguish it from the uncharged metal. *Note:* The group 14 elements, carbon and silicon, do not form simple ions.

Metal ions

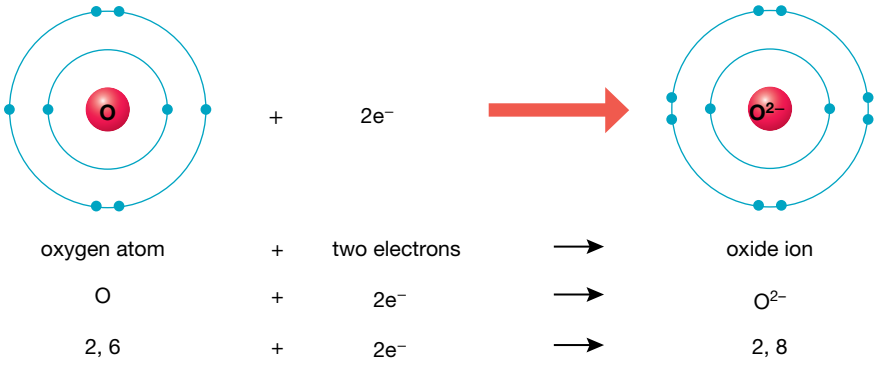
Metals form positive ions.

Non-metallic ion formation

Non-metallic elements are shown on the right side in the purple section of the periodic table in figure 1.16. They *gain* electrons to achieve a noble gas configuration of eight electrons in their outer shells (except for hydrogen).

For example, oxygen in group 16 has six outershell electrons and has the electron configuration 2, 6. It is too difficult to remove all six electrons to achieve a full outer shell, so the oxygen atom gains two electrons instead to become a stable anion, O^{2-}, as shown in figure 1.18 and the simple equation.

FIGURE 1.18 The oxygen atom has six valence electrons, and gains two electrons to form the oxide ion, which has a charge of –2.



An anion has more electrons than a neutral atom of the same element; here, the oxygen ion has eight protons and ten electrons, resulting in a net charge of –2. The electron configuration of the oxygen anion is now that of a neon atom: 2, 8. The oxygen atom has become an oxide ion. (It is a convention in chemistry to indicate the ions of non-metallic elements with the suffix *-ide*.)

Non-metal ions

Non-metals form negative ions.

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SAMPLE PROBLEM 2 Identifying the ions formed from atoms

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

a. **Mg** **b.** **S**

THINK

a. Mg is a metal found in group 2. It has the simplified electron configuration 2, 8, 2. In order to become stable, the Mg atom needs to lose two electrons. The charge of the resultant ion would, therefore, be +2.

b. S is a non-metal found in group 16. It has the simplified electron configuration 2, 8, 6. In order to become stable, the S atom needs two more electrons to complete its outer shell. The charge of the resultant ion would, therefore, be –2.

WRITE

Mg^{2+}
magnesium ion

S^{2-}
sulfide ion

PRACTICE PROBLEM 2

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

a. **Ga**
b. **P**

1.3.5 The quantum-mechanical model of the atom

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

FIGURE 1.19 Schrödinger developed the quantum-mechanical model of the atom.

In 1923 the French scientist Louis de Broglie (1892–1987) proposed, using Albert Einstein's and Max Planck's quantum theory, that electrons show both particle and wave behaviour. In 1926, the Austrian physicist Erwin Schrödinger (1887–1961) made use of the new quantum theory to refine Bohr's model of the atom. He wrote and developed a wave equation describing the location and energy of an electron in a hydrogen atom. From this, he developed the currently accepted quantum-mechanical model of the atom — a complex mathematical model based on particles such as electrons showing wave-like behaviour. This model features electrons arranged in shells, subshells and orbitals within an atom.

According to quantum mechanics, the electron is not considered as moving along a definite path. Instead, the electron is found in a region of space around a nucleus called an **orbital**. An orbital may be visualised as a blurry cloud of negative charge; the cloud is most dense where the probability of finding the electron is large and less dense where the probability of finding the electron is small. As in the Bohr model, the electron is attracted to the nucleus by electrostatic forces and moves in such a way that its total energy has a specific value.

FIGURE 1.20 The quantum-mechanical model showing **(a)** the charge cloud for the 1*s* electron in hydrogen and **(b)** charge clouds for the electrons in the 1*s* and 2*s* subshells.

(a) (b)

In the quantum-mechanical model of the atom:

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

orbital describes a region where an electron is likely to be found, with a maximum capacity of two electrons

principal quantum number, *n* a number related to the energy of an electron and distance from the nucleus

TABLE 1.3 Energy levels within shells of an atom

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

1.3.6 Subshells electron configuration

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

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

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

Notice that the 4*s* subshell is filled before the 3*d* subshell, which is of a higher energy than the 4*s* subshell. Likewise, the 4*d* subshell is higher in energy than the 5*s* subshell, and so on. This is demonstrated in figure 1.22, which shows the order of filling of subshells found by following the arrows from tail to head, starting with the top arrow.

Remember the order of subshells

To remember the subshells letters, use the mnemonic '**S**mart **P**eople **D**on't **F**orget'.

FIGURE 1.21 The energy levels of atomic orbitals

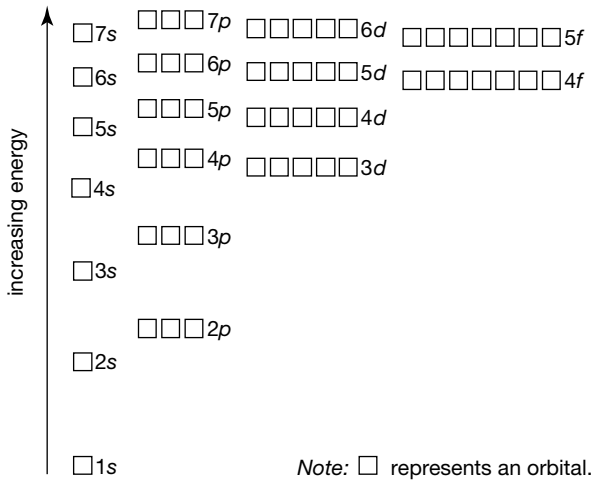
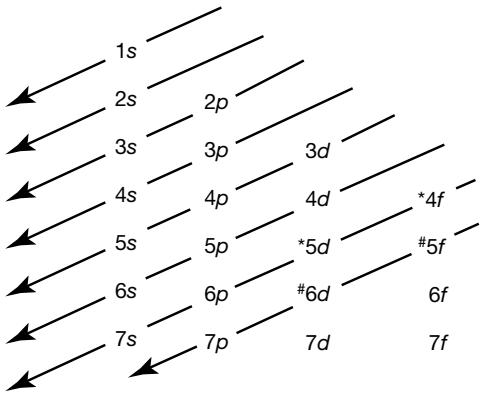


FIGURE 1.22 The order of filling of subshells

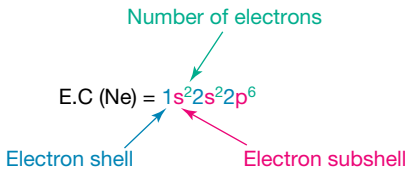


* and #: In both these cases it is found that the two subshells fill at the same time.

Excited states

When an electron moves to a higher energy level than the ground state by absorbing energy, its electron configuration changes. The outermost electron moves to a higher energy level subshell.

For example, neon has two electrons in the first shell and eight in the second shell, E.C. (Ne) = $1s^2 2s^2 2p^6$; when the outermost electron gains energy, this becomes E.C. (Ne*) = $1s^2 2s^2 2p^5 3s^1$. An asterisk is used to indicate and excited atom.



An atom is in an excited state when an electron temporarily occupies an energy state greater than its ground state.

Note that, once the order of filling subshells has been determined, the subshells are written in increasing numerical order, rather than the order of increasing energy.

For titanium, the ground state electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$, which becomes $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^2$.

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

H ($Z = 1$)	$1s^1$
Na ($Z = 11$)	$1s^2 2s^2 2p^6 3s^1$
Sc ($Z = 21$)	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s^2$

The electron configuration for scandium or other larger atoms can also be written in a condensed form using the previous noble gas as a basis. For example, E.C. (Sc) = [Ar]$3d^1 4s^2$, where [Ar] refers to the ground state electron configuration of argon.

The position of an element on the periodic table can also be easily read from the electron configuration. The periodic table is divided into groups (the horizontal rows of the table) and periods (the vertical columns of the table). The significance of the groups and periods, and their relationship to shells and subshells, is discussed in subtopic 1.4.

on Resources

Weblink Interactive periodic table

tlvd-0510

SAMPLE PROBLEM 3 Finding the ground state electron configuration

Find the ground state electron configuration of a fluorine atom.

THINK	WRITE
1. The atomic number of fluorine is 9. Fluorine has nine electrons. According to the order of subshell filling, the 1*s* subshell in a fluorine atom fills first, and contains two of fluorine's nine electrons.	$1s^2$
2. The next energy level is the 2*s* subshell — this holds another two electrons.	$1s^2 2s^2$
3. The 2*p* subshell can hold six electrons; however, since only five electrons remain to be placed (9 – 2 – 2 = 5), one of the 2*p* subshell orbitals is incomplete.	$1s^2 2s^2 2p^5$

PRACTICE PROBLEM 3

State the ground state electron configuration of a potassium atom.

tlvd-0511

SAMPLE PROBLEM 4 Finding the ground state electron configuration of an ion

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

THINK	WRITE
1. An aluminium ion, Al^{3+}, is an aluminium atom ($Z = 13$) that has lost three electrons. Hence $13 - 3 = 10$ electrons need to be placed in shells. The 1*s* shell contains two electrons.	$1s^2$
2. The next energy level is the 2*s* subshell — this holds another two electrons.	$1s^2 2s^2$
3. The 2*p* subshell can hold six electrons and because six electrons remain to be placed ($10 - 2 - 2 = 6$), the 2*p* subshell orbitals is complete.	$1s^2 2s^2 2p^6$

PRACTICE PROBLEM 4

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

$$\mathbf{1}s^2 \mathbf{2}s^2 \mathbf{2}p^6 \mathbf{3}p^1$$

1.3.7 Chromium and copper: atypical electron configurations

A few elements have electron configurations that do not follow the usual pattern. Chromium (atomic number 24) and copper (atomic number 29), for example, may be expected to be written as follows:

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

Copper $1s^2 2s^2 2p^6 3s^2 3p^6 3d^9 4s^2$

The correct electron configurations are:

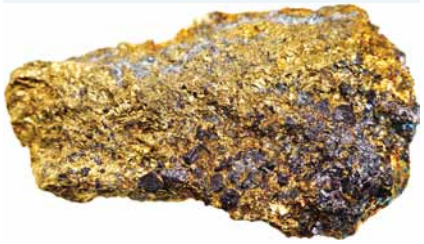
Chromium $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$

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

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

FIGURE 1.23 Chromite is an ore of chromium.

FIGURE 1.24 Chalcopyrite is an ore of copper.



1.3 Activities

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1.3 Quick quiz on | 1.3 Exercise | 1.3 Exam questions

1.3 Exercise

1. Write the full electron configuration of the following elements: lithium, potassium, carbon, chlorine, argon and nitrogen.
2. Use electron shell diagrams and simple equations to show how the following atoms form their corresponding ions.
   a. The metal atoms Ca and Al
   b. The non-metal atoms N and F
3. For each of the following atoms, predict the charge of the ion formed, and write its name and symbol.
   a. Ba
   b. K
   c. P
   d. Cl
   e. S
4. The isotope $^{32}_{15}P$ is used in the treatment of leukaemia.
   a. Write the ground state electron configuration of this isotope.
   b. How does it differ from $^{31}_{15}P$?
5. Name the elements with the following electron configurations.
   a. $1s^2 2s^1$
   b. $1s^2 2s^2 2p^6$
   c. $1s^2 2s^2 2p^6 3s^1$
   d. $1s^2 2s^2 2p^6 3s^2 3p^1$
   e. $1s^2 2s^2 2p^6 3s^2 3p^6$
   f. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
6. For each of the elements in the following two sets of atomic numbers, write the electron configuration.
   a. $Z = 20, 12, 4, 9$
   b. $Z = 5, 6, 8, 16$
7. Distinguish between the terms ‘shell’ and ‘subshell’.
8. Which of the following electron configurations are ground state configurations and which are excited state configurations?
   a. $1s^2 2s^2 2p^6 3s^1$
   b. $1s^2 2s^2 2p^6 3s^2 3p^4 3d^1$
   c. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^2$
9. A neutral magnesium atom has an electron configuration of $1s^2 2s^2 2p^6 3s^1 6s^1$.
   a. How can you tell that the atom is in an excited state?
   b. Describe what would happen if the atom changed its electron configuration to the ground state.
   c. Write the ground state electron configuration of the magnesium ion Mg^{2+}.
10. Identify the following elements with the provided ground state electron configurations.
    a. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^2$
    b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

1.3 Exam questions

Question 1 (1 mark)

MC What is the electron configuration for Ti^{2+} written as?

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

Question 2 (1 mark)

MC Consider the following ground state electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6$.

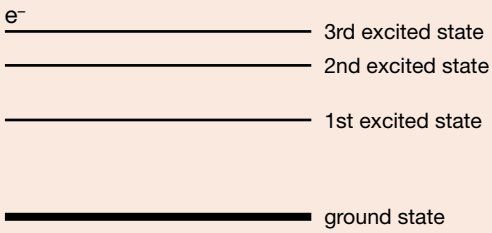
What does this best represent?

A. Ca^{2+}
B. Cl^-
C. Ar
D. All of the above

Question 3 (2 marks)

Each unique electron transition results in a line in the emission spectrum of an element.

a. What type of transition results in an emission spectrum? **(1 mark)**
b. How many lines would arise from an electron in the third excited state as represented by the following diagram? **(1 mark)**



Question 4 (5 marks)

Consider the following particles (shown with abstract symbols).

$$^{40}_{18}Q,\ ^{32}_{16}R,\ ^{38}_{19}T,\ ^{38}_{18}X,\ ^{40}_{20}Z$$

a. Which two particles are isotopes of the same element? **(1 mark)**
b. Which particle has the largest number of neutrons? **(1 mark)**
c. Write the electron configuration for particle *Z*. **(1 mark)**
d. Write the electron configuration for the ion formed by particle *T*. **(1 mark)**
e. Name the particle(s) that is/are filling the *p* subshell. **(1 mark)**

Question 5 (1 mark)

Germanium has an atomic number of 32. In its ground state, how many of its shells are occupied by electrons?

More exam questions are available in your learnON title.

1.4 The periodic table

KEY KNOWLEDGE

- The periodic table as an organisational tool to identify patterns and trends in, and relationships between, the structures (including electronic configurations) of elements

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BACKGROUND KNOWLEDGE: Mendeleev and the periodic table

The periodic table of chemical elements is one of the most significant achievements in science, capturing the essence of not only chemistry, but also physics and biology. It is a unique tool, enabling scientists to predict the appearance and properties of matter on the Earth and in the rest of the universe. The year 1869 is considered the time Dmitri Mendeleev synthesised centuries of previous experiments and constructed the modern Periodic System. In 2019 we saw the 150th anniversary of the Periodic Table of Chemical Elements and so this year was proclaimed the International Year of the Periodic Table of Chemical Elements by the United Nations General Assembly and UNESCO.

Dmitri Mendeleev (1834–1907), a Russian chemist, spent years collecting and sorting information about each of the 63 elements known at the time and constructed a set of data cards (one data card for each element). On each card he noted the atomic mass and other properties of the element and its compounds. He noticed the existence of 'groups' of different elements with similar chemical properties. He then produced a periodic table according to increasing order of relative atomic mass and the periodicity of their properties; that is, similar physical and chemical properties occurring at regular intervals. He even left gaps for elements that, he correctly predicted, had not yet been discovered.

FIGURE 1.25 An early version of Mendeleev's periodic table

1.4.1 The modern periodic table

The periodic table of the elements has become one of the most important icons in science today. A chart of this table hangs on the wall of almost every science classroom and chemical laboratory in schools, universities and research institutions around the world. It is a single document that consolidates much of our knowledge of chemistry and is a vital tool for modern chemists. The modern periodic table now consists of 118 known elements. Elements up to and including uranium (atomic number 92) are naturally occurring. All the elements beyond uranium have been synthesised by chemists since 1940. They are all radioactive elements and are called the transuranium elements. These elements do not occur in nature but have been synthesised in nuclear reactors, which accelerate bombarding particles to very high speeds and shoot them at very heavy target nuclei.

FIGURE 1.26 Particle tracks like this provide evidence that a new particle has been produced. Darmstadtium, Ds, and roentgenium, Rg (elements 110 and 111) were first discovered in 1994 from the fusion of lead and other elements in a heavy ion accelerator.

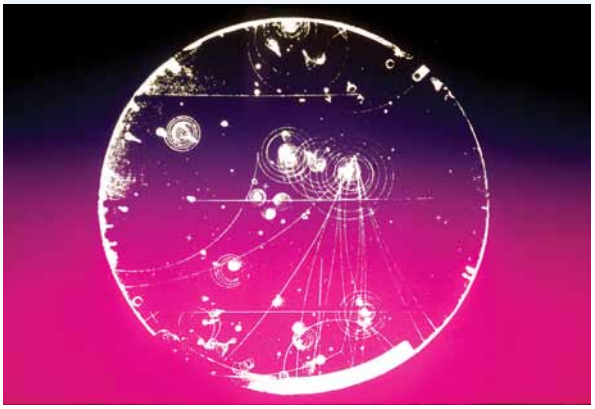
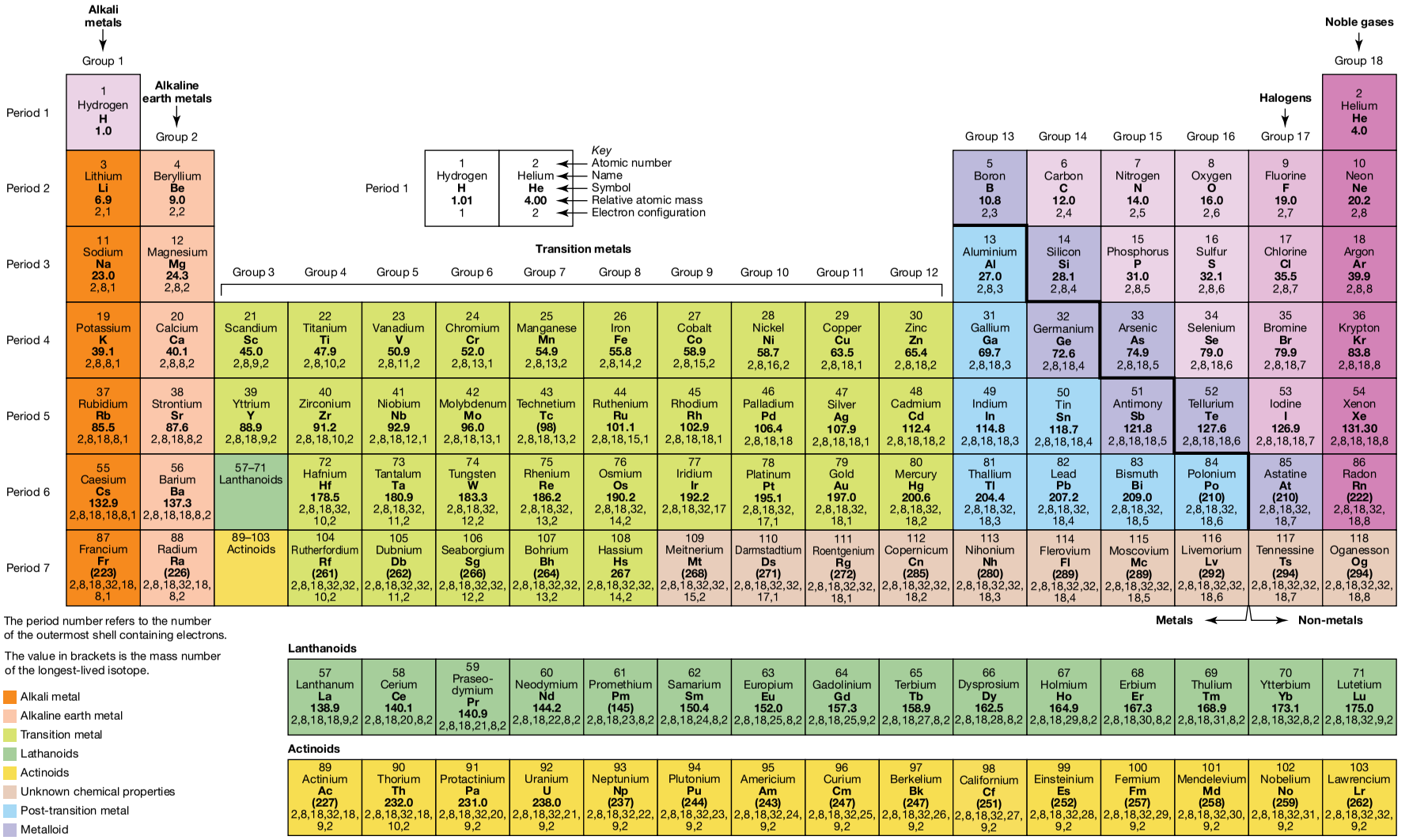


FIGURE 1.27 Periodic table showing relative atomic masses and electron shell configurations



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

The value in brackets is the mass number of the longest-lived isotope.

- Alkali metal
- Alkaline earth metal
- Transition metal
- Lathanoids
- Actinoids
- Unknown chemical properties
- Post-transition metal
- Metalloid
- Reactive non-metal
- Halide
- Noble gas

Lanthanoids

57 Lanthanum La 138.9 2,8,18,18,9,2	58 Cerium Ce 140.1 2,8,18,20,8,2	59 Praseo-dymium Pr 140.9 2,8,18,21,8,2	60 Neodymium Nd 144.2 2,8,18,22,8,2	61 Promethium Pm (145) 2,8,18,23,8,2	62 Samarium Sm 150.4 2,8,18,24,8,2	63 Europium Eu 152.0 2,8,18,25,8,2	64 Gadolinium Gd 157.3 2,8,18,25,9,2	65 Terbium Tb 158.9 2,8,18,27,8,2	66 Dysprosium Dy 162.5 2,8,18,28,8,2	67 Holmium Ho 164.9 2,8,18,29,8,2	68 Erbium Er 167.3 2,8,18,30,8,2	69 Thulium Tm 168.9 2,8,18,31,8,2	70 Ytterbium Yb 173.1 2,8,18,32,8,2	71 Lutetium Lu 175.0 2,8,18,32,9,2

Actinoids

89 Actinium Ac (227) 2,8,18,32,18,9,2	90 Thorium Th 232.0 2,8,18,32,18,10,2	91 Protactinium Pa 231.0 2,8,18,32,20,9,2	92 Uranium U 238.0 2,8,18,32,21,9,2	93 Neptunium Np (237) 2,8,18,32,22,9,2	94 Plutonium Pu (244) 2,8,18,32,23,9,2	95 Americium Am (243) 2,8,18,32,24,9,2	96 Curium Cm (247) 2,8,18,32,25,9,2	97 Berkelium Bk (247) 2,8,18,32,26,9,2	98 Californium Cf (251) 2,8,18,32,27,9,2	99 Einsteinium Es (252) 2,8,18,32,28,9,2	100 Fermium Fm (257) 2,8,18,32,29,9,2	101 Mendelevium Md (258) 2,8,18,32,30,9,2	102 Nobelium No (259) 2,8,18,32,31,9,2	103 Lawrencium Lr (262) 2,8,18,32,32,9,2

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

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

The periodic table in figure 1.27 is based on the IUPAC periodic table and shows relative atomic masses. **IUPAC** is the International Union of Pure and Applied Chemistry and is the world authority on chemical naming and terminology. Relative atomic mass (A_r) is based on the carbon-12 atom, the most common isotope of carbon. This isotope is assigned a mass of exactly 12. On this scale, 1 is, therefore, equal to the mass of a carbon-12 atom. Values in brackets are for the most stable or best-known isotopes. This is explained further in topic 6.

1.4.2 Periods and groups in the periodic table

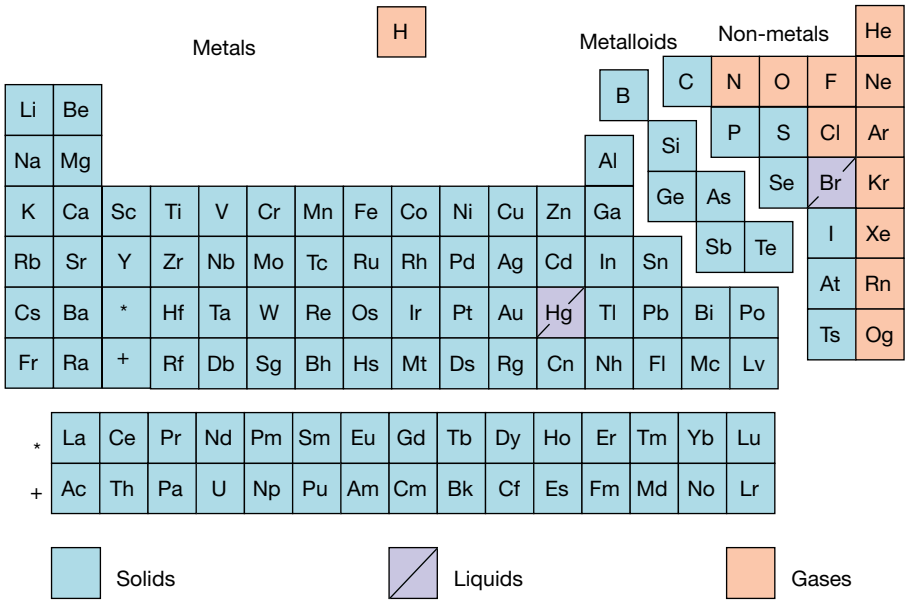
The seven horizontal rows in the periodic table are called **periods**. Each period corresponds to the filling of a shell. The location of an element in a period tells you the number of occupied shells each atom of that element has. Elements in the third period, for example, have three occupied shells.

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

Metals and non-metals in the periodic table

Elements may be classified as **metal** or **non-metal**. In the periodic table, the metals are found towards the left side of the stairway and the non-metals are found towards the right side of the stairway, as shown in figure 1.28. Most elements are metals, although some elements show both metallic and non-metallic characteristics. These elements are known as *metalloids*, and are also shown in figure 1.28.

FIGURE 1.28 Metals, non-metals and metalloids in the periodic table



IUPAC International Union of Pure and Applied Chemistry

period horizontal row of the periodic table

group vertical column of the periodic table

metal ductile, malleable, conducting element

non-metal non-ductile, non-malleable, non-conducting element

The elements in groups 3 to 12 include some familiar metals, such as iron, which is commonly used in construction, copper, which is particularly useful for electrical wiring, and gold, which is used in jewellery. These groups are known as the **transition elements**. Transition elements contain atoms with filling *d* subshells. These elements are metals and have the properties listed in table 1.4 but are less reactive than *s*-block metals (see section 1.4.4), and most form coloured compounds. In addition, most can form ions with different charges; for example, copper can form Cu^{+} ions and Cu^{2+} ions.

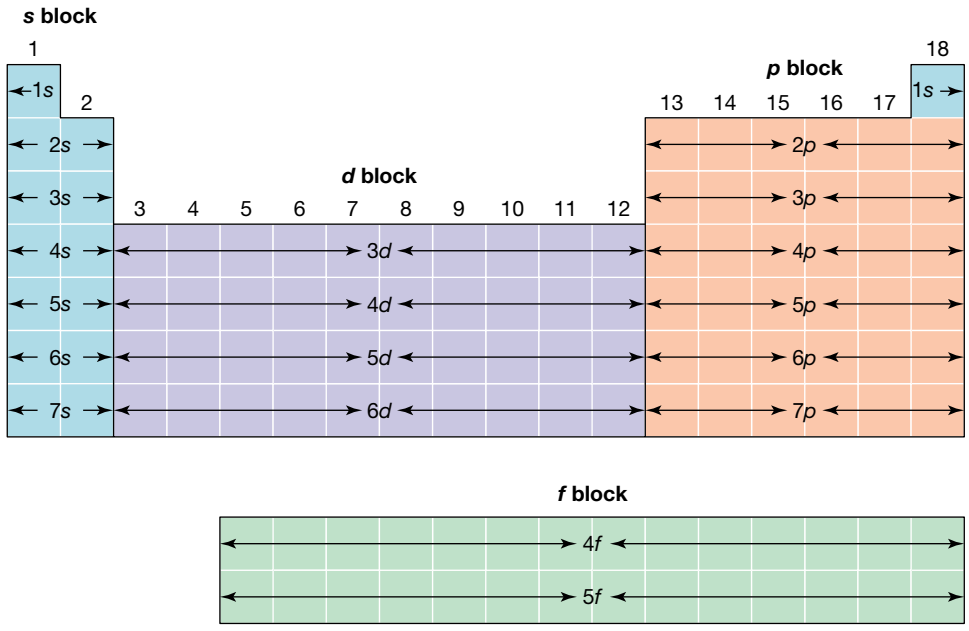
TABLE 1.4 General properties of metals and non-metals

Metals	Non-metals
High melting and boiling points	Low melting and boiling points
Good conductors of heat and electricity	Poor conductors of heat and electricity
Opaque	Transparent in a thin sheet
Shiny appearance	Dull colour
Ductile and malleable	Brittle when solid
Strong	Weak
Form positive ions (cations)	Form negative ions (anions)

1.4.3 Blocks of the periodic table

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

FIGURE 1.29 Each block of the periodic table corresponds to a particular subshell.



s-block elements

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

transition elements metallic elements in groups 3 to 12 with the *d* subshell being filled

p-block elements

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

d-block elements

The *d*-block elements, from group 3 to group 12, are the transition metals or transition elements. All elements in this block fill their electrons subshells in the order $1s < 2s < 2p < 3s < 3p < 4s < 3d$. For example, period 4 transition metal elements fill the 4*s* subshell before the 3*d* subshell. Similarly, electrons are lost from the 4*s* subshell before the 3*d* subshell. Their outermost shell electron configurations are d^1s^2 to $d^{10}s^2$.

f-block elements

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

1.4.4 Using electron shell configuration to identify an element

The best way to identify an element based on its electron shell configuration is to simply add up all of the electrons, because this result is equal to the atomic (proton) number of the element, *Z*. Using the atomic number, you can then find both the period and group. The highest shell number indicates the period number.

To determine the position of an atom, consider the following examples:

- $1s^22s^22p^63s^23p^64s^23d^6$ has 26 electrons, and so 26 protons. The element is iron. From the periodic table, we can see this element in period 4 and group 8.
- $1s^22s^22p^63s^23p^3$ has 15 electrons, and so 15 protons. The element is phosphorus. From the periodic table, we can see this element in period 3 and group 15.

To identify an ion

To determine the position of an ion, add or remove the corresponding number of electrons (depending on whether it is a positive or negative ion) to give the electronic configuration of the atom; this is then equal to the atomic (proton) number of the element. Using the atomic number, you can then again find both the period and group.

For example, consider an ion with a charge of −2 and an electron configuration of $1s^22s^22p^6$. Because the ion has gained two electrons (to become −2), simply subtract two electrons to give the electronic configuration of the atom. Therefore, $1s^22s^22p^6$ becomes $1s^22s^22p^4$, which has eight electrons and, therefore, eight protons, and the element is oxygen (period 3 and group 15). The original ion would be O^{2-} (oxide).

If an ion has a charge of +3 and an electron configuration of $1s^22s^22p^63s^23p^63d^5$, because the ion has lost three electrons to become +3, simply add three electrons to give the electronic configuration of the atom. Remember that the 4*s* subshell fills and empties before the 3*d* subshell. Therefore, $1s^22s^22p^63s^23p^63d^5$ becomes $1s^22s^22p^63s^23p^64s^23d^6$, which has 26 electrons and, therefore, 26 protons, and the element is iron (period 4 and group 8). The original ion would be Fe^{3+}.

tlvd-0512

SAMPLE PROBLEM 5 Identifying the period and group of an atom

An atom has the following electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$. Identify the period, group and name of the element.

THINK	WRITE
1. This is an atom, so the number of protons is equal to the number of electrons. Add the electrons to determine the atomic (proton) number, which can easily be found on the periodic table.	2 + 2 + 6 + 2 + 6 + 2 + 3 = 23 electrons
2. Identify the period and group by recognising the element in the periodic table.	The element is vanadium.The atom belongs to period 4 and group 5.

PRACTICE PROBLEM 5

An ion with a +3 charge has the following electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^3$. Identify the period, block and name of the element.

on Resources

Digital document Investigation: History of the periodic table (doc-30923)

Weblink Electron configuration

1.4 Activities

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1.4 Quick quiz on | 1.4 Exercise | 1.4 Exam questions

1.4 Exercise

1. MC Consider the elements *X* and *Y*, with the following ground electron configurations:

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^3 4s^2$ (*X*) $1s^2 2s^2 2p^6 3s^2$ (*Y*)

In the periodic table, what are these elements in?
   - A. The same group
   - B. The same period
   - C. The same block
   - D. Different groups, periods and blocks
2. Explain why the classification of elements into a periodic table may be useful to chemists.
3. Explain why elements in the same group have similar chemical properties.
4. Why does the periodic table have 10 groups of transition metals?

5. Identify the period and group to which each of the following elements belongs and name the element.
   - **a.** $1s^2 2s^2 2p^4$
   - **b.** $1s^2 2s^1$
   - **c.** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
   - **d.** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
   - **e.** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$
   - **f.** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$
6. Identify the period and group to which each of the following ions belongs and name the element.
   - **a.** An ion with a charge of +7 and the electronic configuration of $1s^2 2s^2 2p^6 3s^2 3p^6$
   - **b.** An ion with a charge of –1 and the electronic configuration of $1s^2 2s^2 2p^6 3s^2 3p^6$
7. Write the ground-state electron configuration for elements with the following atomic numbers, and state the group and period of each element.
   - **a.** 8 **b.** 13 **c.** 20 **d.** 35
8. Iron is a silvery metal and sulfur is a yellow solid.
   - **a.** Write the symbols for these elements.
   - **b.** Describe three differences between these elements.
   - **c.** Explain whether they are metals or non-metals.
   - **d.** State in which blocks of the periodic table they would be found.
9. An element has an electron configuration that ends with $s^2 p^3$. Identify the group this element belongs to.
10. Element *X* is in period 5, group 8. Write the electron configuration of neutral atoms of this element.

1.4 Exam questions

Question 1 (1 mark)

MC What do elements in the same period have in common?

- **A.** They have similar chemical properties.
- **B.** They have the same number of occupied shells.
- **C.** They have the same outershell configuration.
- **D.** They form the same number of isotopes.

Question 2 (10 marks)

- **a.** Describe where you would find metals in the periodic table. **(1 mark)**
- **b.** How many electrons do elements in group 1, group 2 and group 17 have in their outer shell? **(3 marks)**
- **c.** Are metals (in groups 1 and 2) likely to lose or gain electrons? Explain. **(2 marks)**
- **d.** Do metal ions form positive or negative ions? **(1 mark)**
- **e.** Name the three groups. **(3 marks)**

Question 3 (4 marks)

In which block of the periodic table are the following elements?

- **a.** Barium **(1 mark)**
- **b.** Copper **(1 mark)**
- **c.** Uranium **(1 mark)**
- **d.** Neon **(1 mark)**

Question 4 (6 marks)

Identify the following elements.

- **a.** $1s^2 2s^2 2p^6 3s^2 3p^2$ **(1 mark)**
- **b.** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ **(1 mark)**
- **c.** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$ **(1 mark)**
- **d.** [He] $2s^2 2p^4$ **(1 mark)**
- **e.** [Ne] $3s^2 3p^5$ **(1 mark)**
- **f.** [Ar] $3d^6 4s^2$ **(1 mark)**

Question 5 (3 marks)

Explain why the periodic table includes the following.

- **a.** Two elements in the first period **(1 mark)**
- **b.** Eight elements in the second period **(1 mark)**
- **c.** No transition elements in the first three periods **(1 mark)**

More exam questions are available in your learnON title.

1.5 Trends in the periodic table

KEY KNOWLEDGE

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

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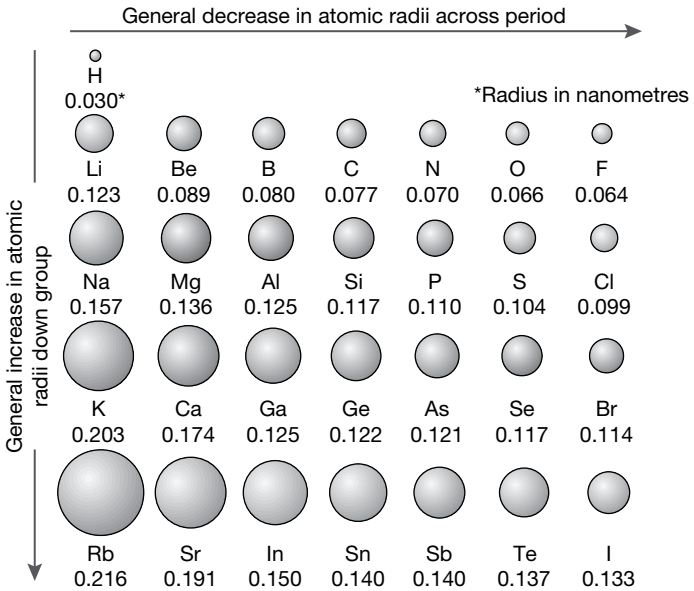
When discussing trends in the periodic table, the concept of **core charge** is relevant. Core charge is the effective nuclear charge experienced by an outer shell (valence) electron. It is calculated by subtracting the number of inner shell electrons from the number of protons and is always positive.

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

1.5.1 Atomic radii

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

FIGURE 1.30 Atomic radii in nanometres of selected elements



Atomic size increases down a group. When going down a group, the number of occupied electron shells increases, which means a greater distance exists between the nucleus and the outer electrons. Along with this comes an increased number of full inner shells and shielding from these electrons, which further decreases the attraction (and therefore closeness) of the outermost electrons to the nucleus.

Atomic size decreases going from left to right across a period. When moving across a period, the total number of occupied shells stays the same because each additional electron goes into the same shell. The effect of electron shielding also stays the same, because there are the same number of inner shell electrons. However, since the number of protons increases, the effective nuclear charge (core charge) increases, which means creates a greater attraction of the outermost electrons to the nucleus.

core charge the attraction of nuclear protons acting on valence electrons, considering electron shielding

In period 3, for example, the effective nuclear charge increases from 1 for sodium to 7 for chlorine (both have 10 shielding electrons, but sodium only has 11 protons compared to chlorine, which has 17 protons). Therefore, the outermost electrons are pulled in closer to the nucleus, resulting in a smaller atomic size (see figure 1.30).

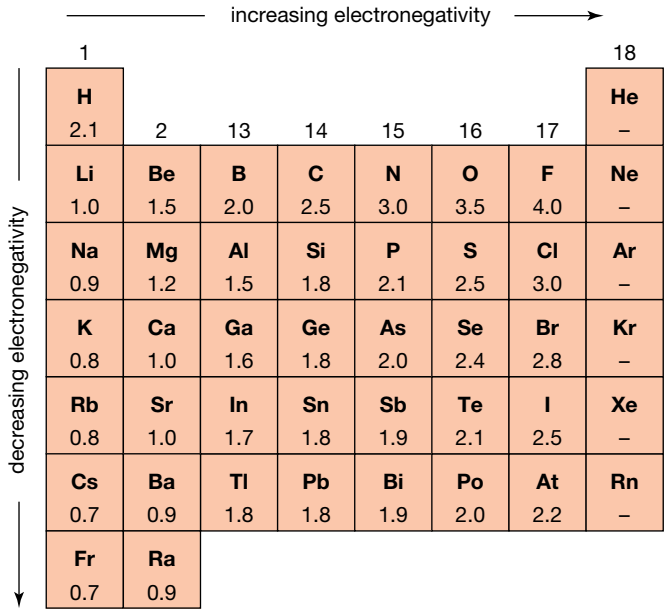
on Resources

Interactivity Periodic table trends (int-6352)

1.5.2 Electronegativity

The **electronegativity** of an element is a measure of the degree to which an atom can attract an electron to itself. This is most evident when it is chemically combined with another element. The extent of attraction is expressed in arbitrary units on the Pauling electronegativity scale (see figure 1.31).

FIGURE 1.31 The Pauling scale of electronegativities



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

The following electronegativity trends may be seen in the periodic table:

- Across a period, left left to right, electronegativities increase. As you move from one element to the next across a period, the nuclear charge increases by one unit, as one electron is added to the outer shell. As the positive charge in the nucleus increases, the atom has an increasing electron-attracting power and, therefore, an increasing electronegativity. This is the same reason the atomic radius tends to decrease across a period, as already described.
- Down a group, electronegativities decrease because the outer electrons are further away from the nucleus and the shielding effect of the inner electrons decreases the electron-attracting power of the atom. Similarly, this results in a general increase in atomic radii down a group.
- Metals generally have lower electronegativities than non-metals.

electronegativity the electron-attracting power of an atom

EXTENSION: Linus Pauling—chemist and peace activist

Linus Carl Pauling (1901–1994) was an American scientist famous for his work on chemical bonding and biochemistry, and his activism for peace. Voted by *New Scientist* as one of the 20th greatest scientists of all time, he was also regarded as one of the two greatest scientists of the 20th century (the other being Einstein). He is widely regarded as the greatest chemist since the founder of the discipline, Antoine Lavoisier, and was the recipient of two Nobel Prizes.

FIGURE 1.32 Linus Carl Pauling received the Nobel Prize for chemistry in 1954, and the Nobel Peace Prize in 1962.

Until the 1930s, chemistry had only been able to describe the properties of substances; it hadn't been able to explain them. Pauling applied the newly developed quantum theory to describe the bonding between atoms and how this explained a substance's properties. In 1931, he published the first of seven papers on the subject; a paper now regarded as one of the most significant works in chemistry. In 1939, following further research, Pauling collated his work up to that date in the landmark book *The nature of the chemical bond* — still used by chemists across the world today.

During the mid-1930s, Pauling became interested in the field of biomolecules and began to apply his knowledge to the structure of the large molecules that exist in living things. He proposed that molecules such as proteins and DNA have a helical structure and laid the groundwork for the ultimate discovery of DNA's structure by Watson, Crick and Franklin in 1953. He was one of the pioneers of model building. He used data from X-ray diffraction to build scale models of the molecules he was studying.

During World War II, Pauling worked for the United States government on several projects associated with the war. In 1948, he was awarded the Presidential Medal for Merit. However, his unease in the postwar period led him to speak out with several other scientists (including Einstein) against the proliferation of nuclear weapons. He became a tireless campaigner against the testing of nuclear weapons.

Pauling was the winner of two Nobel Prizes. The first, for chemistry in 1954, was for his work on the nature of the chemical bond and his second, for peace in 1962, was for his campaigning against the testing of nuclear weapons. He is the only person to have been awarded two unshared Nobel Prizes, and one of only two people to have won them in different fields. Pauling was a popular personality and was as widely known in the general population as he was in the scientific community. Chemistry students around the world recognise him as the developer of the scale of electronegativites that bears his name.

1.5.3 First ionisation energy

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

Removal of one electron from a gaseous metal atom, $M(g)$, results in the formation of a positive ion with a +1 charge:

$$M(g) \rightarrow M^{+}(g) + e^{-}$$

first ionisation energy the energy required to remove one valence electron from a gaseous atom

The energy required to remove this first (outermost) electron is called the first ionisation energy.

To remove the outermost electron from the gaseous +1 ion results in a positive ion with a +2 charge:

$$M^{+}(g) \rightarrow M^{2+}(g) + e^{-}$$

The amount of energy required is called the second ionisation energy, and so on. Table 1.5 shows the first three ionisation energies of the first 20 elements in the periodic table.

TABLE 1.5 Ionisation energies of the first 20 elements (kJ mol^{-1}). The red letters and numbers indicate the elements and first ionisation energies for group 1.

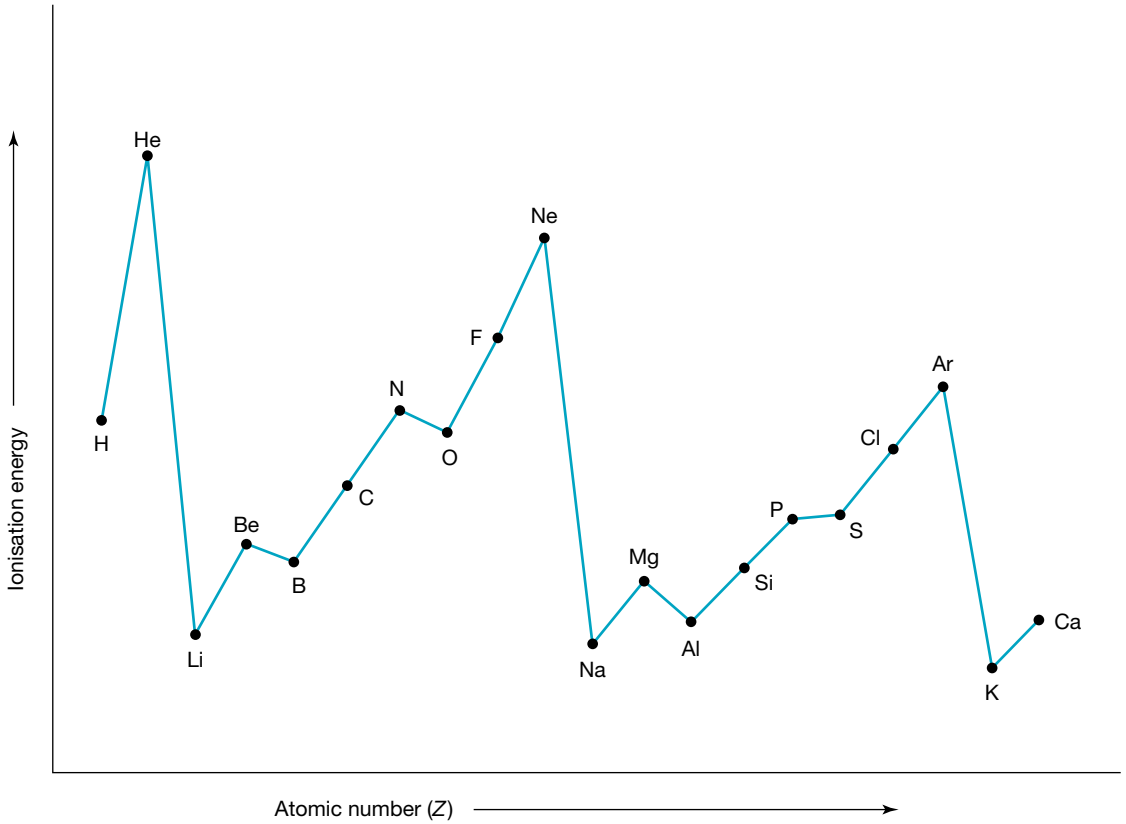
Symbol of element	Ionisation energy (kJ mol^{-1})		
	First	Second	Third
H	1312		
He	2371	5247	
Li	520	7297	11810
Be	900	1757	14840
B	800	2430	3659
C	1086	2352	4619
N	1402	2857	4577
O	1314	3391	5301
F	1681	3375	6045
Ne	2080	3963	6276
Na	495.8	4565	6912
Mg	737.6	1450	7732
Al	577.4	1816	2744
Si	786.2	1577	3229
P	1012	1896	2910
S	999.6	2260	3380
Cl	1255	2297	3850
Ar	1520	2665	3947
K	418.8	3069	4600
Ca	589.5	1146	4941

Some general trends in first ionisation energy include the following:

- Moving down a group, the first ionisation energy generally decreases, since the size of the atom is increasing when moving down a group, and so the outermost electrons are further from the nucleus. The nucleus, therefore, does not hold these electrons as strongly, so they are more easily removed. The atom, therefore, has a lower ionisation energy. Table 1.5 shows the change in ionisation energies for Li, Na and K.
- Across a period, from left to right, the first ionisation energy generally increases. The nuclear charge is increasing, whereas the shielding effect is relatively constant. A greater attraction of the nucleus for the electron, therefore, leads to an increase in ionisation energy. Ionisation energy is high for noble gases such as He, Ne and Ar.

These trends are shown in figure 1.33.

FIGURE 1.33 Trends in ionisation energy of the first 20 elements



on Resources

Weblink Periodic table patterns

1.5.4 Metallic/non-metallic characteristics

In terms of electronic structure, the metallic characteristics of an element are determined by its ease in losing electrons.

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

These metallic/non-metallic characteristics are illustrated in the variation in melting and boiling points (metals typically have high melting and boiling points in comparison with non-metals), as shown in table 1.6.

TABLE 1.6 The melting points (blue) and boiling points (pink) of selected elements in degrees Celsius

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

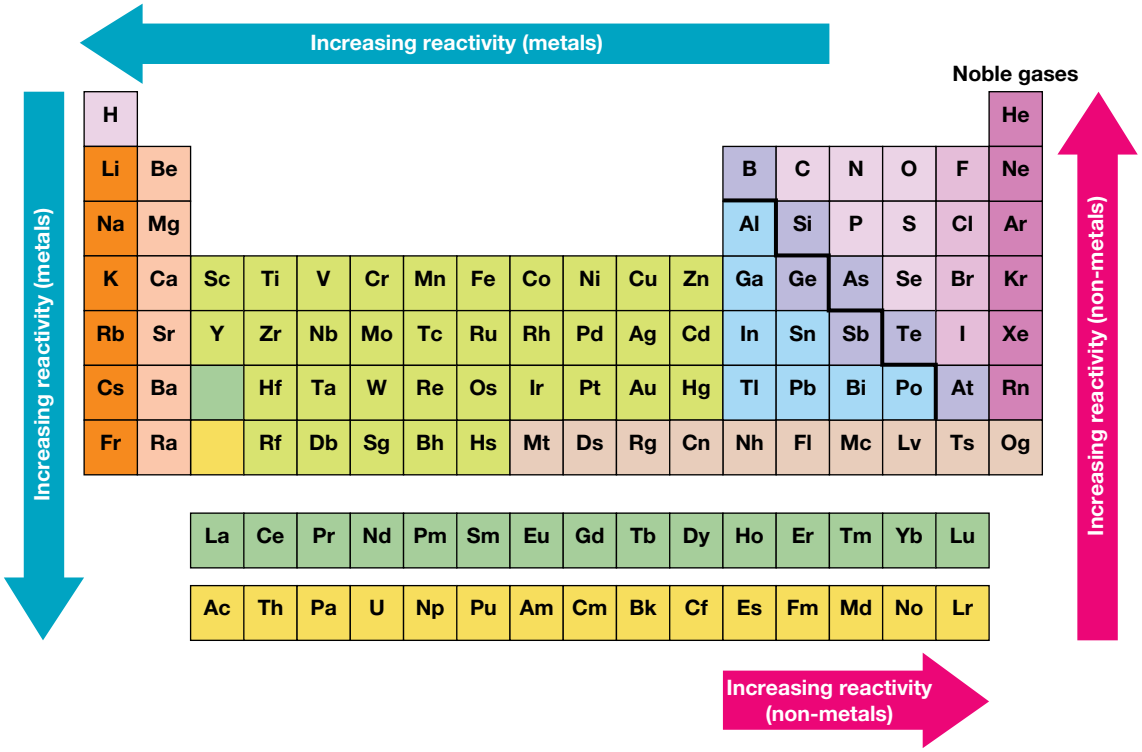
1.5.5 Reactivity of elements

Reactivity refers to the ability of an atom to participate in a chemical reaction, and this is affected by its electronegativity and ionisation energy. Metals and non-metals react differently.

- For metals, the reactivity decreases going across a period (left to right) and increases down a group.
- For non-metals, the reactivity increases going across a period (left to right) and decreases going down a group.

These trends are shown in figure 1.34.

FIGURE 1.34 Reactivity patterns in the periodic table



EXTENSION: Oxidising and reducing strength

The reactivity of elements is related to the ability of the elements to gain or lose electrons.

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

The more readily an element gives up its electrons, the more easily it is oxidised, making it a stronger reducing agent because it has more reducing strength.

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

oxidising strength how readily an element gains electrons

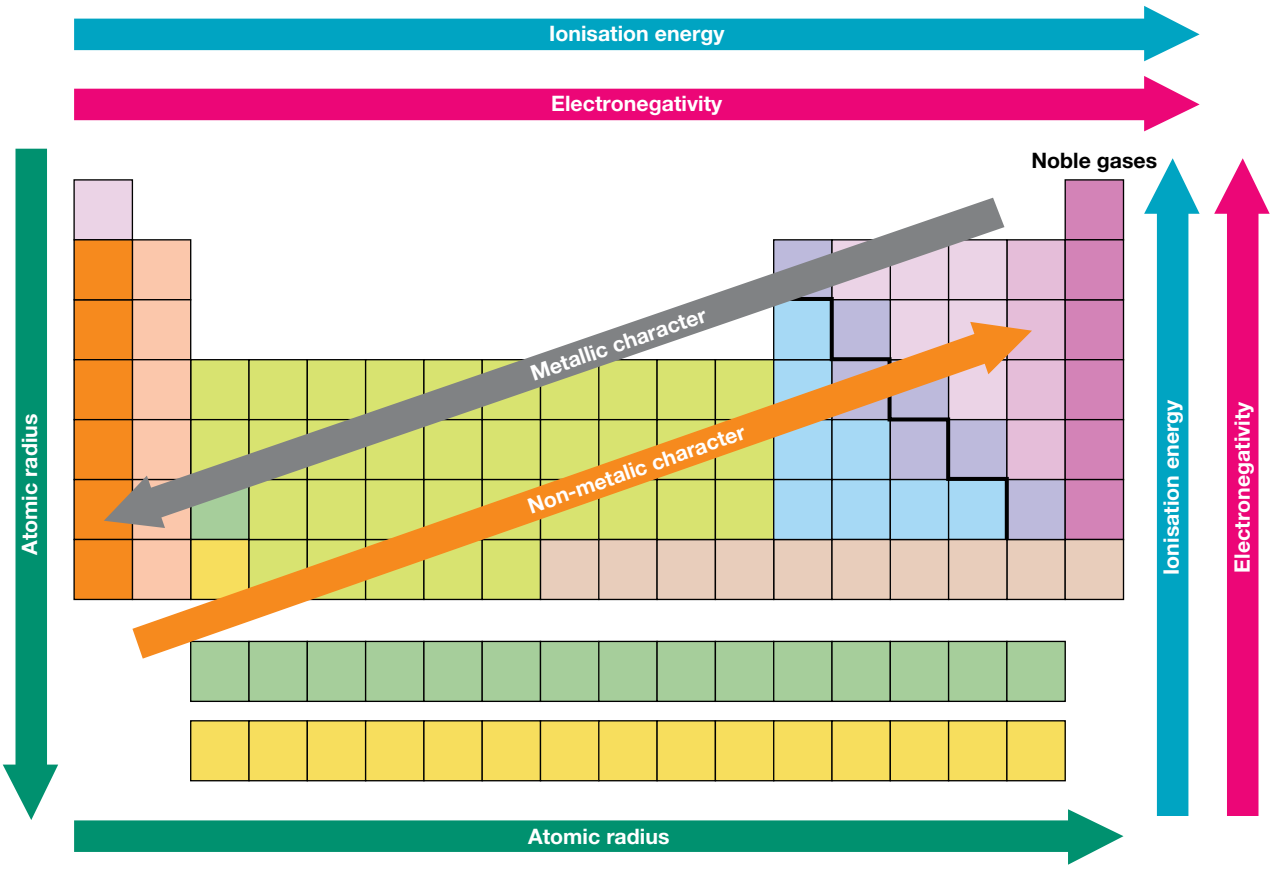
reducing strength how readily an element loses electrons

1.5.6 Summary of trends in the periodic table

TABLE 1.7 Trends in the periodic table

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

FIGURE 1.35 Summary of trends in the periodic table



1.5 Activities

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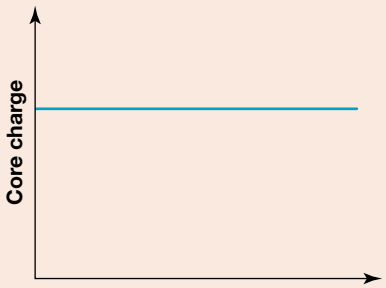
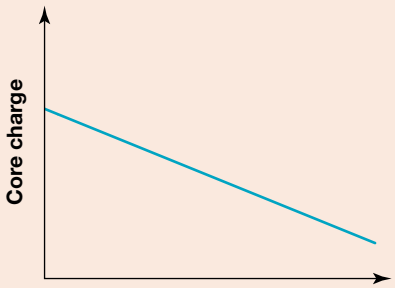
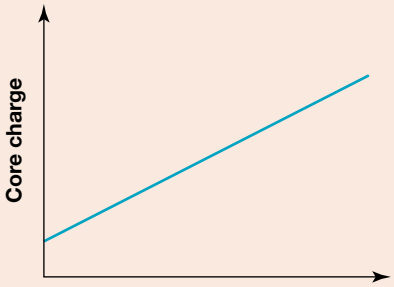
1.5 Quick quiz on | 1.5 Exercise | 1.5 Exam questions

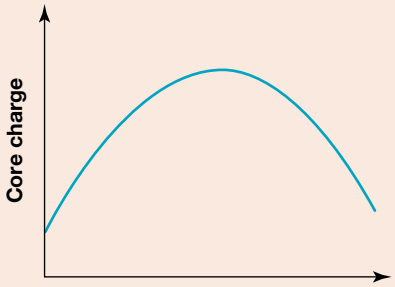
1.5 Exercise

1. MC Which of the following graphs shows how core charge varies going down a group in the periodic table?

A.


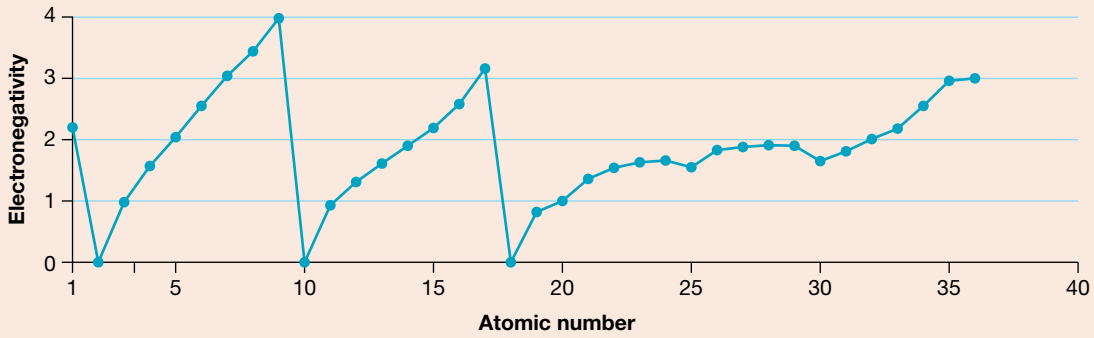
B.


C.


D.


2. MC Which of the following elements would have the strongest attraction between the nucleus and valence electrons?
   - A. Oxygen
   - B. Nitrogen
   - C. Beryllium
   - D. Sulfur
3. MC Going down a group, the atomic radius _____ and going across a period, the electronegativity_____.
   - A. increases, increases
   - B. decreases, increases
   - C. increases, decreases
   - D. decreases, decreases
4. Calculate the core charge for the following.
   - a. Lithium
   - b. Fluorine
5. The first ionisation energy is the energy required to remove an electron from a gaseous atom.
   - a. Would it require more, less or the same energy to remove a second electron? Explain your answer.
   - b. Explain the relationship between ionisation energy and reactivity of metals.
   - c. Name the group of elements in the periodic table that would have the lowest ionisation energy. Explain your answer.
   - d. Why is it difficult for an atom with a high ionisation energy to form a positive ion?

6. The following graph shows the electronegativities of the first 36 elements.



Use the graph to respond to the following.

a. Identify the elements at the peaks and explain why they are at the top of the different sections of the graph.
b. Circle the elements in the second period of the periodic table.
c. Identify the elements at the troughs and explain their position on the graph.
d. How does the graph support the idea that elements are arranged in a periodic way in the periodic table?

7. Explain the relationship between metallic character and electronegativity. Describe these trends going down a group to illustrate your answer.
8. Describe and explain the general trends in the following.
   a. Electronegativity across the periodic table
   b. The atomic radii of elements down a group
   c. The reducing strength of elements across a period
9. For each of the following pairs of elements, state which element is the more electronegative.
   a. K, Ca
   b. Be, Ca
   c. Cl, Br
10. Suggest a reason gold nuggets are found on the Earth's surface but sodium nuggets are not.

1.5 Exam questions

Question 1 (1 mark)

MC Which of the following statements correctly describes the trend moving across a period in the periodic table?

A. The electronegativity increases due to increasing core charge.
B. The electronegativity increases due to increasing atomic radius.
C. The electronegativity decreases due to increasing core charge.
D. The electronegativity decreases due to increasing atomic radius.

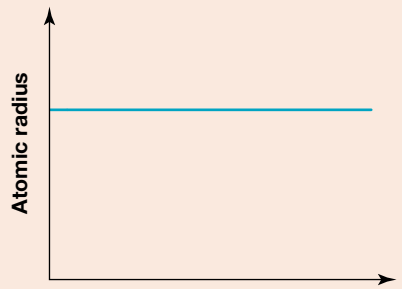
Question 2 (3 marks)

a. Define the term 'electronegativity'. **(1 mark)**
b. Of the list of elements K, Al, S, F, Fe and C, give an example of an element that has the following properties.
   i. High electronegativity **(1 mark)**
   ii. Low electronegativity **(1 mark)**

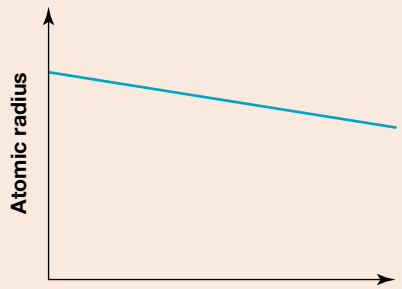
Question 3 (1 mark)

MC Which of the following graphs shows how atomic radius varies as atomic number increases from 3 to 9?

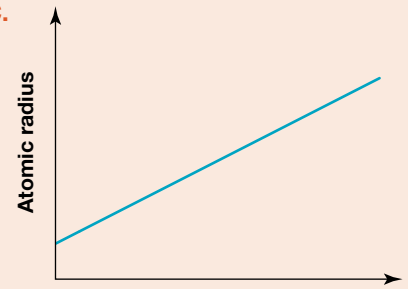
A.



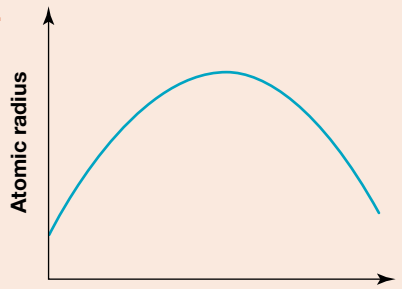
B.



C.

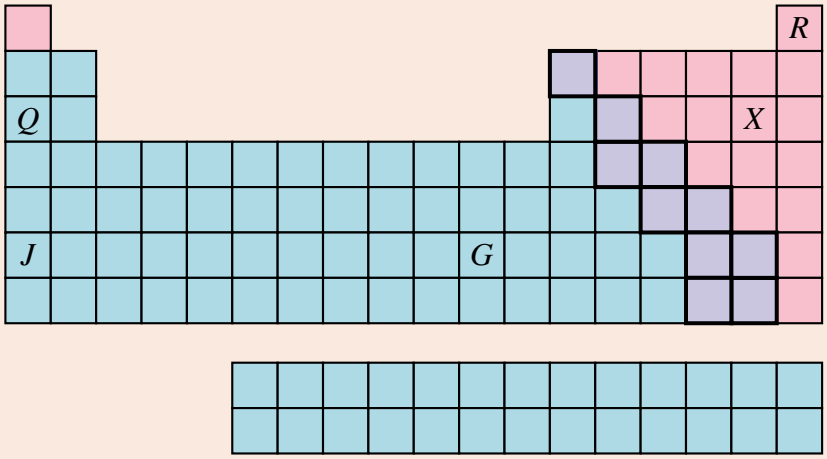


D.



Question 4 (7 marks)

Consider the following elements (shown with abstract symbols).



a. Which two elements have the same number of valence electrons? **(1 mark)**
b. Which element is in the *d* block? **(1 mark)**
c. Which element has an incomplete *p* subshell? **(1 mark)**
d. Which element is lowest in reactivity? **(1 mark)**
e. Which element would have the largest atomic radius? **(1 mark)**
f. Define the term 'ionisation energy'. **(1 mark)**
g. Which element has the lowest ionisation energy? **(1 mark)**

Question 5 (3 marks)

Explain how trends in metallic character, both vertically and horizontally in the periodic table, can be linked to the core charge and atomic radius of an element.

More exam questions are available in your learnON title.

1.6 Critical elements

KEY KNOWLEDGE

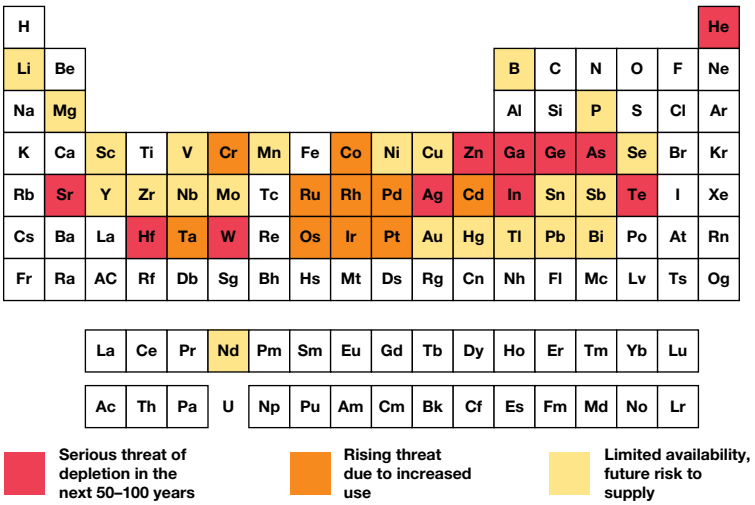
- Critical elements (for example, helium, phosphorus, rare-earth elements and post-transition metals and metalloids) and the importance of recycling processes for element recovery

Source: VCE Chemistry Study Design (2023–2027) extracts © VCAA; reproduced by permission.

1.6.1 Critical elements

Just two generations ago, only about 30 of the elements of the periodic table were familiar and in regular use. Today, however, with the advancement of technology more than 60 elements are in regular use. A mobile phone, for example, contains about 70 elements. The escalating demand for electronic devices is putting a strain on resources, as do the increasing developments in biotechnology and renewable energy technology. Another difficulty is that some elements are found in a small number of countries that sometimes place restrictions on trade. Only small supplies of valuable resources are available in different countries, making it uneconomical to extract them. The tendency now is to keep looking for new supplies of elements and dispose of e-waste at the end of life into landfills.

FIGURE 1.36 Abundance and availability of natural elements



In chemical reactions, the atoms in the reactants are rearranged to form the products. Mass is conserved and no atoms are consumed. So, in theory, elements cannot be lost. **Critical elements** are considered vital for the world's economy but could be in short supply in the coming years. The reason is not necessarily because they are being used up or destroyed, but because, after production and use, they are too difficult and too expensive to recover and so are left in an unusable form. Only two elements are actually being lost: helium and uranium. Helium escapes the Earth's atmosphere and uranium undergoes radioactive decay to form other elements.

Critical elements

Critical elements are considered vital for the world's economy but are in short supply.

Our society is highly dependent on the use of chemicals to produce not only technological devices but also a huge variety and quantity of materials used in our daily lives. This production results in waste materials that are detrimental to individuals and communities. The world's population is increasing, resources are being rapidly consumed, water shortages exist in many parts of the world along with extensive poverty and income inequality. This is not **sustainable**; to be sustainable, chemical processes must meet current human needs without affecting the needs of future generations. Sustainable chemistry demands that we use innovative practices to enhance human health and protect the environment by limiting our use of resources and the creation of the wastes generated in manufacturing processes.

critical elements elements considered vital for the world's economy but in short supply

sustainable developments that meet current demands without affecting the needs of future generations

Sustainability

Sustainability refers to developments that meet current demands without having an impact on the needs of future generations.

To foster sustainable development, in September 2020 the United Nations produced a film entitled *Nations United: Urgent Solutions for Urgent Times*, which set out the requirements for addressing the world's major problems. The film marked the UN's 75th anniversary, as well as the fifth anniversary of the Sustainable Development Goals (SDGs). Of the 17 SDGs, eight are relevant to VCE Chemistry:

- Goal 6: clean drinking water and effective sanitation must be available to all.
- Goal 7: clean and inexpensive energy is needed for activities such as cooking, heating, health care and communications.
- Goal 9: innovation and technological progress in industries is essential to solving many economic and environmental challenges.
- Goal 11: make cities sustainable and safe by decreasing carbon emissions.
- Goal 12: develop responsible consumption and production by reducing food and industrial waste.
- Goal 13: take urgent action against climate change to prevent adverse effects on our environment.
- Goal 14: conserve aquatic environments through protection from absorbing CO_2 and dumped rubbish that threatens life under the sea and dependent species.
- Goal 15: sustainably manage ecosystems to protect biodiversity by recycling, not wasting resources and effective management.

on Resources

Weblink Urgent solutions for urgent times

1.6.2 Green chemistry

In 1991, Paul Anastas and John Warner developed the 12 Principles of Green Chemistry, which provide guidelines for achieving more sustainable manufacture of materials. This is the concept of **green chemistry**. It has been established to encourage creative design in the manufacture and use of materials. By modifying chemical processes so that more materials are recycled, waste is minimised, less energy is used and the impact on living organisms and the environment is diminished. These measures mean a more sustainable future is possible.

FIGURE 1.37 Renewable biofuel product derived from corn

Of Anastas and Warner's 12 principles, the following seven are relevant to the study of VCE Chemistry:

1. *Atom economy:* waste as little material as possible. As many of the reactant atoms as possible should end up in the final product. If a process has only a 50 per cent atom economy, it means that

$$\text{atom economy} = \frac{\text{molecular weight of products}}{\text{molecular weight of reactants}} \times 100\% = 50\%.$$

2. *Use catalysts:* catalysts speed up reactions and allow reactions to occur using less energy without being consumed in the reaction.

green chemistry the design of chemical products and processes that decrease or eliminate the use or creation of hazardous wastes and the impact on the environment

3. *Design for degradation:* the proper disposal of chemical products at the end of their use should be part of the design process. The products should ideally be able to break down into non-toxic materials and not remain in the environment.
4. *Design for energy efficiency:* in the designing of the chemical synthesis process, energy efficiency should be considered. Reducing energy requirements by conducting processes at room temperature and pressure reduces costs and lowers the impact on the environment.
5. *Design safer chemicals:* chemical reactants and products that function well but are non-toxic will create fewer medical and environmental problems.
6. *Prevention of waste:* it is better to prevent waste than to treat it or clean it up after it forms.
7. *Use renewable feedstocks:* starting materials for a chemical product should, where possible, be renewable resources. For example, considerable research has gone into creating biodiesel derived from farm crops such as soybeans and sunflower seeds. Another example is the production of adipic acid (needed to produce nylon, polyurethane, lubricants and plasticisers) from glucose instead of the cancer-causing chemical benzene.

Source: Adapted from VCE Chemistry Study Design (2023–2027).

Green chemistry

Green chemistry involves the design and manufacture of efficient chemical processes and products that minimise the use or generation of harmful substances so that impacts on human health and the environment are reduced.

1.6.3 Linear and circular economies

To achieve the desired sustainable development, it is necessary to transfer from a **linear economy** to a more innovative **circular economy**. A linear economy is an unsustainable, traditional 'take-make-dispose' model where materials are manufactured from raw materials, used and thrown away afterwards.

In the circular economy, resources are more carefully managed; processes are restorative and designed more efficiently to decrease the use of energy (and preferably use renewable energy) and minimise any wastage. Damage to our health and our environment is limited as a result. Manufactured goods have a life cycle and are repaired, repurposed or recycled wherever possible. After all, we should emulate nature, which continually cycles water, oxygen, carbon and nitrogen throughout the Earth's environment. Applying a circular economy will have a positive effect on soil, water and air, and, ultimately, on climate change. Products are used and not consumed. The circular economy is not just about recycling; it also seeks to avoid creating any waste.

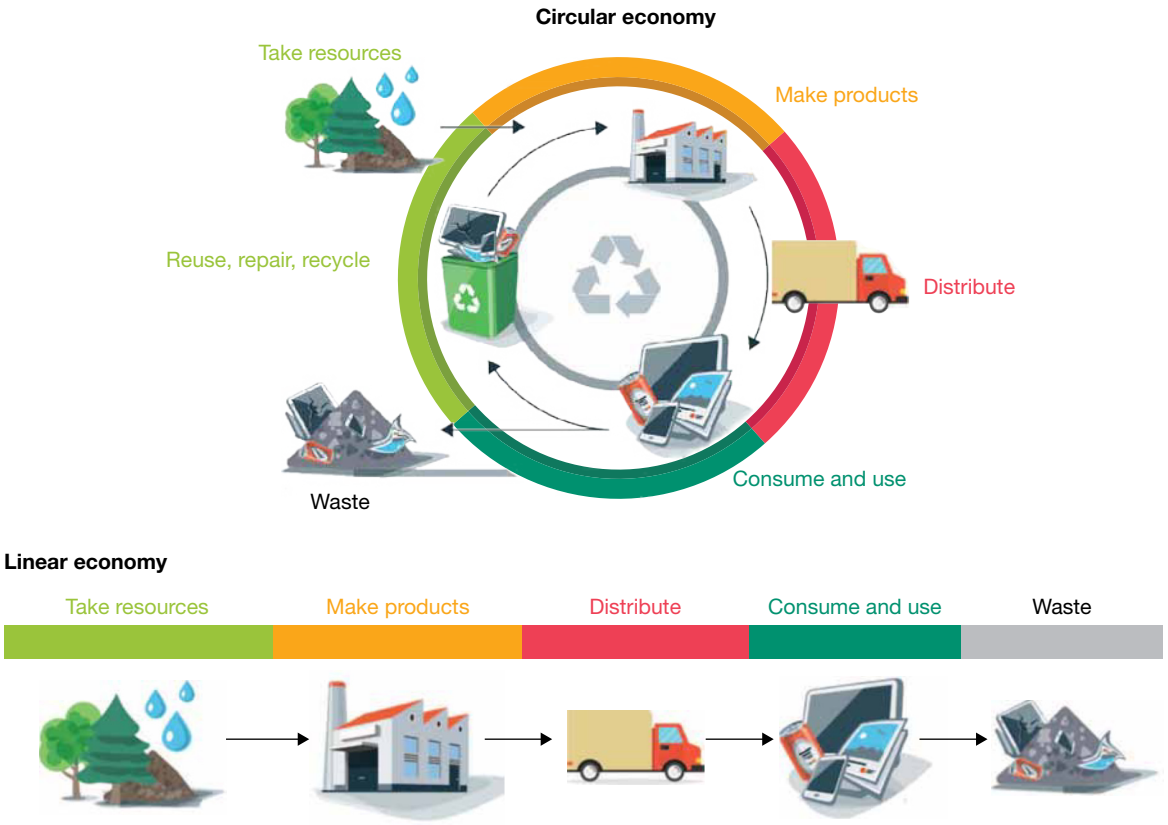
linear economy raw materials are converted into products that are used and then discarded as waste

circular economy economy in which waste and pollution is eliminated, products at the end of life are reused, repaired or recycled, and the environment is not harmed

Circular economy

A circular economy model intends to design out waste by repairing, reusing or recycling materials.

FIGURE 1.38 Circular economy versus linear economy



1.6.4 Elements on the critical list

Helium

When thinking about helium, people often have images of party balloons and making squeaky voices. Helium, however, has far more significant scientific uses. This is why, even though it is the second most abundant element in the universe after hydrogen, it is a concern that it is becoming scarce on Earth. This noble gas is an odourless, colourless, non-toxic gas that is also the second lightest element. It is one element that is actually escaping Earth's atmosphere. Helium does not react with other elements and is lighter than air so it can easily escape the atmosphere.

Helium is a by-product of natural gas extraction and is formed deep underground as a result of the natural decay of some radioactive elements such as uranium-238. (The half-life of U-238 is about 4.5 billion years.) Production of helium is costly and, although Australia has relatively good supplies, only a decreasingly small amount is produced. Sourcing helium from natural gas is problematic because natural gas use involves the production of carbon dioxide and the trend is to use renewable energy resources.

Uses of helium

Helium has various important applications. It can be in liquid form to an extremely cold temperature — colder than any other element — and so is used as a coolant. Superconducting magnets used in MRI scanners need liquid helium to keep them cool, as do some nuclear reactors. Rocket engines use helium to cool and clean them out, and so space exploration could not happen without it. Deep-sea divers use a mixture of helium and oxygen to prevent getting the 'bends'. It is also used in cryogenics and the manufacture of liquid crystal displays.

FIGURE 1.39 (a) Blowing up a helium balloon (b) Adding helium to an MRI scanner

(a)

(b)

Helium recovery and supply management

Although a small amount of helium is continually produced by radioactive decay, finding new methods of preserving helium is essential. Some possibilities include:

- finding more economical methods of production
- finding more economical methods of recycling
- collecting and recycling helium from MRI scanners
- not using helium for non-industrial purposes — for example, no helium balloons
- using alternative elements where possible — for example, hydrogen
- finding more supplies — possibly on the moon!

Indium

The relatively rare element indium is not an element that is well known but it is essential and seen by everyone daily. Indium is found in group 13 period 5, and is rarely found as an element or as its own mineral. It is obtained mainly as a by-product of zinc mining, but as the supply of zinc is declining so is the supply of indium. In this process, usually only a relatively small proportion of the indium ore is recovered. Despite its importance, no specific indium mines exist. China is the main source of indium but Australia has good deposits. The metal indium is a soft, toxic, silvery-white metal that is stable in air and water. It has a low melting point and is resistant to corrosion.

FIGURE 1.40 Mobile screens containing indium tin oxide ready to be recycled

Uses of indium

Touch screens, and TV and computer screens use indium in the compound indium tin oxide (ITO) because it conducts electricity well, is transparent and adheres to the glass. This is by far the main use of indium. ITO is also used as a reflective coating in architectural and automotive glasses, and in photovoltaic cells for solar panels. Solar cells are an important source of renewable energy.

Other applications of indium compounds include light-emitting diodes (LEDs), solders and batteries.

Indium recovery and supply management

At present, the cost of recovering the small amount of indium in devices is uneconomical, so only a small quantity is recycled and more yield is obtained from waste from mining. Possible supply solutions include:

- finding an alternative substance to use in certain applications — for example, carbon nanotubes (large molecules consisting of single layers of hexagonally arranged carbon atoms in cylindrical sheets)
- developing more economical ways of recycling indium
- increasing the efficiency of the production process

- upgrading devices such as mobile phones less frequently
- collecting and recycling more unused devices
- using bioleaching to recover indium from waste (this method will be discussed in topic 3).

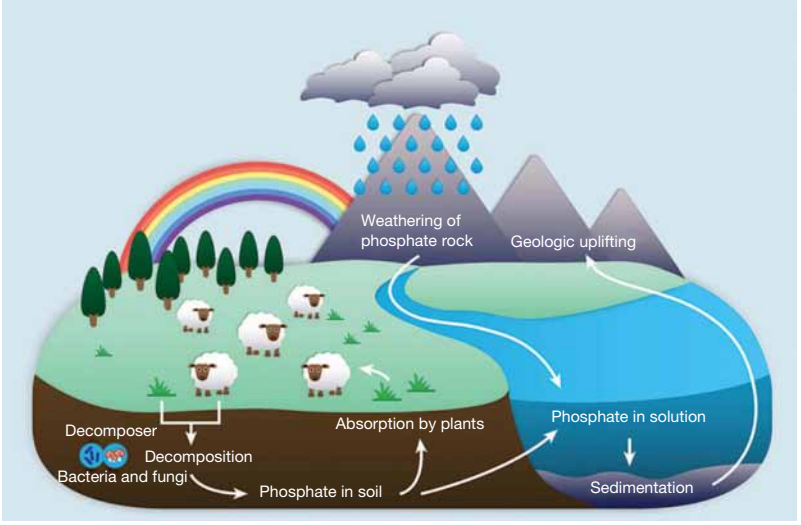
Phosphorus

Another element that is on the critical list is the little known but important element phosphorus. Phosphorus is essential to all life functions. It forms part of the sugar-phosphate backbone of DNA, is involved in energy transport, and is present in all cell membranes and bones. After calcium, it is the second most plentiful mineral in the body. We obtain phosphorus from the food that we eat. Phosphorus occurs in nature in the form of phosphate salts, which are obtained from sedimentary rocks.

Uses of phosphorus

The main use of phosphorus is as a fertiliser. Originally human waste and food waste was spread over food crops as fertiliser but now copious supplies of artificial fertilisers are applied to provide enough food for the increasing world population. Fertilisers are required to provide both nitrogen and phosphorus. While nitrogen is plentiful in the atmosphere, the supply of phosphorus is not sustainable.

FIGURE 1.41 The phosphorus cycle



One reason that the supply is limited is that phosphate resources are concentrated in certain areas such as Morocco, but the more significant issue is that much of the fertilisers that are being used are washed into waterways. This promotes the rapid growth of algal blooms, which use the oxygen present during the decay process. This causes problems for the survival of living organisms present there. This process is called **eutrophication**. Phosphate that reaches the ocean settles and forms insoluble rock. Another issue is the phosphorus 'loss' in the enormous amount of wasted food in developed countries, and in the human waste at the end of the digestion process that is eventually released into the environment. Some European municipalities treat sewage to extract phosphorus. Other treatments use algae to remove phosphate or magnesium oxide to precipitate the phosphate. Animal manure can also be used as fertiliser.

Phosphorus recovery and supply management

Phosphorus has no substitute, so it needs to be conserved. Ways to preserve phosphorus include developing innovative and economical procedures to:

- recycle from streams with high phosphate content
- regularly monitor soils to ensure appropriate amounts of fertiliser are applied
- develop time-release fertilisers
- recover phosphorus from human and animal waste
- recycle industrial waste
- develop plants that are better able to absorb phosphorus
- increase the efficiency of phosphate fertiliser manufacture from phosphate rock.

Individuals can contribute by:

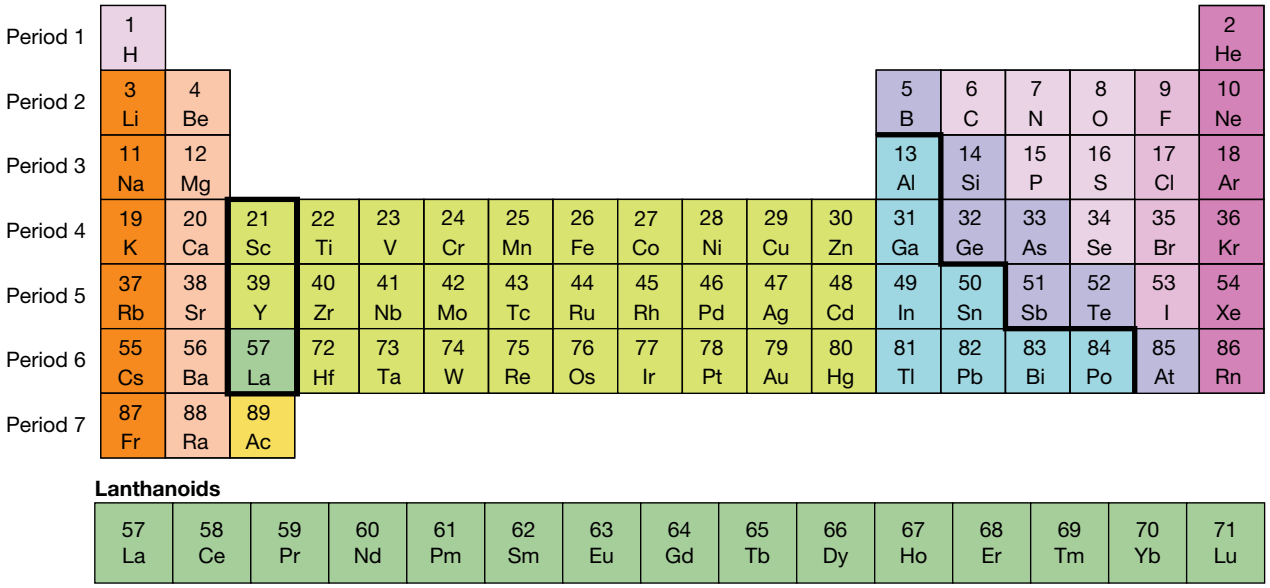
- wasting less food
- eating less meat and dairy to reduce the number of livestock eating plants.

eutrophication form of water pollution involving excess nutrients, such as nitrogen and phosphorus, leaching from soils, typically resulting in excessive growth of algae

Rare earth elements

Rare earth elements (REEs) are not rare but recycling them is. The REE cerium, for example, is more abundant than copper. The REEs include the 15 elements in the lanthanoid series together with yttrium and scandium. They are considered rare because they do not exist in pure form and are difficult to mine, so it is costly to extract them. They are grouped because they have similar properties, and are silvery-white soft metals that are corrosion resistant. They also have some unique and valuable catalytic, nuclear, electrical, magnetic and luminescent properties that make them very useful.

FIGURE 1.42 The rare earth elements include the lanthanoids as well as yttrium and scandium.



Uses of REEs

REEs are used in hundreds of different applications, from LEDs, TVs, batteries and mobile phones to electric vehicles and wind turbines (see figure 1.43). Even though REEs are used in these 'green technologies', mining for REEs is quite destructive to the environment and the toxic waste produced is detrimental to human health. This means recycling is a preferable option, and potential exists for some innovative design here. Recycling can currently be uneconomical because generally only a small amount of REEs is used in each device; however, it is possible. France, for example, has a recycling plant to isolate REEs. Most REEs come from China, but Australia does have supplies.

REE recovery and supply management

Possible solutions for the supply and recycling issues relating to REEs include:

- designing products to use smaller amounts or alternative elements — for example, using ionic liquids instead of strong acids to extract metals from ores
- designing economical recycling processes
- developing more efficient manufacturing processes
- designing products that are more easily recycled
- collecting more products to be recycled.

Recycling REEs

REEs in consumer products are converted into forms that are not easily recyclable. Innovative techniques need to be developed to extract these elements economically and sustainably from waste and consumer products.

FIGURE 1.43 Uses of rare earth elements in electric vehicle manufacture

Source: Data from UNEP (2013) *Metal Recycling: Opportunities, Limits, Infrastructure, A Report of the International Resource Panel* (www.unep.org), figure 14, p. 60.

1.6 Activities

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1.6 Quick quiz on | **1.6 Exercise** | **1.6 Exam questions**

1.6 Exercise

1. Match one or more of the critical elements helium (He), indium (In), phosphorus (P) and the rare earth elements (REEs) with the provided descriptions.

Description	Element(s)
Second most abundant element	
Part of DNA	
Used to manufacture touch screens	
Often magnetic	
Used in 'green technology'	
Can be responsible for eutrophication	
Used in MRI scanners	
Can be obtained from waste organic matter	
Stays liquid at extremely low temperature	
Includes elements filling the 4*f* subshell	

2. MC Which of the following is not a Sustainable Development Goal (SDG) relevant to chemistry?
   - A. Clean drinking water and sanitation
   - B. Readily available and inexpensive energy
   - C. End to racism
   - D. Action against climate change
3. MC What is the main source of phosphorus?
   - A. A by-product of metal extraction
   - B. Sedimentary rocks
   - C. Waterways
   - D. Decaying dead animals and plants
4. a. How is helium formed in the earth?
   b. How is helium extracted for human use?
5. MC How does phosphorus naturally become part of an ecosystem?
   - A. From the atmosphere
   - B. Decaying organic matter
   - C. Radioactive decay
   - D. From acid rain
6. MC Where is most indium metal sourced from?
   - A. Sedimentary rocks
   - B. As the pure metal as it is unreactive
   - C. As a result of radioactive decay
   - D. As a by-product of zinc mining
7. a. Identify the process by which the excess growth of aquatic plants removes oxygen from waterways during decay.
   b. Explain which critical element is involved in this process.
8. What is the main use of indium? What properties of indium make it suitable for this purpose?
9. Describe how three of the SDGs highlighted in the film *Nations United: Urgent Solutions for Urgent Times* relate to the issue of declining phosphorus availability.
10. a. From where is the element helium sourced?
    b. List three uses of helium.

1.6 Exam questions

Question 1 (1 mark)

MC What is development that meets current needs without affecting the ability of future generations to meet their own needs called?

- A. Green chemistry
- B. Recycling
- C. Circular economy
- D. Sustainability

Question 2 (1 mark)

MC Why is phosphorus a critical element that is essential for the body?

- A. It is used in energy transport in cells.
- B. It is part of all cell membranes.
- C. It contributes to the structure of DNA.
- D. All of the above

Question 3 (3 marks)

Explain why helium is considered a critical element when it is the second most abundant element in the universe.

Question 4 (3 marks)

Helium is originally produced from the radioactive decay of uranium-238, with mass number 238 and atomic number 92, to form thorium-234 and helium.

The nuclear equation is:

$$^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He$$

Write the nuclear equation for the decay of thorium-232 to radium-288 and helium.

Question 5 (7 marks)

Describe how the following principles of green chemistry could be applied to the manufacturing, use and recycling of indium in touch screens.

a. Atom economy **(1 mark)**
b. Use of catalysts **(1 mark)**
c. Design for degradation **(1 mark)**
d. Design for energy efficiency **(1 mark)**
e. Design for safer chemicals **(1 mark)**
f. Prevention of waste **(1 mark)**
g. Use renewable feedstocks **(1 mark)**

More exam questions are available in your learnON title.

1.7 Review

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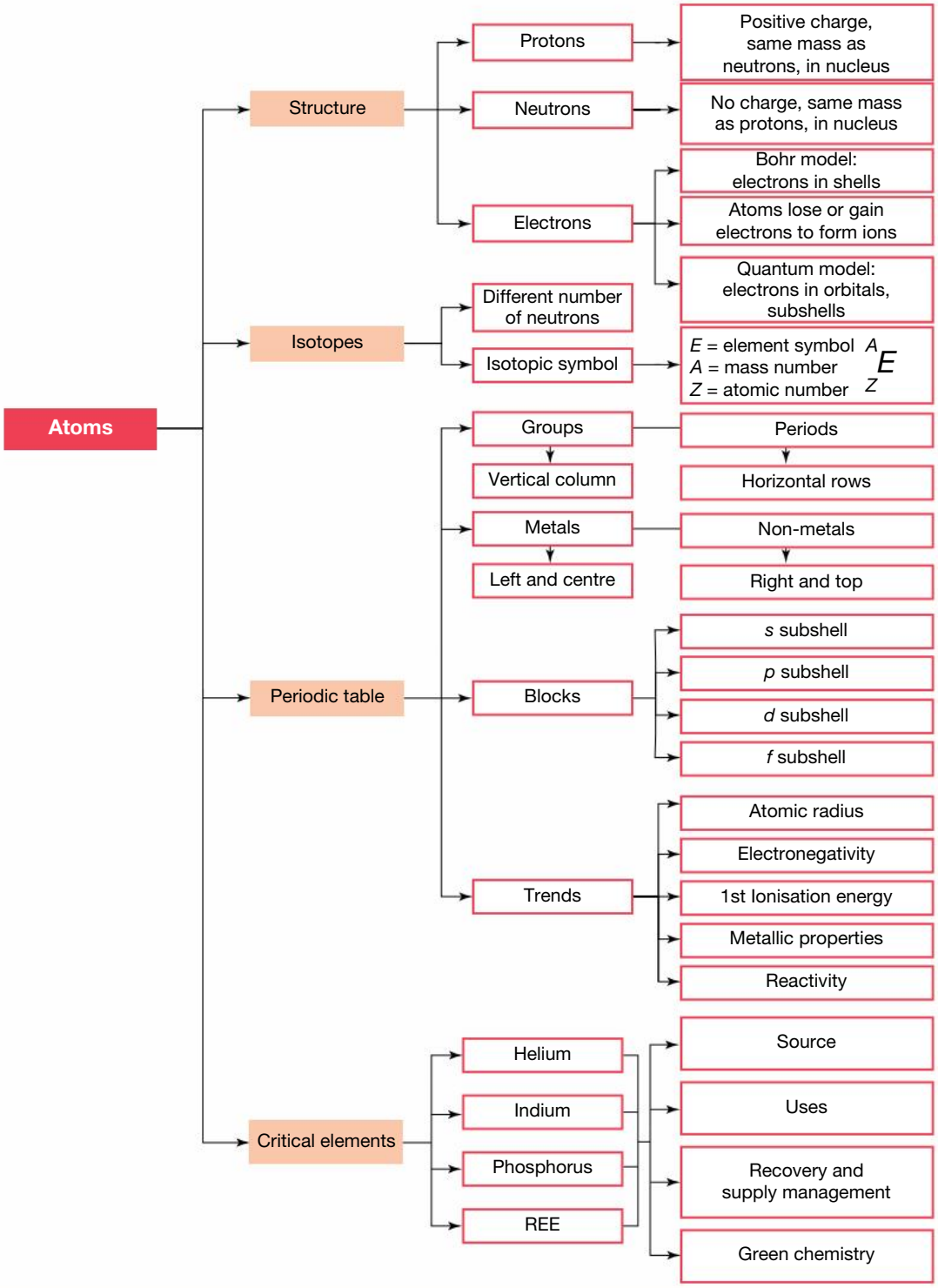
1.7.1 Topic summary



1.7.2 Key ideas summary

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1.7.3 Key terms glossary

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on Resources

Solutions	Solutions — Topic 1 (sol-0800)
Practical investigation eLogbook	Practical investigation logbook — Topic 1 (elog-1606)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 1 (doc-37067) Key idea summary — Topic 1 (doc-37068)
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1.7 Activities

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1.7 Review questions

1. Complete the following table.

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

2. From the information in the table, complete the following.

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

a. Identify those species that are isotopes of the same element.
b. Write the isotopic symbol for each atom.

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

4. Chromium (atomic number 24) and copper (atomic number 29) do not follow the usual pattern of subshell electron configuration.
   a. Why does this occur?
   b. Write the subshell electron configuration for chromium.
   c. Write the subshell electron configuration for copper.
5. Use the list of elements to answer the questions that follow: Cl, Ni, N, Mg, I, B, Na.
   a. Choose the two elements that have the most similar chemical properties and explain your choice.
   b. Which of the elements is the most reactive metal? Explain why.
   c. Which element is a transition metal?
   d. Which of the elements is the most reactive non-metal?
   e. Which element has the highest electronegativity? Explain why.
6. Predict which pair of elements in each of the following sets would have the greatest similarities in chemical properties and which would have the greatest differences.
   a. Na and Cl, Na and K, Na and Ca
   b. Cl and I, Cl and S, Cl and Mg
7. Consider the element phosphorus ($Z = 15$).
   a. What is its ground state electron configuration?
   b. In what group is it?
   c. In what period would you find it?
   d. In what block would you find it?
   e. Would it have a higher or lower electronegativity than chlorine?
   f. Would it have a larger or smaller atomic size than nitrogen?
   g. Explain why phosphorus is considered a critical element and suggest two ways we can preserve or protect of supply of this element.
8. a. Explain what is meant by the term ‘nuclear charge’.
   b. Which has the higher nuclear charge: potassium or calcium?
9. How is an element’s outer electron configuration related to its position in the periodic table? Give three examples that illustrate your answer.
10. Use your knowledge of trends in the periodic table to explain the following.
    a. Strontium has a lower first ionisation energy than magnesium.
    b. A beryllium atom has a larger atomic radius radium than a nitrogen atom.
    c. Fluorine is the most electronegative atom in the periodic table.
    d. Potassium is a strong reducing agent.

1.7 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC Which of the following determines the chemical properties of an element?

A. Mass number
B. Number of protons
C. Number of electrons
D. Number of neutrons

Question 2

MC Different isotopes of the same element have different what?

A. Atomic numbers
B. Numbers of neutrons
C. Numbers of protons
D. Numbers of electrons

Question 3

MC How can the periodic table of the elements be used?

A. To find out the order in which the elements were discovered
B. To predict how atoms of different elements will combine
C. To identify an element's physical and chemical properties
D. To determine how many isotopes an element has

Question 4

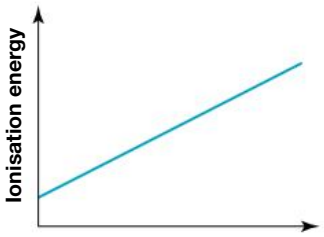
MC If the lowest two electron shells of atom are full, what is its atomic number?

A. 20 **B.** 16 **C.** 14 **D.** 10

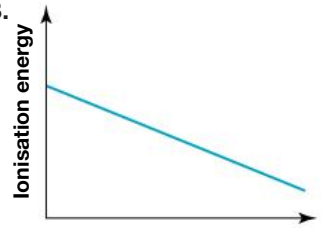
Question 5

MC Which of the following graphs shows how ionisation energy generally varies as atomic number increases across period 2?

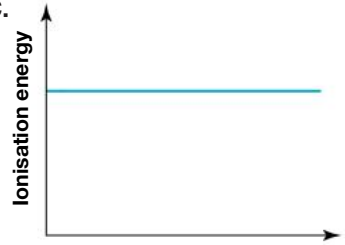
A.



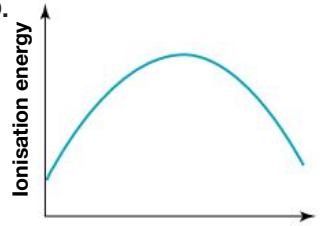
B.



C.



D.



Question 6

MC Consider the following electron configuration:

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

For which of the following would this *not* apply?

A. Ar **B.** Mg^{2+} **C.** S^{2-} **D.** K^+

Question 7

MC What is the effective nuclear charge of the element in period 2, group 14?

A. −4 **B.** +4 **C.** +6 **D.** +14

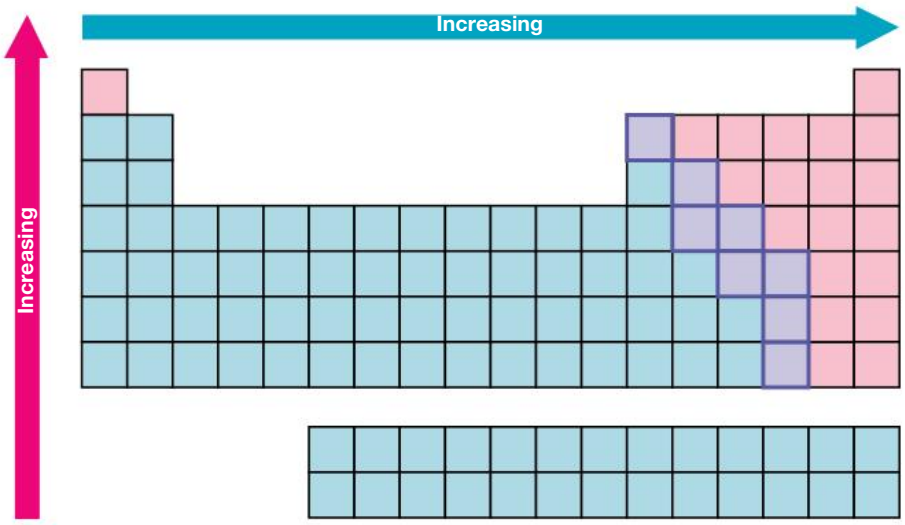
Question 8

MC The ground state electron configuration of a heavy metal atom ends with $6s^2 6p^3$. What must the element be?

A. Antimony **B.** Bismuth **C.** Cadmium **D.** Lead

Question 9

MC Consider the following trends.



To which of the following do these trends *not* apply?

A. Ionisation energy
B. Electronegativity
C. Non-metallic character
D. Atomic radius

Question 10

MC Where in the periodic table is the critical element indium found?

A. Period 4, group 3
B. Period 4, group 13
C. Period 5, group 3
D. Period 5, group 13

Section B — Short answer questions

Question 11 (1 mark)

An isotope of tungsten, tungsten-180, has 74 protons. For the isotope tungsten-183, state the number of neutrons present.

Question 12 (6 marks)

a. Write the ground state electron configuration of each of the following species.

i. Arsenic atom, $Z = 33$

ii. Calcium ion, Ca^{2+}, $Z = 20$ **(2 marks)**

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

Question 13 (5 marks)

Consider the following elements of the periodic table.

$$^{7}_{3}J,\ ^{31}_{15}P,\ ^{195}_{78}X,\ ^{48}_{22}L,\ ^{40}_{18}Z$$

Which of these element(s) is/are most likely to lose its valence electron/s?

Question 14 (7 marks)

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

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

Selecting only from the elements labelled on the periodic table shown, write the letter (*A, B, D, E, L, G, J, M, Q, R, T* or *X*) corresponding to the element with the following.

a. A non-metal with four electrons in its outer shell **(1 mark)**

b. In group 14, period 2 **(1 mark)**

c. A transition element **(1 mark)**

d. A halogen with two occupied shells **(1 mark)**

e. In period 2 and with a total of three electrons **(1 mark)**

f. A halogen **(1 mark)**

g. An alkaline earth metal **(1 mark)**

Question 15 (5 marks)

Describe how the idea of a circular economy could be applied in the life cycle of indium. State the benefits of using the circular economy.

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2 Covalent substances

KEY KNOWLEDGE

In this topic you will investigate:

Covalent substances

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

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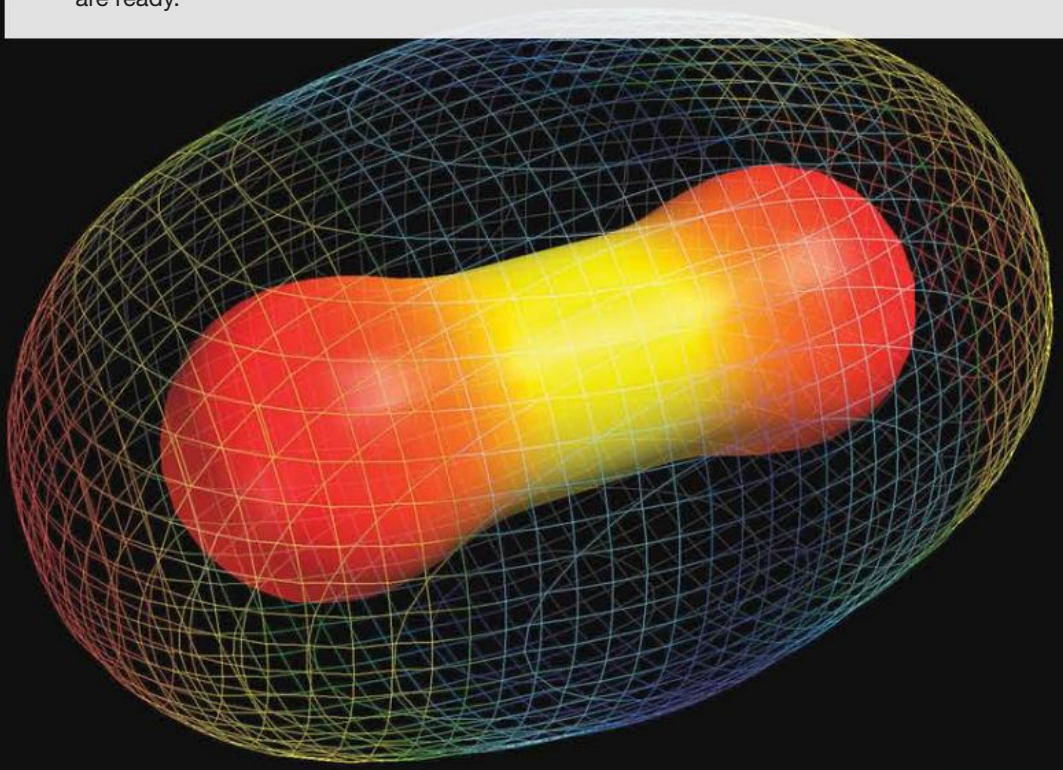
PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of VCE Chemistry. Experiments and investigations, supported by a **practical investigation eLogbook** and **teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

EXAM PREPARATION

Access exam-style questions and their video solutions in every lesson, to ensure you are ready.



2.1 Overview

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2.1.1 Introduction

Our Earth is a complex and fascinating world of chemicals. The air we breathe consists mainly of the gases nitrogen and oxygen, which are examples of covalent molecular elements. Covalent molecular substances are all around us. Water is the most common molecule in our body and is a covalent molecular compound. Other essential molecules range from small oxygen molecules to DNA, which is a very large molecule or macromolecule. DNA carries the genetic information to make new proteins for our cells.

All these molecular substances are composed of non-metals and behave very differently from the metallic substances we will study next. This different behaviour is due to the unique structure that results when non-metal atoms are held together by a covalent bond. Covalent compounds are mainly gases or liquids at room temperature and pressure, and are usually poor conductors of electricity. How can substances that are different sizes and look so different and behave in such different ways have the same type of bonding between atoms? Why are substances such as carbon dioxide and nitrogen gases, and others such as water and ethanol liquid?

FIGURE 2.1 The macromolecule DNA provides the code for the synthesis of every protein molecule in the body.

This topic explains how the properties of covalent compounds are based on the structures and arrangement of molecules. It explores how to predict the way atoms of non-metals join to make molecules and the methods that chemists use to represent the molecules that are formed.

LEARNING SEQUENCE

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2.2 Representing molecules 61
2.3 Shapes of molecules 73
2.4 Comparing intramolecular bonding and intermolecular forces 86
2.5 Physical properties of molecular substances 95
2.6 Structure and bonding of diamond and graphite 99
2.7 Review 107

on Resources

Solutions	Solutions — Topic 2 (sol-0802)
Practical investigation eLogbook	Practical investigation logbook — Topic 2 (elog-1608)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 2 (doc-37071) Key ideas summary — Topic 2 (doc-37072)
Exam question booklet	Exam question booklet — Topic 2 (eqb-0084)

2.2 Representing molecules

KEY KNOWLEDGE

- The use of Lewis (electron dot) structures, structural formulas and molecular formulas to model the following molecules: hydrogen, oxygen, chlorine, nitrogen, hydrogen chloride, carbon dioxide, water, ammonia, methane, ethane and ethene

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2.2.1 Covalent bonding

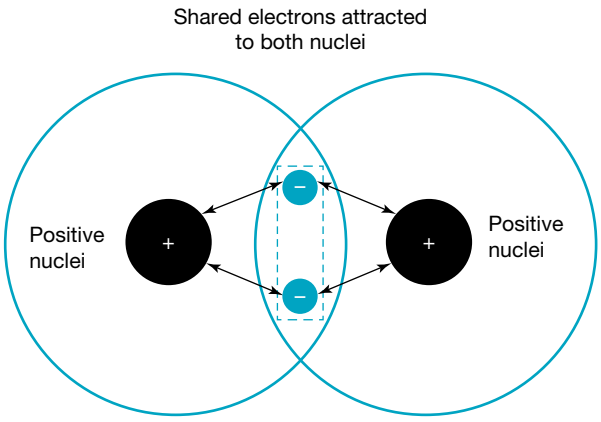
We have observed that atoms of elements with a stable outer shell, such as the noble gases helium, neon and argon, are particularly stable. Other atoms tend to achieve noble gas configuration — that is, eight electrons in the outer shell, apart from helium — by losing or gaining electrons to form ions, as discussed in topic 1, or *sharing* electrons to form **covalent bonds**. The hydrogen atom is different; it needs two electrons in its outer shell for stability. Most of the elements in the second and third periods of the periodic table need eight outershell electrons for stability (see 'The octet rule' in section 2.2.2). This *sharing* of electron *pairs* between two non-metallic atoms produces the covalent bond. The atoms bond to reach a more stable state.

Covalent bonds

A covalent bond occurs between non-metallic atoms and involves sharing electron pairs to attain a noble gas configuration.

The type of bonds formed between atoms is affected by their **electronegativity**. As discussed in topic 1, the electronegativity of an element is a measure of the degree to which an electron can be attracted to the positively charged nucleus of another atom. This is most evident when it is chemically combined with another element (figure 2.2).

FIGURE 2.2 The shared electrons in a covalent bond are attracted to both nuclei



Electronegativity

Electronegativity is the relative attraction that an atom has for shared electrons in a covalent bond.

covalent bonds sharing of electrons between nuclei that bonds them together in a molecule

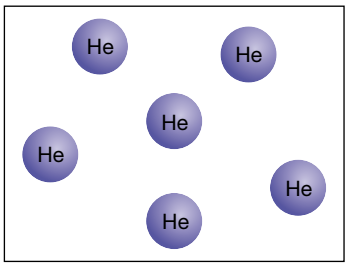
electronegativity the electron-attracting power of an atom

2.2.2 Atoms and molecules

As described earlier, the noble gases all exist as **discrete** atoms (figure 2.3). This means they do not share electrons. Since they have stable outer shells, they are unreactive and few of their compounds are known.

When a fixed number of atoms combine and share electrons in covalent bonds, a **molecule** is formed. Molecules can consist of the same elements or a combination of different elements.

FIGURE 2.3 Discrete *atoms* of helium gas



Molecules

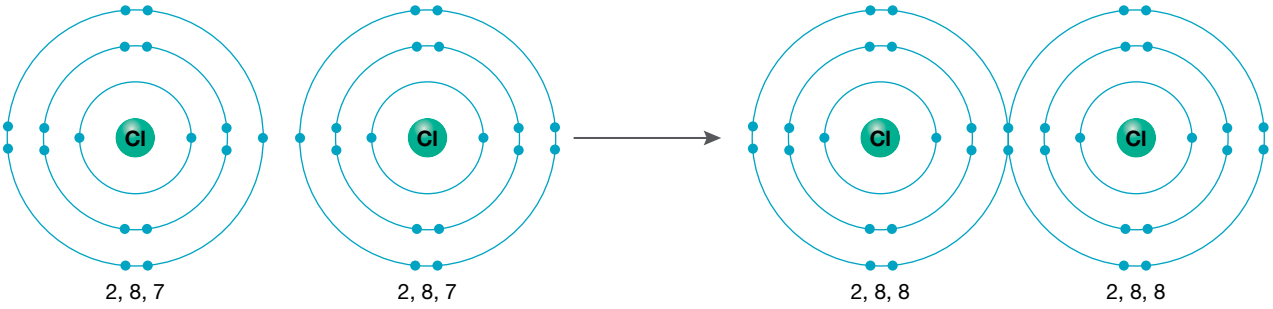
A molecule is formed when two or more non-metallic atoms join using covalent bonds. The atoms are combined in a fixed ratio and are electrically neutral.

Molecules of elements

Consider the two atoms of the element chlorine (figure 2.4). Chlorine (and all other members of group 17) exists in nature as **diatomic molecules**; that is, two atoms covalently bonded together. Each chlorine atom has seven valence electrons. To attain a noble gas configuration (an outer shell of eight electrons), each chlorine atom needs one more electron.

By sharing the remaining unpaired electron (six are already paired up), both atoms can obtain eight in their outer shells, as shown in figure 2.4. A covalent bond results from the simultaneous attraction of the two positive chlorine nuclei to the same shared pair of electrons. This bond holds the two atoms together. A molecule of chlorine forms and is represented by the symbol Cl_2.

FIGURE 2.4 Two atoms of chlorine sharing a pair of electrons to create a covalent bond and the chlorine molecule Cl_2



Hydrogen – a special case

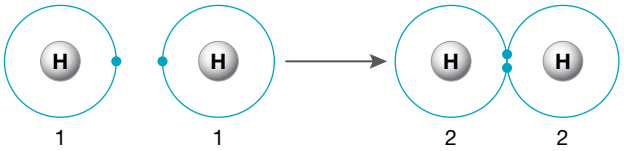
The covalent molecular element hydrogen also exists as a gas made up of discrete diatomic molecules. The hydrogen atom, however, is a special case because it is such a small atom and only requires two electrons to fill its outer shell. Each of the two hydrogen atoms contain one electron in its first (and outer) shell. They both need another electron to attain the stable electron configuration, 2, of the nearest noble gas, helium. This is achieved by each atom gaining a share in the electron of the other and forming a covalent bond, as shown in figure 2.5. In this way, the diatomic hydrogen molecule, H_2, is formed.

discrete separate, distinct; not in an infinite array or lattice

molecule group of atoms bonded together covalently

diatomic molecules substance containing two atoms only

FIGURE 2.5 The formation of a hydrogen molecule



A **covalent molecular element** is made up of identical atoms held together in discrete groups by covalent bonds. Hydrogen and chlorine are two of several non-metallic elements that are made up of diatomic molecules (see figures 2.6 and 2.7). Other molecular elements include oxygen, O_2; nitrogen, N_2; phosphorus, P_4; fluorine, F_2; and sulfur, S_8. These molecules are described as discrete because each molecule is separate and distinct from the others.

FIGURE 2.6 The molecules of hydrogen gas are diatomic.

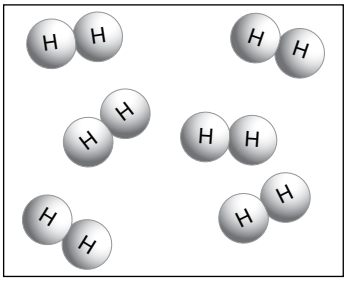
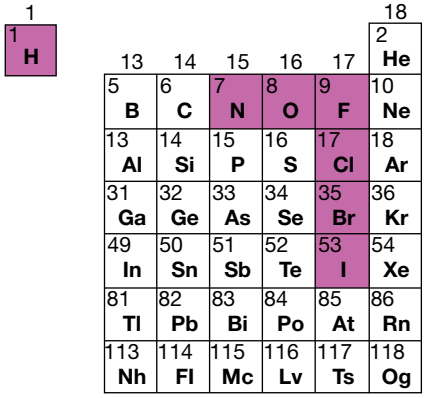


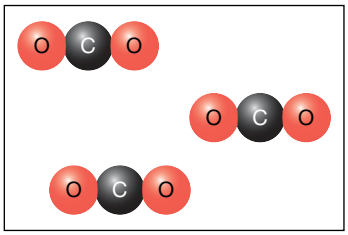
FIGURE 2.7 Remembering the seven diatomic molecules



Molecules of compounds

In addition to these elements, many compounds also exist as molecules. In a molecular compound, atoms of *different* non-metallic elements share electrons with each other. These compounds are often called **covalent molecular compounds** because they contain covalent bonds. Water, carbon dioxide and methane are examples of these compounds (figure 2.8).

FIGURE 2.8 Carbon dioxide is a covalent molecular compound formed from atoms of carbon and oxygen.



A theory for the structure of covalent molecular substances

The major features of the current theory of the structure of covalent molecular substances are as follows:

- The basic units of covalent molecular substances are groups of atoms called molecules. All molecules within a pure substance are identical.
- Adjacent atoms within a molecule share electrons in order to achieve a stable outer shell.
- Electrostatic attraction between the nuclei of adjacent atoms and the shared electrons causes the atoms in a molecule to be held together. This force of attraction is called covalent bonding.
- The overall charge on each molecule is zero and so adjacent molecules are only held together by weak intermolecular forces (see subtopic 2.4).

covalent molecular element
element made up of identical atoms held together by covalent bonds

covalent molecular compound
a molecular compound in which atoms of different elements share electrons with each other

The octet rule

As we have already seen, the driving force behind all bonding is to obtain a more stable electron arrangement for the atoms involved. This arrangement in most cases is that of the nearest noble gas. Because all the noble gases (except helium) have an outer shell of eight electrons, this tendency is often referred to as the **octet rule**. In other words, if an atom does not already have eight electrons in its outer shell, it will bond with other atoms in such a way as to achieve this.

The octet rule should not be confused with the concept of a 'full outer shell'. Many of the molecules you will come across involve period two atoms and, because these atoms have outer shells with a maximum capacity of eight, effectively no difference exists between these terms. However, once period three or higher atoms are involved, outer shells can contain more than eight electrons. In cases such as these, it is inaccurate to state that they are bonding with other atoms *'to obtain a full outer shell'*. The alternative, that they bond to *'achieve eight electrons in their outer shell'*, or they bond to *'achieve a noble gas configuration'*, is more accurate.

As with many generalisations, exceptions always exist. The most common of these is hydrogen. As discussed earlier, because hydrogen's outer shell is stable with only two electrons, the octet rule obviously cannot apply. Hydrogen atoms will always bond in such a way as to obtain *two* electrons in their outer shell.

2.2.3 Electron dot diagrams

To simplify drawings of atoms, **electron dot diagrams** (also called electron dot formulas) may be used to represent the outershell electrons. These can be drawn for atoms and molecules and are a useful way to understand the sharing of electrons in covalent bonds.

Electron dot diagrams for atoms

In an electron dot diagram:

- the atom's nucleus and all innershell electrons are replaced by its element symbol. (*Note*: The innershell electrons do not participate in the bonding.)
- the outershell electrons (**valence electrons**) are represented by dots or small crosses around the symbol in a square arrangement. These dots are arranged in pairs if more than four outer electrons are present.

In the electron dot diagram for nitrogen, the two dots arranged as a pair represent **non-bonding electrons** or **lone pairs**. The three that appear as single dots (unpaired) are *available for sharing* and are called **bonding electrons**.

FIGURE 2.9
Nitrogen has five outershell electrons

Elements of the same group in the periodic table have electron dot diagrams with the same number of dots or crosses. The halogens (group 17), for example, all have seven dots around their element symbols, representing seven outershell (valence) electrons.

Table 2.1 shows example electron dot diagrams for non-metals in the second period.

Electron dot structures within groups

Atoms in the same groups have the same number of valence electrons and so have the same electron dot structure.

octet rule a generalisation that works for many (but not all) atoms, stating that atoms will donate or share electrons in order to achieve eight electrons in their outer shells

electron dot diagram representation where the atom's nucleus and all inner shell electrons are replaced by its element symbol and the outer shell electrons are represented by dots around the symbol in a square arrangement

valence electrons electrons in the outermost shell of an atom; largely determine chemical properties of an element and contribute to chemical bond formation

non-bonding electrons electrons that are not involved in bonding

lone pairs pairs of outershell electrons which are not involved in bonding (non-bonding electrons)

bonding electrons the electrons shared as a pair to form a covalent bond

TABLE 2.1 Electron dot diagrams for non-metals in the second period of the periodic table

Group	Element	Valence electrons	Electron dot diagram	Lone pairs	Bonding electrons
14	C	4		0	4
15	N	5		1	3
16	O	6		2	2
17	F	7		3	1
18	Ne	8		4	0

Electron dot diagrams for molecular elements

Figure 2.4 in section 2.2.2 showed how two chlorine atoms can combine, by sharing two electrons, to each achieve a noble gas configuration (eight electrons in their outer shells). Electron dot diagrams can be used to represent this process in a simpler and clearer way (see figure 2.10).

FIGURE 2.10 The formation of a chlorine molecule, shown using electron dot representation

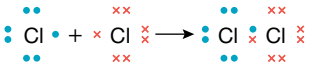


FIGURE 2.11 The formation of a hydrogen molecule, shown using the electron dot representation

In these representations, dots (•) and crosses (×) are used to represent the electrons that come from each atom. This does not mean different types of electrons exist — all electrons are the same. These diagrams are just a convenient way of tracking the origins of the electrons.

In the case of chlorine, each atom shares a pair of electrons to form a single covalent bond. Three pairs of unused electrons remain on each chlorine atom, and these form lone pairs. For hydrogen (figure 2.11), there are no unused electrons and, therefore, no lone pairs.

on Resources

Video eLesson Dot diagrams (eles-2474)

tlvd-0526

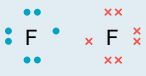
SAMPLE PROBLEM 1 Drawing electron dot diagrams of molecules with single covalent bonds

Draw an electron dot diagram for a molecule of fluorine, F_2.

THINK

1. Draw the dot diagrams for the atoms involved.
   Fluorine is in group 17 and has seven outershell electrons.
   Draw the two fluorine atoms dot diagrams with three pairs and a single electron in a square pattern.
2. Draw atoms close together and rearrange electrons so that each atom has eight electrons and there is a shared pair of electrons.
   TIP Check that each atom has eight electrons.

WRITE



PRACTICE PROBLEM 1

Draw an electron dot diagram for a molecule of iodine, I_2.

Multiple covalent bonds

Many molecular substances are held together by **multiple bonds**; that is, bonds formed when two atoms share two or more pairs of electrons.

- **Double bonds** occur when two atoms share two pairs of electrons in a covalent bond. Double bonds are found in molecules such as oxygen, O_2.
- **Triple bonds** occur when two atoms share three pairs of electrons, as in the case of nitrogen, N_2.

tlvd-0527

SAMPLE PROBLEM 2 Drawing electron dot diagrams of molecules with multiple covalent bonds

Draw an electron dot diagram for a nitrogen molecule, N_2.

THINK

1. Draw the dot diagrams for each of the nitrogen atoms that are present.
   Nitrogen has five electrons in its outer shell. It needs three more. It will share three of its electrons with three from another N atom. This leaves two unused on each atom.
2. Draw the diagrams with three electrons on one side and two on the other.
   Draw atoms close together so that the sharing of the electron pairs is obvious.
   TIP Check that each atom has eight electrons.

WRITE

PRACTICE PROBLEM 2

Draw an electron dot diagram for an oxygen molecule, O_2.

Valence structures and Lewis structures

Electron dot diagrams themselves may be further simplified; for example:

- A **valence structure**: substitutes a dash (–) for each pair of electrons. These dashes may represent bonding electron pairs (covalent bonds) or lone pairs (non-bonding electrons), depending on where they are located.
- A **Lewis structure** uses dashes for bonding pairs and dots for lone pairs. The terms Lewis structure, Lewis dot diagram and electron dot diagram are often used interchangeably.
- A **structural formula** omits the lone pairs, and thus shows the way the atoms in a molecule are connected. Although it is not necessary for a structural formula to indicate the shape of a molecule, it is possible to use them for this. This is discussed later in this topic.

multiple bond bond formed when two atoms share two or more pairs of electrons

double bond strong bond between two atoms formed by two pairs of electrons that are shared by the two nuclei

triple bond very strong bond between two atoms formed by three pairs of electrons that are shared by the two nuclei

valence structure a diagrammatic representation of the outershell electrons in a molecule; similar to an electron dot diagram but replaces each pair of electrons with a dash (–)

Lewis structure diagram that shows how outer shell electrons are arranged in a molecule using dashes for bonding pairs and dots for lone pairs

structural formula a diagrammatic representation of a molecule showing every bond, but not showing lone pairs of electrons

FIGURE 2.12 Different ways of representing the chlorine molecule

Molecular formula | Electron dot diagram | Lewis structure | Valence structure | Structural formula

Electron dot diagrams, Lewis structures, valence structures and structural formulas are different ways that molecules may be represented 'on paper'. The method chosen will often depend on the context in which the diagram is to be used.

on Resources

Weblink Covalent bonding tutorial

Electron dot diagrams for molecular compounds

The procedure for drawing electron dot diagrams for molecular elements is easily extended to molecular compounds. However, these can quickly become complicated as molecular size and the number of different elements increases. For this reason, the electron dot method tends to be used for smaller and simpler molecules.

Drawing the electron dot formula for a molecular compound

1. Place the atom with the most bonding electrons (the **central atom**) in the centre so that the other atoms can be arranged around it.
2. Draw dot diagrams for each of the atoms in the molecule. Remember that lone pairs do not participate in the bonding.
3. Attach all atoms in such a way that each atom is surrounded by eight electrons (except hydrogen).

Remember that if you have trouble achieving eight in an outer shell, you might have to consider multiple bonding.

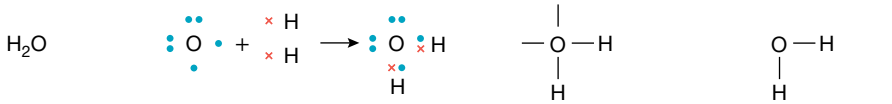
Water

The formula for water is H_2O. In each water molecule, an oxygen atom shares electrons with two hydrogen atoms so that they all achieve complete outer shells.

Notice that the central oxygen atom in a water molecule has two pairs of unbonded electrons (lone pairs). Although they do not participate in the covalent bond, they affect the shape of the molecule, and are shown in the valence structure by the two 'unbonded' dashes.

central atom the atom in a molecule with the most bonding electrons

FIGURE 2.13 Different ways of representing the water molecule

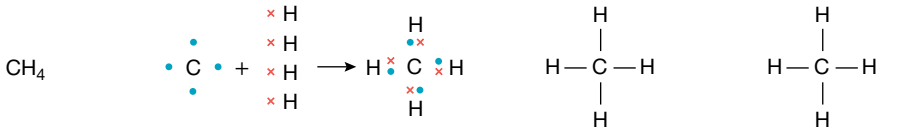


Molecular formula | Electron dot diagram | Valence structure | Structural formula

Methane

The formula for methane is CH_4. In each methane molecule, carbon is the central atom and forms a single bond with each of the four hydrogen atoms. All atoms have achieved complete outer shells.

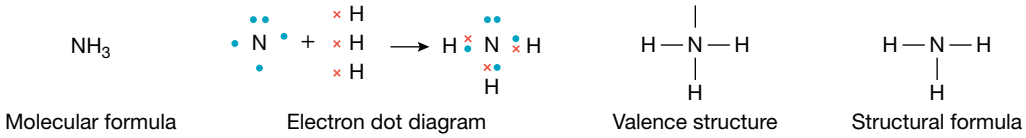
FIGURE 2.14 Different ways of representing the methane molecule



Ammonia

The formula for ammonia is NH_3. In each ammonia molecule, nitrogen is the central atom and forms a single bond with each of the three hydrogen atoms in addition to having one lone pair. This arrangement achieves a stable outer shell for each atom.

FIGURE 2.15 Different ways of representing the ammonia molecule



tlvd-0528

SAMPLE PROBLEM 3 Drawing element dot diagrams and valence structures for compounds

For hydrogen chloride, draw the following:

a. Electron dot diagram

b. Valence structure

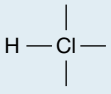
THINK

a. 1. Draw the electron dot diagrams for each atom
Hydrogen has one electron in its shell, and so needs one more. The chlorine atom has seven electrons in its outer shell, so it needs one more. Arrange atoms so that hydrogen has two electrons and chlorine has eight electrons.

2. Draw the diagrams together, showing overlap for sharing.
TIP Check that hydrogen has two electrons in its outer shell and eight electrons for chlorine.

b. Draw valence structure by replacing each pair of electrons with a dash.

WRITE



PRACTICE PROBLEM 3

For hydrogen fluoride, draw the following:

a. Electron dot diagram

b. Valence structure

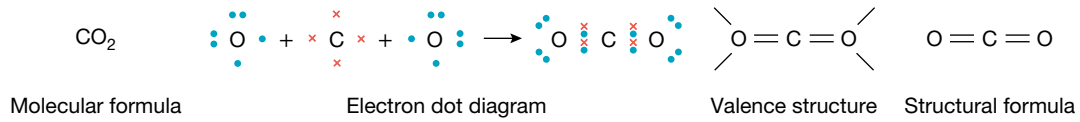
Electron dot diagrams for compounds with multiple bonds

Carbon dioxide

The formula for carbon dioxide is CO_2. To satisfy the octet rule in each carbon dioxide molecule, carbon forms a double bond with each of the oxygen atoms.

The central carbon atom in a carbon dioxide molecule now has eight valence electrons, and so does each oxygen atom. There are four bonding pairs and four lone pairs of electrons.

FIGURE 2.16 Different ways of representing the carbon dioxide molecule

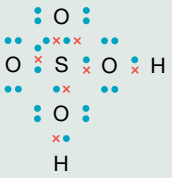


EXTENSION: Exceptions to the octet rule

The octet rule is a very useful *rule of thumb* for chemists when working out bonding within molecules and molecular shape.

However, many molecules do not obey this rule. In some instances, molecules bond with fewer than eight electrons in their outer shell, and in many cases have more. Obviously, the latter cannot occur for period two elements due to the capacity of the second shell. However, for periods three and beyond, the outer shell has room for more than eight electrons. Although many compounds involving period three and beyond still obey the octet rule, some don't. Some of these are quite common; for example, sulfuric acid, the structure of which is shown in figure 2.17.

FIGURE 2.17
Electron dot diagram of sulfuric acid showing 12 electrons on the sulfur atom



If you count the electrons in the outer shell of the sulfur atom, you will notice that there are 12! Some other molecules that also exhibit this feature are PCl_5, SF_6, IF_5, SO_3, H_3PO_4 and $HClO_4$.

Look up the structures for these molecules (or even try to work them out first). Also look for any other examples, and whether you can find some examples of molecules that have fewer than eight outershell electrons around their central atom.

This exercise is a good illustration that many of the tools and techniques used by chemists have their limitations. The bonding models discussed in this topic are a further example of this.

This does not mean that such ideas are not useful. The main thing to remember is that they serve a purpose and, so long as their limitations are recognised, no reason exists to not keep using them.

2.2.4 Physical representations

The representations discussed so far have one obvious disadvantage: they are two dimensional. Physical models allow us to visualise molecules in three dimensions. As we shall see later in this topic, the three-dimensional shape of a molecule can have a significant effect on its properties. The two main types of three-dimensional model used by chemists to represent molecules are **ball-and-stick models** and **space-filling models**. These can be used to inform **computer-generated models**, and you will also use them in topic 4 with ionic compounds.

ball-and-stick models representation of a molecule in which the atoms are shown as balls and the bonds as sticks

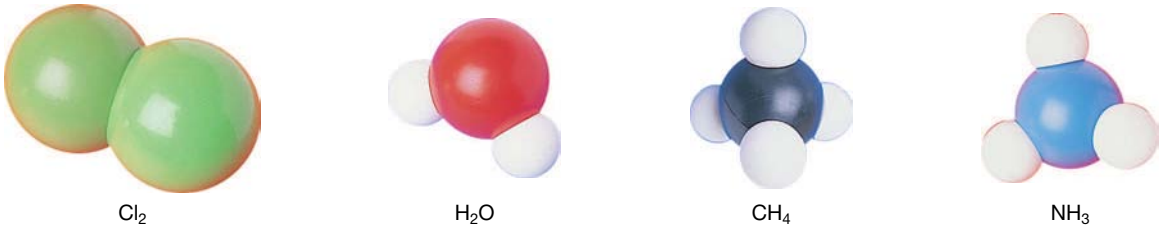
space-filling model three-dimensional representation of a molecule that shows the relative sizes of atoms within the molecule and the distances between them

computer-generated model representation of a molecule that is produced by a computer

Space-filling models

Space-filling models use spheres (usually made of plastic) to represent atoms. These are colour coded to represent different types of atoms. Typical examples are black (carbon), white (hydrogen), red (oxygen), blue (nitrogen) and green (chlorine). The hydrogen spheres are usually smaller than the others to represent the much smaller relative size of the hydrogen atom. These are attached to each other so as to resemble the shape of the molecule as closely as possible.

FIGURE 2.18 Space-filling models for (left to right) chlorine, Cl_2, water, H_2O, methane, CH_4, and ammonia, NH_3



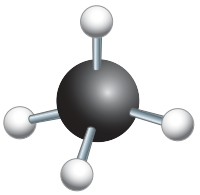
Cl_2 H_2O CH_4 NH_3

Ball-and-stick models

Ball-and-stick models are similar to space-filling models except that the spheres are separated and joined by sticks. Each stick represents a covalent bond. Ball-and-stick models have the advantage that they can show multiple bonds (see table 2.2).

In summary, various methods can be used to represent molecules (see table 2.2).

FIGURE 2.19 Ball-and-stick model for methane, CH_4.

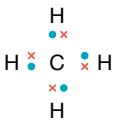


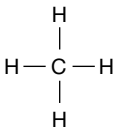
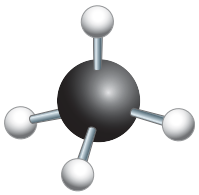
2.2.5 Limitations of these representations

Why are different methods used to represent molecules? The answer is that chemists use different ways for different purposes. Each representation has its own strengths and limitations. The method chosen will often depend on the context in which it is to be used. If the bonding between atoms is being emphasised, a ball-and-stick model is better than a space-filling one. While both electron dot and valence structures will also highlight bonding between atoms, the former may lose clarity among a maze of dots representing the electrons. If shape is important, both ball-and-stick and space-filling may be adequate, with both being superior to electron dot and valence structures. In many situations, especially in organic chemistry, a structural formula is an excellent way to represent molecules.

Table 2.2 lists some of the limitations of each of these methods. You may be able to think of more.

TABLE 2.2 Methods of molecular representation with some of their limitations

Method	Limitations
Electron dot diagrams (and Lewis structures) 	• Two dimensional • Does not indicate shape • Tedious and slow to produce • Not useful for large molecules
Valence structure 	• Two dimensional • Does not necessarily indicate shape* • Can be tedious and slow to write for larger molecules • Important lone pairs may be difficult to distinguish from non-important ones

Structural formula 	• Two dimensional • Does not necessarily indicate shape* • Can be tedious and slow to write for larger molecules
Space-filling 	• Does not show the bonding between atoms • Relative sizes of atoms involved are only approximate
Ball-and-stick 	• Can be physically difficult to produce for larger molecules • Kits to produce can be expensive

* These methods can be adapted to give an indication of shape by using shape diagrams (see subtopic 2.3).

2.2 Activities

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2.2 Quick quiz on | **2.2 Exercise** | **2.2 Exam questions**

2.2 Exercise

1. MC Which of the following lists contains molecular compounds only?
   - **A.** CO_2, Cl_2 and HCl
   - **B.** LiBr, Br_2 and $KBrO_4$
   - **C.** H_2O, Al_2O_3 and NO_2
   - **D.** $ZnSO_4$, H_2SO_4 and SO_3
2. MC Which of the following elemental substances will not be bonded by covalent bonds?
   - **A.** N_2
   - **B.** C
   - **C.** I_2
   - **D.** Al
3. MC Which of the following shows the electron dot structure for hydrogen cyanide, HCN?
   - **A.**
   - **B.**
   - **C.**
   - **D.**

4. Draw an electron dot diagram for an atom of the following.
   a. Tellurium (Te)
   b. Arsenic (As)
5. Complete the following table.

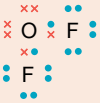
Group	Element	Valence electrons	Electron dot diagram	Lone pairs	Bonding electrons
	S				
	Si				
	P				
	Cl				
	Br				
	Se				
	Ar				

6. Draw electron dot diagrams and structural formulas for the following molecules.
   a. CS_2
   b. C_2F_4
   c. C_3H_8
   d. HCN
7. Non-metals often exist as diatomic gases at room temperature. Why does neon not exist as Ne_2?
8. Draw the electron dot structure for phosphine, PH_3.
9. Name one advantage that electron dot structures have over other models.
10. a. Draw a dot diagram and structural formula for HBr.
    b. Label the bonding pairs and lone pairs.

2.2 Exam questions

Question 1 (1 mark)

MC Which of the following correctly shows the electron dot structure for OF_2?

A. 
B.
C.
D.

Question 2 (1 mark)

MC 'Covalency' is a term for the number of covalent bonds an element makes when bonding.

Which of the following elements has the highest covalency?

A. Nitrogen
B. Silicon
C. Oxygen
D. Chlorine

Question 3 (1 mark)

Which of the following substances contain discrete diatomic molecules? Justify your response.

a. Magnesium oxide
b. Iodine
c. Aluminium
d. Neon
e. Phosphorus
f. Carbon dioxide

Question 4 (4 marks)

Complete the following table.

Molecule	Electron dot structure	Valence structure
H_2S		
CH_3Br		

Question 5 (4 marks)

Carbon combines with chlorine to form a covalent molecule.

a. How many chlorine atoms will combine with carbon in the molecule? **(1 mark)**
b. Draw the Lewis diagram for this molecule. **(1 mark)**
c. How many bonding pairs of electrons are present? **(1 mark)**
d. How many non-bonding pairs of electrons are present? **(1 mark)**

More exam questions are available in your learnON title.

2.3 Shapes of molecules

KEY KNOWLEDGE

- Shapes of molecules (linear, bent, pyramidal and tetrahedral, excluding bond angles) as determined by the repulsion of electron pairs according to valence shell electron pair repulsion (VSEPR) theory
- Polar and non-polar character with reference to the shape of the molecule

Source: VCE Chemistry Study Design (2023–2027) extracts © VCAA; reproduced by permission.

2.3.1 Molecules are three-dimensional

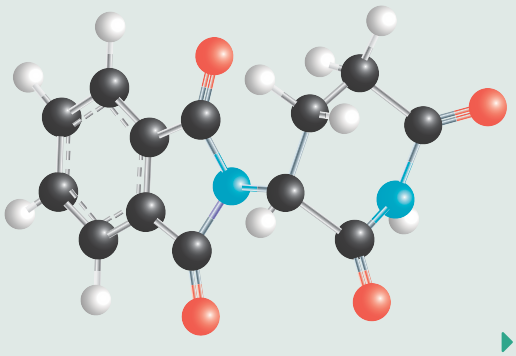
The shape of a molecule can have an important effect on its properties and how it behaves. For this reason, it is important to know not only how atoms are joined together to make a molecule, but also in what orientation they are with respect to one another. This is especially important for biological molecules and drugs where a relatively small change in shape can have dramatic differences in effect.

CASE STUDY: Thalidomide and the effect of atom orientation

Thalidomide, $C_{13}H_{10}N_2O_4$, is a drug that was first produced in the 1950s. It became popular as an over-the-counter drug sold to combat morning sickness in pregnant women. However, it soon became apparent that it was producing a devastating side effect: children were being born with severely deformed limbs. Over 10 000 children were born before the drug was withdrawn. Today the use of the drug is strictly controlled and prescribed for treating leprosy and certain types of cancers.

Notice the usual colour convention in figure 2.20: hydrogen atoms are white, oxygen atoms are red, nitrogen atoms are blue and carbon atoms are black.

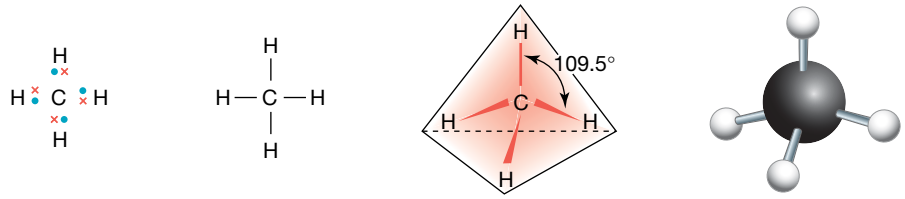
FIGURE 2.20 Ball-and stick-model of thalidomide ($C_{13}H_{10}N_2O_4$)



Thalidomide exists in two forms that are nearly identical. The only difference is a small change in orientation around one of the carbons. One form *did* cure morning sickness. However, the other form caused the birth defects. To make matters worse, the body can convert the harmless form into the harmful one, making it almost impossible to prevent the drug's serious side effects.

All discrete molecules have a definite three-dimensional shape. Electron dot diagrams and structural formulas fail to represent the three-dimensional shapes of molecules. For example, the electron dot diagram and structural formula of methane, CH_4, show the molecule in only two dimensions. In reality, methane molecules exist in three dimensions. The hydrogen atoms are arranged in a tetrahedral shape around the central carbon atom, as shown by the ball-and-stick diagram included in figure 2.21.

FIGURE 2.21 The electron dot diagram and structural formula of methane do not show its three-dimensional structure, which is tetrahedral. This structure is better represented using a ball-and-stick model.



2.3.2 The valence shell electron pair repulsion (VSEPR) theory

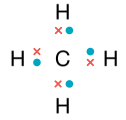
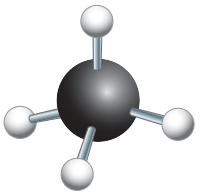
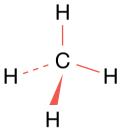
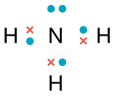
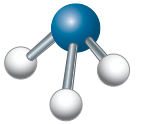
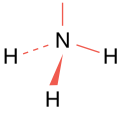
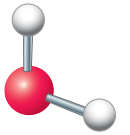
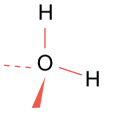
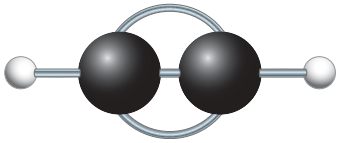
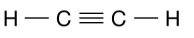
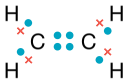
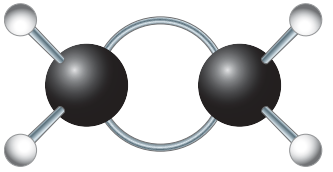
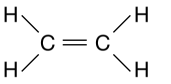
The **valence shell electron pair repulsion (VSEPR) theory** provides a relatively simple basis for understanding and predicting molecular geometry. The theory requires only that the number of outershell electrons for each atom in the molecule be known. The electron pairs in the molecule repel each other and take up positions as far from one another as possible. When determining the shapes of molecules, the electron pairs of a multiple bond count as only one pair of electrons for prediction purposes. Lone pairs should also be considered, especially when they are on a central atom. Examples of the five most common molecular shapes are given in table 2.3. Note that, in the shape diagrams, a dotted line indicates that the bond is directed into (or behind) the plane of the paper and the wedge indicates that the bond is directed out (or in front) of the plane of the paper. The solid lines represent bonds that are in the plane of the paper. The use of ball-and-stick-modelling is an excellent way to understand this theory.

valence shell electron pair repulsion (VSEPR) theory model used to predict the shape of a molecule based on minimising electron repulsions around a given atom in a molecule

VSEPR theory

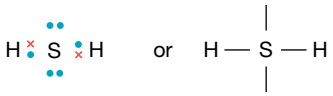
VSEPR theory states the electron pairs in the molecule repel each other and take up positions as far from one another as possible.

TABLE 2.3 Shapes of some common molecules

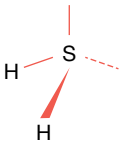
Compound	Electron dot diagram	Number of lone pairs around central atom	Number of bonding electron groups (pairs) around central atom	Shape	Ball-and-stick model	Shape diagram
Methane $CH_4(g)$		0	4	Tetrahedral (around C atom)		
Ammonia $NH_3(g)$		1	3	Pyramidal (around N atom)		
Water $H_2O(l)$		2	2	Bent, or V-shaped (around O atom)		
Ethyne $C_2H_2(g)$		0 for each C atom	2 for each C atom	Linear (around each C atom)		
Ethene $C_2H_4(g)$		0 for each C atom	3 for each C atom	Planar (around each C atom)		

For example, consider hydrogen sulfide, H_2S, and silane, SiH_4.

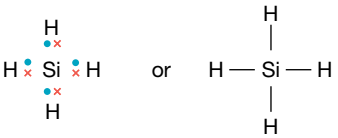
To predict the shape of H_2S, first draw either its electron dot diagram or valence structure.



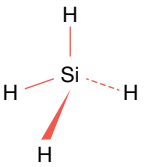
Next, count all the electron pairs (both bonding and non-bonding) around the central atom (in this case, sulfur) and arrange them as far from each other as possible. Arrange the four electron pairs in three dimensions (this orientation is a tetrahedral one). The shape diagram may now be drawn. Once drawn, the shape is then named by the positions of only the bonding pairs (or atoms). Therefore, the shape of hydrogen sulfide is bent, just like water (see table 2.3).



A similar procedure yields the shape of the SiH_4 molecule. Once again, four pairs of electrons are around the central atom, and they therefore adopt a tetrahedral orientation.



Note that the structure of silane, SiH_4, is the same as the structure for methane, CH_4, because both silicon and carbon are in group 14 and so have four electrons in their outer shell.



Tips to help recognise the shape of a molecule

- Determine the arrangement of electron pairs around the central atom.
- Decide on the basic structure (for example, four pairs of electrons will be tetrahedral).
- To decide on the final shape, look at the arrangement of the bonding pairs only. (For example, if only two bonding pairs are in the tetrahedral arrangement, the shape is described as bent.)

on Resources

Video eLessons Molecular shapes (eles-2475)
VSEPR — Sulfur dioxide (eles-2488)

tlvd-0529

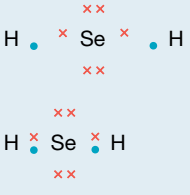
SAMPLE PROBLEM 4 Drawing shape diagrams for molecules, H_2Se

Using its electron dot diagram and its valence structure, draw the shape diagram for a molecule of H_2Se.

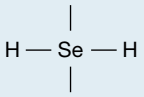
THINK

1. Draw the electron dot diagram for the atoms. H must only show one electron. Se is in group 16 so it must show six electrons in the outer shell.
2. Arrange H atoms around the Se with an overlap to show sharing. (Remember that electron sharing may be clearer using a valence structure.)

WRITE



3. Draw the valence structure.



4. Draw the 3D arrangement, which is tetrahedral. Count the pairs of electrons around the central atom to identify four pairs. Two lone pairs of electrons and two pairs of bonding electrons around the central atom correspond to a V-shaped (or bent) arrangement.
TIP Selenium is in group 16, so this is the same arrangement as H_2O and H_2S.

PRACTICE PROBLEM 4

Using its electron dot diagram and its valence structure, draw the shape diagram for a molecule of phosphine, PH_3.

If when drawing structures single bonds are not enough to use all the electrons available, consider inserting double or triple bonds. Many organic compounds, in particular, contain one or more double or triple carbon to carbon bonds. Ethene and ethyne are the simplest examples.

tlvd-0530

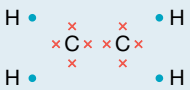
SAMPLE PROBLEM 5 Drawing shape diagrams for molecules, C_2H_4

Using its electron dot diagram and its valence structure, draw the shape diagram for a molecule of ethene, C_2H_4.

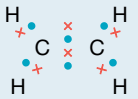
THINK

1. H atoms are always arranged around the outside of the central atom, so the two C atoms must be in the centre.

WRITE



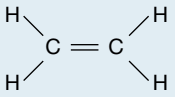
2. Draw each H showing one electron and each C showing four electrons. Arrange H atoms around the C with an overlap to show sharing.
TIP If sharing is difficult to show with an electron dot diagram, try forming multiple bonds. Electron sharing in this molecule works with a double bond!



3. Replace pairs of electrons with dashes. Draw the molecule showing four electrons between each C as a double bond.

   Three areas of electrons are around each C (double bonds count as one) so the shape is planar around each carbon, with bonds 120° apart.

   This planar structure makes the whole molecule flat.



PRACTICE PROBLEM 5

Using its electron dot diagram and its structural formula, draw the shape diagram for a molecule of C_2H_2.

The shape of a molecule is not necessarily the same as the distribution of electron pairs around the central atom. The shape refers to the three-dimensional orientation of the *atoms* in space.

VSEPR summary

- A molecule will assume the shape that minimises electron pair repulsions.
- Electron pairs around the central atom will be as far apart as possible.

In general:

- Molecules with no lone pairs around the central atom are tetrahedral, unless a double or triple bond creates a linear or planar structure (depending upon the distribution of atoms around the central atom).
- Molecules with a single lone pair around the central atom are pyramidal.
- Molecules with two lone pairs around the central atom are bent (V-shaped).

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elog-1621

EXPERIMENT 2.1 online only

Building molecular models

Aim

To construct molecular models and then use these to determine molecular shapes

2.3.3 Polar and non-polar bonds

All bonding involves the attractions set up between positive charges and negative charges. The attraction is electrostatic in nature and electrons play a critical role. Even though molecules are electrically neutral in an overall sense, the possibility exists of regions within a molecule where small, localised imbalances may occur. **Polarity** is the term that chemists use to describe this situation.

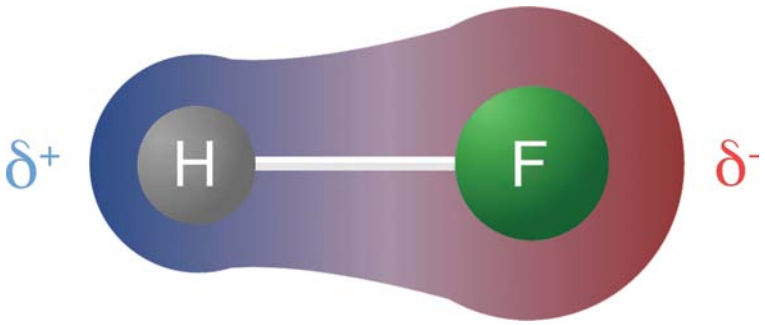
Polar covalent bonds

The reason for this imbalance of charge is due to the different electronegativities of the atoms in the covalent bond. The shared pair of electrons moves closer to the more electronegative atom. If one atom has a lower electronegativity than the other, there will be a slight deficiency of electrons in a region (and, therefore, a slight positive charge). This is indicated by the symbol $\delta+$ (taken from the lowercase Greek letter for delta). Likewise, a slight excess of electrons is indicated by $\delta-$. The result of this unequal sharing of electrons is a **polar covalent bond**.

polarity localised imbalances in electric charges within a molecule resulting in a negatively charged end and positively charged end

polar covalent bond formed when two atoms that have different electronegativities share electrons unevenly

FIGURE 2.22 A polar covalent bond forms between hydrogen and fluorine as the electrons are attracted to the highly electronegative fluorine atom.



The electronegativities of atoms can be viewed on a scale developed by Linus Pauling in which the most electronegative atom, fluorine, is assigned a value of 4.0 (see figure 2.23). Fluorine attracts electrons almost twice as well as hydrogen, which is given an electronegativity value of 2.1. No values are assigned for the noble gases. Electronegativity can be likened to a 'tug of war', where the two teams are like the atoms on each end of the bond trying to pull the electrons towards them.

FIGURE 2.23 The Pauling scale of electronegativities

Increasing electronegativity →

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

TIP: The closer an atom is on the periodic table to the most electronegative element, fluorine, the greater is its electronegativity.

Summary of the Pauling scale of electronegativities

The following electronegativity trends may be seen in the periodic table:

- Electronegativities increase from left to right within a period.
- Electronegativities decrease from top to bottom within a group.
- Metals generally have lower electronegativities than non-metals.

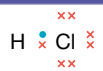
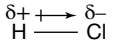
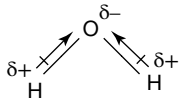
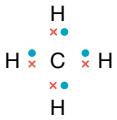
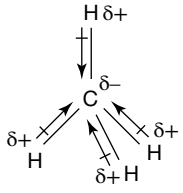
Polar covalent bonds can be said to have a charge separation, or a **bond dipole**. The dipole is represented by an arrow pointing to the more electronegative end of the molecule (showing the direction the shared electrons are pulled towards) and resembling a plus symbol at the more positive end.

bond dipole separation of charge in a polar covalent bond

Polar covalent bonds

Polar covalent bonds are covalent bonds in which the bonding electrons are unequally shared and, therefore, unsymmetrically distributed between the nuclei of two bonded atoms.

TABLE 2.4 Polar covalent bonds

Molecule	Electron dot diagram	Structural formula and polarity
Hydrogen chloride		
Water		
Methane		

Non-polar covalent bonds

Non-polar covalent bonds are covalent bonds in which the bonding electron pair is shared equally and uniformly distributed between the nuclei of two bonded atoms. Such bonds can result only when two atoms of the same element or two atoms of equal electronegativity simultaneously attract a shared pair of electrons.

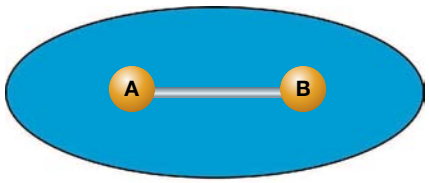
non-polar covalent bonds formed between atoms with the same electronegativity

A polar covalent bond is not purely covalent since there is not an *equal* sharing of electrons. The relationship between electronegativity and bond type is shown in table 2.5.

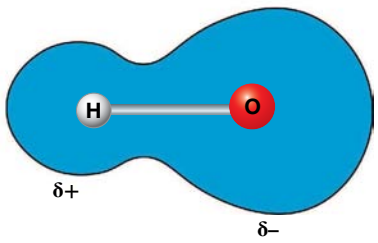
TABLE 2.5 Relationship between electronegativity and bond type

Difference in electronegativity between bonding atoms	Type of bond formed
Zero; e.g. H (2.1) and H (2.1)	Non-polar covalent
Medium; e.g. H (2.1) and Cl (3.0)	Polar covalent

FIGURE 2.24 The difference between polar and non-polar bonds



Non-polar bond
Atoms A and B have equal electronegativity, so share electrons equally (blue region).



Polar bond
The oxygen atom (O) has a higher electronegativity than the hydrogen atom (H) so has a greater share of electrons. This unequal share of electrons forms a dipole with one end of the molecule slightly negative (δ–) and the other end slightly positive (δ+).

TABLE 2.6 Non-polar covalent bonding in molecules

Molecule	Electron dot diagram	Structural formula
Hydrogen		H — H
Chlorine		Cl — Cl
Oxygen		O = O
Carbon disulfide		S = C = S

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SAMPLE PROBLEM 6 Determining if bonds are polar or non-polar

Refer to the Pauling scale in figure 2.23 and state whether the following bonds are polar or non-polar.

a. Oxygen to oxygen bonds
b. Hydrogen to phosphorus bonds
c. Oxygen to chlorine bonds

THINK	WRITE
A bond will be polar if a difference in electronegativity exists between the two atoms on each end.	
a. No difference exists for O (3.5) to O (3.5).	Oxygen to oxygen bonds are non-polar.
b. No difference exists for H (2.1) to P (2.1).	Hydrogen to phosphorus bonds are non-polar.
c. A difference does exist for O (3.5) to Cl (3.0).	Oxygen to chlorine bonds are polar.

PRACTICE PROBLEM 6

Refer to figure 2.23 and state whether the following bonds are polar or non-polar.

a. Chlorine to chlorine bonds
b. Phosphorus to chlorine bonds

2.3.4 Polar and non-polar molecules

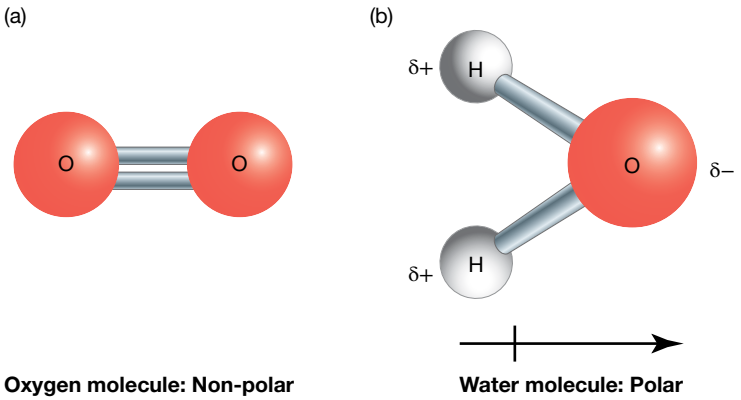
Non-polar molecules contain only non-polar bonds, such as O_2.

In a **polar molecule**, one end of the molecule is slightly negative, and one end is slightly positive. Such a molecule is sometimes called a dipolar molecule, or molecular dipole. All polar molecules contain polar bonds. However, some molecules with polar bonds are non-polar. This occurs when the individual dipoles are arranged in such a way that they *cancel each other out*. So having polar bonds in a molecule does not necessarily mean that the molecule is polar overall. The polarity or non-polarity of the molecule depends, rather, on the *direction* of the bond dipoles in the molecule, which is determined by the molecule's *shape*. Figure 2.25 shows a non-polar and a polar molecule.

non-polar molecule molecule that does not have permanent dipoles or is symmetrical

polar molecule a molecule which, due to its polar bonds *and* its asymmetric shape, has an overall imbalance in the distribution of its electrons

FIGURE 2.25 (a) A non-polar oxygen molecule and (b) a polar water molecule

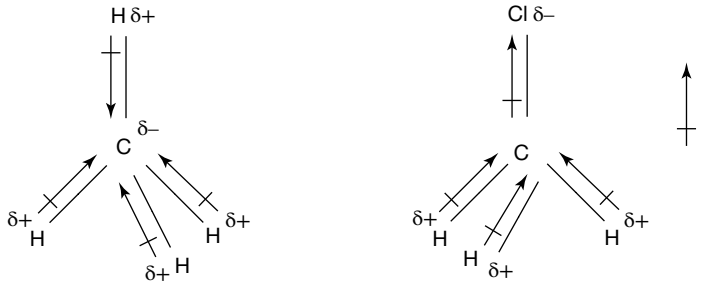


Oxygen molecule: Non-polar **Water molecule: Polar**

In the non-polar oxygen molecule, O_2, the oxygen atoms have equal electronegativity and share electrons equally. There is no more positively or negatively charged end of the molecule so it is non-polar (a). Water molecules are polar because the oxygen atom has greater electronegativity than the hydrogen atoms, so it becomes the more negatively charged end of the molecular dipole (b).

Another example is methane, CH_4, and chloromethane, CH_3Cl. Methane is non-polar because the difference in electronegativity is small and the bond dipoles cancel each other out; CH_3Cl, however, is polar because of the more electronegative chlorine atom that is not balanced out.

FIGURE 2.26 Methane is a non-polar molecule and chloromethane is polar.

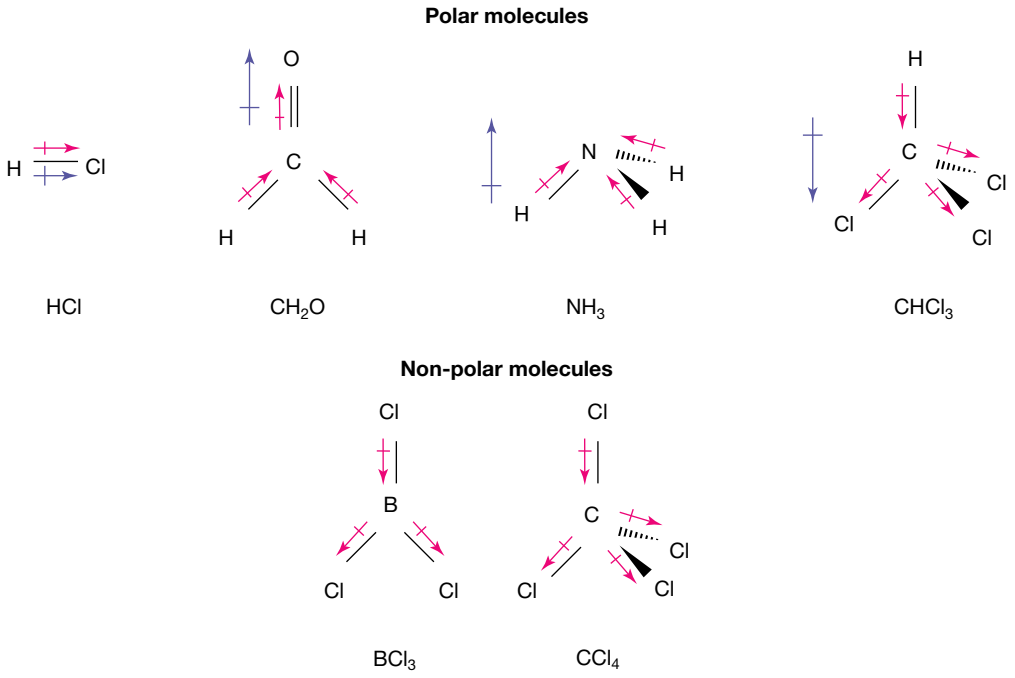


Methane, CH_4 **Chloromethane, CH_3Cl**

To determine the polarity of a molecule, the following steps should be followed:

1. Draw an electron dot diagram of the molecule.
2. Apply the VSEPR rules to draw a shape diagram of the molecule.
3. Use electronegativities to determine bond dipoles.
4. Use the shape diagram along with bond dipoles to determine whether the molecule is polar or non-polar. If bond dipoles cancel each other out, the molecule is non-polar. If bond dipoles do not cancel each other out, one side of the molecule attracts electrons more than another side and the molecule is polar.

FIGURE 2.27 All of these molecules have polar bonds but only some are polar molecules.



Tips for identifying polar and non-polar molecules

- Linear, planar and tetrahedral molecules with equal polar bonds that cancel each other out are non-polar.
- If the bond dipoles are not equal or do not cancel each other out, the molecule is polar. For example, CH_4 is non-polar whereas CH_3Br is polar because Br is more electronegative and the resulting C—Br bond dipole is not cancelled by the C—H bond dipoles.
- Bent or pyramidal molecules are polar because their polar bonds do not cancel each other out.

tlvd-0532

SAMPLE PROBLEM 7 Predicting the polarity of molecules

Predict the polarity of the following molecules.

a. Water

b. Carbon dioxide

THINK

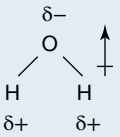
Recall that a molecule being polar requires two things: it must have polar bonds and its shape must be such that the dipoles from these bonds do not cancel out (that is, it must be asymmetrical).

Therefore, check electronegativities first and then draw the shape diagrams.

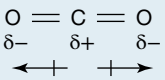
a. Water, H_2O, is a V-shaped (bent) molecule because of the two lone pairs on O. The two O—H bond dipoles are equal but since they are not at an angle of 180° to one another (that is, aligned in opposite directions) they do not cancel each other out. A water molecule is, therefore, polar.

b. Carbon dioxide, CO_2, is a linear molecule with two polar bonds. However, the bond polarities cancel each other out because they are in opposite directions. Electrons are not attracted preferentially to one side or the other of the molecule. Carbon dioxide is, therefore, a non-polar molecule.

WRITE



Water has polar bonds and is a polar molecule.



Carbon dioxide has polar bonds but is a non-polar molecule.

PRACTICE PROBLEM 7

Predict the polarity of the following molecules.

a. Iodine

b. Hydrogen bromide

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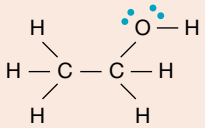
2.3 Quick quiz on | 2.3 Exercise | 2.3 Exam questions

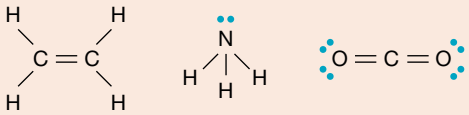
2.3 Exercise

1. MC The molecule shown can best be described as



   A. polar with all bonds polar.
   B. non-polar due to symmetry.
   C. polar overall with a mixture of polar and non-polar bonds.
   D. non-polar due to lack of polar bonds.

2. MC Consider the structures given. Which is the most polar bond present in these structures?



   A. C=C
   B. N—H
   C. C=O
   D. C—H

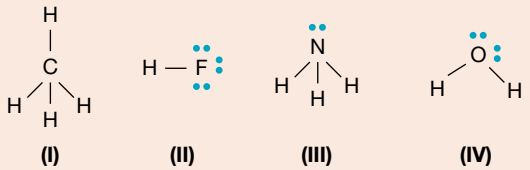
3. Explain the importance of non-bonding pairs in the molecular shape.
4. Identify the diatomic molecules in the following list that are polar.
   HF, F_2, FCl, HCl, H_2
5. Using its electron dot diagram or its valence structure, draw the shape diagram for a molecule of AsH_3.
6. State whether the following bonds are polar or non-polar. Justify your answer.
   a. Sulfur to oxygen bonds
   b. Bromine to bromine bonds
7. Explain whether the following molecules are polar or non-polar.
   a. Chlorine (Cl_2)
   b. Ammonia
8. Use the electronegativity trends in the periodic table (shown in figure 2.23 in section 2.3.3) to predict whether the following bonds are polar or non-polar.
   Show the bond dipoles where they are present using $\delta+$ and $\delta-$.
   a. H—F
   b. O—H
   c. C—H
   d. N—H
   e. C—C

9. Sketch the shapes of the following molecules and use your knowledge of electronegativity to draw bond dipoles for each to predict whether it is polar or non-polar.
   a. CH_4
   b. CH_3Cl
   c. CH_2Cl_2
   d. $CHCl_3$
   e. CCl_4
10. A molecule has a tetrahedral shape. Does this mean that it is non-polar? Explain.

2.3 Exam questions

Question 1 (1 mark)

MC Which of the following lists the shapes of the molecules shown in the diagram below in order from I to IV?



(I) (II) (III) (IV)

A. Pyramidal, linear, tetrahedral, bent
B. Tetrahedral, planar, pyramidal, bent
C. Pyramidal, planar, tetrahedral, bent
D. Tetrahedral, linear, pyramidal, bent

Question 2 (1 mark)

MC Which of the following has a different shape from the other three molecules?
A. CO_2
B. H_2O
C. SF_2
D. H_2S

Question 3 (5 marks)

Hydrogen peroxide, H_2O_2, can be used as mild antiseptic.
a. Draw a dot diagram for this molecule. **(1 mark)**
b. How many lone pairs are present? **(1 mark)**
c. Show a possible structure and label any polar or non-polar bonds. **(3 marks)**

Question 4 (6 marks)

Fluoromethane, CH_3F, is a colourless flammable gas. Complete the following for this molecule.
a. Draw the electron dot diagram. **(1 mark)**
b. Draw the structural formula. **(1 mark)**
c. State the shape of the molecule. **(1 mark)**
d. Identify the type(s) of bond present. **(2 marks)**
e. Explain whether this molecule is polar or non-polar. **(1 mark)**

Question 5 (9 marks)

The formulas for the molecules BF_3 and NF_3 are similar, yet their shapes are different.
a. Draw the structural formulas for each molecule. **(2 marks)**
b. State what is unusual about the electron arrangement in BF_3. **(1 mark)**
c. State the shapes of the molecules. **(2 marks)**
d. Explain why the shapes are different. **(2 marks)**
e. Explain if these molecules are polar or non-polar. **(2 marks)**

More exam questions are available in your learnON title.

2.4 Comparing intramolecular bonding and intermolecular forces

KEY KNOWLEDGE

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

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2.4.1 Intermolecular attractions

The covalent bonding between atoms in molecules discussed so far in this topic, where electrons are shared, is described as **intramolecular bonding**. This is relatively strong bonding compared to the weak interactions that hold atoms and molecules together in the liquid and solid states, which are described as **intermolecular forces** or attractions.

- Intramolecular bonding is the covalent bonds *within* molecules.
- Intermolecular forces or intermolecular attractions exist *between* molecules and hold molecules to each other. These attractive forces are much weaker than covalent bonds, but they determine whether a molecular compound exists as a solid, liquid or gas. The temperature at which a molecular substance melts or boils, therefore, depends on the strength of the intermolecular forces. The stronger the intermolecular forces, the higher the melting or boiling point.

FIGURE 2.28 Intermolecular forces hold water molecules together in liquid water.

Effect of intermolecular forces

Melting and boiling points depend on intermolecular forces, which are forces *between* atoms or molecules.

Three types of intermolecular forces are possible. In some molecular substances, more than one force may be operating between the molecules.

- The weakest attractions between molecules are **dispersion forces**.
- Slightly stronger are the **dipole–dipole interactions**.
- The third and strongest intermolecular force is called **hydrogen bonding**. These forces are weak but critical to our existence; the numerous covalent compounds in our body including proteins and the strands of our DNA are held together by these intermolecular forces.

intramolecular bonding internal bonds within a molecule

intermolecular forces forces that exist between molecules

dispersion force the bond between adjacent molecules formed by instantaneous dipoles; this weak non-directional bonding is also known as van der Waals force

dipole–dipole interactions weak bonding caused by the positive end of one dipole attracting the negative end of another dipole

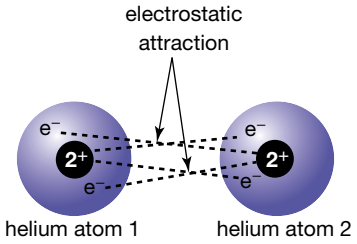
hydrogen bonding the bond between a hydrogen atom covalently bonded to an atom of F, O or N and another molecule that also contains an atom of H, F, O or N

Dispersion forces

In a sample of molecules, the nuclei of atoms in one molecule can attract the electrons of atoms in neighbouring molecules (in addition to attracting their own electrons in forming a covalent bond). All electrons are attracted by all neighbouring nuclei.

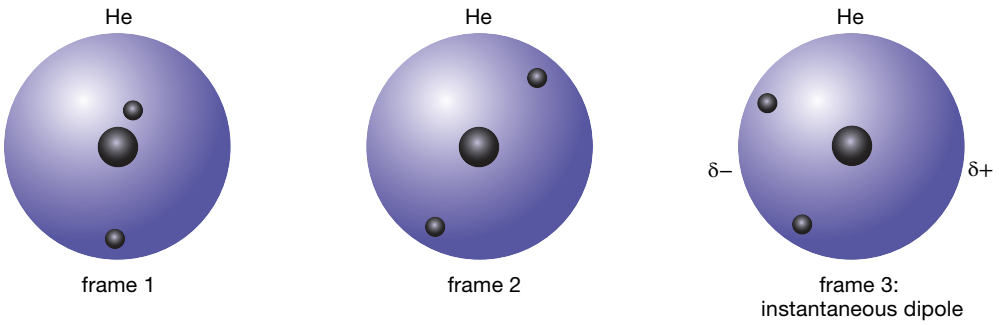
Dispersion forces are the weakest of the intermolecular forces. To be condensed to a liquid, the noble gas helium must be cooled to –269 °C so that the atoms are travelling slowly enough for the dispersion forces to have a better chance of pulling these two atoms together. The helium nucleus in one atom attracts the electrons in a neighbouring atom. If cooled to –272 °C, at pressures of 25 atm, liquid helium solidifies.

FIGURE 2.29 Below 269 °C helium is a liquid because atoms are very weakly attracted to each other by the electrostatic attraction of one nucleus to the electrons of other atoms.



Dispersion forces also occur between instantaneous dipoles. Electrons are moving constantly. At any point in time, a chance exists that the electrons may be found on one side of an atom or molecule, and it can develop an 'instantaneous dipole' (frame 3 in figure 2.30). This temporary dipole can cause a shift in the distribution of electrons in neighbouring atoms or molecules, resulting in 'induced dipoles' (figure 2.31). Very small forces of attraction result between the particles.

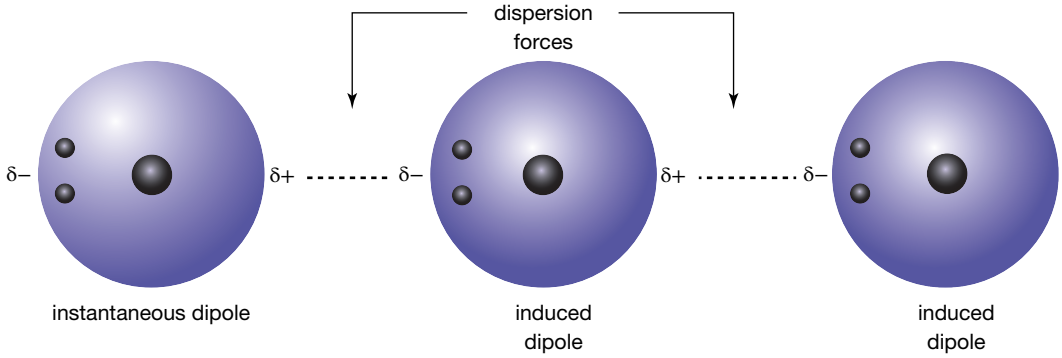
FIGURE 2.30 The electrons in this helium atom are moving all the time; an instantaneous dipole is created.



Two factors influence the strength of dispersion forces:

1. The number of electrons in the molecules. In general, the more electrons the molecules of a substance have, the stronger the dispersion forces between them.
2. The shapes of the molecules. Shape affects how closely the molecules may approach each other in the solid and liquid states. The closer the molecules can get, the stronger the attraction is. All covalent molecular substances have dispersion forces between their molecules.

FIGURE 2.31 The atom on the left has caused a temporary induced dipole in the middle atom, which in turn has caused another temporary induced dipole in the atom on the right.



Consider the noble gases He and Ne, or small molecules such as F_2, H_2 and CH_4. These atoms or molecules have relatively few electrons and so have very weak dispersion forces. Such substances have low boiling points and exist as gases at room temperature. Larger compounds, such as octane, C_8H_{18} (a component of petrol), exist as liquids. However, candle wax, $C_{25}H_{52}$, is a solid with a low melting point.

Dispersion forces and molecule size and position

Dispersion forces increase as the size of the molecule increases, and/or if the molecules can be closer together.

This trend can be illustrated using the halogens, as shown in table 2.7. As you go down the group, the molecules become larger and a progression occurs from a gaseous to liquid and ultimately solid state. Boiling and melting points of halogens show a corresponding increase.

TABLE 2.7 Properties of group 17 elements

Name	Formula	State	Melting point (°C)	Boiling point (°C)
Fluorine	F_2	Gas	–220.3	–188
Chlorine	Cl_2	Gas	–101	–34
Bromine	Br_2	Liquid	–7.3	59
Iodine	I_2	Solid	113	184

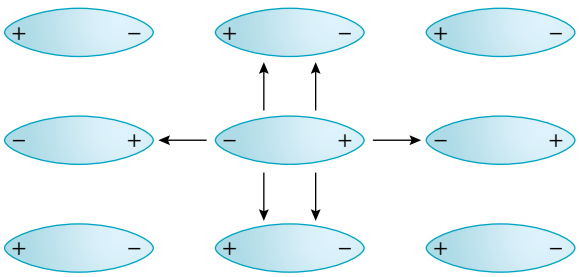
Dispersion forces

Dispersion forces operate between all atoms and molecules, since they all contain electrons that can be attracted to nuclei of other atoms or molecules.

Dipole–dipole interactions

If the molecules in a sample are polar, the presence of molecular dipoles causes simultaneous intermolecular attraction. The positive side of one molecule attracts the negative side of another molecule, which attracts the next, and so on to the limits of the sample.

FIGURE 2.32 In dipole–dipole interactions, the central polar molecule is attracted to other polar molecules around it. They, in turn, are attracted by their neighbours.



Polar molecules have two forces of attraction operating *between* their molecules. If two molecules are of similar size — that is, have a similar number of electrons — the dispersion forces acting on them are similar. However, if one of the molecules is polar, it will also be affected by dipole–dipole attractions. This results in a stronger overall intermolecular force and a higher boiling point.

Consider Ar (boiling point −186 °C) and HCl (boiling point −83.7 °C). Both substances have 18 electrons, yet HCl has a much higher boiling point. This is because HCl is a polar molecule and so has an extra intermolecular force of dipole–dipole interactions. A polar compound is generally more likely to exist as a liquid or solid at room temperature than a non-polar compound of similar size.

Dipole–dipole forces

Dipole–dipole forces operate between polar molecules.

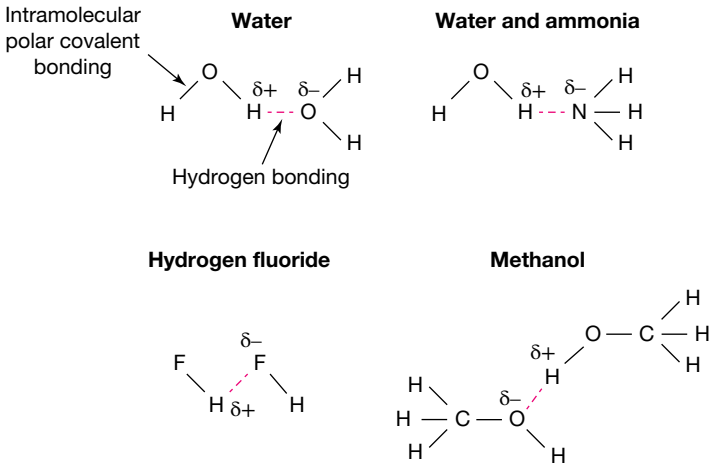
Hydrogen bonding

Hydrogen bonding is a special case of dipole–dipole attraction. Hydrogen is an atom containing one proton in the nucleus and one electron revolving around it.

When hydrogen bonds to a more electronegative atom such as nitrogen, oxygen or fluorine, its electrons move toward that atom. This causes the positively charged hydrogen nucleus to be exposed or unshielded. Hence, the molecule that forms has a strong dipole.

A lone pair at the negative end of one of these molecules is attracted to an unshielded hydrogen at the positive end of another, as shown in figure 2.33.

FIGURE 2.33 Examples of hydrogen bonding showing the lone pairs in one molecule are attracted to the unshielded hydrogens in another molecule.



Because the hydrogen (positive end) is unshielded, the other dipole can approach far more closely. The closer the dipoles get, the stronger the bond that forms. The bond between the dipoles is called a hydrogen bond. These bonds are represented in figure 2.33 by dotted lines, indicating that they are a weaker bond type than a covalent bond. Hydrogen bonds form only *between* molecules and only when hydrogen has been bonded to fluorine, oxygen or nitrogen. In fact, while they are called hydrogen *bonds,* they are really hydrogen *interactions* because they are relatively weak (but the strongest of the weak!) compared to other types of bonding.

Hydrogen bonds are much stronger than other dipole–dipole bonds and result in materials with higher melting and boiling points than would otherwise be expected.

Hydrogen bonding

Hydrogen bonds only operate between molecules when hydrogen is bonded to fluorine, oxygen or nitrogen.

tlvd-0533

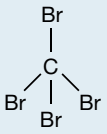
SAMPLE PROBLEM 8 Naming the types of intermolecular forces between molecules

Name the type(s) of intermolecular forces that exist between molecules in the following substances.

a. **CBr_4** **b.** **CH_3OH**

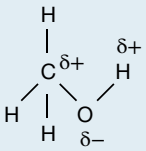
THINK

a. Draw the molecule.
Observe that the polar bonds are symmetrically placed around the carbon atom and so the molecule is non-polar even though the bonds are polar.



CBr_4 is a non-polar tetrahedral molecule.

b. Draw the molecule.



CH_3OH has a tetrahedral shape around the carbon atom and a V-shape around the second central atom, oxygen, resulting in a polar molecule. It also has a highly electronegative atom, oxygen, bonded to a hydrogen atom.

WRITE

Only dispersion forces exist between non-polar molecules.

Dispersion forces and hydrogen bonding exist between these molecules.

PRACTICE PROBLEM 8

Explain the type(s) of intermolecular forces that exist between molecules in the following substances.

a. **CI_4**

b. **CH_3NH_2**

on Resources

Weblink Intermolecular forces

2.4.2 Strengths of intramolecular bonding and intermolecular forces

The energy required to break bonds between a pair of atoms gives an indication of how strongly each atom is joined to the other atom, the intramolecular bond. For example, to separate one mole (you will find out about moles in topic 6) of hydrogen molecules into hydrogen atoms is 436 kJ.

$$H_2\,(g) \rightarrow 2H\,(g) \quad \text{bond energy} = 436\,\text{kJ mol}^{-1}$$

Bond dissociation energy

Bond dissociation energy measures how difficult it is to break the covalent bonds within a molecule in the gaseous state and separate it into its component atoms. The higher the value, the stronger the bonds.

Melting and boiling points indicate the strength of *intermolecular* forces. Table 2.8 lists some properties of the hydrogen halides (compounds between hydrogen and a halogen, group 17). The patterns and trends within this table illustrate some of the factors that determine bond strengths.

TABLE 2.8 Some physical properties of the hydrogen halides

Hydride	Formula	Melting temperature (°C)	Boiling temperature (°C)	Bond dissociation energy* (kJ mol^{-1})	Bond length (picometres) (pm)
Hydrogen fluoride	HF	+19.5	–83	568	92
Hydrogen chloride	HCl	–85	–115	432	127
Hydrogen bromide	HBr	–67	–87	366	141
Hydrogen iodide	HI	–35	–51	298	161

*Bond dissociation energy measures how difficult it is to break the bonds within a molecule in the gaseous state and separate it into its component atoms. The higher the value, the stronger the bonds.

It can be observed that, for these compounds, the strength of the covalent bonds decreases as the molecules become larger and, apart from hydrogen fluoride, the dispersion forces increase as the molecules increase in size. Hydrogen fluoride has an additional type of bonding between its molecules: hydrogen bonding.

Going down table 2.8 (which corresponds to going down group 17), the values for bond dissociation energy decrease, meaning the covalent bond between the hydrogen and the halide atom is becoming weaker. This is because the electrons involved from the halogen atom are one electron shell further out each time. This also means the length of the bond longer (as shown in the last column in table 2.8). The bond is weaker because the two electrons involved in it are more spread out and further away from the nucleus. This trend is obvious in nearly all other molecules — the longer the length of the bond, the weaker it is.

Multiple bonds are stronger than single bonds; that is, double carbon to carbon bonds are shorter and stronger than single carbon to carbon bonds.

In conclusion, considerably more energy is required to separate the atoms where covalent bonds are involved. For example, separating the molecules in one mole of liquid HCl when converting it into a gas requires only 17 kJ, compared to the 432 kJ needed to split one mole of HCl molecules into hydrogen and chlorine atoms.

Relative bond strengths summary

Molecular substances may experience up to four different types of interaction.

- All molecules are held together internally by covalent bonding. Because this acts within a molecule, it is classified as an *intramolecular* force. It is a very strong force and considerably greater than any of the other forces that may be experienced.
- All molecules in the liquid and solid states also experience attraction to other molecules from dispersion forces. These are the very weak forces that originate due to the movement of electrons within the molecule.
- Other forces that may be experienced between molecules are dipole–dipole attractions and hydrogen bonding. These two are due to the polarity that exists within some molecules.

All forces between molecules may be classified as *intermolecular* forces.

Relative strength of bonds

The decreasing order of strength between the four types of bonding is as follows, starting with strongest.

Covalent bonding >>>>> hydrogen bonding >> dipole–dipole attraction > dispersion forces

tlvd-0534

SAMPLE PROBLEM 9 Determining the bond strength within and between molecules

Consider the following hydrides from group 16: H_2O, H_2S, H_2Se.

a. Which will have the strongest bonding within its molecules?
b. Which will have the strongest dispersion forces between its molecules?
c. Which will have the overall strongest bonding between its molecules?
d. Which will have the highest boiling point?

THINK	WRITE
a. Recall that bond length gets longer down a group, and that the longer the bond, the weaker it is. H_2O has the shortest bonds and, therefore, the strongest intramolecular bonding.	H_2O has the strongest bonding within its molecules.
b. Recall that dispersion forces depend on the number of electrons in the molecule — the more electrons, the stronger the dispersion forces. H_2Se is the largest molecule and has the greatest number of electrons. It will, therefore, have the strongest dispersion forces between its molecules.	H_2Se has the strongest dispersion forces between its molecules.
c. Recognise that hydrogen bonding exists in H_2O but in none of the others. This is stronger than the dispersion forces that are (always) present. H_2O, therefore, has the strongest overall intermolecular bonding.	H_2O has the strongest overall bonding between its molecules.
d. Recall that the stronger the forces between molecules, the higher the boiling point. H_2O, therefore, has the highest boiling point.	H_2O has the highest boiling point.

PRACTICE PROBLEM 9

Consider the following hydrides from group 17: HF, HCl, HBr, HI.

a. Which will have the strongest bonding within its molecules?
b. Which will have the strongest dispersion forces between its molecules?
c. Which will have the overall strongest bonding between its molecules?
d. Which will have the highest boiling point?

2.4 Activities

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2.4 Exercise

1. MC What is the correct order of the bond distances in HBr, HI and HCl?
   - A. HCl > HBr > HI
   - B. HBr < HI < HCl
   - C. HCl < HBr < HI
   - D. HBr > HI > HCl
2. MC Consider the bonds C – N, C = N, and C ≡ N. The C ≡ N bond is _____.
   - A. strongest/longest
   - B. weakest/longest
   - C. weakest/shortest
   - D. strongest/shortest
3. MC Which of the following statements is *incorrect*?
   - A. The bond between carbon and oxygen in carbon dioxide is covalent.
   - B. The bond between two chlorine atoms in Cl_2 is intramolecular.
   - C. The bond between two chlorine molecules is intermolecular.
   - D. The bond between two carbon dioxide molecules is covalent.
4. MC Which of the following factors explains why a substance has a high melting point?
   - A. Strong intermolecular forces
   - B. Strong covalent bonds
   - C. Double covalent bonds
   - D. Polar bonds in the molecule
5. Consider the following list of compounds: CF_4, CCl_4, CBr_4.
   - a. Does hydrogen bonding exist in any of these substances?
   - b. Which has the strongest bonding within its molecules?
   - c. Which has the strongest dispersion forces between its molecules?
   - d. Which has the highest boiling point?
6. How would you expect the boiling points of argon, Ar, and fluorine, F_2, to compare? Explain.
7. CH_3F has a boiling point of –78 °C, whereas CH_3OH has a boiling point of 65 °C. Explain this difference.

2.4 Exam questions

Question 1 (1 mark)

MC Organic compounds can contain the following carbon to carbon bonds: C–C, C=C and C ≡ C.

The correct order of bond dissociation energy and bond length is:

- A. C ≡ C > C=C > C–C; and C ≡ C > C=C > C–C.
- B. C ≡ C < C=C < C–C; and C ≡ C < C=C < C–C.
- C. C ≡ C > C=C > C–C; and C ≡ C < C=C < C–C.
- D. C ≡ C < C=C < C–C; and C ≡ C > C=C > C–C.

Question 2 (1 mark)

MC The table provided shows the boiling temperatures of some molecular substances.

Substance	Carbon dioxide	Carbon monoxide	Methane	Carbon disulfide
Molecular formula	CO_2	CO	CO_4	CS_2
Boiling temperature (°C)	–57.0	–191.5	–161.5	46.3

Which of the following lists the strength of the intermolecular bonding from weakest to strongest?

A. CO_2, CO, CH_4, CS_2
B. CO, CH_4, CO_2, CS_2
C. CO, CH_4, CS_2, CO_2
D. CS_2, CO_2, CH_4, CO

Question 3 (1 mark)

Why is it that water boils at 100 °C but needs to be heated to over 2000 °C before it decomposes?

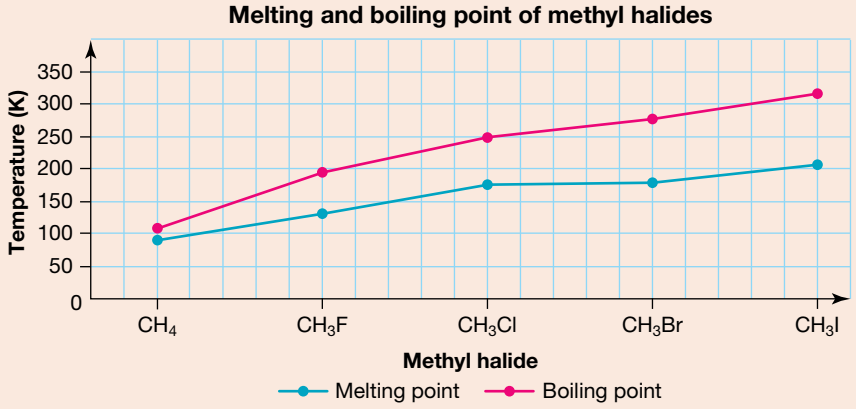
Question 4 (1 mark)

Nitrogen, N_2, has a triple covalent bond between the nitrogen atoms yet it has a lower boiling temperature (–195.8 °C) than that of fluorine, F_2, which has only a single covalent bond and a boiling temperature of –188.1 °C.

Explain this observation.

Question 5 (10 marks)

Consider the following list of compounds: CH_4, CH_3F, CH_3Cl, CH_3Br and CH_3I.



a. Draw the shapes of the first four of these molecules. **(4 marks)**
b. State the intramolecular bonding occurring. **(1 mark)**
c. State the intermolecular bonding occurring. **(1 mark)**
d. What is the physical state of these compounds at room temperature (298 K)? **(1 mark)**
e. Explain the trend in melting and boiling points of these compounds. **(3 marks)**

More exam questions are available in your learnON title.

2.5 Physical properties of molecular substances

KEY KNOWLEDGE

- Physical properties of molecular substances (including melting points and boiling points and non-conduction of electricity) with reference to their structure

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2.5.1 Explaining physical properties

Although the covalent bonding holding the atoms together in a molecule is strong, the forces between molecules are usually weak. The physical properties of a compound depend on the type of bonding it displays. Molecular substances may exist as solids, liquids or gases. A great variety of physical properties occurs among molecular substances due largely to the differences in strength of the intermolecular attractions. Even so, a few generalisations can be made.

- Molecular substances do not conduct electricity in the solid or molten form because the molecules are electrically neutral, so no ions or delocalised electrons are present. Carbon, in the form of graphite, is an exception because it does have delocalised electrons and so conducts.
- If dissolved in water without reacting with it, molecular substances do not conduct electricity, again because the molecules are electrically neutral. For example, when sugar is dissolved in water.
- Some molecules dissolve in water to produce ions. These molecules are said to ionise when they are dissolved in water and can conduct electricity owing to the movement of the ions produced. For example, hydrogen chloride gas dissolved in water.
- Molecular substances vary in their solubility in water and other solvents. Generally, polar compounds are soluble in polar solvents such as water but insoluble in non-polar solvents such as tetrachloromethane. Non-polar solutes tend to be more soluble in non-polar solvents than polar solvents. This is sometimes described as the 'like-dissolves-like' rule.
- Molecular compounds have low melting and boiling points since the forces between the molecules are weak, and relatively little energy is required to break them. Many molecular substances are gases or liquids at room temperature.

FIGURE 2.34 The softness of wax together with its insolubility in water make it ideal to apply to surfboards.

- Most molecular substances are soft and easily scratched. Once again this is due to the weak forces of attraction between the molecules that mean molecules can be easily removed from the surface. An example of an easily scratched covalent substance is wax.

Note that in all of these cases in the preceding list (except for the third point), the molecules themselves remain unchanged. Although they may be separated from other molecules by processes such as boiling, dissolving and scratching, the molecules themselves remain intact. This is because the *intra*molecular force of covalent bonding is very strong.

Note also that these are properties of covalent molecular compounds. The properties of covalent lattice compounds, graphite and diamond, are quite different because of their lattice structures, these are examined in subtopic 2.6.

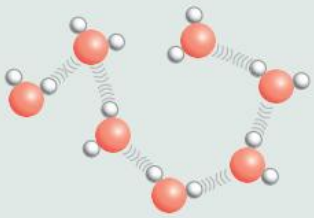
CASE STUDY: The effect of hydrogen bonding on the properties of water

Hydrogen bonding significantly affects the physical properties of water. The existence of hydrogen bonding between water molecules results in:

- the relatively high melting and boiling points of water compared with other substances. More heat is required to enable the molecules to gain sufficient kinetic energy to break free of the hydrogen bonds, which are stronger than dispersion forces alone. For example, H_2O is a liquid at room temperature with a boiling point of 100 °C, whereas H_2S, which has more electrons, is a gas at room temperature and boils at –61 °C. The behaviour of H_2O as it changes state between liquid water and solid ice is also governed by the hydrogen bond.
- the expansion of water upon freezing. Generally, as substances are heated, they expand. However, water behaves a little differently. The density of water is greatest at 4 °C. (Density refers to how much mass is in a given volume; for example, oil is less dense than water so it floats on top of water.) Water does expand when heated from 4 °C but, unusually, it also expands as the temperature decreases from 4 °C. As the temperature continues to decrease, the denser water at 4 °C sinks and eventually the temperature at the surface becomes 0 °C and freezes. Ice forms in an open, hexagonal crystalline lattice that places the water molecules further apart than occurs in the liquid state. Because water expands on freezing, it is less dense as a solid. Ice, therefore, floats on water.

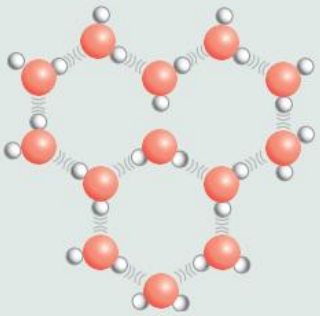
FIGURE 2.35 Hydrogen bonding in water when (a) water molecules are in liquid form and (b) water molecules in solid form (ice). In ice, the extensive hydrogen bonding holds the molecules in a more orderly arrangement.

(a)



Water molecules move too quickly when in liquid form to form regular structures; however, they still remain attached.

(b)



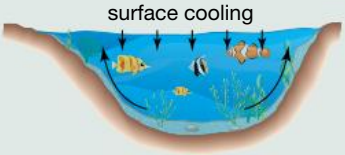
When in solid form (ice), water molecules join together to form hexagonal structures.

If ice were denser than water, some bodies of water would freeze solid during winter and the aquatic life would die. When ice forms on the surface of water, it acts as an insulator, preventing the water below from freezing. This means that aquatic life can survive, even in sub-zero conditions. Icebergs form when sections of the frozen icecap break off and float in the sea.

FIGURE 2.36 The density of water is greatest at 4 °C.

cold air

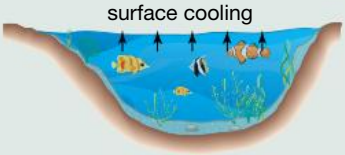
surface cooling



At 4 °C, water sinks and is replaced by water over 4 °C.

cold air

surface cooling



Water below 4 °C rises to the surface.

cold air

ice at surface insulates

Water at the surface freezes, but deeper water is 0–4 °C so aquatic life survives.

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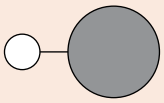
2.5 Quick quiz on | 2.5 Exercise | 2.5 Exam questions

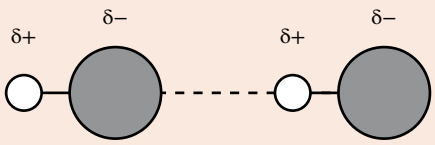
2.5 Exercise

1. MC Which of the following statements about dispersion forces is *incorrect*?
   - A. Dispersion forces increase as the size of the molecule increases.
   - B. Only non-polar molecules exhibit dispersion forces between molecules.
   - C. Dispersion forces are the weakest form of intermolecular bonding.
   - D. Non-polar molecules exhibit dispersion forces only between molecules.
2. Which of the following is a correctly labelled diagram of the interaction between two HCl molecules? (*Note:* Only the strongest interaction is marked.)

Key: H —— Cl

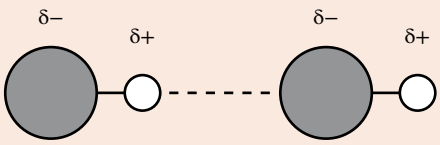


A.



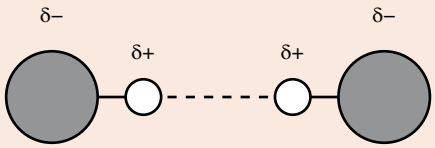
dipole–dipole interaction

B.



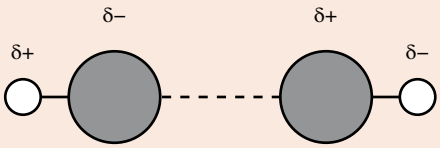
hydrogen bonding

C.



hydrogen bonding

D.



dipole–dipole interaction

3. MC Which of the following contains only compounds that display hydrogen bonding?
   - A. Hydrogen fluoride, HF, and hydrogen, H_2
   - B. Hydrochloric acid, HCl, and hydrogen fluoride, HF
   - C. Water, H_2O, and carbon dioxide, CO_2
   - D. Ammonia, NH_3, and ethanoic acid, CH_3COOH
4. MC Which of the following molecules will bond to like molecules by dispersion forces only?
   - A. HF, HCl, HBr
   - B. F_2, CO_2, CH_4
   - C. NH_3, HF, CH_3OH
   - D. C_2H_4, C_2H_5OH, CH_3COOH
5. Explain the type(s) of intermolecular forces that exist between molecules in the following substances.
   - a. CCl_4
   - b. $CH_3CH_2NH_2$
   - c. Cl_2
   - d. CH_3OH
   - e. CH_3Cl
6. Explain why water, H_2O, has a higher boiling point than hydrogen sulphide, H_2S, despite it being a smaller molecule with fewer electrons.

7. The boiling points of HCl, HBr and HI are, respectively, –85 °C, –67 °C and –35 °C. Explain the difference in these boiling points with reference to the forces that exist between the molecules of each substance.
8. Br_2 and ICl have the same number of electrons, yet Br_2 melts at –7.2 °C, and ICl melts at 27.2 °C. Explain this in terms of intermolecular forces.
9. HCl has more electrons than HF so we would expect it to have the higher boiling point. However, this is not the case; the boiling points of HCl and HF are –83.7 °C and 19.4 °C respectively. Explain what factors could account for this reversal in trend.
10. Explain why F_2, O_2 and N_2 are all gases at room temperature.

2.5 Exam questions

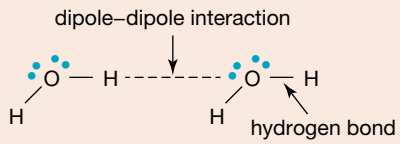
Question 1 (1 mark)

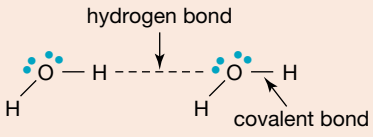
MC Which of the following molecules will bond to like molecules by dispersion forces only?

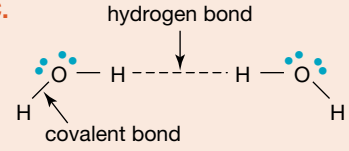
A. Hydrogen bromide, HBr
B. Nitrogen, N_2
C. Ammonia, NH_3
D. Ethanol, C_2H_5OH

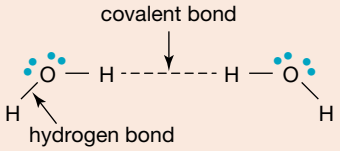
Question 2 (1 mark)

MC Which diagram showing intermolecular forces in water is correct?

A.


B.


C.


D.


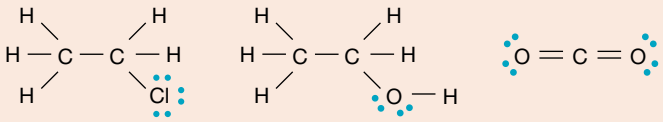
Question 3 (1 mark)

MC Which bonds are broken when a sample of liquid butane, C_4H_{10}, is allowed to boil?

A. Covalent bonds between all atoms are broken.
B. Covalent bonds between carbon atoms only are broken.
C. Covalent bonds between carbon and hydrogen atoms only are broken.
D. Dispersion forces between different molecules are broken.

Question 4 (9 marks)

Consider the following molecules.



a. Write the molecular formula for each molecule. **(1 mark)**
b. Classify each molecule as either polar or non-polar. **(1 mark)**
c. Which is the most polar bond among the three molecules? **(1 mark)**
d. Name the type of intramolecular bonding shared by all three molecules. **(1 mark)**
e. Name all types of intermolecular bonding that would operate between molecules of
   i. chloroethane. **(1 mark)**
   ii. ethanol. **(1 mark)**
   iii. carbon dioxide. **(1 mark)**
f. Which of the three compounds would be expected to have the highest boiling point? Explain your reasoning. **(2 marks)**

Question 5 (10 marks)

a. Complete the provided table for the listed following molecular substances. **(7 marks)**

Compound	Ammonia, NH_3	Ethyne, C_2H_2	Fluoromethane, CH_3F
Valence structure (including nonbonding electrons)			
Shape			
Is the molecule polar or non-polar?			
Type/s of intramolecular bonding			
Type/s of intermolecular bonding			

b. Which of the compounds listed in the table would be expected to have the highest boiling point? Explain your choice. **(1 mark)**

c. Which of the compounds listed in the table would be expected to have the weakest dispersion forces? Explain your choice. **(1 mark)**

d. Which of the compounds listed in the table would be expected to conduct electricity? **(1 mark)**

More exam questions are available in your learnON title.

2.6 Structure and bonding of diamond and graphite

KEY KNOWLEDGE

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

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2.6.1 Carbon

What do you think is the most important element? Carbon should probably be at the top of the list. It has been known from ancient times and yet there are still many exciting discoveries of new materials based on carbon. Carbon, with its four bonding electrons, is present in all forms of life and is known to form millions of compounds. We have seen that this element forms simple, small molecular compounds such as methane, CH_4, and carbon dioxide, CO_2. We breathe out carbon dioxide, and it is an essential reactant in photosynthesis, the process by which plants make food. Carbon dioxide produced as a result of human activity is increasing and so contributing to climate change.

FIGURE 2.37 While diamonds are composed of the same atoms as graphite (carbon), their arrangement causes their vastly different physical properties.

Topic 7 focuses on **organic chemistry**, which is a branch of chemistry that is devoted entirely to compounds of carbon from the small to very large molecules. Topic 8 introduces the incredible versatile materials that are generally called plastics, but scientists call polymers, which are large molecules with carbon backbones. Organic compounds are obtained from fossil fuels that have been formed from the decayed remains of plants and formed over millions of years. Fossil fuels also provide us with energy for households, transportation and manufacturing. Carbon in its different forms is a major part of our lives.

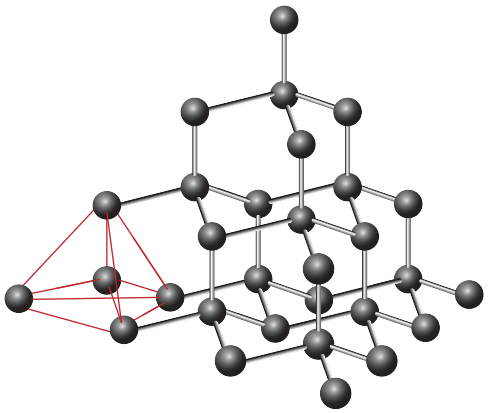
organic chemistry the study of carbon-containing compounds and their properties

How can both the hardest material, diamond, and one of the softest materials, graphite, consist of the same element? How can one material be an insulator and another a conductor? It is all about the arrangement of the atoms. Diamond and graphite are naturally occurring **allotropes** of carbon; 'allotrope' is the term used to describe different forms of the same element. In this subtopic, carbon's ability to form both network lattices and layer lattices is discussed.

2.6.2 Covalent network lattices

Some non-metals form giant structures in which no individual molecules exist. They consist of countless numbers of atoms covalently bonded to each other, forming a *three-dimensional* **network lattice**. Common examples of covalent network lattices include diamond, silicon carbide, silicon dioxide (quartz) and tungsten carbide.

FIGURE 2.38 Diamonds are prized for their brilliance, durability and hardness. These properties are due to their rigid internal structure.



Diamond — the hardest known substance

Diamond is the hardest substance known because of the bonding pattern of its carbon atoms. Their hardness and brilliance are the consequence of the ordered and rigid internal structure in which each carbon atom bonds with four single covalent bonds to four other carbon atoms. Millions of carbon atoms can combine in a three-dimensional covalent network lattice to form a diamond crystal (figure 2.38).

Hardness is measured on the **Mohs scale** (see table 2.9). The Mohs scale gives the order of hardness, but this is not linear; the difference between corundum and diamond is only one on this scale but diamond is many times harder than corundum.

A diamond is really one 'giant' molecule. It is made up of carbon atoms bonded together by strong covalent bonds. Each carbon atom has four covalent bonds around it and is surrounded by four other carbon atoms. The covalent bonding is three-dimensional since the bonds are arranged tetrahedrally around each atom. To scratch a diamond, the carbon atoms need to be separated. This is very difficult to do because they are held together by four covalent bonds.

allotrope different forms of an element

network lattice a substance that displays covalent bonding in three dimensions without the formation of discrete molecules

Mohs scale a scale used to measure the hardness of substances

TABLE 2.9 Mohs hardness scale

Hardness	Mineral	Scratch test	Other materials
1	Talc	Very easily scratched by fingernail	Sulfur, sodium, chalk
2	Gypsum	Scratched by fingernail	Lead, **graphite**
3	Calcite	Very easily scratched with knife	Gold, silver, copper
4	Fluorite	Easily scratched with knife	Platinum, nickel
5	Apatite	Scratched with knife	Tooth enamel, glass (5.5)
6	Orthoclase	Scratches glass	Manganese, steel
7	Quartz	Scratches glass easily	Vanadium
8	Topaz	Scratches glass very easily	Cubic zirconia
9	Corundum	Cuts glass	Tungsten carbide, ruby, sapphire
10	**Diamond**	Scratches all other materials	

Uses of diamonds

Diamonds are used in industrial situations where a very hard substance is required for a particular purpose. They are typically used as an abrasive for sawing, cutting and grinding hard substances such as glass, stone and porcelain. The tips of drilling bits used on oil rigs and other similar rigs that need to bore through hard rock are coated with diamonds. All of these uses are explained by the strong covalent bonding that exists throughout the crystals and which makes diamond exceedingly hard and suitable for these purposes.

In jewellery, cut diamonds are prized for their beauty. Diamonds have a very high **refractive index** and are cut at angles and polished so that the light entering them is refracted to make them sparkle. As can be imagined, due to diamond's extreme hardness this is no easy task and specialised techniques and equipment are required.

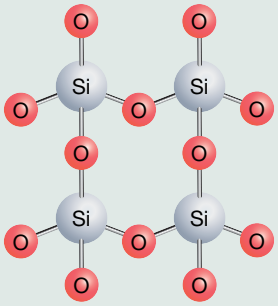
EXTENSION: Silicon dioxide

Another covalent network lattice is silicon dioxide, SiO_2. Silicon dioxide is found as quartz in rocks such as granites and sandstones, and very pure varieties weather to white sand beaches, as shown in figure 2.39(a). Quartz, SiO_2, is a three-dimensional covalent network lattice shown in figure 2.39(b). Each silicon atom is bonded to four oxygen atoms, and each oxygen atom is bonded to two silicon atoms. The resulting substance is a hard, crystalline rock.

FIGURE 2.39 (a) Whitehaven beach in Queensland is renowned for its ultra-white sand, which consists of 98 per cent pure silicon dioxide (quartz). (b) The structure of silicon dioxide

(a)

(b)



Properties of covalent network lattices

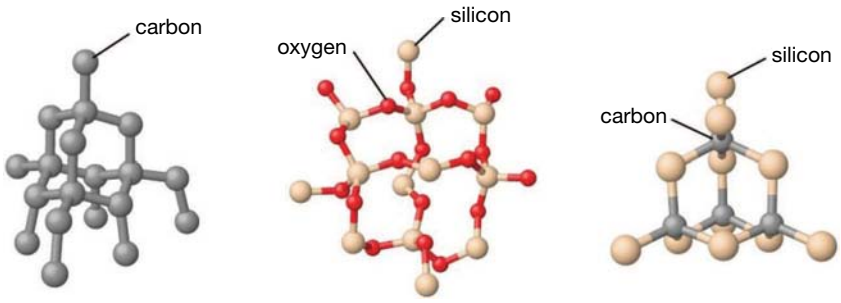
Covalent network lattices have the following general properties:

- Since strong covalent bonding extends through the crystal structures of covalent network lattices, the particles are held rigidly. These substances are, therefore, very hard, difficult to scratch and have high melting points and boiling points.
- Since no free ions or electrons are in the structure, covalent network lattices are usually non-conductors of electricity in the solid and liquid states.
- Covalent network lattices are brittle and must be cut in a specific way or they can shatter. When the covalent bonds break, the lattice is distorted.
- Covalent network lattices are chemically inert and are insoluble in water and most other solvents.

refractive index a measure of how much a substance refracts (bends) light when it enters it

- Covalent network lattices such as diamond are good conductors of heat. The conduction of heat is a mechanical process due to the collision of particles. A collision at one end of a diamond will be effectively transmitted through the crystal structure because the atoms are so rigidly bonded together — there is little 'give' in the bonds to soak up the energy of the collision.

FIGURE 2.40 Examples of covalent network structures



diamond | silicon dioxide | silicon carbide

tlvd-0535

SAMPLE PROBLEM 10 Determining the type of bonding in covalent network lattices

Pure germanium (Ge) forms crystals in the same way as diamond.

a. Describe the bonding that occurs in a crystal of germanium with respect to type and orientation.

b. Describe one difference between the bonding in germanium compared to diamond.

THINK	WRITE
a. Recall that in diamond each carbon has four covalent bonds that are bonded to four other C atoms, and so on. These bonds are arranged tetrahedrally. No discrete molecules are present.	Germanium atoms will be covalently bonded to four other germanium atoms, which in turn will be bonded to four more each, and so on. The arrangement of bonds around each germanium atom will be tetrahedral.
b. Germanium is in period 4 compared to carbon in period 2, so its bonding electrons are further from the nucleus than carbon, which makes the bond length longer.	The bonds are longer in germanium than in diamond.

PRACTICE PROBLEM 10

Pure silicon (Si) forms crystals in the same way as diamond.

a. Describe the bonding that occurs in a crystal of silicon with respect to type and orientation.

b. Describe one difference between the bonding in silicon compared to diamond.

2.6.3 Covalent layer lattices

Covalent layer lattices consist of countless atoms held strongly together in planes by covalent bonds. Covalent bonding is, therefore, present in *two dimensions only*. These layers are held together by weaker dispersion forces. Graphite is a common example of a covalent layer lattice.

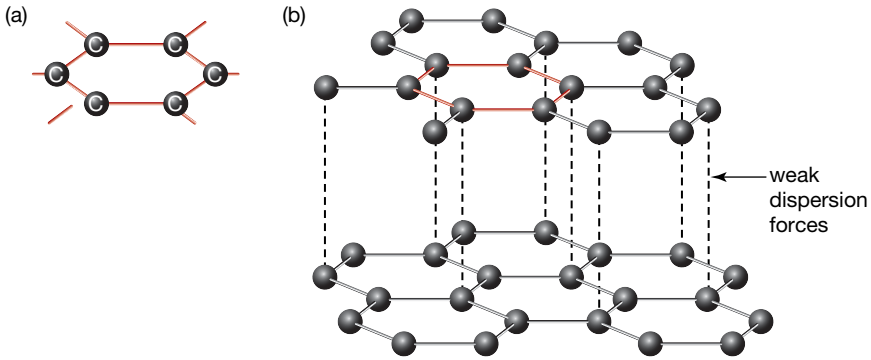
covalent layer lattice a substance that displays covalent bonding in two dimensions without the formation of discrete molecules

Graphite

Graphite is an oily, black, opaque solid with a metallic sheen. Like diamond, graphite is made solely of carbon atoms.

Graphite, however, looks and behaves quite differently from diamond, and this can be explained by its structure. Each carbon atom in graphite forms three covalent bonds with three other carbon atoms within the same plane. This forms layers of hexagonal rings (six atoms) with strong covalent bonds forming flat sheets in two dimensions. Since only three of the four outershell electrons are used in the covalent bonds, one outershell electron remains. This electron from each atom becomes **delocalised** and can move *across* the graphite layers. Layers are stacked on top of each other in the crystal lattice and are held together by weak dispersion forces.

FIGURE 2.41 Graphite is made of flat sheets of carbon atoms. Each carbon atom covalently bonds to three other carbon atoms, producing rings of six atoms that join to form flat sheets, as shown in (a). These sheets of carbon atoms lie on top of each other, held together by dispersion forces, as shown in (b).



Properties and uses of graphite

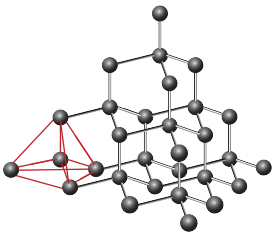
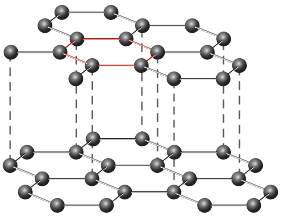
- Graphite is a good conductor of electricity because it contains charged particles that are free to move. These are the delocalised electrons left over from the bonding into the hexagonal rings that form its sheets.
- Graphite is a good conductor of heat also due to its delocalised electrons. The energy of collisions in one location can easily be transmitted through further collisions between these electrons and others to a different location.
- Graphite is soft and feels slippery. Although the carbon atoms in the layers in graphite are held together by strong covalent bonds, only weak dispersion forces exist between the layers. The layers, therefore, slide over each other very easily. This makes it feel slippery and, in powdered form, this feeling is further enhanced. The most familiar use of graphite is as 'lead' in pencils. Here it is mixed with varying proportions of clay to make the different grades of pencil. The higher the proportion of graphite, the softer and darker the pencil will be. The weakly bonded layers slide easily over each other and are left behind as marks on the paper. This property also explains its use as a dry lubricant for machine parts and in locks.
- Graphite is inert (it can withstand chemical change) due to its stable hexagonal rings and delocalised electrons. It is used extensively as an inert electrode material in batteries and electrolysis applications due to its electrical conductivity and chemical inertness. It is also used as a moderator in nuclear reactors, where it absorbs fast-moving neutrons. It is also used in the making of graphene and synthetic diamonds.

delocalised describes electrons that are not bound to any one atom but are free to move throughout a lattice

2.6.4 Comparison of allotropes of carbon

Table 2.10 compares different allotropes of carbon.

TABLE 2.10 Various allotropes of carbon

Name	Structure	Property	Use
Diamond	• Covalent network lattice: each carbon atom bonds with four single covalent bonds to four other carbon atoms, creating tetrahedral structure. 	• Hardest known substance due to rigid crystal structure • High refractive index • High melting and boiling points • Non-conductors of electricity due to lack of free ions or electrons in the lattice • Brittle • Chemically inert • Insoluble in water • Good conductor of heat due to collision of particles through crystal structure • Formed at very high pressures	• Sawing, cutting, drilling • Jewellery
Graphite	• Covalent layer lattice: layers of hexagonal carbon rings with strong covalent bonds forming two-dimensional flat sheets, with sheets held together by weak dispersion forces. • Remaining outershell electron is delocalised and can move between graphite layers. 	• Soft • Low refractive index (opaque) • High melting and boiling points • Good conductor of electricity due to free moving delocalised electrons • Slippery feel because layers can slide over each other easily • Chemically inert • Good conductor of heat due to free moving delocalised electrons	• Dry lubricant • Pencils • Inert electrodes • Moderators in nuclear reactors
Amorphous carbon	• No consistent structure: impure forms of carbon including coal and soot are produced from incomplete combustion of organic material.	• Reactive • Conductive • No permanent crystal structure	• Charcoal fuel • Printer (toner) ink • Activated charcoal for water filtration, and treatment of poisoning and drug overdoses

on Resources

Weblinks Diamond and silicon dioxide
Graphite

2.6 Activities

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2.6 Quick quiz on | 2.6 Exercise | 2.6 Exam questions

2.6 Exercise

1. Why is carbon able to form such a variety of compounds?
2. The density of diamond is 3.5 g cm^{-3} and that of graphite is 1.9–2.3 g cm^{-3}. Explain why they are different.
3. Explain why the following substances are used in the way described.
   a. Graphite in pencils
   b. Diamond in oil drilling
4. Graphite is used to make electrodes in batteries. Explain why graphite is a good conductor of electricity and diamond is not.
5. Gemstones that are similar in appearance and structure to diamonds can be produced from a form of silicon carbide (SiC). These are called *synthetic moissanites*.
   a. Describe the bonding that occurs in a crystal of synthetic moissanite with respect to type and orientation.
   b. Describe one difference between the bonding in synthetic moissanite compared to diamond.

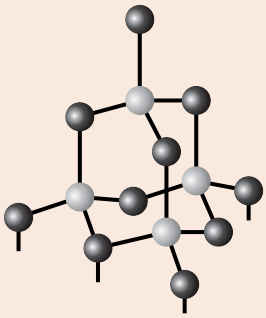
2.6 Exam questions

Question 1 (1 mark)

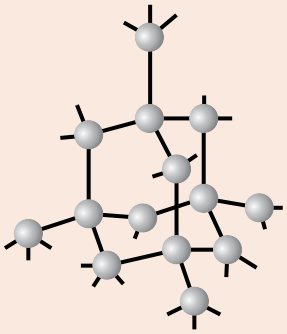
MC Silicon carbide, like diamond, forms a covalent network lattice of silicon and carbon atoms.

Which of the following diagrams represents the covalent network lattice of silicon carbide, SiC?

A.



B.



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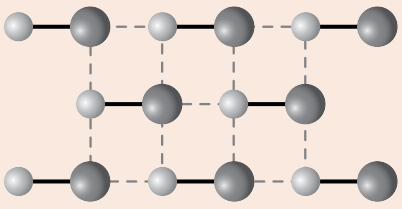
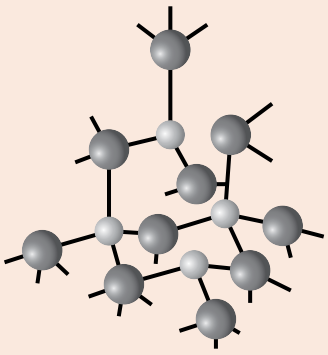


D.



Question 2 (1 mark)

Diamond does not melt but instead it sublimes. Write an equation, including symbols of state, to represent this process.

Question 3 (2 marks)

Why is graphite both shiny and able to conduct electricity?

Question 4 (1 mark)

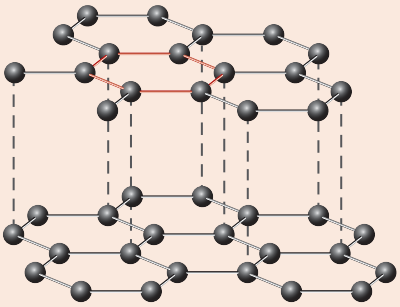
MC Boron nitride, BN, forms covalent lattices of different types. Cubic boron nitride (c-BN) forms a lattice like diamond, while hexagonal boron nitride (h-BN) forms a lattice like graphite. Neither lattice, however, conducts electricity.

Which of the following statements is *incorrect*?

A. c-BN would be expected to be a very hard substance.
B. h-BN would be expected to contain delocalised electrons.
C. c-BN would be expected to have a very high melting point.
D. h-BN would be expected to have a flaky nature.

Question 5 (4 marks)

Graphite (shown) and graphene consist of continuous rings of carbon atoms. Why are their properties so different?



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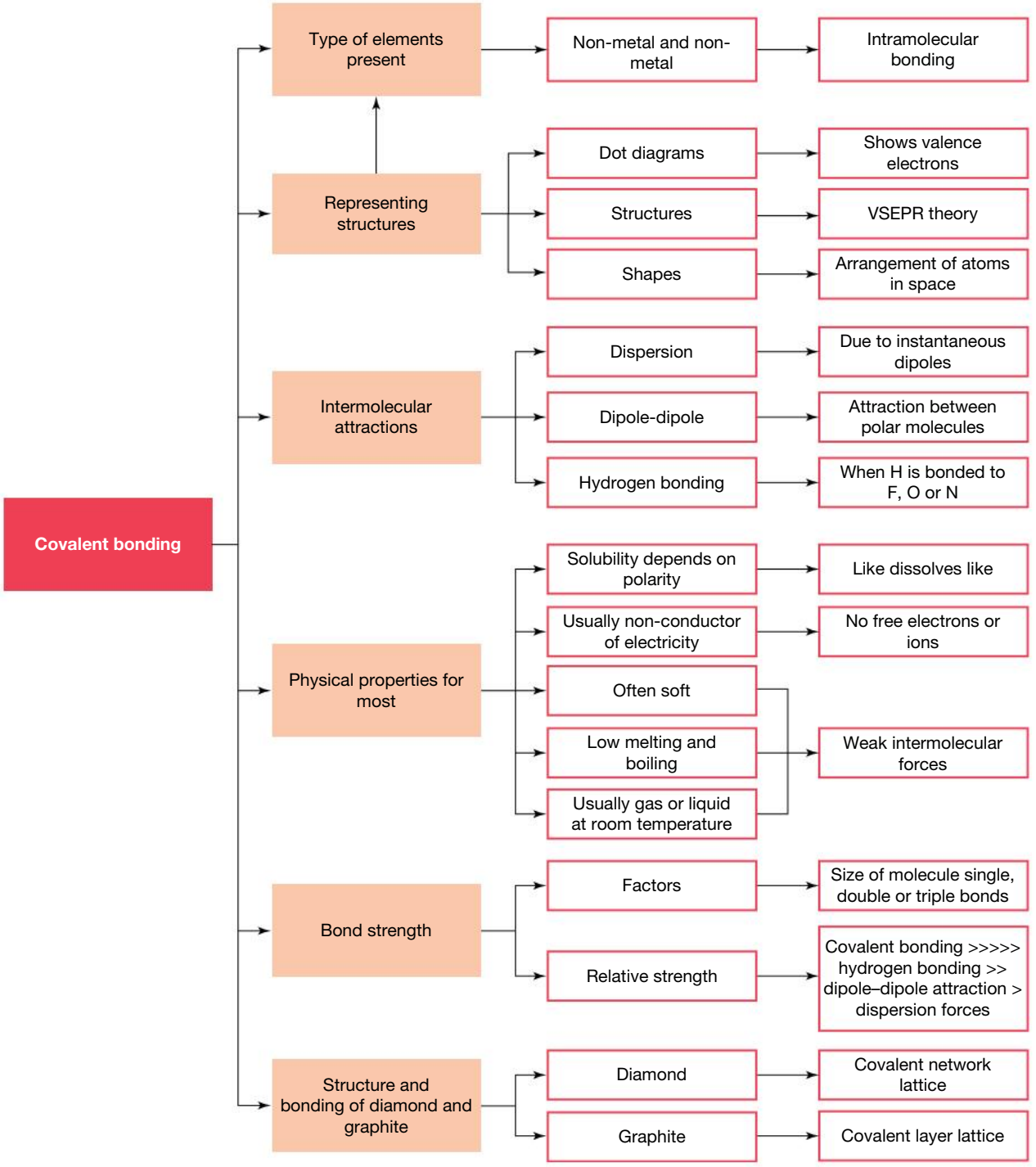

2.7.2 Key ideas summary

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2.7.3 Key terms glossary

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on Resources

Solutions	Solutions — Topic 2 (sol-0802)
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Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 2 (doc-37071) Key ideas summary — Topic 2 (doc-37072)
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2.7 Activities

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2.7 Review questions

1. What is the best model to illustrate the shape of a molecule?
2. What does the sharing of electrons in bond formation *always* involve?
3. From this list of compounds, which is *not* a covalent compound?

$$CaO,\ SO_3,\ NH_3,\ CH_3Cl$$

4. List three typical properties of covalent molecular substances.
5. Give the group number of the element that each of the following electron dot diagrams represents.
   a. ·M·
   b. :N:
   c. :X·
   d. ·A·
6. Use electron dot diagrams to draw the structural formulas of the following.
   a. H_2S
   b. NCl_3
   c. CH_4
   d. Br_2
   e. HOCl
   f. H_2CO

7. Write the structural formulas for the following and name the shape of each molecule.
   a. NH_3
   b. CH_3OH
   c. HCN
   d. PCl_3
   e. N_2
8. State whether the following molecules are polar or non-polar: NH_3, CH_3OH, HCN, PCl_3, N_2, H_2S, NF_3, C_2Br_4.
9. Explain why the ammonia molecule is polar but carbon dioxide molecules are not.
10. Which of the following exhibits hydrogen bonding? Justify your response.
    a. CF_3H
    b. CH_3OCH_3
    c. NH_3
    d. HF

2.7 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC A molecule has three bonding pairs of electrons and a lone pair around its central atom. What would the shape of this molecule be described as?

A. Tetrahedral
B. Pyramidal
C. Bent
D. Linear

Question 2

MC Which one of the following electron dot diagrams best represents the bonding in the nitrogen molecule?

A. :N⋮⋮N:
B. :N⋮⋮N:
C. :N⋮N:
D. :N:

Question 3

MC Which of the following is a characteristic property of a covalent molecular compound?

A. Relatively low melting point
B. Malleable and ductile
C. High melting point
D. Conducts electricity when molten but not when solid

Question 4

MC What is the shape of C_2H_2?

A. Linear
B. Bent
C. Planar
D. Pyramidal

Question 5

MC Which of the following best describes the molecular shape, polarity of bond and molecular polarity of the H_2S molecule?

A. Linear, polar covalent, non-polar
B. Linear, polar covalent, polar
C. Bent, ionic, polar
D. Bent, polar covalent, polar

Question 6

MC Which of the following substances might you expect to exhibit the weakest intermolecular forces?

A. HCl
B. He
C. NH_3
D. H_2O

Question 7

MC At room temperature, methane is a gas while tetrachloromethane is a liquid. Choose the statement that best explains this.

A. There is appreciable hydrogen bonding in tetrachloromethane but not in methane.
B. The tetrachloromethane molecule is polar, while the methane molecule is non-polar.
C. Tetrachloromethane has an appreciably higher molecular mass than methane.
D. The bonds in tetrachloromethane are polar covalent.

Question 8

MC What is the correct order of decreasing attractive strength, from highest to lowest, for weak intermolecular forces?

A. Hydrogen bonding, dipole–dipole interaction, dispersion forces
B. Hydrogen bonding, dispersion forces, dipole–dipole interaction
C. Dipole–dipole interaction, hydrogen bonding, dispersion forces
D. Dispersion forces, dipole–dipole interaction, hydrogen bonding.

Question 9

MC Which one of the following substances forms hydrogen bonds?

A. HCOOH
B. CH_3CN
C. SiH_4
D. CH_3OCl.

Question 10

MC Which of the following *best* explains the relatively low melting point of covalent molecular substances?

A. Covalent molecular materials depend on weak electrostatic forces holding the ions together.
B. The intermolecular forces between the molecules are weaker than covalent bonds.
C. The metals involved create uneven bonding with the non-metals.
D. The similar electronegativity of the atoms causes repulsions between the molecules.

Section B — Short answer questions

Question 11 (4 marks)

a. The boiling point of iodine is 184 °C, whereas fluorine has a boiling point of –188 °C. Explain this difference. **(2 marks)**
b. What states do fluorine and iodine assume at room temperature? Explain. **(2 marks)**

Question 12 (6 marks)

The amount of CO_2 in the atmosphere is increasing.

a. Use electron dot diagrams to determine the structure and shape of a CO_2 molecule. **(1 mark)**
b. What is the shape of the CO_2 molecule? **(1 mark)**
c. How many lone pairs are in a molecule of carbon dioxide? **(1 mark)**
d. Are the bonds in the molecule polar or non-polar? **(1 mark)**
e. Is the molecule polar or non-polar? **(1 mark)**
f. What type of interactions would you expect between the molecules in solid carbon dioxide? **(1 mark)**

Question 13 (5 marks)

Ammonia, NH_3, is a gas at room temperature. It has a characteristic pungent odour.

a. Show how the shape of the molecule is determined. **(2 marks)**
b. Discuss the polarity of the molecule. **(1 mark)**
c. Explain in terms of the intermolecular bonding why ammonia is a gas at room temperature. **(2 marks)**

Question 14 (9 marks)

The electron dot diagrams for six elements are shown with their corresponding electronegativities in brackets. The usual symbols have been replaced by the letters A to F.

A (2.1) B (3.0) C (3.5)

D (3.0) E (2.8) F (2.5)

a. Describe the position of D and E relative to each other on the periodic table. **(1 mark)**

b. Name a possible element that C might represent. **(1 mark)**

c. Element A forms covalent bonds with B, C, D, E and F. Which bonds will be the most polar? **(1 mark)**

d. Draw electron dot diagrams for the following. **(3 marks)**

- **i.** A bonding with B
- **ii.** A bonding with D
- **iii.** A bonding with F

e. Which of the molecule(s) from part (d) will be the following? **(2 marks)**

- **i.** Polar
- **ii.** Non-polar

f. What is the shape of the molecule formed between A and B? **(1 mark)**

Question 15 (7 marks)

When diamond is heated to high temperature, it burns to form carbon dioxide. If oxygen is excluded, however, it changes to graphite, which then sublimes at about 3500 °C. Three other substances, silicon, a form of silicon carbide, SiC, called moissanite and germanium are also able to resist very high temperatures.

a. Describe the type of bonding that you would expect in each of the substances mentioned. **(1 mark)**

b. Would you expect germanium to be hard or soft? Explain. **(2 marks)**

c. Samples of these four substances are labelled A, B, C and D. In an experiment, they are subjected to extreme heat in the absence of oxygen. It is observed that D melts before B. At the conclusion of the experiment it is noted that A has not melted and C has turned a dark grey colour. Identify which substance is which and explain your reasoning. **(4 marks)**

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3 Reactions of metals

KEY KNOWLEDGE

In this topic you will investigate:

Reactions of metals

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

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of VCE Chemistry. Experiments and investigations, supported by a **practical investigation eLogbook** and **teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

EXAM PREPARATION

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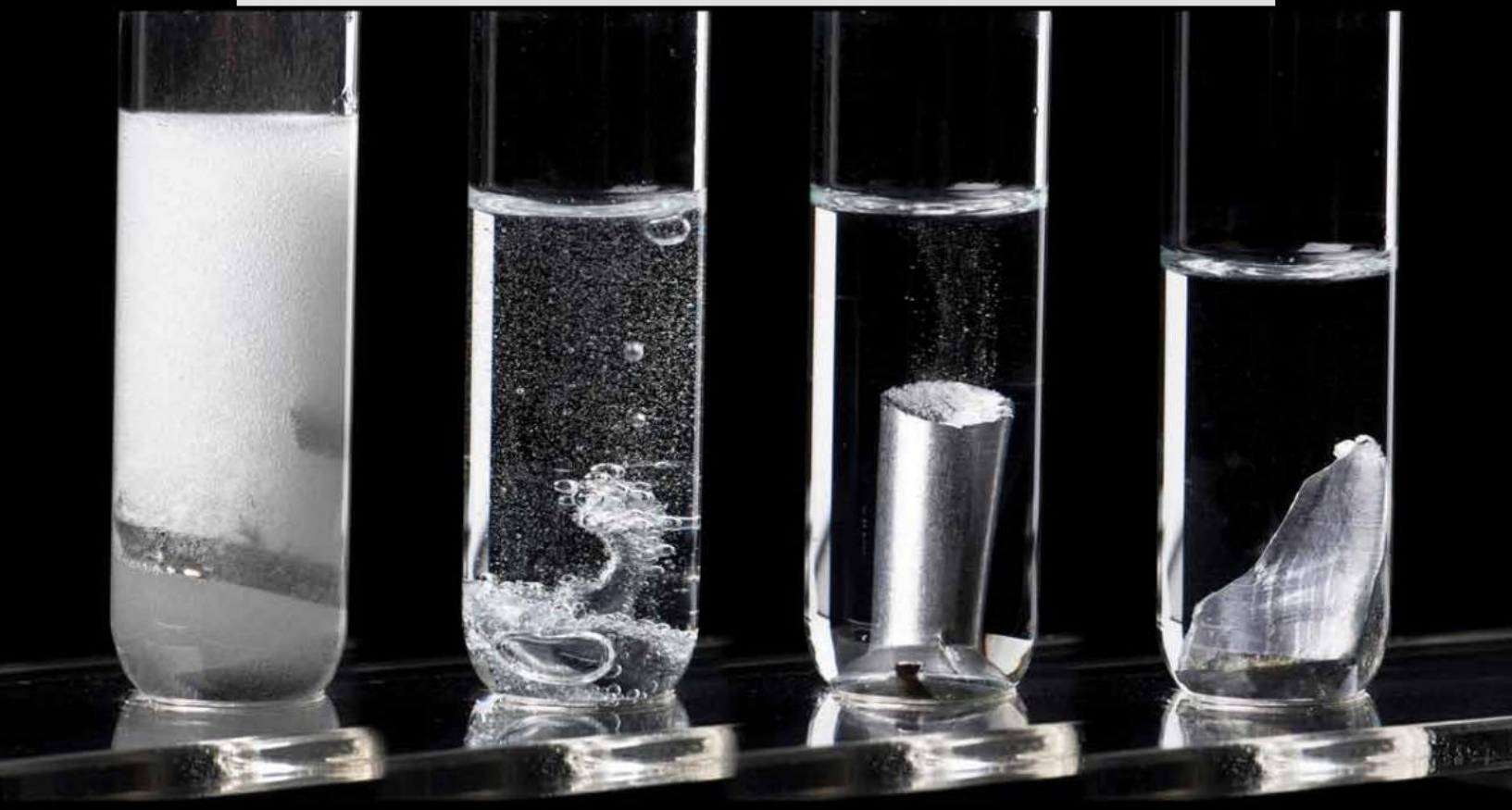

3.1 Overview

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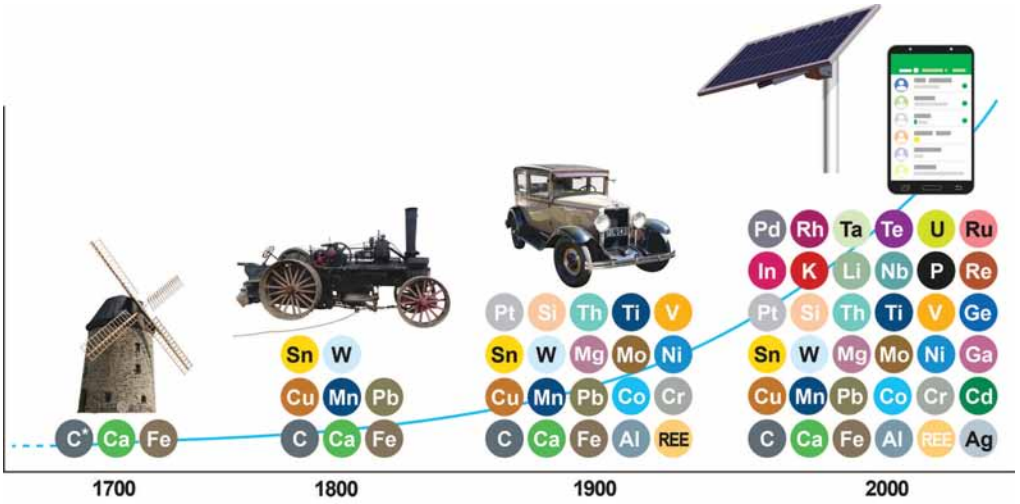
3.1.1 Introduction

A limited number of metallic elements has been in use for many centuries, but currently, the number of different metals in use has increased significantly, either as pure metals or combined with other elements as alloys. Your mobile phone, for example, contains over 60 metals each chosen for its specific properties. The quantity of metals required by industry is increasing and some are becoming scarce, so we need to review how we use metals. The good news is that metals can be recycled and still retain their properties, but a process that is innovative, economical and circular is required.

FIGURE 3.1 The number of different metals we use has increased over time.



*Carbon is not a metal itself, but is present in the metals.

LEARNING SEQUENCE

3.1 Overview 114
3.2 Properties of metals 115
3.3 Reactivity of metals 121
3.4 Recycling metals 129
3.5 Review 139

on Resources

Solutions	Solutions — Topic 3 (sol-0801)
Practical investigation eLogbook	Practical investigation logbook — Topic 3 (elog-1607)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 3 (doc-37069) Key idea summary — Topic 3 (doc-37070)
Exam question booklet	Exam question booklet — Topic 3 (eqb-0083)

3.2 Properties of metals

KEY KNOWLEDGE

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

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3.2.1 The physical properties of metals

Since more than 80 metal elements exist, it is not surprising that some of their properties vary widely.

At room temperature, for example, mercury is a liquid, but magnesium is a solid; zinc is brittle, iron is hard and sodium (figure 3.2) can be cut with a knife; lead has a high density (11.3 g/cm^3) and sodium has a low density (0.97 g/cm^3). Look around you and identify where metals are used and suggest why each metal was chosen. Most metals, however, have the following properties in common.

FIGURE 3.2 Sodium, as with other group 1 and 2 metals, has a low density and can be cut with a knife.

FIGURE 3.3 Steel can be pressed and rolled into sheets, as at this factory.

Metals:

- are lustrous (have a shiny surface when polished)
- are malleable (can be hammered, bent or rolled into sheets or other shapes)
- are ductile (can be drawn out into wires)
- are good conductors of heat
- are good conductors of electricity
- generally have a high density (mass per unit volume)
- have a range of melting points (but most are quite high)
- are often hard (have high resistance to denting, scratching and bending).

3.2.2 Crystal nature of metallic bonding

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

X-ray studies show that metals occur as crystal lattices. Chemists imagine these **metallic lattice structure** as being made up of a patterned array of cations (see figure 3.4). The electrons from each metallic atom overlap each other forming a sea of mobile electrons that can flow between all the cations. They are referred to as **delocalised** electrons (figure 3.5).

Electrostatic forces of attraction between the cations and the negatively charged valence electrons occur in all directions and this holds the lattice together. This type of non-directional bonding is known as **metallic bonding**. This means that metal atoms are hard to separate but relatively easy to move.

cation a positively charged atom or group of atoms

metallic lattice structure crystal lattice made up of an array of cations

delocalised describes electrons that are not bound to any one atom but are free to move throughout a lattice

electrostatic force attractive or repulsive force between particles that is due to their electric charges

metallic bonding positively charged metal cations arranged in a lattice with delocalised valence electrons being able to flow around them

FIGURE 3.4 This box of oranges gives a representation of a repeated pattern array of cations in a lattice structure.

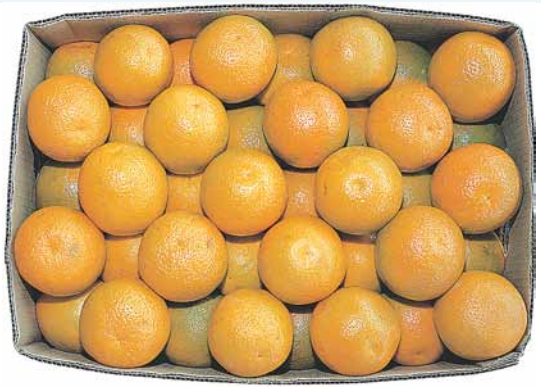
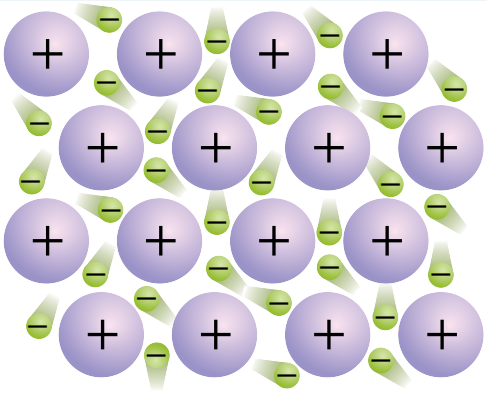


FIGURE 3.5 Metal cations vibrate about a fixed position surrounded by a sea of delocalised electrons.

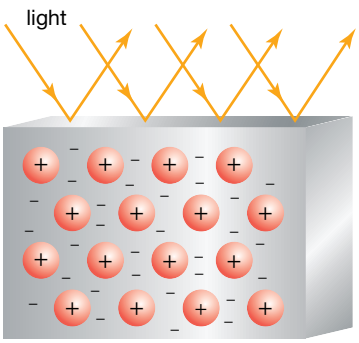
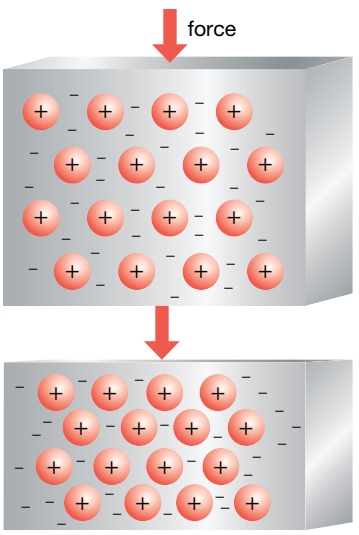


3.2.3 The structure and properties of metallic crystals

The lattice structure of metals may be used to explain many of their properties. These properties give metals their many applications in our society, as described in table 3.1.

lustrous having a shiny surface
malleable able to be hammered or pressed into shape without breaking or cracking
ductile able to be drawn out into a thin wire

TABLE 3.1 How the properties of metals relate to structure

Property of metal	Diagram
Lustre The **lustrous** appearance of a metal is due to the mobile electrons within the lattice being able to reflect light back into your eye, causing the metal to look shiny.	
Malleability and ductility Metals are **malleable** and **ductile**, rather than brittle, as a result of the non-directional nature of metallic bonds. The attractive forces exerted by the cations for the mobile electrons occur in all directions. This means that layers of atoms can move past one another without disrupting the force between the cations and the negative sea of electrons. The nature of the metal does not change when the metal becomes thinner.	

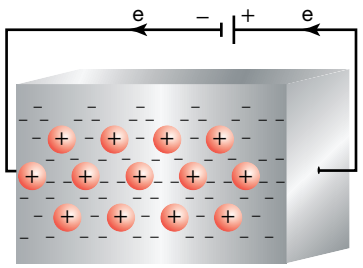
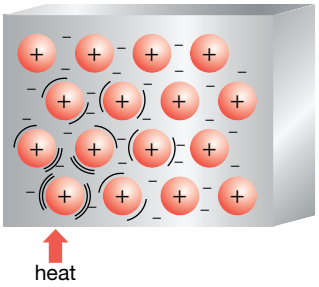
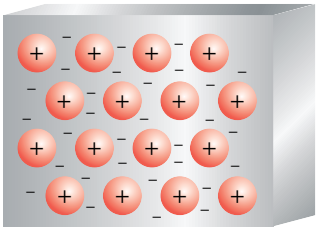
Property of metal	Diagram
Electrical conductivity Metals are good conductors of electricity. When an electric field is applied to a metal, one end of the metal becomes positive and the other becomes negative. All the electrons experience a force towards the positive end. This flow of charge (movement of electrons) is what we call an electric current.	
Heat conductivity Metals are good conductors of heat. Electrons gain kinetic energy in hotter areas of the metal and quickly transfer it to other colder parts of the metal lattice due to the electron's freedom of movement. The heat causes the electrons to move faster, and the 'bumping' of these electrons against each other and the protons transfers the heat throughout the metal.	
Density Most metals have relatively high densities because metallic lattices are close-packed. **Melting points and hardness** The generally high melting points and hardness of metals indicate that metallic bonding is quite strong (although not as strong as covalent or ionic bonding). Melting points and hardness increase with an increase in the number of outer shell electrons, since this greater number means a greater attractive force exists between the cations and the electrons.	

TABLE 3.2 Examples of the variety in each of these properties of metals and non-metals

Element	Metal/metalloid/ non-metal	Density ($g\ mL^{-1}$)	Hardness (Moh's scale)	Thermal conductivity ($W\ m^{-1}K^{-1}$)	Electrical conductivity (S/m)	Melting point (°C)
Aluminium	Metal	2.7	2.8	235	3.77×10^{7}	660
Copper	Metal	9.0	3.0	400	5.96×10^{7}	1085
Silver	Metal	10.5	2.5	430	6.30×10^{7}	961
Iron	Metal	7.9	4.0	80	1.00×10^{7}	1535
Silicon	Metalloid	2.3	6.5	150	4.35×10^{-4}	1410
Boron	Non-metal	2.3	9.4	27	10^{-4}	2300
Phosphorus	Non-metal	1.8	0.5	0.24	1.0×10^{-10}	44
Sulfur	Non-metal	2.1	2.0	0.205	10^{-16}	115
Oxygen	Non-metal	0. 0014	0	0.0277	N/A	–223

on Resources

Weblinks Structure of metals
Structure and bonding
Metallic crystals

EXTENSION: Limitations of models

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

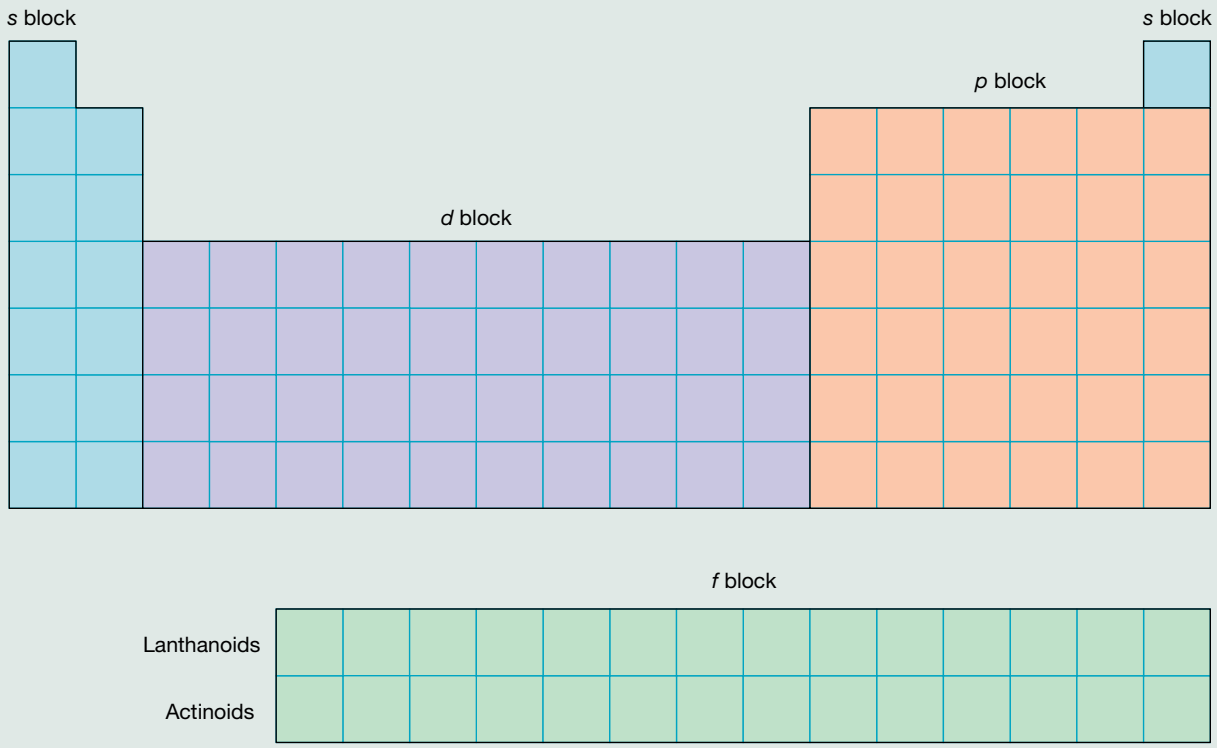
- Why is iron magnetic and copper non-magnetic?
- Why is platinum 20 times as dense as sodium?
- Why is lead malleable whereas iron is tough?
- Why does such a difference exist between the hardness of the different metals?
- Why is mercury a liquid and most other metals solid at room temperature?

As with many models the 'cations in a sea of electrons' model provides a simplified representation of the structure of metals to allow greater understanding, and to explain the majority of metal properties. Other factors also contribute to the differences in properties, and these include the number of valence electrons and atomic mass, and how atomic size impacts on how well the atoms fit together. Metal properties can be changed by heat treatment when they are manufactured and by combining them with other metals or non-metals to form alloys.

EXTENSION: Properties of main group and transition metals

Metals are widely used in our society. Their different properties make them useful for particular purposes. The different blocks of the periodic table were introduced in topic 1, and are shown in figure 3.6.

FIGURE 3.6 The blocks of the periodic table



Properties of the *s*-block metals

The *s*-block metals include the alkali metals (group 1) and the alkali earth metals (group 2). Their properties are related to having only one (alkali metals) or two (alkaline earth metals) electrons in their outer shell.

Properties of alkali metals	Properties of alkaline earth metals
• Very reactive and are not found in elemental form. Most uses involve their compounds.	• Reactive metals, usually found as compounds. Less reactive than alkali metals.
• Low densities. The atoms are larger than other atoms in the same period, and they are not as efficiently packed together.	• Harder and denser than alkali metals.
• Relatively low melting points.	• Higher melting points than alkali metals.
Uses: Sodium is used as a coolant in nuclear reactors, in the manufacture of titanium and for street lights. Lithium is used in rechargeable batteries and, because of its lightness, is combined with other metals to make alloys.	Uses: Magnesium is used in alloys for aircraft and guided missile parts where lightness and high tensile strength are essential. The compound calcium carbonate is used to make calcium oxide for neutralising soils, glassmaking and as a component of cement.

Properties of the *p*-block metals

The *p*-block of the periodic table contains metals such as aluminium, tin and lead, as well as metalloids and non-metals.

- Aluminium is the most abundant metal in the Earth's crust and is very versatile. When exposed to air it develops a coating of aluminium oxide, Al_2O_3, which prevents further reactions with oxygen, and hence further corrosion. Aluminium is strong and light, and is used in transport, building, and food storage and preparation such as drink cans and saucepans. The good electrical conductivity and low density of aluminium make it ideal for use in overhead wires.
- Tin does not corrode easily, and tin cans are highly resistant to the corrosive properties of acidic foods, such as tomatoes. The alloys solder, bronze and pewter all contain tin.
- Lead has been used since Roman times because it is easily worked and resistant to corrosion. Lead was used for water pipes, in pottery glazes and paint pigments, in jewellery and as a petrol additive. Due to health hazards, these uses for lead have significantly declined in recent decades. Current uses include car batteries, roofing, exercise weights and bullets.

Properties of the *d*-block metals

The *d*-block metals are transition metals. These elements have a number of interesting properties.

- Many (except scandium and zinc) form coloured compounds, and are used in stained glass.
- The magnetic metals — iron, cobalt and nickel — are transition metals.
- They can form ions of different charges and, therefore, a wide variety of compounds. For example, iron forms two compounds with chlorine — $FeCl_2$, which is a green compound, and $FeCl_3$, which is an orange compound.
- As with other metals, most *d*-block metals are used in the form of alloys. Iron is widely used in construction, often in alloy form as steel, which is less brittle and more resistant to corrosion than elemental iron.
- They are useful catalysts. A catalyst is a substance that increases the rate of a chemical reaction without itself being chemically changed. Iron is used as a catalyst in the production of ammonia.
- They are good conductors of electricity, with copper used in saucepans and gold in electrical wiring.

FIGURE 3.7 Transition elements form coloured solutions.

3.2 Activities

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3.2 Quick quiz on | 3.2 Exercise | 3.2 Exam questions

3.2 Exercise

1. MC What is the metallic bonding model best described as?
   - A. An electrostatic attraction between metal atoms and delocalised electrons
   - B. The tendency of metal atoms to lose electrons when bonding with other elements
   - C. The electrostatic attraction between layers of positive metal ions and negative metal ions
   - D. Layers of metal cations held electrostatically by surrounding delocalised electrons
2. MC Which of the following *cannot* be explained by the metallic bonding model?
   Property I: Magnetism
   Property II: Ductility
   Property III: Variations in melting temperatures
   - A. All three properties can be explained by the metallic bonding model
   - B. Property I only
   - C. Properties I and III only
   - D. All three properties are unable to be explained by the metallic bonding model
3. Lithium exists as a crystalline solid at temperatures lower than 180 °C.
   - a. Write the electron configuration of a lithium atom.
   - b. Write the symbol and electron configuration of its cation.
   - c. Describe how lithium atoms are bonded together in the solid.
4. Explain the following properties of iron.
   - a. Hard
   - b. Malleable and ductile
   - c. Able to conduct electricity in solid and molten form
5. Magnesium is produced from the mineral dolomite, $CaMg(CO_3)_2$.
   - a. Describe the bonding in magnesium metal crystals.
   - b. Magnesium crystals are silver-white and shiny. Explain this property of lustre in terms of structure and bonding.
6. a. The model of metals that describes them as consisting of 'cations in a sea of electrons' does not explain all properties of metals. Why is it used?
   - b. State two properties of metals that are not explained by this model.
7. Explain why aluminium is used for drink cans whereas 'tin' cans are used for other foods.
8. Give three reasons silver and gold are used for jewellery.
9. Choose a metal that would be best suited for fabrication of the following.
   - a. Car bodies
   - b. Beer cans
   - c. Household electrical wiring
   - d. Fishing sinkers

   Justify your choices.
10. Electricity transmission lines for industry are usually made from aluminium even though copper is a better electrical conductor. Suggest two reasons aluminium is used.

3.2 Exam questions

Question 1 (1 mark)

MC Which of the following best describes the metallic bonding model?

A. A lattice of metal molecules connected by delocalised electrons
B. A lattice of metal cations held electrostatically to a 'sea' of moving electrons
C. A lattice of metal cations and non-metal anions held by strong ionic bonds
D. A lattice of metal atoms closely held together by electrostatic forces

Question 2 (1 mark)

MC A sample of element *Z* can be hammered out into sheets.

Which of the following describes other properties of element *Z*?

A. Non-conductor of electricity, low melting point
B. Non-conductor of electricity, high melting point
C. Conductor of electricity, low melting point
D. Conductor of electricity, high melting point

Question 3 (2 marks)

Some properties of four elements *M*, *N*, *O* and *P* are as follows.

M melts at 2200 °C, is hard and is a non-conductor.
N has a high density and can be drawn into wires.
O is liquid at room temperature and can conduct electricity.
P is lustrous, has a melting point of 1410 °C and is a non-conductor of electricity.
State which elements are metals.

Question 4 (5 marks)

List two metallic properties required for each of the following objects.

a. Plane **(1 mark)**
b. Saucepan **(1 mark)**
c. Mobile phone cover **(1 mark)**
d. Bridge **(1 mark)**
e. Outdoor furniture **(1 mark)**

Question 5 (3 marks)

Copper has far better electrical conductivity than aluminium and is used for household wiring.

a. Explain why aluminium is used instead of copper for overhead wiring. **(1 mark)**
b. Describe why the structure of these substances allow them to conduct electricity. **(1 mark)**
c. Name the property that refers to the ability to be drawn out into wires. **(1 mark)**

More exam questions are available in your learnON title.

3.3 Reactivity of metals

KEY KNOWLEDGE

- Experimental determination of a reactivity series for metals based on their relative ability to undergo oxidation with water, acids and oxygen

Source: VCE Chemistry Study Design (2023–2027) extracts © VCAA; reproduced by permission.

3.3.1 Reactivity and the periodic table

Most of the elements in the periodic table are metallic but, just as the elements vary in their physical properties, their chemical properties can show extremes. Sodium and potassium, for example, react with the oxygen in the air so they are stored in oil. Gold and silver, however, are unreactive and retain their metallic lustre.

The reactivity of metals is related to how readily they lose electrons to form positive ions (cations). Metals usually have one, two or three electrons in their valence shells, which can be donated in chemical reactions.

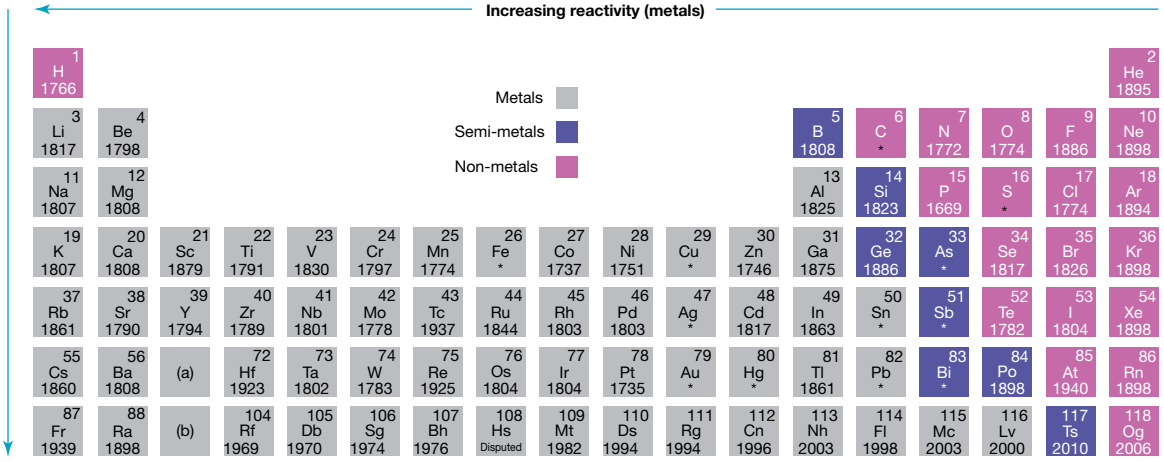
In chemical terms, when a substance loses electrons, it is described as an oxidation reaction. The elements sodium and potassium are in group 1, and readily lose an electron to achieve a stable noble gas electron configuration. This ability to donate an electron increases going down the periodic table, as the force of attraction to the valence electrons by the atom's nucleus decreases as the number of shells increases.

Reactivity of metals

The reactivity of a metal is dependent on its ability to lose electrons to form cations — in other words, how readily it is oxidised.

This tendency to form cations is confirmed by observing the changes in ionisation energy of elements, as described in topic 1. This chemical property distinguishes metals from metalloids and non-metals because they do not form cations.

FIGURE 3.8 General trends in reactivity in the metallic elements



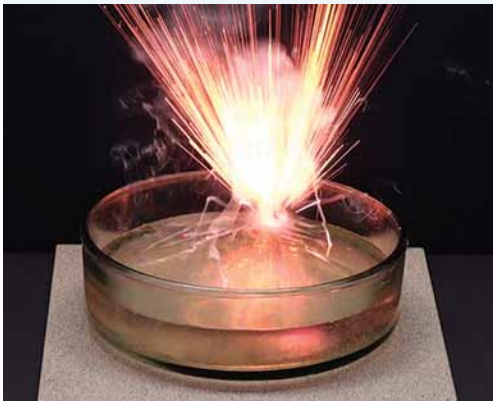
Determining qualitatively the reactivity of samples of different metals is possible by investigating experimentally how readily they react with oxygen, cold water, steam and dilute acids (a summary of the reactions can be found in topic 1).

3.3.2 Relative reactivity of metals with water

Potassium reacts very vigorously with water to produce hydrogen gas and potassium hydroxide (see figure 3.9). The heat given off in this exothermic reaction is sufficient to ignite the hydrogen gas produced, and allows it to burn in air with a bright lilac flame. But not all metals react with water; fortunately, stainless steel is unreactive, for example, and so water can be stored safely in metallic drink bottles.

As figure 3.8 indicates, group 1 (alkali metals) and group 2 (alkaline earth metals) are highly reactive, and most of these metals react with water at room temperature to produce alkalis and hydrogen gas. Alkalis are metal hydroxides that are soluble in water and form solutions that neutralise acids. This is why highly reactive metals must be stored under oil to prevent their reaction with oxygen or atmospheric water vapour.

FIGURE 3.9 Although the lattice structure of metals is strong, they can still react — as potassium does on contact with water.



	Group 1 metal	+	**water**	→	**group 1 metal hydroxide**	+	**hydrogen gas**
Example:	potassium	+	water	→	potassium hydroxide	+	hydrogen gas
	2K(s)	+	$2H_2O(l)$	→	2KOH(aq)	+	$H_2(g)$

	Group 2 metal	+	**steam**	→	**group 2 metal hydroxide**	+	**hydrogen gas**
Example:	magnesium	+	steam	→	magnesium hydroxide	+	hydrogen gas
	Mg(s)	+	$2H_2O(g)$	→	$Mg(OH)_2(aq)$	+	$H_2(g)$

In each of the reactions shown in the equations, the solid metal reactant is converted into a metal ion in solution. In effect, the metal ions have given away electrons to form metal ions. The electrons are accepted by the water and hydrogen gas is formed. This is an oxidation reaction. Oxidation reactions and the accompanying reduction reactions are discussed in more detail in topic 12.

on Resources

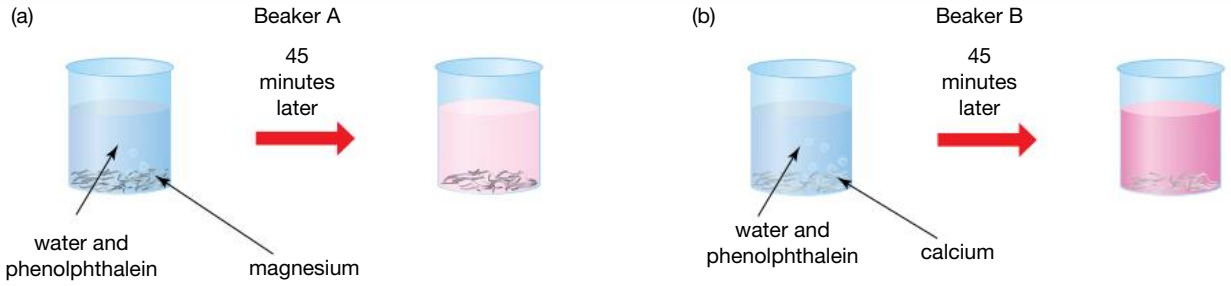
Video eLesson Reaction of potassium with water (eles-3209)

Weblinks Introduction to the periodic table: alkali metals
Alkali metal reactivity

Relative reactivity down a group

By reacting the different metals with water, we can compare the reactivity. The experiment illustrated in figure 3.10 shows how the relative reactivity of magnesium and calcium can be observed. When equal amounts of freshly sanded magnesium ribbon and strips of oil-free calcium react with water in beakers A and B as shown, bubbles of hydrogen gas and a hydroxide compound are formed. The presence of the hydroxide compound is shown by the indicator, phenolphthalein. After 45 minutes, beaker B has a more intense pink colour, suggesting that more hydroxide compound has been produced by the reaction. We can deduce qualitatively from this that calcium is more reactive with water than magnesium.

FIGURE 3.10 Demonstration of relative reactivity of (a) magnesium and (b) calcium



In general terms, if a metal reacts with cold water, a hydroxide and hydrogen gas is formed, whereas if they only react with steam, they tend to form metal oxides because the high temperatures decompose the hydroxide into the oxide. Magnesium reacts more readily with steam; aluminium, zinc and iron will also react with steam. Most other metals, however, do not show any reaction with water.

Reactions between metals and water

Reactive metal + water → metal hydroxide + hydrogen

tlvd-3045

SAMPLE PROBLEM 1 Writing an equation for the reaction of a metal and water

Barium, a group 2 metal, reacts readily with cold water and considerable bubbling is observed in the solution. When phenolphthalein indicator is added, the solution becomes an intense pink colour. Write a balanced equation for this reaction.

THINK	WRITE
1. Write out the formulas of the reactants: barium and water.	1. $Ba + H_2O$
2. Write out the formulas of the products: barium hydroxide and hydrogen. Remember that barium is in group 2 and so its ion is Ba^{2+}, while the hydroxide ion is OH^-. Hydrogen is a diatomic gas, H_2.	2. $Ba(OH)_2 + H_2$
3. Write out the equation.	3. $Ba + H_2O \rightarrow Ba(OH)_2 + H_2$
4. Balance the equation.	4. $Ba + 2H_2O \rightarrow Ba(OH)_2 + H_2$
5. Add states.	5. $Ba(s) + 2H_2O(l) \rightarrow Ba(OH)_2(aq) + H_2(g)$

PRACTICE PROBLEM 1

The density of lithium, a group 1 metal, is about half that of water so it floats. Lithium reacts with the water producing some bubbles and a clear and colourless solution. Write a balanced equation for this reaction.

3.3.3 Relative reactivity of metals with acids

Hydrogen is also produced when reactive metals react with acids. Usually, the acids used for these tests are hydrochloric or sulfuric acid. The simple pop test can be used to confirm that the gas evolved is hydrogen. A burning wooden splint is placed into the test tube of collected hydrogen, which then makes a 'pop' sound. The other product is a metal salt. A **salt** is a neutrally charged chemical compound made of cations (positive ions), which are usually metallic, bonded to anions (negative ions), which are usually non-metallic.

Besides observing the formation of gas bubbles during these experiments, heat is released and the container may feel warm. The more reactive metal reacts faster and produces more bubbles. In these reactions, electrons are also transferred in an oxidation reaction. The many H^+ ions, provided by the acid, readily accept the electrons, making the formation of hydrogen gas more feasible than in the reaction with water.

Elements in group 1 react violently with acids; magnesium, aluminium, zinc, iron all react progressively less vigorously with acids; tin and lead react very slowly.

FIGURE 3.11 Reactivity of metals with hydrochloric acid; from left to right: calcium, magnesium, zinc, copper

salt a neutrally charged chemical compound made of cations (positive ions), which are usually metallic, bonded to anions (negative ions), which are usually non-metallic

	Group 1 metal	+	**acid**	→	**salt**	+	**hydrogen**
Example:	lithium	+	hydrochloric acid	→	lithium chloride	+	hydrogen
	2Li(s)	+	2HCl(aq)	→	2LiCl(aq)	+	$H_2(g)$

	Group 2 metal	+	**acid**	→	**salt**	+	**hydrogen**
Example:	magnesium	+	sulfuric acid	→	magnesium sulfate	+	hydrogen
	Mg(s)	+	$H_2SO_4(aq)$	→	$MgSO_4(aq)$	+	$H_2(g)$

Note that electrons have also been donated by the metal atoms in these reactions in order to form metal ions. The metals are described as being oxidised in an oxidation reaction.

TABLE 3.3 Salts produced when metals react with acids

Acid	Salt formed
Hydrochloric (HCl)	Metal chloride; e.g. NaCl
Nitric (HNO_3)	Metal nitrate; e.g. $NaNO_3$
Sulfuric (H_2SO_4)	Metal sulfate; e.g. Na_2SO_4

Reactions between metals and acid

Reactive metal + acid → metal salt + hydrogen

tlvd-3046

SAMPLE PROBLEM 2 Writing an equation for the reaction of a metal and an acid

When a small piece of potassium is added to hydrochloric acid, a very vigorous reaction occurs — the metal immediately ignites and burns with a lilac flame. The products are a solution and hydrogen gas. Write a balanced equation for this reaction.

THINK	WRITE
1. Write out the formulas of the reactants: potassium and hydrochloric acid.	1. $K + HCl$
2. Write out the formulas of the products: potassium chloride and hydrogen gas. Remember that potassium is in group 1 and so its ion is K^+, while chlorine is in group 17 so chloride ion is Cl^-. Hydrogen is a diatomic gas, H_2.	2. $KCl + H_2$
3. Write out the equation.	3. $K + HCl \rightarrow KCl + H_2$
4. Balance the equation.	4. $2K + 2HCl \rightarrow 2KCl + H_2$
5. Add states.	5. $2K(s) + 2HCl\ (aq) \rightarrow 2KCl(aq) + H_2(g)$

PRACTICE PROBLEM 2

The products formed when magnesium reacts with nitric acid depend on the concentration of the nitric acid. A small spatula of magnesium powder was added to a large test tube that contained dilute nitric acid solution, HNO_3. Bubbles were initially observed. When bubbles were no longer observed, more magnesium powder was added and the reaction again produced bubbles. This was repeated until magnesium powder remained in the beaker and no further bubbles were observed. When a flame was placed at the mouth of the test tube, the gas burned with a squeaky pop.

Answer the following.

a. Identify the gas produced.
b. Give the name of the colourless solution formed in this reaction.
c. Write a balanced equation for the reaction.
d. Suggest why the reaction stopped.

3.3.4 Relative reactivity of metals with oxygen

Many metals react with oxygen to form metal oxides. Very reactive metals such as potassium and sodium (group 1) react very vigorously with pure oxygen (see figure 3.12) and rapidly with air, which is why they need to be stored in oil. Generally, reactive metals form metal oxides when they react with oxygen.

FIGURE 3.12 Sodium reacting violently with pure oxygen

Iron will burn in air but only if in powder form or a thin wire. Copper will not burn, but if heated very strongly a layer of black copper oxide will form. Gold and silver will not react with oxygen even after strong heating. An oxide layer can form on some metals and cause the metals to appear dull, but when this layer is removed or the metal is cut, the shiny surface appears.

The thin oxide layer on aluminium protects the metal underneath from further reaction and so prevents corrosion. Unlike aluminium oxide, when iron rusts, the oxide forms a permeable layer that allows air and moisture to pass through and continue to corrode the metal. Therefore, iron usually needs some form of protection to prevent corrosion.

	Group 1 metal	**+**	**oxygen**	**→**	**metal oxide**
Example:	sodium	+	oxygen	→	sodium oxide
	4Na(s)	+	O_2(g)	→	$2Na_2O$(s)

	Group 2 metal	**+**	**oxygen**	**→**	**metal oxide**
Example:	calcium	+	oxygen	→	calcium oxide
	2Ca(s)	+	O_2(g)	→	2CaO(s)

The metals in these reactions donate electrons to the oxygen and become metal ions in an oxidation reaction.

Reactions between metals and oxygen

Reactive metal + oxygen → metal oxide

tlvd-0513

SAMPLE PROBLEM 3 Understanding reactions of metals

Why is it necessary for magnesium to be freshly sanded for a reaction to occur with cold water?

THINK

Magnesium, and some other metals, appear to be unreactive, but this is not always the case. Consider what happens to the magnesium when it is exposed to air that may affect the reaction with cold water.

WRITE

Metals like magnesium (and especially aluminium) react readily with atmospheric oxygen, forming a protective coating of the metal oxide. This metal oxide coating acts as a barrier, preventing water from coming into direct contact with the metal so no reaction or a slow reaction occurs. Sanding the magnesium allows the water to come into contact with the pure metal, allowing a reaction to occur.

PRACTICE PROBLEM 3

Explain why calcium is stored in oil and why is it necessary for it to be oil-free for a reaction to occur.

3.3.5 The reactivity series of metals

By comparing and observing experiments showing reactions of metals with water, acids and oxygen, a list from most reactive to least reactive metal can be established. These reactions demonstrate how readily metals release electrons in an oxidation reaction to other substances. Displacement reactions confirm the general order of metals in the **reactivity series of metals** (see topic 12).
A summary of the order of reactivity of metals is shown in table 3.4.

reactivity series of metals
an ordered list of how readily the metals react with oxygen, water, steam, dilute acids and salt solutions

TABLE 3.4 Reactivity series of metals

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

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

elog-1670

tlvd-3047

EXPERIMENT 3.1 online only

Reactivity of metals

Aim

To investigate the reactivity of different metals through reactions with water and acid

elog-1669

EXPERIMENT 3.2 online only

Metallic trees

Aim

To produce metal crystals of silver or lead

on Resources

Weblinks Metals in aqueous solution
Metals and reactivity series

3.3 Activities

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3.3 Quick quiz on | 3.3 Exercise | 3.3 Exam questions

3.3 Exercise

1. MC Which of the following will not react with water?
   A. Sodium **B.** Calcium **C.** Silver **D.** Copper
2. MC Name the products of the reaction of zinc and sulfuric acid.
   A. Zinc sulfate and oxygen gas **B.** Zinc sulfate and hydrogen gas
   C. Zinc chloride and oxygen gas **D.** Zinc chloride and hydrogen gas
3. MC Complete the following equation: $4Fe\ (s) + 3O_2\ (g) \rightarrow$
   A. $4Fe_3O_2$ **B.** $4Fe_2O_3$ **C.** $2Fe_2O_3$ **D.** 2FeO
4. MC Which four numbers are required to balance the equation?

$$\mathbf{a}\ Li(s) + \mathbf{b}\ H_2SO_4(aq) \rightarrow \mathbf{c}\ Li_2SO_4(aq) + \mathbf{d}\ H_2(g)$$

   A. 2, 1, 1, 2 **B.** 1, 2, 1, 1 **C.** 2, 2, 1, 2 **D.** 2, 1, 1, 1
5. MC As we move across a period from left to right, what happens to the reactivity of metals with oxygen?
   A. Increases **B.** Decreases
   C. Stays the same **D.** Increases then decreases
6. When a sample of calcium metal reacts with water, the equation for the reaction is as follows.

$$Ca(s) + H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$$

   Explain why this is described as an oxidation reaction.
7. Write fully balanced equations, including states, for the following reactions.
   a. Potassium + oxygen
   b. Caesium + water
   c. Tin + nitric acid, HNO_3
8. Why must sodium and potassium be stored in an oil?
9. Magnesium metal reacts vigorously with a solution of hydrochloric acid, producing a gas that gives a trademark 'pop' sound with a lighted match. An ionic salt is formed at the same time. Write the balanced equation for this reaction including symbols of state.
10. You have samples of three metals, *X*, *Y* and *Z*. Describe a possible method to find out the order of reactivity, including any chemicals and equipment required, and the reactants formed.

3.3 Exam questions

Question 1 (1 mark)

MC Which of the following metals would be expected to react most vigorously with an acid solution?

A. Copper
B. Calcium
C. Aluminium
D. Magnesium

Question 2 (1 mark)

MC What are the products of the reaction of magnesium with water?

A. Magnesium hydroxide and oxygen gas
B. Magnesium oxide and oxygen gas
C. Magnesium hydroxide and hydrogen gas
D. Magnesium oxide and hydrogen gas

Question 3 (1 mark)

MC The equation for the burning of magnesium metal in oxygen gas to form a white powder is as follows.

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

In this reaction:

A. Magnesium has gained electrons and forms Mg^{2+} cations.
B. Magnesium has lost electrons and forms Mg^{2+} cations.
C. Oxygen has gained electrons and forms O^{2-} cations.
D. Oxygen has lost electrons and forms O^{2-} cations.

Question 4 (5 marks)

The following is a section of the reactivity series of metals: K Ca Mg Pb Cu Ag

Choose a metal from the list for each of the following.

a. This metal is used extensively for household wiring because it is a good conductor. **(1 mark)**
b. This metal is found as an element in nature. **(1 mark)**
c. Under the same conditions, when reacted with acid this metal produces the greatest volume of hydrogen in 5 minutes. **(1 mark)**
d. This metal is in group 2 and is more reactive than magnesium. **(1 mark)**
e. This metal is very malleable and dense. **(1 mark)**

Question 5 (6 marks)

Sodium and calcium are reactive metals.

a. Describe two similarities between the reactions of these metals and water. **(2 marks)**
b. Describe two differences between the reactions of these metals and water. **(2 marks)**
c. How could you confirm what the products of these reactions are? **(2 marks)**

More exam questions are available in your learnON title.

3.4 Recycling metals

KEY KNOWLEDGE

- Metal recycling as an example of a circular economy where metal is mined, refined, made into a product, used, disposed of via recycling and then reprocessed as the same original product or repurposed as a new product

Source: VCE Chemistry Study Design (2023–2027) extracts © VCAA; reproduced by permission.

3.4.1 Use and re-use of metals

Metals are used for an incredible range of products, from minute circuits in electronic devices to gigantic skyscrapers. When no longer useful, these metals are often disposed of as waste. As seen in topic 1, this is no longer sustainable. Reserves of many materials are becoming scarce and the damage to the environment to extract more resources is considerable. By following a **circular economy** model, materials should be designed and used to ensure little or no waste — including during both manufacture and at the end of a product's useful life, where metals should be reused or recycled. The circular economy also requires that no harm comes to the health of individuals or the environment during any of these processes.

circular economy economy in which waste and pollution is eliminated, products at the end of life are reused, repaired or recycled, and the environment is not harmed

CASE STUDY: Environmentally friendly Olympic medals

About 5000 medals were required for the 2020 Tokyo Olympics (held in 2021 due to COVID-19). The officials decided that sourcing these in an environmentally friendly and sustainable manner was important. So, remarkably, the gold, silver and bronze were extracted from used electronic devices, 79,000 tonnes of them, including more than 6 million used mobile phones. This meant that this waste did not go into landfill. This is beneficial, not only from a waste-management perspective, but also because electronic waste is highly toxic. As consumers, we should all be aware that many of the resources used in electronic devices are quickly being depleted.

FIGURE 3.13 Tokyo Olympic medals

3.4.2 Mining metals

The benefits to society of metal production include not only providing essential materials but also offering employment in mining, production, transport and sales. Exporting metal ores substantially contributes to the economy and, therefore, the wealth of Australia. These benefits must be balanced with the costs of obtaining and transporting the raw materials, and the energy requirements for the process. The cost to the environment is also considerable. Mine waste has caused substantial damage to the environment with destruction of the landscape, damage to ecosystems and pollution of air and land.

Mining

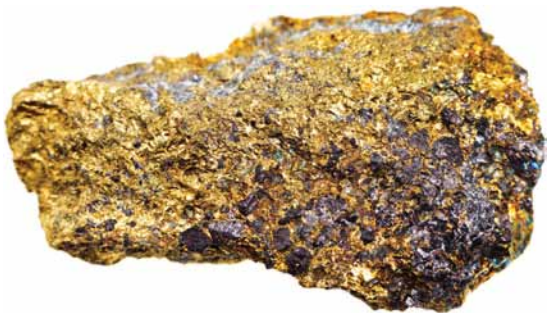
Mining is the process in which metals are extracted from ores in the Earth. **Ores** are rocks that are mixtures of different compounds and from which it is economically viable to separate out the metal. Copper, for example, which ranks second only to iron in its importance in human history, is extracted from the ore chalcopyrite (see figure 3.14).

mining the process of extracting metal ores from the earth

ore a rock containing minerals from which a valuable metal can be removed for profit

FIGURE 3.14 (a) Copper can be extracted from the ore chalcopyrite. (b) Today, the main use of copper is in electrical cables, pipes and wires, shown here ready for recycling.

(a)



(b)

TABLE 3.5 Examples of economically valuable metals, and their reactivity and extraction technique

Reactivity	Metal	Extraction
Very reactive	K, Na, Ca, Mg, Al	Electrolysis of molten compound, using electricity to cause a chemical reaction
Reactive	Zn, Fe, Sn, Pb	Heating the metal ore with carbon
Less reactive	Cu, Hg, Ag	Heating the metal ore in air
Unreactive	Ag, Pt, Au	Can occur as free elements

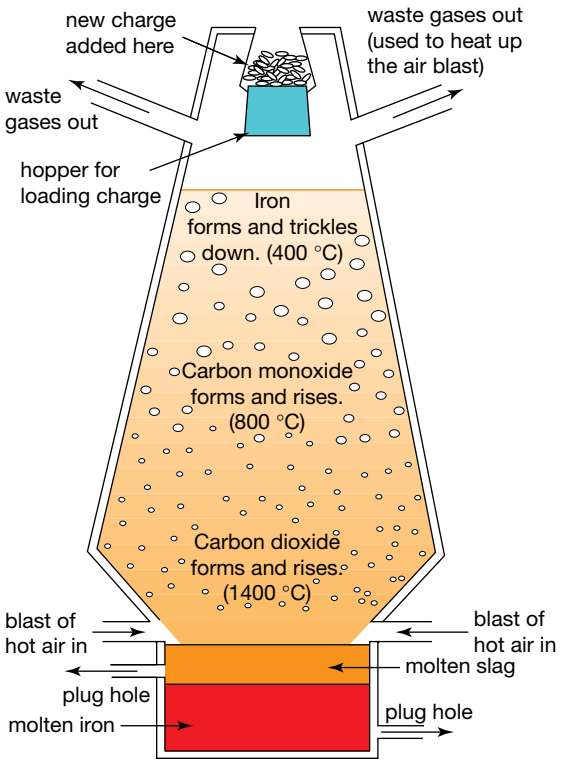
The two main forms of mining are **open-cut** (surface) and **underground**. The choice of method depends on the type of ore that is to be mined, how far it is below the surface and whether the cost of extraction will result in a profit for the company. Often, an enormous amount of rock at the surface must be removed to reach the ore and, once the ore is reached, waste rock must be separated from the ore. The ore is then crushed into small pieces and then finely ground with water, producing even more waste called **tailings**.

Metal extraction

The reactivity of a metal influences how easily it is extracted from its ore (see table 3.5). Silver and gold, for example, are unreactive and can often be found naturally in their elemental state. This is why they have been known and used since ancient times. Gold, for example, can be found through the process of panning in stream beds. Today, however, economically viable deposits of gold and silver require extraction from the **host rock**, because large occurrences of elemental gold and silver have already been found.

More reactive metals such as iron must be extracted from their ores. Our society is highly dependent on iron for transport, construction, machinery, containers and appliances. Because iron is a metal of medium reactivity, it is extracted from its ores by reaction with carbon (charcoal) in a blast furnace in a **reduction** reaction (see figure 3.15). Impurities are removed from the molten metal as a glassy by-product called **slag**. Iron ore is composed of several different minerals (iron-rich compounds), from which iron is extracted. The most used iron ore is haematite, Fe_2O_3.

FIGURE 3.15 Iron production in a blast furnace



Overall reaction: $Fe_2O_3\,(s) + 3CO\,(g) \rightarrow 2Fe\,(l) + 3CO_2\,(g)$

In general: metal oxide (from ore) $\xrightarrow{\text{reduction}}$ metal

The most reactive metals are extracted from their ores by **electrolysis**, a process that requires an electric current to breakdown the ore. Most metals can be obtained from their ores by electrolysis, but this requires a huge amount of electricity, so it is only used if other methods are not suitable. Aluminium is a reactive metal that can be obtained from the ore, bauxite. It is the most abundant metal on Earth but there was a time when aluminium was more valuable than gold because of the difficulty in obtaining it. This changed in 1882, when scientists discovered how to produce aluminium metal using electricity. In their process, alumina (Al_2O_3) is heated to about 960 °C. A large amount of electricity is passed through this molten material and aluminium is formed at the base of the container. Electrolysis, including modern innovations, is examined in unit 3.

Refining metals

Metals formed in the processes described are sometimes not particularly pure and need to be refined after the initial main steps are completed and before the metal can be made into a commercial product. The method of **refining metals** depends on the metal and the type of impurity. Typical impurities include other metals, unreduced oxides of the metal, and non-metals such as carbon, silicon, phosphorus or sulfur. For example, the iron produced in the blast furnace is very brittle with low tensile strength. Further processing is required to remove impurities. The process of electrolysis is also used to refine copper — that is, to obtain pure copper from a sample of impure copper.

open-cut a surface mining technique that removes ore and waste rock from an open pit from the surface

underground a variety of mining techniques that use tunnels and shafts to extract ore and waste rock from under the ground; this process minimises waste rock removal when the ore body is deep below the surface

tailings waste material remaining after the removal of the ore from the host rock

host rock the rock surrounding or containing the ore minerals

reduction a gain of electrons (decrease in oxidation number)

slag waste material from the smelting or refining of ore

electrolysis the decomposition of a chemical substance (in solution or the molten state) by the application of electrical energy

refining metals purifying a metal

3.4.3 Metals and the circular economy

The proportion of pure metal in ores is very small; for example, the percentage of copper in copper ore is about 1 per cent. So a considerable amount of waste material results from metal manufacturing. According to the principles of a circular economy, waste material must be limited if it cannot be eliminated, energy should be conserved, damage to health and environment must be controlled, and products must have a life cycle.

Applying the principles of the circular economy

Table 3.6 summarises examples of the current challenges and opportunities in metal production.

TABLE 3.6 Challenges and opportunities in metal production and moving towards a circular economy model

Consideration	Challenge	Circular economy opportunity
Human health	• Mining can be hazardous, both due to accidents and long-term illness. • Local communities can be subjected to pollution.	• Health and safety regulations should protect the health of workers and communities. • More effective, efficient and safe design processes can be introduced.
Raw materials	• Many resources on Earth are finite. • Populations are increasing so demand for products is also increasing.	• Raw materials must be minimised, reused or used elsewhere. • Manufactured goods should be repaired, repurposed or recycled wherever possible. For example, 'urban mining' can be used to recycle components of technical devices.
Mineral processing: waste generation	• Mining produces large quantities of waste material.	• Waste solids can be used for backfilling (refilling holes), landscaping and road construction. • Some waste material can be reprocessed to obtain other metals; e.g. historical gold tailings can be reprocessed to remove remaining gold using newer, more efficient techniques.

Mineral processing: water use	• Water is used during extracting and processing the metal. Wastewater is often acidic and can be contaminated with hazardous chemicals containing arsenic, cobalt, copper, cadmium, lead, silver, mercury and/or zinc. These can find their way into water sources, affecting the health of humans and wildlife.	• Mine water can be:
		  • recycled through the plant
		  • used for dust suppression or as a coolant
		  • monitored and treated to remove harmful chemicals by raising the pH through the addition of limestone or caustic soda (which also results in precipitate forming that either settles or can be filtered out)
		  • reprocessed to remove trace metals.

Impacts on the environment	• Mining has detrimental effects on ecosystems and land use, including contamination, deforestation, erosion and potential loss of biodiversity.	• Processing at mine sites can be more regulated (e.g. dust and hazardous substance air filtration), and regeneration can occur during mining and at the end of the mine's life. • Gases released can be collected and used in other industries; e.g. sulfur dioxide can be used to make sulfuric acid.

Consideration	Challenge	Circular economy opportunity
Energy use	• Initial manufacture using raw materials uses significant amounts of energy.	• Recycling materials requires less energy than the initial manufacture; e.g. aluminium recycling uses 95% less energy than production from ore. • Renewable energy should be used wherever possible; e.g. solar and battery storage.
Product life	• Few products are currently designed considering recycling opportunities at the end of the product's life. • Products are often designed to be replaced rather than repaired.	• Manufactured goods should be repaired, repurposed, or recycled to conserve energy and resources, and also built to last where possible. • Regulations should be changed to ensure product design includes end-of-life recycling. • Consumer behaviour can change to purchasing higher quality (often more expensive) goods that last longer and/or can be recycled.
Ethical considerations	• Often mining occurs in land covered by native title.	• Regulation and dialogue should occur between the mining/processing companies and indigenous land holders. • Care must be taken to respect cultural sites. • Indigenous peoples' concerns about the impact of mining on their landscape, cultural heritage and traditional way of life should be considered.
Economic considerations	• Mining and mineral processing is very expensive, making raw materials required for product development expensive.	• Metal recycling processes should be financially beneficial. • Economic benefits come through the recycling industry saving raw materials, using less energy, reducing damage to the environment and providing employment.

The opportunity exists for more innovative design of manufacturing processes to achieve sustainable goals. Recycling metals can reduce some of these problems. Less energy is used to recycle metals than the original process, and resources are conserved, resulting in lower costs and less pollution. Collecting waste gases and focusing on revegetation after mining can also reduce some of the environmental impact.

3.4.4 Recycling metals

Metals have distinctive structures and properties, as described using the metallic model of bonding. Unlike plastics, metals can undergo heating, melting and cooling without affecting this basic structure, so metals can be continually recycled, with their properties fully restored. The process, however, is not always easy or economically favourable.

Difficulties with recycling include the following:

- Collection and sorting waste can be difficult and uneconomical.
- In some waste, as in electronic devices, only very small quantities of metals are available, which makes the metal expensive to extract.
- Metal alloys require more sophisticated and expensive technology for separation.

- Metal waste can be contaminated with hazardous waste, which can be detrimental to workers in the industry.
- Toxic metals such as lead, mercury and radioactive metals are too dangerous to recycle.

Metals can be divided into two types: ferrous and non-ferrous.

- Ferrous metals include iron and its alloy, steel.
- Non-ferrous metals include aluminium, copper and gold.

Steel is 100 per cent recyclable and is the most recycled metal. Sorting these two types of metals is relatively easy due to iron's magnetic property. Iron products include cans, containers, whitegoods, bicycles, steel or alloy wheels and many other items including the kitchen sink. Larger steel structures include ships, railroad tracks, farm machinery and, especially, vehicles.

Steps in metal recycling

The process of metal recycling depends on the metal being recycled; figure 3.16 outlines the general steps involved.

FIGURE 3.16 Generalised steps in metal recycling

Some precious metals such as gold, silver, platinum and palladium can be extracted from discarded electronic devices. This process of obtaining metals from waste and metal scrap is described as **urban mining**. While about 75 per cent of gold is obtained from mine production, demand requires that recycled gold is used. Using recycled gold means reduced environmental damage and using less chemicals. Interestingly, one tonne of mobile phones contains more gold than one tonne of gold ore.

urban mining obtaining metals from e-waste and other materials already in our cities

3.4.5 Using microbes to extract metals from waste

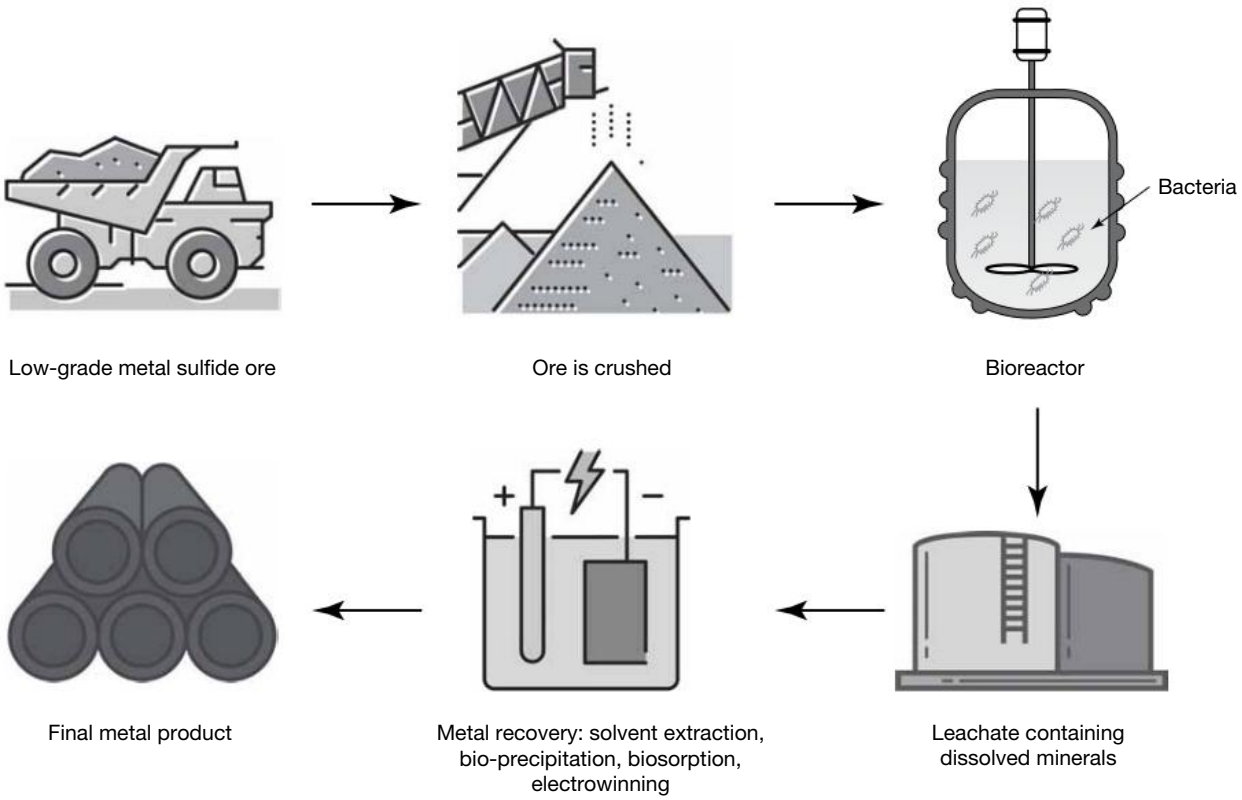
Bioleaching (biomining) is a new technology that uses specific bacteria or fungi to extract valuable metals from metal ores, tailings, e-waste or contaminated water. The microorganisms efficiently catalyse the breakdown of minerals and gain energy in an oxidation reaction process. The bacteria are sourced from the mine environment. The reaction takes place in large, closed tanks (**bioreactors**) that are continually stirred. Low-grade ores can be treated to produce an acidic solution containing metal ions. Copper is one of the metals commonly obtained but other metals such as nickel, silver, cobalt and gold can also be extracted using this process.

This method is simpler and more sustainable than smelting, and helps ensure a more circular economy. The process is also more economical and uses less energy because high temperatures are not required. Although toxic acid is produced, this can be used in other industries and sulfur dioxide emissions are reduced. If used in mining, environmental damage is also reduced, and fewer health problems arise for miners. Manufacture of metals using bioleaching takes a longer time and only a few plants are currently using the technology on an industrial scale. Research is continuing to improve the efficiency of the process.

bioleaching (biomining) the process of extracting metals from ores or waste by using microorganisms

bioreactor industrial device that allows a biological process to occur

FIGURE 3.17 Low-grade metal sulfide ore can be processed using bioleaching before the metal recovery phase.



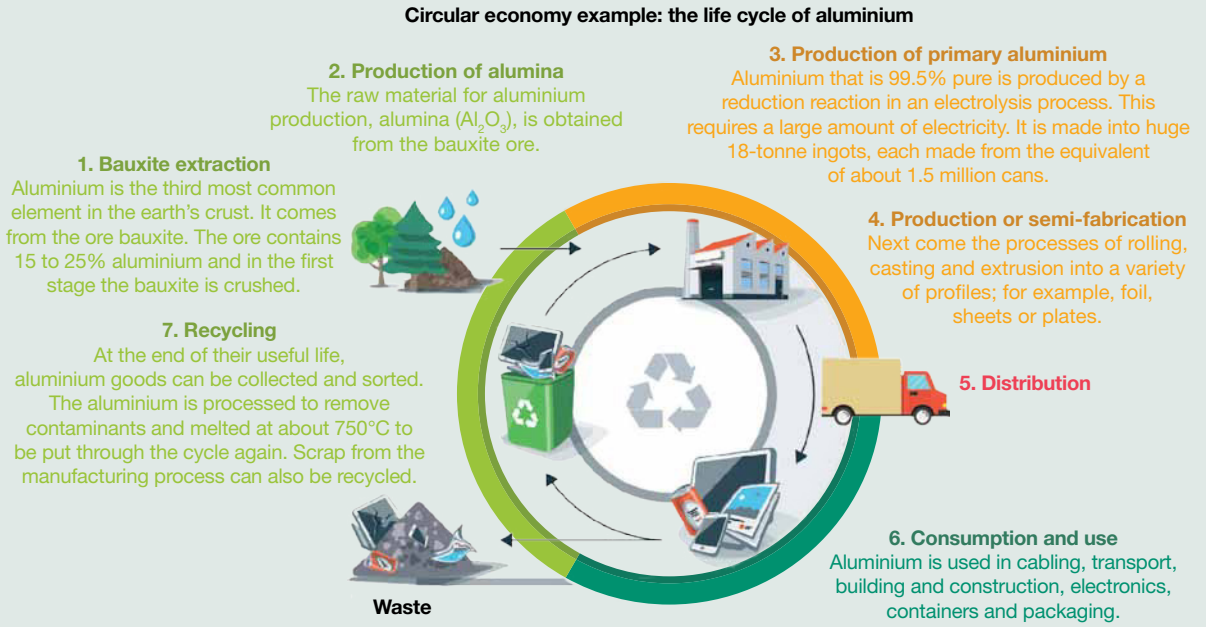
CASE STUDY: Life cycle of aluminium

An ideal metal for the circular economy is aluminium. It is lightweight, strong and corrosion resistant. Aluminium is used not just in cans, but also in construction, packaging, aviation, shipping and rail industries. The demand for aluminium is growing annually; for example, aluminium use in cars is increasing because it is lighter than steel — meaning the cars require less fuel. Recycling aluminium can help meet this increased demand. Aluminium is 100 per cent recyclable, and recycling saves 95 per cent of the energy required to manufacture the aluminium from bauxite ore. Indeed, 75 per cent of all aluminium produced is still in use, compared to 9 per cent of plastics (although plastic is cheaper).

Some other aspects that need to be considered include cost of collection, how easily products can be disassembled, and disposal of waste created in the recyling process. Good design can be beneficial. Firstly, good design can help reduce the use of unnecessary joining materials. Secondly, materials being built to last will reduce the need for recycling.

When aluminium products are melted, impurities are separated to form 'salt cake'. This not-so-delicious cake contains aluminium oxides, aluminium, carbides, nitrides, sulfides and phosphides, and is highly toxic so must be disposed of carefully. Engineers are trying to find innovative ways to use this, possibly in asphalt and cement. Energy is required for recycling and some gas pollutants are released; however, these are considerably less than in the initial production of aluminium. So make sure that used can is placed in the recycling bin!

FIGURE 3.18 Life cycle of aluminium



CASE STUDY: Titanium bicycles

Materials research has become an important component of competition preparation in many sports. Track and field athletes look for running shoes that provide stability, support and track grip. Tennis players seek the ultimate racquet that will give them a larger 'sweet spot' (the optimum position on the mesh from which to hit the ball), greater power and more control. Bicycle frames have until recently been built from steel, aluminium and carbon fibre composites. A new material has now emerged as the preferred frame material — titanium. (Titanium bike frames were first made in 1986 but only in recent years has the cost of these frames dropped sufficiently to make them a viable alternative for cyclists.)

FIGURE 3.19 Titanium mountain bike frames are also becoming popular.

Each frame material has its advantages and disadvantages, as shown in table 3.7.

Titanium is used in bicycle frames due to its combination of toughness, elasticity and resilience. Bicycles that are built stiffly, so that they resist pedalling forces, often provide a harsh, uncomfortable ride. Although titanium frames are very stiff against high pedalling forces, they seem to transmit much less road shock than bicycles made of other materials. Steel, for example, has a significantly higher density than titanium, but shock waves travel faster in steel than

in titanium. Despite the fact that titanium's high vibration dampening properties are not yet fully understood, we do know that titanium provides three things crucial to cyclists: low weight, stiffness and a smooth ride.

TABLE 3.7 Materials used in bicycle frames

Material	Advantages	Disadvantages
Aluminium	• Light • Stiff	• Low resistance to repeated stresses
Steel	• Strong • Economical • Easily shaped to required specifications	• Rusts • Heavy
Carbon fibre composites	• High strength-to-mass ratio • Good shock-dampening power • Good vibration-dampening power	• Expensive • Can't be recycled if damaged
Titanium	• Less than half the density of steel • Excellent resistance to metal fatigue • High corrosion resistance	• Expensive • Difficult to shape

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

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

3.4 Activities

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3.4 Exercise

1. Explain why some metals are found as metals and others are found combined with other elements in minerals.
2. What is an ore? Give an example of a metal and its ore.
3. MC Which of the following is not an environmental benefit of recycling metals?
   - **A.** It reduces the amount of waste created.
   - **B.** Less energy is required than extracting from an ore.
   - **C.** The cost of the metal is less.
   - **D.** More resources remain in the ground.
4. Metals are used in a variety of ways, from carrying oxygen in our blood to batteries and building bridges.
   - **a.** What are two advantages of metals to society?
   - **b.** Mining metals has serious implications for the environment. Describe how each of the following is involved in manufacturing metals and how their detrimental effects can be minimised.
     - **i.** Waste rock
     - **ii.** Raw materials
     - **iii.** Land use
5. **a.** What factors of metals and the metallic bonding model make it easier for metals to be recycled?
   b. What factors make it harder or more expensive for metals to be recycled?
6. **a.** Explain what bioleaching is.
   b. In what circumstances can it be used?
   c. What are two advantages of this process?
   d. What are two disadvantages?

3.4 Exam questions

Question 1 (1 mark)

MC When is a mineral described as an ore?

A. If a metal is present in the mineral
B. If a metal can be produced from it
C. If it is magnetic
D. If a metal can be profitably obtained from it

Question 2 (1 mark)

MC Recycling metals is often thought of as a better way to provide more metals. Which of the following is NOT an advantage of the recycling process?

A. It is more economical than extraction from the ore.
B. Fewer ecosystems are destroyed.
C. Natural resources are saved.
D. More landfill space is required.

Question 3 (3 marks)

a. Describe what 'urban mining' is. **(1 mark)**
b. Provide an example of urban mining. **(1 mark)**
c. Why is urban mining important? **(1 mark)**

Question 4 (6 marks)

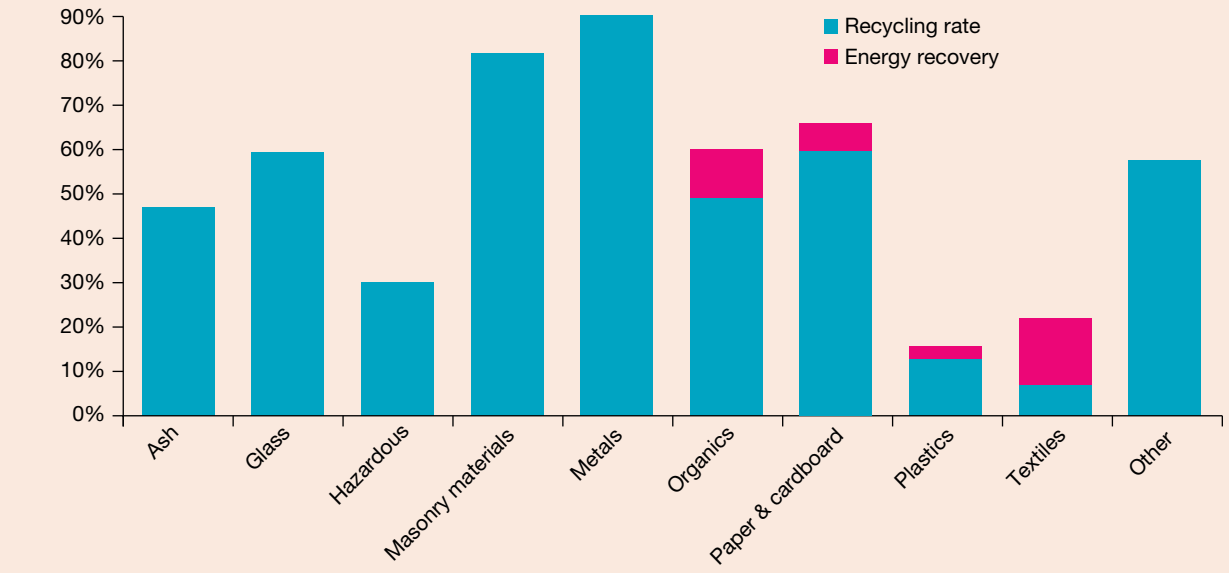
Aluminium is a widely used metal and is manufactured using electricity in the process of electrolysis from bauxite ore.

a. Using an aluminium can as an example, show flow charts indicating the difference between a linear and circular economy. **(4 marks)**
b. State two advantages of using a circular economy rather than a linear economy. **(2 marks)**

Question 5 (5 marks)

The following graph shows the recovery and recycling rates by material category in Australia, 2018–2019.

	Ash	Glass	Hazardous	Masonry materials	Metals	Organics	Paper & cardboard	Plastics	Textiles	Other
Recovery rate	47%	59%	30%	82%	90%	60%	66%	15%	22%	58%
Recycling rate	47%	59%	30%	82%	90%	49%	60%	13%	7%	58%



a. Compare the information provided for metals with that provided for plastics. **(4 marks)**
b. The main metals recycled in Australia are iron, copper and aluminium. Why do you think these metals are more likely to be recycled? **(1 mark)**

More exam questions are available in your learnON title.

3.5 Review

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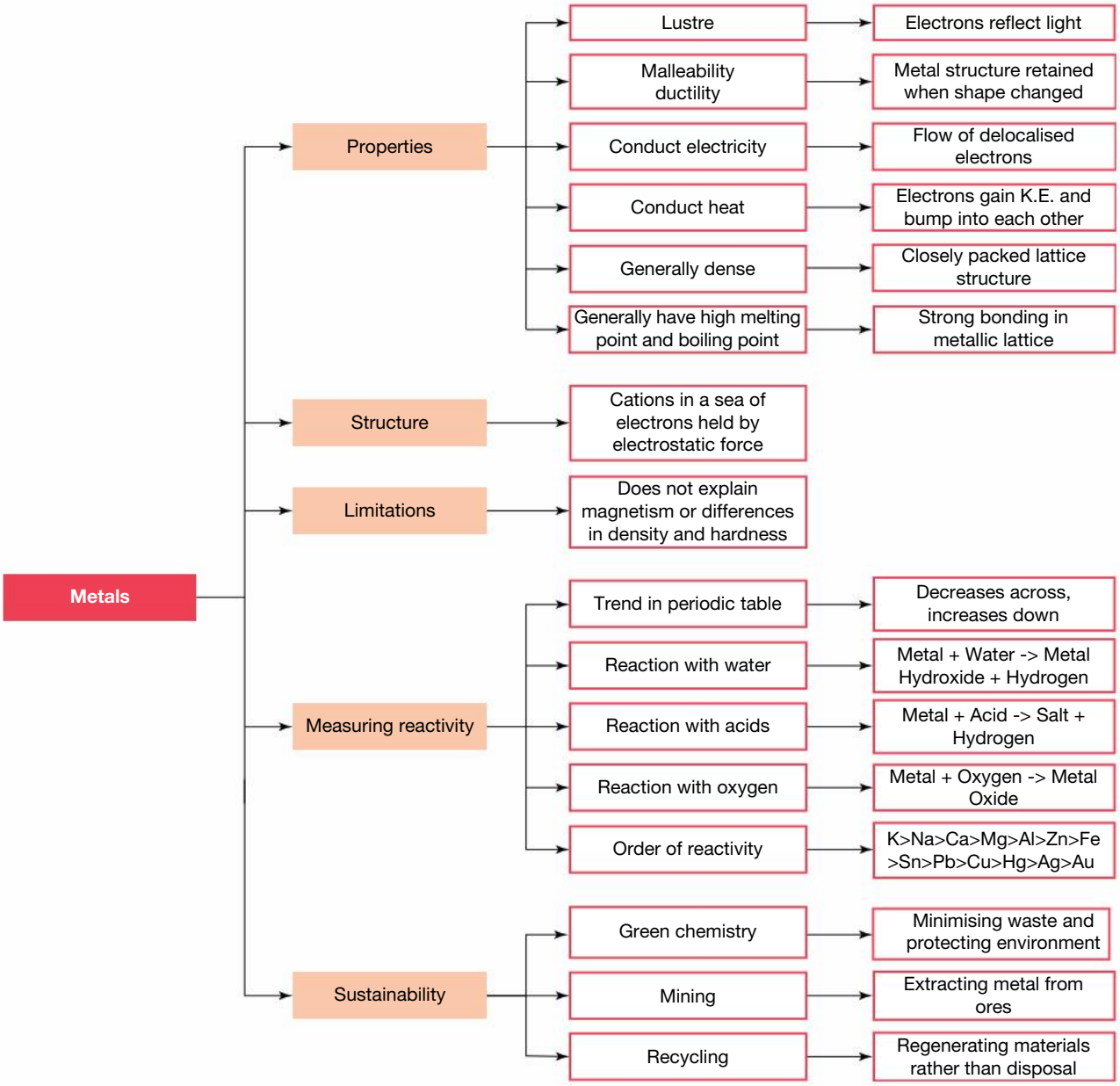

3.5.2 Key ideas summary

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3.5.3 Key terms glossary

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on Resources

Solutions	Solutions — Topic 3 (sol-0801)
Practical investigation eLogbook	Practical investigation logbook — Topic 3 (elog-1607)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 3 (doc-37069) Key idea summary — Topic 3 (doc-37070)
Exam question booklet	Exam question booklet — Topic 3 (eqb-0083)

3.5 Activities

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3.5 Review questions

1. What single feature in the metallic bonding model gives rise to the following properties: thermal conductivity, lustre, malleability and electrical conductivity?
2. Why does aluminium, a fairly reactive metal, not react with oxygen gas in air?
3. Consider the following:
   - Metal *G* slowly reacts with cold water.
   - Metal *H* is generally shiny.
   - Metal *E* readily reacts with oxygen.
   - Metal *F* produces hydrogen when it reacts with acid.

   What is their reactivity order, from the least to most reactive?

   A. $H > G > F > E$ **B.** $H > E > F > G$
   C. $F > H > G > E$ **D.** $H > F > G > E$
4. Which of the following substances contain metallic bonding?

   a. Lithium **b.** Sulfur **c.** Sodium bromide
   d. Mercury(II) fluoride **e.** Calcium **f.** Argon
5. The electron configurations of elements *A, B*, and *C* are as follows.
   - *A*: 2, 6
   - *B*: 2, 8, 1
   - *C*: 2, 8, 5

   Which of these elements is a metal?
6. Silver is a transition metal used for jewellery because it shows very high reflectivity; sodium is an alkali metal found in group 1 of the periodic table.
   a. State two differences about the chemical or physical properties of these metals.
   b. State two similarities about the chemical or physical properties of these metals.

7. Calcium is a reactive metal. When it burns in oxygen, a white powder is formed. It reacts slowly with cold water and rapidly with dilute hydrochloric acid.
   a. State the group and period of the periodic table where calcium is found.
   b. State the electron configuration of calcium.
   c. Write balanced chemical equations with states for the following.
      i. Reaction with oxygen
      ii. Reaction with cold water
      iii. Reaction with dilute hydrochloric acid
   d. How could you demonstrate the products formed in the reaction with water?
8. Describe three ways the environment is affected by mining.
9. List the steps involved in recycling metals.
10. a. Draw a flow chart showing the bioleaching process.
    b. Explain how bioleaching might be part of a circular economy for removing gold from mobile phone e-waste.

3.5 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC How can the structure and bonding in solid sodium metal at room temperature be described?

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

Question 2

MC How can the physical properties of solid metals best be explained?

A. Each metal atom is held in the crystal lattice by covalent bonds.
B. Positive metal ions are arranged in an orderly way, with delocalised electrons able to move freely throughout the crystal lattice.
C. Positive and negative metal ions are arranged in an orderly way, with delocalised electrons able to move freely through the crystal lattice.
D. Each metal atom is surrounded by a variable number of valence electrons, which complete a 'noble gas' electronic structure in the crystal lattice.

Question 3

MC What is the property of metals that enables them to conduct an electric current?

A. The outer electrons of metal atoms are not firmly bound to the atom.
B. Metal atoms are better suited as charge carriers than non-metal atoms.
C. Ions in the metal can move freely through the metallic lattice.
D. Metal atoms are not as firmly bonded to each other as are non-metallic atoms.

Question 4

MC Which is formed by the reaction of zinc with hydrochloric acid?

A. Zinc hydroxide + hydrogen
B. Zinc sulfate + chlorine
C. Zinc chloride + hydrogen
D. Zinc chloride + water

Question 5

MC What is a disadvantage of extracting aluminium from aluminium ore?

A. The aluminium ore contains many impurities that must be removed.
B. The aluminium obtained corrodes readily.
C. The aluminium obtained from the blast furnace is impure.
D. The aluminium ore is a finite resource.

Question 6

MC Place the elements magnesium, lead, calcium and silver in order from the most to least reactive.

A. Lead, potassium, magnesium, silver
B. Magnesium, potassium, lead, silver
C. Potassium, magnesium, lead, silver
D. Potassium, magnesium, silver, lead

Question 7

MC What are metals categorised as?

A. A renewable resource
B. A non-renewable resource
C. A biodegradable resource
D. A degradable resource

Question 8

MC Why are metals recyclable?

A. They retain their structure at the end of the process.
B. They are malleable.
C. They are ductile.
D. They are hard.

Question 9

MC Iron oxide reacts with carbon monoxide in a blast furnace to produce iron. What are the correct coefficients to balance the equation?

$$Fe_2O_3 + CO \rightarrow Fe + CO_2$$

A. 1, 2, 3, 2 **B.** 2, 6, 2, 3 **C.** 1, 3, 2, 3 **D.** 2, 4, 6, 4

Question 10

MC Three metals *X*, *Y* and *Z* are added individually to test tubes containing hydrochloric acid and then to test tubes containing water. The observations were recorded in the table below.

Metal	Hydrochloric acid	Water
X	Bubbles	No reaction observed
Y	Bubbles	Bubbles
Z	No reaction observed	No reaction observed

List the metals in order from least reactive to most reactive

A. X, *Y*, *Z** **B.** Z, *Y*, *X* **C.** X, *Z*, *Y* **D.** Z, *X*, *Y*

Section B — Short answer questions

Question 11 (8 marks)

Metals can be obtained from their ores in different ways. Consider the following metals.

Metal *A* is extracted from its oxide by heating with carbon.

Metal *B* occurs naturally in the ground.

Metal *C* is obtained from its oxide by the process of electrolysis.

a. What is an ore? **(1 mark)**

b. List these metals in order of decreasing reactivity, giving reasons for your answer. **(3 marks)**

c. Explain which metal would be most suitable to use for jewellery. **(1 mark)**

d. Why are metals relatively easy to recycle? **(1 mark)**

e. Give two reasons it is better to recycle these metals rather than dispose of them as waste after use. **(2 marks)**

Question 12 (8 marks)

The reactivity of metals can be determined by their reaction with water, acids and oxygen. Four metals were tested and some of the observations are as follows.

Manganese reacts readily with dilute hydrochloric acid to form $MnCl_2$ and bubbles of gas.

Barium (Ba) is in group 2 and reacts moderately with cold water and vigorously with steam.

Silver does not react with water, acid or oxygen.

Nickel (Ni) reacts slowly with sulfuric acid.

a. List the metals in order of increasing reactivity. **(1 mark)**

b. Write a balanced equation with states for the reaction of manganese and dilute hydrochloric acid. **(2 marks)**

c. In the periodic table, calcium is in group 2 period 4, and barium is in period 6. Explain if you would expect calcium to be more reactive or less reactive than barium. **(2 marks)**

d. Write a balanced equation for the reaction of barium and water. **(2 marks)**

e. Explain if you would expect silver to be found naturally or need to be extracted from its ore. **(1 mark)**

Question 13 (8 marks)

Design an experiment to see which is the more reactive metal out of zinc or lead.

a. Which reactions would you use and why? **(2 marks)**
b. State the chemicals and equipment required. **(3 mark)**
c. Describe how the reactivity is to be measured. **(1 mark)**
d. What are two safety precautions that would be necessary? **(2 mark)**

Question 14 (8 marks)

A student investigated the reactivity of three metals by reacting the metals with dilute sulfuric acid.

The results observed were recorded in the table.

Metal	**Metal *M***	**Metal *N***	**Metal *O***
Observation	A few bubbles	No observable change	Many bubbles

a. State the independent variable. **(1 mark)**
b. State the dependent variable. **(1 mark)**
c. What factors must be controlled in this experiment? **(1 mark)**
d. How do you know that sodium is not one of the metals? **(3 marks)**
e. Which metal is most readily oxidised? **(1 mark)**
f. Which metal would be most likely to show a reaction with steam? **(1 mark)**

Question 15 (8 marks)

The average temperature on Mars is about –27 °C, much lower than here on Earth because Mars is further from the sun. The atmosphere contains more than 95 per cent carbon dioxide and less than 1 per cent oxygen. Gravity on the moon is only about 38 per cent of that on Earth. Its surface is rocky, with canyons, volcanoes, dry lake beds and craters, and contains a variety of minerals. Some of the minerals found on Mars contain iron, magnesium and aluminium. An iron ore containing hematite, Fe_2O_3, is very common on Mars and is what gives Mars its familiar red colour.

a. The purity of the iron and other ores is not high. If you were given the opportunity to be part of an expedition to Mars, explain why you might take a selection of microbes from specific mining environments. What other items would be needed to assist in this process? **(2 marks)**
b. Consider the properties of iron and suggest an easy method of separating particles of iron from the leachate. **(1 mark)**
c. Native iron from meteorites can also be found. Explain what native iron might be and why it would not corrode on Mars. **(2 marks)**
d. Explain why construction using iron on Mars would be easier than using iron on Earth. **(1 mark)**
e. What is a possible advantage of extracting iron rather than extracting aluminium from Martian ores? **(2 marks)**

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4 Reactions of ionic compounds

KEY KNOWLEDGE

In this topic you will investigate:

Reactions of ionic compounds

- the common properties of ionic compounds (brittleness, hardness, melting point, difference in electrical conductivity in solid and molten liquid states) with reference to the nature of ionic bonding and crystal structure
- deduction of the formula and name of an ionic compound from its component ions, including polyatomic ions (NH_4^+, OH^-, NO_3^-, HCO_3^-, CO_3^{2-}, SO_4^{2-}, PO_4^{3-})
- the formation of ionic compounds through the transfer of electrons from metals to non-metals, and the writing of ionic compound formulas, including those containing polyatomic ions and transition metal ions
- the use of solubility tables to predict and identify precipitation reactions between ions in solution, represented by balanced full and ionic equations including the state symbols: (s), (l), (aq) and (g).

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of VCE Chemistry. Experiments and investigations, supported by a practical investigation eLogbook and teacher-led videos, are included in this topic to provide opportunities to undertake investigations and communicate findings.

EXAM PREPARATION

Access exam-style questions and their video solutions in every lesson, to ensure you are ready.

4.1 Overview

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4.1.1 Introduction

How many chemical compounds are on our planet? One estimate is that humans have discovered or made nearly 50 million. This estimate is based on a data base maintained by the American Chemical Society. When one considers that chemical compounds are all made from a mere 90 naturally occurring elements, the estimated number is staggering.

Central to this amazing fact, of, course, is the concept of chemical bonding. Two mechanisms for doing this have already been seen. These depend on whether metal atoms or non-metal atoms are joining together. Covalent bonding was discussed in topic 2 and metallic bonding in topic 3. This topic introduces a third type of bonding called ionic bonding. This occurs when metal atoms join together with non-metal atoms.

FIGURE 4.1 Pink salt lake in Western Australia. The colour is due to the presence of algae.

Ionic substances have been known and used for thousands of years. They contain some of the most familiar compounds that we come across and use each day. What is commonly called salt, but chemically called sodium chloride, is the most obvious example. To a chemist, salts are a group of compounds of which sodium chloride is merely one example. All of these are ionic in nature. Ionic compounds are also found in fluoridated toothpaste, food preservatives, gunpowder, mortar and cement, antacids, plaster, drain cleaners, washing powders and in improved salt formulas, to name just a few.

Like all the bonding discussed in the previous topics, the driving force behind ionic bonding is the desire of atoms to become more stable by obtaining a noble gas configuration for their electrons. In this topic, however, you will learn that the mechanism is different. It will involve the donation of electrons from one atom to another.

LEARNING SEQUENCE

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4.3 Formation of ionic compounds 157
4.4 Precipitation reactions 165
4.5 A review of bonding 176
4.6 Review 182

on Resources

Solutions	Solutions — Topic 4 (sol-0803)
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Exam question booklet	Exam question booklet — Topic 4 (eqb-0085)

4.2 Structure and properties of ionic substances

KEY KNOWLEDGE

- The common properties of ionic compounds (brittleness, hardness, melting point, difference in electrical conductivity in solid and molten liquid states) with reference to the nature of ionic bonding and crystal structure

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4.2.1 What are ionic substances?

As has been seen in previous topics, the electronic configuration of the atoms in an element is important in determining its reactivity. This comes from the desire of all atoms to achieve the most stable arrangement of their electrons that is possible. Often this arrangement is that of the closest noble (group 18) gas. In most cases, atoms have to join (bond) to other atoms to achieve this stability. Sometimes they might join together with atoms of the same type (for example, O_2, Fe), but in many cases the atoms are different.

The mechanism by which this stability is achieved depends on the nature of the atoms that are joining together. As has already been seen:

- If only metallic atoms are joining together, **metallic bonding** results. This can be pictured as a 'sea' of mobile, delocalised electrons that move amid a lattice of positive ions. This model explains many of the properties of metals (refer to topic 3).
- If only non-metal atoms join together, an electron sharing model is best. This leads to **covalent bonding**. This model can then be used to explain the properties of non-metallic elements and compounds (refer to topic 2).

The third possible way in which atoms may bond together is when metallic and non-metallic atoms join. This is called **ionic bonding**.

Ionic compounds have been known for thousands of years. They often form attractive crystals and have long been valued for their decorative properties and for use in jewellery. They are also among the most common and important substances on Earth. Common salt (sodium chloride) is an example of such a substance. It is important for life and as a raw material to make other substances.

metallic bonding positively charged metal cations arranged in a lattice with delocalised valence electrons being able to flow around them

covalent bonding sharing of electrons between nuclei that bonds them together in a molecule

ionic bonding bonding that forms between cations and anions

ionic compound compound containing cations and anions in an ionic lattice structure or an ionic liquid, depending on ion size

4.2.2 Properties of ionic substances

Many of the properties of ionic substances differ from the metallic and covalent substances discussed in previous topics.

Key features and properties of ionic substances include:

- They form between metal atoms and non-metal atoms. (Some exceptions to this exist, which are mentioned later in this topic.)
- They have high melting points.
- They conduct electricity in the liquid state, but not in the solid state.
- Many (but not all) are soluble in water, and the resulting solution is able to conduct electricity.
- They form crystals. Ionic crystals are hard (not easily scratched) and brittle. If a localised force is applied, they may shatter.

These properties will be explained once the nature of ionic bonding is understood.

FIGURE 4.2 The main component of the mineral fluorite is the ionic compound calcium fluoride, CaF_2. Its beautiful crystals are often cubic in shape and appear in a variety of colours.

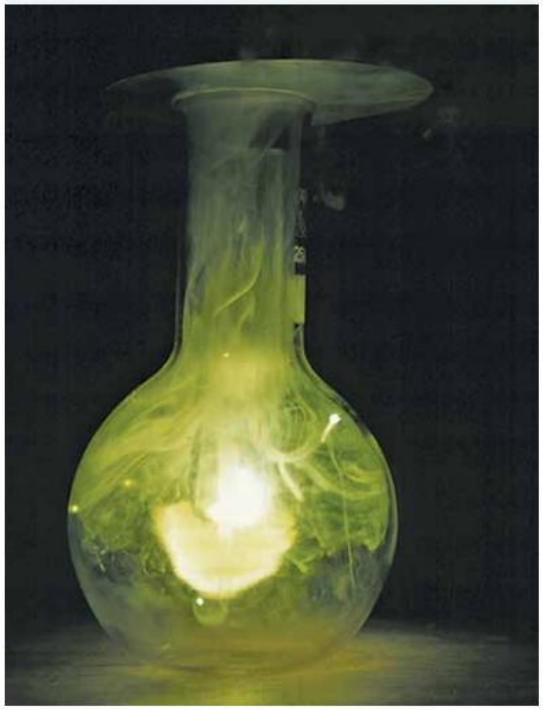
4.2.3 The nature of ionic bonding

Metals and non-metals can react together because the electrons lost by the metals can be taken up by the more electronegative non-metals. Recall that **electronegativity** is the electron-attracting power of an atom, and this transfer of electrons results in an ionic bond between the metal cations and the non-metal anions.

Consider the chemical reaction shown in figure 4.3. A small piece of sodium metal is added to a flask containing chlorine gas. The sodium ignites, and the solid ionic compound sodium chloride forms. This occurs because:

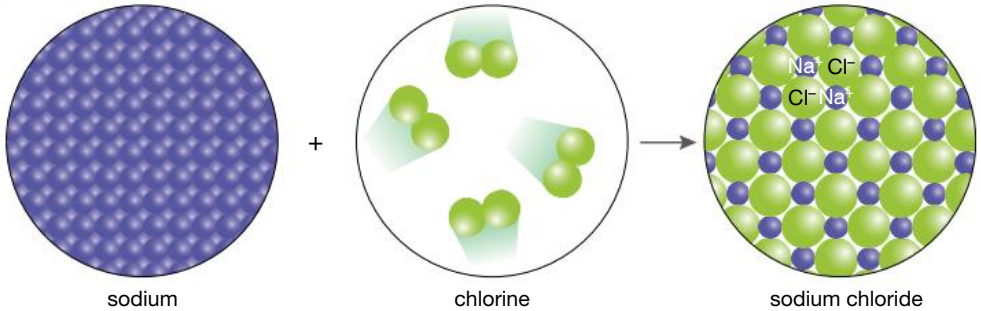
- One electron is transferred from each sodium atom to each chlorine atom. The ways in which we can represent this process are shown in the solution to sample problem 1.
- This transfer of electrons results in **ion** formation:
  - sodium atoms lose an electron and become positively charged sodium ions (**cations**)
  - chlorine atoms gain an electron and become negatively charged chloride ions (**anion**).
- These oppositely charged ions attract each other by **electrostatic attraction**. This strong force of attraction between positive and negative ions is what produces an **ionic bond**.

FIGURE 4.3 Molten sodium metal reacts violently with chlorine gas to form sodium chloride crystals.



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

FIGURE 4.4 The formation of sodium chloride from solid sodium and gaseous chlorine



electronegativity the electron-attracting power of an atom

ion an atom that has lost or gained electrons and so has a charge

cation a positively charged atom or group of atoms

anion a negatively charged atom or group of atoms

electrostatic attraction strong force of attraction between positive and negative charges

ionic bond bond between cations and anions in an ionic compound

salt ionic compound consisting of a metal ion and a non-metal ion, except oxides and hydroxides

tlvd-0516

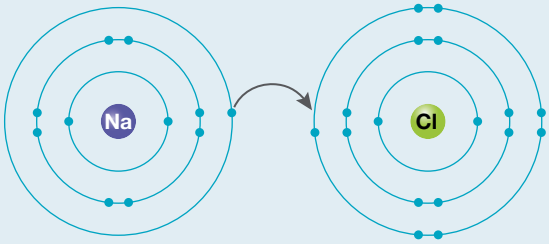
SAMPLE PROBLEM 1 Displaying ionic bonding using electron shells and equations

Consider the reaction between sodium and chlorine and explain it using the following.

a. Electron shell diagrams **b. A simple equation**

THINK

a. Sodium and chlorine can be represented as follows:

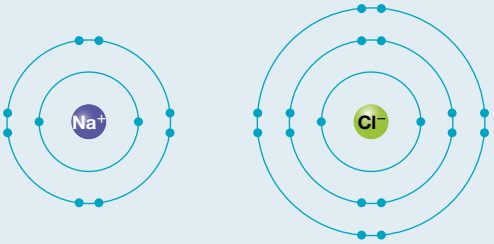


2, 8, 1 2, 8, 7

When a sodium atom forms an ionic bond with a chlorine atom, the sodium atom loses one electron (its entire outer shell) to become a positive ion, or cation. The chlorine atom gains one electron in its outer shell. It is now a negative ion, or anion. The resulting ions have a stable, noble gas configuration.

b. The sodium and chloride ions form the substance called sodium chloride.

WRITE



2, 8 2, 8, 8

$Na^+ + Cl^- \rightarrow NaCl$

PRACTICE PROBLEM 1

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

a. Magnesium oxide

b. Magnesium fluoride

Tip for drawing electron shell diagrams

Using dots and crosses is sometimes useful when drawing shell diagrams because this makes it easier to see the origin of the transferred electron(s).

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Weblinks Reaction of sodium with chlorine
Ionic and covalent bonding

4.2.4 Sodium chloride: a typical ionic compound

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

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

Ionic lattices

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

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

Sodium chloride is a typical ionic compound. Its ionic structure and appearance are shown in figures 4.5 and 4.6. Its crystals are cubic, reflecting the regular arrangement of the ions in the lattice. Each sodium ion is surrounded by six chloride ions and each chloride ion is surrounded by six sodium ions. In this arrangement, each ion is strongly attracted to each of its neighbours. The large attractive forces result in a very stable structure.

FIGURE 4.5 Sodium chloride lattice

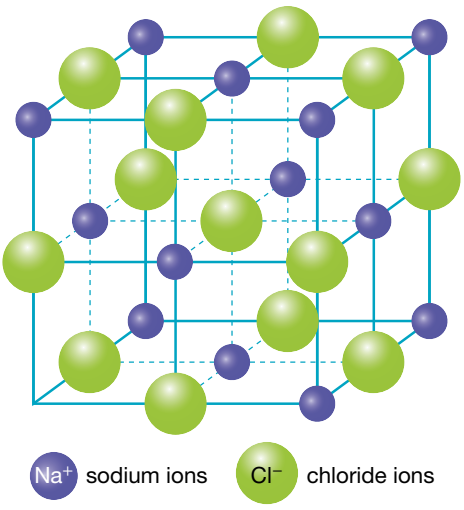

FIGURE 4.6 Each sodium ion is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions. This gives the crystals of sodium chloride their characteristic cubic shape.

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

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

empirical formula chemical formula that shows the simplest numerical ratio in which atoms are combined

EXTENSION: What affects the arrangement of ions in an ionic lattice?

When ions come together to form ionic solids, they do so in arrangements designed to:

- maximise the attractive forces between oppositely charged ions
- minimise the repulsive forces between like charged ions.

This arrangement on the atomic scale will be reflected in the shape of the ionic crystal on the macroscopic scale.

A large number of arrangements are possible. The arrangement for a particular ionic solid depends on the following:

- the relative numbers of cations and anions in the formula
- the relative sizes of the cations and anions and the physical barrier that this presents as to how many oppositely charged ions can be arranged around a particular ion.

An example of the size factor is the different arrangements for sodium chloride and caesium chloride. It might be suspected that the lattice structures of these would be similar because both are group 1 chlorides. As has been seen, the structure of sodium chloride displays 6 : 6 co-ordination. That is, each sodium ion is surrounded by six chloride ions and visa-versa. However, in caesium chloride the larger caesium ions mean that more room is available to fit more chloride ions around them. Caesium chloride displays 8 : 8 co-ordination.

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Investigating calcite crystals

Aim

To observe the nature of crystal structure using calcite crystals

on Resources

Interactivity Pass the salt (int-0675)

Weblink Animation of NaCl structure

4.2.5 Connecting properties of ionic compounds to structure

BACKGROUND KNOWLEDGE: What makes a substance conduct electricity?

For a substance to conduct electricity, it must have charged particles that are able to move. These two requirements must be met simultaneously.

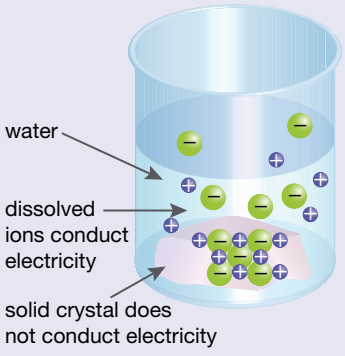
As was seen in topic 3, metals are able to conduct electricity because they have negatively charged electrons that are able to move freely through the crystal structure.

Ionic substances can conduct electricity in two ways.

They can be molten (melted), which means that the ions they are made from can move freely in the now liquid state. A second way is through dissociation. This is where the substance dissolves in water and the ions from the lattice are pulled apart by the water molecules and are able to move freely through the solution. This free movement of ions in water forms a conducting solution called an electrolyte.

Some molecular substances also conduct electricity. This is because their molecules react with water and make ions. This process is termed ionisation.

FIGURE 4.7 Dissociation of sodium chloride to form an electrolyte

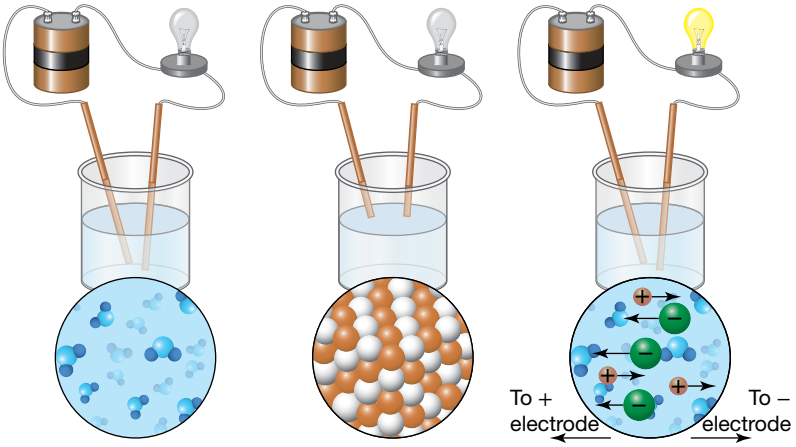


Ionic compounds are composed of two or more different kinds of ions that have opposite charges. The electrostatic forces of attraction holding them together are strong. As a result, ionic compounds:

- are usually crystalline solids, owing to the arrangement of ions in repeating three-dimensional patterns
- have high melting and boiling points, because a large amount of energy is needed to separate the ions. Each ion is held strongly in place by the oppositely charged ions that surround it.
- do not conduct electricity in the solid form, because the charged particles (ions) are not able to move. *(To conduct electricity, a substance must contain mobile charged particles.)*
- are hard, since the surface of the crystal is not easily scratched due to the strong ionic bond holding the ions together
- often dissolve in water to form ions (dissociate), because water molecules are able to move between ions and free them by disrupting the rigid crystal structure
- conduct electricity when molten or in aqueous solution (dissolved in water). In the molten form, ions are able to slide past one another and can, therefore, conduct electricity. When an ionic substance is dissolved in water, the ions separate from the lattice and can move freely to conduct an electric current. The solution is called an **electrolyte** (see figures 4.7 and 4.8).
- are brittle, since distortion of the crystal causes ions of like charge to come close together and the repulsion between these ions cleaves or shatters the crystal, as shown in figure 4.9. If the distortion of the crystal is caused by a sharp blow along the plane of ions, cleavage (a clean split) along the plane occurs. If the blow is along different planes of ions, the crystal lattice shatters.

electrolyte solution or liquid that can conduct electricity

FIGURE 4.8 Electrical properties of distilled water, solid ionic compounds and ionic solutions

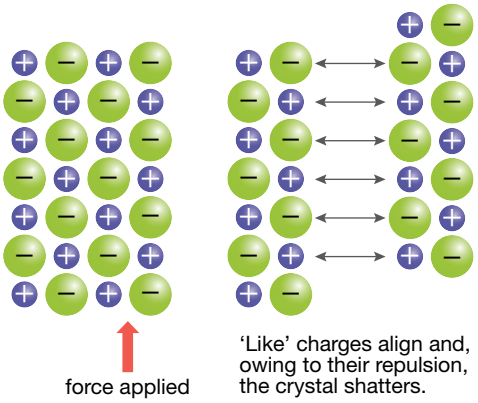


A current cannot flow through distilled water.

A current cannot flow through a solid ionic compound.

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

FIGURE 4.9 The brittle nature of an ionic compound, which cleaves or shatters when distorted.



'Like' charges align and, owing to their repulsion, the crystal shatters.

Conductivity in ionic substances

When explaining conductivity in ionic substances, remember that it is not the electrons that move. Only the ions are mobile in solution and when molten, not the electrons.

FIGURE 4.10 Old alkaline batteries often leak potassium hydroxide, which reacts with carbon dioxide in the air to form crystals of potassium carbonate. Both potassium hydroxide and potassium carbonate are ionic substances.

elog-1751

EXPERIMENT 4.2 online only

Ionic models

Aim

To construct a model of sodium chloride and use it to simulate some properties of sodium chloride

4.2.6 Limitations of ionic models

As is the case with metallic and covalent bonding, a variety of models are used to simulate the bonding and to explain the properties of the substances that result. Each model has its strengths and limitations, and this is also the case with the models used for ionic bonding.

Electron dot diagrams

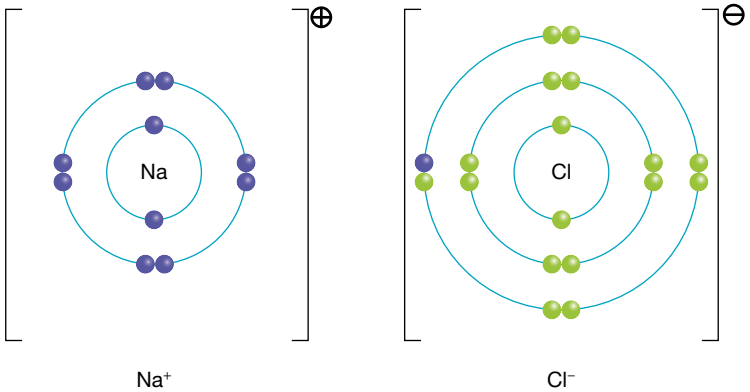
An electron dot diagram shows:

- how the ionic bonds are formed
- the ratio in which the atoms react.

However, it does not show how the ions are arranged in space.

The electron dot diagram for the sodium chloride ions shown in figure 4.11 suggests that it is made up of one pair of sodium and chloride ions. This is not the case. Sodium chloride has a giant ionic lattice structure with equal numbers of sodium and chloride ions. In a 1 g sample of pure sodium chloride, for example, will be roughly 5×10^{21} of each ion.

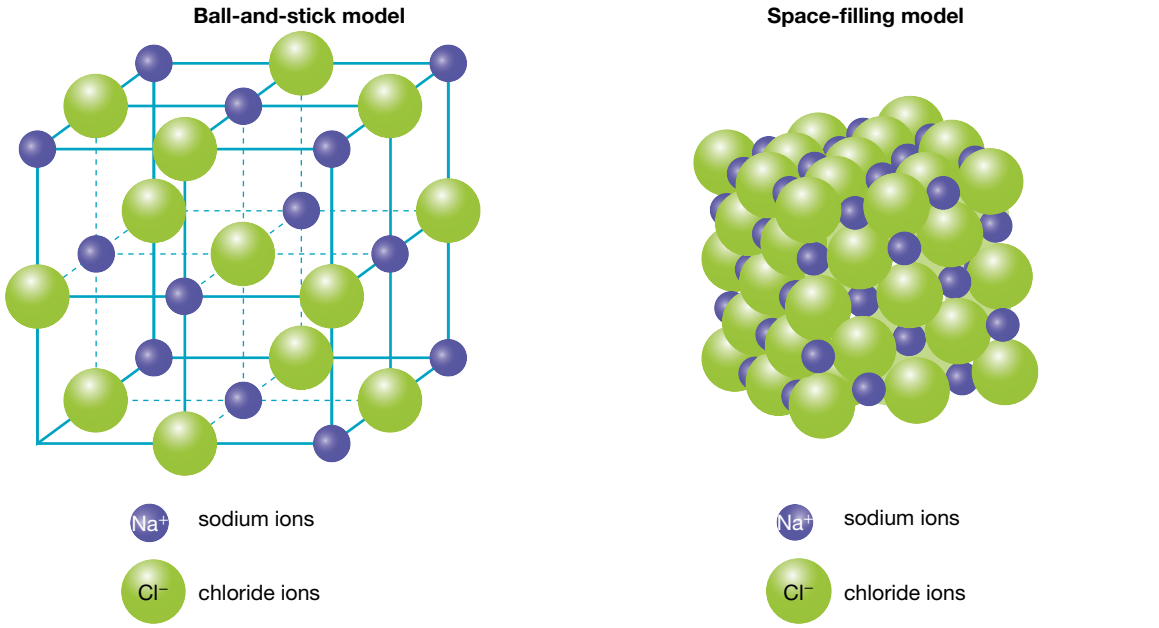
FIGURE 4.11 In electron dot diagrams, square brackets are often used to indicate that the charge is spread over the whole ion.



Three-dimensional models of ionic compounds

A three-dimensional (3D) model shows how the ions are arranged in a lattice structure. These models usually have coloured balls to represent the ions and use sticks to show the ionic bonds (the ball-and-stick model), or they can be space-filling models (see figure 4.12 for both options).

FIGURE 4.12 Three-dimensional representations of NaCl using the ball-and-stick model and space-filling model



A 3D model still has the following limitations:

- It is not to scale.
- It gives no information about the forces of attraction between the ions.
- The movement of electrons to form the ions cannot be shown.
- It only gives an indication of the relative sizes of the ions.

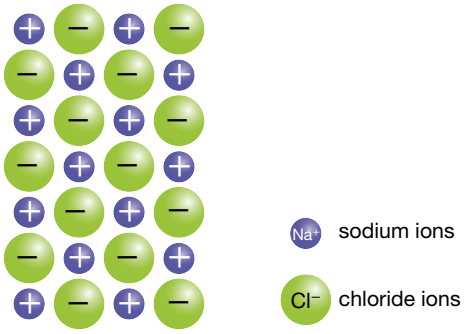
Two-dimensional models of ionic compounds

A two-dimensional (2D) representation is similar to the three-dimensional ball-and-stick model, but is an easier model to draw.

In addition to the 3D model limitations, the 2D model has further limitations:

- It only shows the arrangement of one layer of ions.
- It does not show where the ions are located on the other layers or the relationship between them.
- It does not show that many different arrangements of ions are possible.

FIGURE 4.13 A 2D model for the ionic lattice in sodium chloride



4.2.7 Uses of ionic compounds

Ionic compounds have a wide range of uses — particularly in the body, as can be seen in table 4.1. Our bodies can function only because of the presence of particular ions. Many metals are obtained from ionic compounds that have been extracted from different ores. Sodium hydrogen carbonate is used in cooking and also as an environmentally friendly cleaning agent. Ammonium nitrate is used to manufacture fertilisers and, as such, plays a crucial role in the growing of food crops. It is also used in explosives. Compounds containing nitrate, nitrite and sulfite ions are used to preserve food.

TABLE 4.1 Common ions needed in the body

Name	Symbol	Functions in the body
Sodium	Na^+	Regulates fluid balance Involved in transmission of nerve impulses Controls blood pressure by controlling blood volume Involved in acid–base balance of blood
Potassium	K^+	Controls the level of body fluids Involved in transmission of nerve impulses and muscle contraction, including heartbeat Important in reactions within cells
Calcium	Ca^{2+}	Needed for building teeth and bones Involved in blood clotting Involved in transmission of nerve impulses and muscle contraction
Magnesium	Mg^{2+}	Needed for bone formation Assists in energy production Involved in functioning of muscle and nerve tissue Component of enzymes
Chloride	Cl^-	Involved in acid–base balance of blood Assists in fluid balance in the body Involved in formation of hydrochloric acid in stomach

4.2 Activities

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4.2 Exercise

1. Explain why atoms such as Ca and Al attain a noble gas configuration and form ions.
2. Draw electron shell structures and write simple equations to show the ionic bonding between the following.
   a. Calcium and oxygen
   b. Beryllium and chlorine
   c. Lithium and fluorine
   d. Aluminium and sulfur
   e. Sodium and nitrogen
   f. Magnesium and sulfur
3. Magnesium chloride, $MgCl_2$, and potassium chloride, KCl, are typical ionic compounds.
   a. Describe how their ionic lattice structure is formed from their respective atoms.
   b. Show why their empirical formulas have ratios of 1 : 2 and 1 : 1 respectively.
   c. Explain what holds the ions close together in their lattice structures.
4. Calcium chloride is a crystalline substance at room temperature. Predict whether it
   a. has a low or high melting point.
   b. shatters when pressure is applied.
   c. conducts electricity in the solid or liquid state.
   Justify your predictions.

5. **a.** Compare the ratio of positive to negative ions in the lattice of the following.
   i. Potassium chloride
   ii. Calcium oxide
   b. Which of these compounds would you expect to have the higher melting point? Give a reason for your choice.
6. A solid substance has a high melting point, conducts electricity only in molten form and cleaves when struck with a sharp knife. However, it does not readily dissolve in water. Is this enough data to predict the chemical bonding in this substance? Explain in terms of the limitations of the bonding model.

4.2 Exam questions

Question 1 (1 mark)

MC Which of the following lists contains a compound that is *not* ionic?

A. $MgBr_2$, $AgNO_3$ and $PbCO_3$
B. ZnO, CO and MgO
C. NH_4I, AgI and PbI_2
D. $LiNO_3$, $Cu(NO_3)_2$ and $Fe(NO_3)_3$

Question 2 (1 mark)

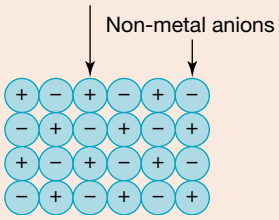
MC Which of the following properties is *not* typical of ionic compounds?

A. Brittleness
B. High melting temperature
C. Lustre
D. Hardness

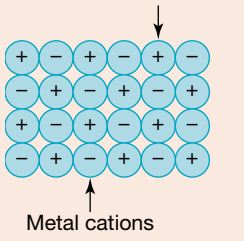
Question 3 (1 mark)

MC Which of the following diagrams represents the ionic bonding model?

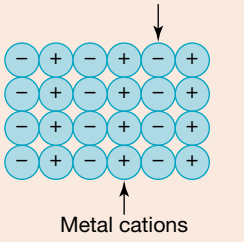
A.



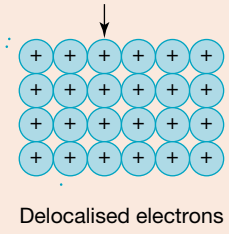
B.



C.



D.



Question 4 (2 marks)

List two factors that influence the crystal structure of an ionic compound.

Question 5 (2 marks)

Explain why an ionic solid breaks into small pieces when hit with a hammer.

More exam questions are available in your learnON title.

4.3 Formation of ionic compounds

KEY KNOWLEDGE

- The formation of ionic compounds through the transfer of electrons from metals to non-metals, and the writing of ionic compound formulas, including those containing polyatomic ions and transition metal ions
- Deduction of the formula and name of an ionic compound from its component ions, including polyatomic ions (NH_4^+, OH^-, NO_3^-, HCO_3^-, CO_3^{2-}, SO_4^{2-}, PO_4^{3-})

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4.3.1 Electron transfer and ionic compounds

As noted in the previous subtopic, ionic compounds are formed through the transfer of electrons from metal atoms to non-metal atoms. The positive and negative ions that are formed through this process are then held together by electrostatic attraction.

For example, the alkali metals, lithium, sodium and potassium, are very reactive and are normally stored under oil to keep them from reacting with the oxygen in the air (see figure 4.14). If exposed to air, the alkali metals lose electrons to form positively charged ions. These electrons are taken up by oxygen atoms, which consequently form negatively charged oxide ions. The surface of the metal quickly dulls due to the formation of metal oxides.

FIGURE 4.14 Lithium (Li), sodium (Na) and potassium (K) are highly reactive and are stored under oil. Lithium has a low density and floats on the oil.

4.3.2 Naming ionic compounds

An ionic compound is named using the ions it is made from. The positive ion is specified first, followed by the negative ion. For example, when iodine reacts with sodium, iodide ions (I^-) and sodium ions (Na^+) are formed. The resulting compound is called sodium iodide (NaI).

Chemists have developed a set of rules for naming ionic compounds, and these rules allow us to predict their formulas. You will notice that metal ions retain the name of the element they are formed from. For simple non-metal ions (ions made from just one atom) the suffix *-ide* replaces the -ine ending of the element it is formed from. Some compounds that have been known for many years also have common names, as listed in table 4.2.

TABLE 4.2 Common ionic compounds

Common name	Systematic name	Formula	Used to manufacture
Salt	Sodium chloride	NaCl	Chlorine
Soda ash	Sodium carbonate	Na_2CO_3	Glass
Baking soda	Sodium hydrogen carbonate	$NaHCO_3$	Cake raising agent
Lime	Calcium oxide	CaO	Mortar
Limestone	Calcium carbonate	$CaCO_3$	Cement
Potash	Potassium nitrate	KNO_3	Gunpowder
Milk of magnesia	Magnesium hydroxide	$Mg(OH)_2$	Laxative or antacid
Gypsum	Calcium sulfate	$CaSO_4$	Plaster
Caustic soda	Sodium hydroxide	NaOH	Soap or drain cleaner

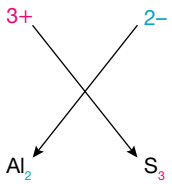
FIGURE 4.15 Sodium hydrogen carbonate, commonly known as baking soda

4.3.3 Formulas of binary ionic compounds

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

- Write the symbol for the cation first, followed by the symbol for the anion.
- Determine the lowest whole number ratio of ions that provides a net charge of zero.
- The formula is then written using the simplest whole number ratio as subscripts.

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



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

binary ionic compound describes ionic compounds made up of only two elements

tlvd-0517

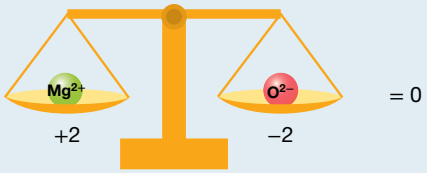
SAMPLE PROBLEM 2 Writing formulas for ionic compounds

Write formulas for the following compounds.

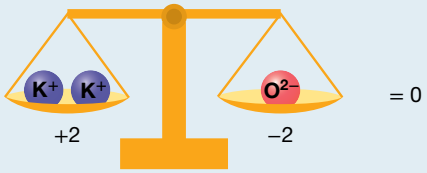
a. Magnesium oxide **b. Potassium oxide** **c. Aluminium sulfide**

THINK

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



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



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

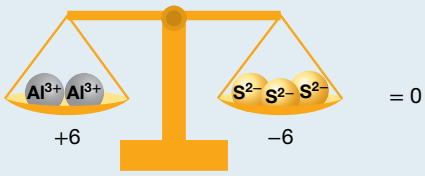
WRITE

MgO

K_2O

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

Al_2S_3



PRACTICE PROBLEM 2

Name the compound SrF_2.

on Resources

Weblink Balanced ionic compounds quiz

4.3.4 Transition metals ions: ions with a variable charge

As the atomic number increases, atoms become larger and have more electrons. Some transition metals have quite complicated arrangements of electrons and may be able to form more than one type of ion. Such ions have different charges. This is very significant because ions of the same element with different charges have different physical and chemical properties.

For example, a solution of Cr(III) is yellow, whereas a solution of Cr(VI) is orange (see figure 4.16).

FIGURE 4.16 When chromium reacts with other chemicals, often Cr(III) (chromium in the +3 oxidation state; yellow beaker) changes to Cr(VI) (orange beaker). The colour change can indicate that a reaction has occurred.

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

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

tlvd-0518

SAMPLE PROBLEM 3 Writing formulas for compounds with ions with a variable charge

a. Name the compound CuO.

b. Give the formula for iron(II) chloride.

THINK	WRITE
a. In order to determine the charge on the metal ion, we need to work backwards from the known charge on the anion. The charge on the oxide ion is 2–. Therefore, for a net charge of zero, the charge on the Cu ion must be 2+. The compound is, therefore, named copper(II) oxide.	Copper(II) oxide
b. Iron(II) means that the Fe^{2+} ion is present in the compound. The chloride ion has a charge of 1–. For a net charge of zero, the 2+ of the Fe ion must be balanced with two chloride ions.	$FeCl_2$

PRACTICE PROBLEM 3

a. Name the compound of SrI_2.

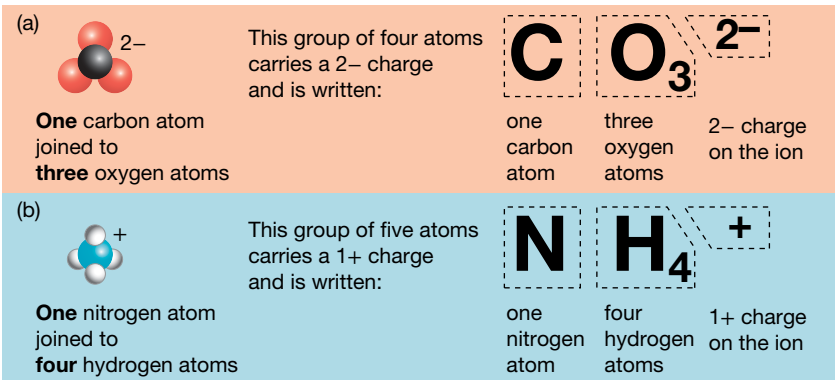
b. Give the formula for vanadium(V) oxide.

4.3.5 Polyatomic ions

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

polyatomic ion charged ion composed of two or more atoms

FIGURE 4.17 Two polyatomic ions: (a) the anion carbonate (CO_3^{2-}) and (b) the cation ammonium (NH_4^+).



Ionic compounds consist of arrays of cations and anions. A table of charges may be used to help write the empirical formula of an ionic compound (see table 4.3).

TABLE 4.3 Common anions and cations

Cations					
1^+		2^+		3^+	
Lithium	Li^+	Magnesium	Mg^{2+}	Aluminium	Al^{3+}
Sodium	Na^+	Calcium	Ca^{2+}	Chromium(III)	Cr^{3+}
Potassium	K^+	Barium	Ba^{2+}	Iron(III)	Fe^{3+}
Caesium	Cs^+	Iron(II)	Fe^{2+}		
Silver	Ag^+	Nickel	Ni^{2+}		
Copper(I)	Cu^+	Chromium(II)	Cr^{2+}		
Ammonium	NH_4^+	Copper(II)	Cu^{2+}		
		Zinc	Zn^{2+}		
		Tin(II)	Sn^{2+}		
		Lead(II)	Pb^{2+}		
		Manganese(II)	Mn^{2+}		
		Mercury(II)	Hg^{2+}		
		Strontium	Sr^{2+}		
Anions					
1^-		2^-		3^-	
Hydride	H^-	Oxide	O^{2-}	Nitride	N^{3-}
Fluoride	F^-	Sulfide	S^{2-}	Phosphide	P^{3-}
Chloride	Cl^-	Sulfate	SO_4^{2-}	Phosphate	PO_4^{3-}
Bromide	Br^-	Carbonate	CO_3^{2-}		
Iodide	I^-	Sulfite	SO_3^{2-}		
Hydroxide	OH^-	Dichromate	$Cr_2O_7^{2-}$		
Nitrate	NO_3^-	Chromate	CrO_4^{2-}		
Hydrogen carbonate	HCO_3^-	Thiosulfate	$S_2O_3^{2-}$		
Hydrogen sulfate	HSO_4^-	Hydrogen phosphate	HPO_4^{2-}		
Chlorate	ClO_3^-				
Hydrogen sulfite	HSO_3^-				
Nitrite	NO_2^-				
Permanganate	MnO_4^-				
Hypochlorite	OCl^-				
Dihydrogen phosphate	$H_2PO_4^-$				
Cyanide	CN^-				

Tips for remembering anions

- Anions with *-ite* endings have one fewer oxygen atom than *-ate*, but the same charge (e.g. nitrite, NO_2^-, and nitrate, NO_3^-).
- Anions with *-ide* endings mean a single atom (e.g. sulfide, S^{2-}, versus sulfite, SO_3^{2-}, and sulfate, SO_4^{2-}).
- Adding a hydrogen decreases the negativity of the charge by one (e.g. sulfate, SO_4^{2-} and hydrogen sulfate, HSO_4^-).

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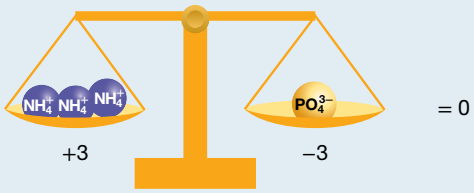
SAMPLE PROBLEM 4 Writing formulas for compounds with polyatomic ions

Write the formula for the compound ammonium phosphate.

THINK

1. The method for writing the formula for this compound is the same as that for a binary ionic compound. Write the symbol for the cation followed by the symbol for the anion.
2. Since the net charge must be zero, we need three NH_4^+ ions for every PO_4^{3-} ion.



Place brackets around the entire ammonium ion so that the subscript '3' applies to all of it.
Note that brackets need to be used whenever more than a single polyatomic ion is needed to balance a formula. This is the only time they are used.

WRITE

Cation: NH_4^+
Anion: PO_4^{3-}

$(NH_4)_3PO_4$

PRACTICE PROBLEM 4

Write the formula for the compound sodium dichromate.

In sample problem 4, no metal atoms are present. This is an example of the exception to the metal/non-metal rule for ionic bonding. Because the structure of ammonium phosphate is held together by the attractions between positively charged and negatively charged ions, it is still regarded as an example of ionic bonding.

4.3.6 Hydrated ionic compounds

Hydrated ionic compounds also contain water molecules bonded within the crystal.

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

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

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

These compounds are discussed further in topic 16.

hydrate any compound that absorbs water in the form of H_2O molecules into their chemical structure

4.3 Activities

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4.3 Quick quiz on | 4.3 Exercise | 4.3 Exam questions

4.3 Exercise

1. Write the formula for the ionic compound formed between each of the following sets of ions.
   - **a.** Aluminium and chlorine
   - **b.** Barium and oxygen
   - **c.** Sodium and sulfur
   - **d.** Magnesium and phosphorus
2. **a.** Complete the following table by writing the formulas for the compounds formed when each cation is bonded to each anion.

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

   b. Name the nine compounds formed in part **a**.
3. Name the following compounds.
   a. KCl **b.** Ag_2S **c.** Na_3N
4. Name the following compounds.
   a. FeS **b.** Fe_2S_3 **c.** $CuCl_2$
5. Name the following compounds.
   a. SnO **b.** Cu_3P **c.** $PbBr_2$
6. Give the formula for each of the following compounds.
   - **a.** Tin(IV) fluoride
   - **b.** Lead(II) sulfide
   - **c.** Mercury(II) oxide
   - **d.** Iron(III) nitride
   - **e.** Copper(I) sulfide
   - **f.** Tin(II) oxide
7. Refer to table 4.3. Write the formula and name of each of the polyatomic ions that contains sulfur.
8. Using the charges provided in table 4.3, complete the following table.

Name of ionic compound	Valency of cation	Valency of anion	Empirical formula
Silver chloride			
Potassium sulfide			
Magnesium oxide			
Aluminium bromide			
Iron(III) carbonate			
Barium phosphate			
Ammonium sulfate			

9. Give the formulas for the following compounds.
   a. Sodium sulfite
   b. Ammonium nitrite
   c. Copper(II) hydrogen carbonate
10. Name the compound formed when each of the following pairs of ions is bonded, and write its formula.
   a. Al^{3+} and CO_3^{2-}
   b. Sodium and nitrate
   c. Hg^{2+} and PO_4^{3-}
   d. Lead(II) and sulfate

4.3 Exam questions

Question 1 (1 mark)

MC Potassium dichromate has the formula $K_2Cr_2O_7$. If the dichromate ion forms a salt with an unknown metal, *M*, with the formula $M_2(Cr_2O_7)_3$, then what would *M* also form?

A. MCO_3
B. MO_2
C. MPO_4
D. M_2SO_4

Question 2 (1 mark)

MC In the ionic compound with the formula $Os(SO_4)_2$, what does the osmium ion have a charge (or oxidation state) of?

A. +2
B. –2
C. +4
D. +8

Question 3 (1 mark)

MC Ammonium molybdate has the formula $(NH_4)_2MoO_4$. The molybdate ion forms another salt with an unknown metal, *M*, with the formula $M(MoO_4)_2$. What is the charge on the metal *M* ion?

A. M^+
B. M^{2+}
C. M^{3+}
D. M^{4+}

Question 4 (8 marks)

Answer the following.

a. Write the formulae for the following ionic compounds.
   i. Sodium carbonate **(1 mark)**
   ii. Barium sulfate **(1 mark)**
   iii. Iron(III) carbonate **(1 mark)**
   iv. Ammonium nitrate **(1 mark)**
   v. Copper(I) sulfide **(1 mark)**

b. Write the names for the following ionic compounds.
   i. $Ca_3(PO_4)_2$ **(1 mark)**
   ii. VBr_3 **(1 mark)**
   iii. $FeSO_4$ **(1 mark)**

Question 5 (3 marks)

a. What are the components of a lattice of calcium fluoride? **(1 mark)**
b. Pick one of the components from part **a** and determine the following.
   i. What will it have as its closest neighbour in the lattice structure? **(1 mark)**
   ii. What will it have as its second nearest neighbour in the lattice structure? **(1 mark)**

More exam questions are available in your learnON title.

4.4 Precipitation reactions

KEY KNOWLEDGE

- The use of solubility tables to predict and identify precipitation reactions between ions in solution, represented by balanced full and ionic equations including the state symbols: (s), (l), (aq) and (g)

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4.4.1 When substances dissolve

To understand what precipitation reactions are, it is first necessary to understand the process by which ionic substances dissolve and the solutions that they form.

Dissolving occurs when one substance is pulled apart by another or moves into the spaces within another.

- The **solute** is the substance that is dissolved and the **solvent** is the liquid that the solute is dissolved in.
- Together the solute and the solvent form a homogeneous solution. This means that it has uniform composition and properties throughout the whole solution. If the solvent is water, the solution is known as an **aqueous** solution. When salt is dissolved in water, for example, an aqueous solution is formed. Our survival is dependent on reactions that occur in aqueous solutions inside and outside of the cells in our body.

FIGURE 4.18 When salt is dissolved in water, an aqueous solution is formed.

Solubility is the *extent* to which a solute dissolves in a solvent. This can be measured in quantitative terms and this is discussed further in topic 13. It can also be described in qualitative terms as follows:

- A **soluble** substance is one that readily dissolves in a particular solvent.
- An **insoluble** substance is one that does not dissolve.

These terms are sometimes further modified by descriptions such as *moderately soluble* or *sparingly soluble*.

The most common solvent is water and water is particularly good at dissolving ionic compounds. This is due to the polar nature of the bonds within water. These areas of charge are able to form attractions with the ions in an ionic lattice and this helps to pull them apart. The positive ions are attracted to the slightly negative end of the water molecules, and the negative ions are attracted to the slightly positive end of the water molecules. These interactions are called **ion–dipole bonds**. This separation from the lattice is called **dissociation** and can be represented by equations such as the following:

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

While many ionic substances dissolve well in water, some do not (i.e. they are insoluble). Whether or not a substance dissolves depends on the balance between three different forces and the energy changes involved. In the case of ionic substances dissolving in water, these are:

- the forces between the water molecules themselves
- the forces between the ions in the solid lattice structure
- the forces between the water molecules and the ions in the lattice.

For some substances, the energy required to break the bonds in the ionic lattice is too great and dissociation does not occur. Such substances are insoluble. In silver chloride, AgCl, for example, not enough energy is released when the ion–dipole interactions are formed to provide the energy needed to break the bonds between the ions.

solute dissolved substance in a solution

solvent substance in which the solute is dissolved

aqueous solution with water as the solvent

solubility grams of a substance that will dissolve in an amount of solvent at a particular temperature

soluble will dissolve

insoluble will not dissolve

ion–dipole bond bond between an ion and a polar molecule

dissociation the process by which ions separate when an ionic substance dissolves

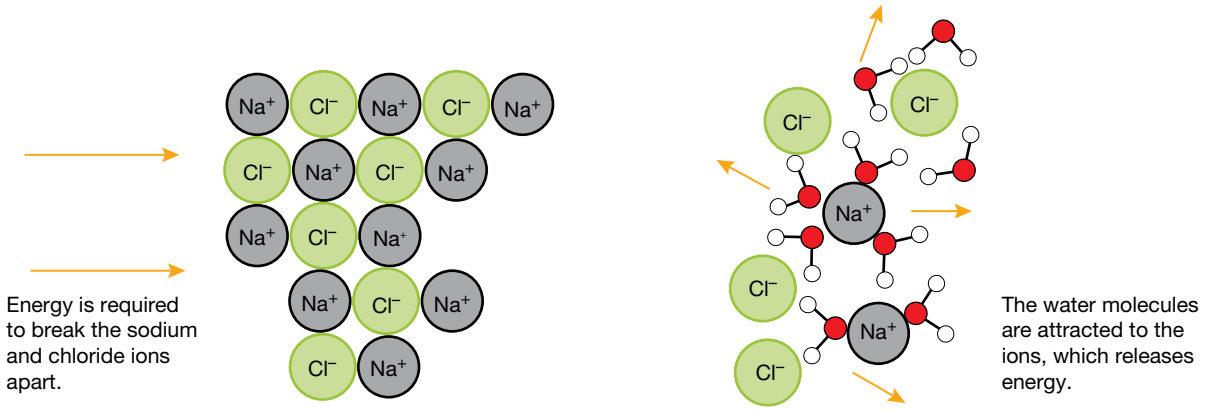
It should be noted that ionic solids are not the only substances that can dissolve in water. Many polar molecular solids can as well, as do a number of liquids and gases. Some of these also react with water to form ions. Acids and bases are a common example of such substances. For example, hydrogen chloride reacts with water as shown by the following equation:

$$HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

The solution formed is more commonly known as hydrochloric acid. This process is called **ionisation**.

ionisation a process in which ions are formed

FIGURE 4.19 Sodium chloride dissolving in water



Why water is a good solvent

Water is a good solvent for many ionic substances due to the polarity of its bonds.

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SAMPLE PROBLEM 5 Describing the process when ionic compounds dissolve

Describe the process by which potassium chloride dissolves in water and give an equation for this process.

THINK	WRITE
1. Determine the type of bonding present. Because it is an ionic compound, bonds will be broken between ions.	Potassium chloride is an ionic compound because it consists of a metal and non-metal. Potassium ions and chloride ions are dissociated by the polar water molecules.
2. Interactions occur between the ions and water, and the ions present become hydrated.	The positive potassium ions are attracted to the negative ends of the water molecules by ion–dipole interactions. The negative chloride ions are attracted to the positive ends of the water molecules by ion–dipole interactions.
3. The ions are surrounded by water molecules.	The ions are separated and hydrated.
4. Write the equation.	$KCl(s) \xrightarrow{H_2O} K^+(aq) + Cl^-(aq)$

PRACTICE PROBLEM 5

Determine the type of bonding in rubidium bromide and hence describe the process by which it dissolves in water.

EXTENSION: Ionic compounds and world food production

One of the largest and most important worldwide uses of ionic compounds is as fertilisers. While a number of important fertilisers are organic in nature, the extensive growth of plant food production over the past century has been driven by artificially produced, often ionic compounds. Cultivated croplands across the world are essential to feeding the world's population.

Plants need a number of nutrients to grow and thrive. Some of these are required in only minor amounts (micronutrients), while others such as nitrogen, phosphorus and potassium are required in much larger amounts (macronutrients). Unfortunately, the worldwide demand for food is such that natural sources are insufficient to meet these needs, resulting in fertilisers being critical to this supply. Table 4.4 lists some of the most commonly used fertilisers and the compounds they contain.

TABLE 4.4 Some common ionic compounds used in fertilisers

Nutrient supplied	Compound	Formula
Nitrogen	Ammonium nitrate	NH_4NO_3
	Ammonium sulfate	$(NH_4)_2SO_4$
Phosphorous	Calcium dihydrogenphosphate	$Ca(H_2PO_4)_2$
Potassium	Potassium chloride	KCl
	Potassium sulfate	K_2SO_4
	Potassium carbonate	K_2CO_3
Multiple	Ammonium phosphate	$(NH_4)_3PO_4$
	Ammonium monohydrogenphosphate	$(NH_4)_2HPO_4$
	Ammonium dihydrogenphosphate	$NH_4H_2PO_4$
	Potassium nitrate	KNO_3

While the industrial preparation and use of artificial fertilisers has been a boon for agriculture, they are not without their problems. Such problems include the following:

- *Run-off from excess fertiliser:* if plants cannot use the applied fertiliser quickly enough, water soluble components can dissolve and run off into rivers and lakes. This is especially true for nitrogen-containing components. The level of fertiliser that accumulates in these bodies of water can build up to the level where it stimulates out of control growth of plants such as algae. These plants then use up much of the dissolved oxygen in the water, resulting in the death of fish and other aquatic species. This is called *eutrophication*.
- *Increased production of greenhouse gases from excess fertiliser:* microbes in the soil can break down excess fertiliser if it remains in the soil for too long a period of time and produce gases that add significantly to the enhanced greenhouse effect.
- *Problematic production processes:* chemicals such as sulfuric acid, nitric acid and ammonia are used extensively in the manufacture of artificial fertilisers. These chemicals are currently made using processes that are both energy-intensive and contribute significantly to the enhanced greenhouse effect.

The United Nations has the use of climate friendly, sustainable food production methods as one of its 17 Sustainable Development goals. The challenge is to maintain and improve current crop yields while minimising and even removing the disadvantages mentioned. Current responses to this situation are focused on the following:

- Developing procedures and policies that encourage producers to use less amounts of fertiliser without sacrificing crop yields. This might be done by evaluating the best time to apply fertiliser and developing more slow-release forms. The role of artificial intelligence — for example, using sensors to monitor how quickly nutrients are being absorbed by plants — is also being investigated. And governments are moving to enact polices to disincentivise farmers using more and more fertiliser and instead encourage them to do the opposite.
- Using the principles of green chemistry to manufacture the fertilisers themselves in a more sustainable and environmentally friendly fashion. Already significant progress in this area is being made with respect to the production of ammonia and nitric acid.

4.4.2 What are precipitation reactions?

Precipitation reactions occur when two solutions combine and a compound of low solubility in water is formed. This compound is formed by a new combination of ions from the solutions that were mixed. In nearly all such cases, some ions will also be leftover that remain in solution. These are called **spectator ions**. The low solubility compound forms as solid particles that eventually settle. It is called a **precipitate**. To predict whether a precipitate will form, we must know which substances are soluble in water and which substances are insoluble.

FIGURE 4.20 Precipitate of lead iodide

Precipitation reactions

Precipitation occurs when two solutions combine and a new compound of low solubility in water is formed.

Visual evidence is the most obvious sign that a precipitation reaction has occurred. However, not all precipitates look the same. Most are a milky white colour, although examples of coloured precipitates are also possible, such as the lead iodide shown in figure 4.20. The nature of the precipitate can also vary. In some cases, it will be obvious that a solid has formed due to the solid particles being relatively large and quickly settling under gravity to the bottom. However, in other cases the solid particles are very small and may settle very slowly, or even remain in the bulk of the liquid as a suspension. Such precipitates will produce a 'cloudy' appearance that will be difficult to see through. These precipitates can often be forced to settle by using a centrifuge.

Another type of precipitate is a gelatinous precipitate. These form when water molecules get trapped in the precipitate as it forms, leading to a 'jelly-like' appearance in the precipitate.

In most cases, a precipitate can be separated from the mixture by filtration, using a funnel, filter paper and a conical flask (see figure 4.22).

The precipitate remains in the filter paper and is referred to as the **residue**, and the clear solution in the conical flask is the **filtrate** and contains the spectator ions. The filter paper with the residue can be placed in an oven at 110 °C to evaporate the water.

precipitation reaction reaction where two solutions are mixed and an insoluble solid is formed

spectator ion ion that is present but takes no part in a reaction

precipitate solid with low solubility formed in a liquid

residue substance remaining in filter paper after the process of filtration

filtrate liquid that has passed through a filter

FIGURE 4.21 A centrifuge can be used to make a precipitate settle faster.

FIGURE 4.22 Filtration apparatus (funnel, filter paper and a conical flask) used to separate precipitate from solution

4.4.3 Rules for solubility

Chemists use solubility tables such as table 4.5 to predict whether or not a compound is soluble. If such a compound is formed when solutions are mixed, it will form a precipitate. For example, if solutions of sodium nitrate and lead(II) chloride are mixed, sodium ions might combine with chloride ions to make sodium chloride, and lead(II) ions might combine with nitrate ions to make lead(II) nitrate. However, both sodium chloride and lead(II) nitrate are soluble according to table 4.5. Therefore, no precipitate will be present and all ions will remain in solution. However, if sodium nitrate is replaced by sodium sulfate, the possible new combinations become sodium chloride and lead(II) sulfate. Because lead(II) sulfate is insoluble, we can predict that a precipitate will form.

TABLE 4.5 Solubility in water of compounds of common ions

Name of ion	Symbol	Soluble compounds of ion	Insoluble compounds of ion
Group I ions	Li^+, Na^+, K^+, Rb^+, Cs^+, Fr^+	All	None
Ammonium	NH_4^+		
Hydrogen	H^+		
Nitrate	NO_3^-		
Nitrite	NO_2^-		
Chlorides Bromides Iodides	Cl^- Br^- I^-	Most	Ag^+, Pb^{2+}, Hg^{2+} ($PbCl_2$ is moderately soluble in hot water.)
Sulfates	SO_4^{2-}	Most	Ba^{2+}, Pb^{2+} (Ag_2SO_4 and $CaSO_4$ are slightly soluble.)
Carbonates Phosphates	CO_3^{2-} PO_4^{3-}	Na^+, K^+, NH_4^+	Most
Sulfides	S^{2-}	Na^+, K^+	Most (MgS, CaS, BaS, Al_2S_3 and Fe_2S_3 decompose in water.)
Hydroxides	OH^-	Na, K^+, Ba^{2+} (NH_4OH and $(NH_4)_2O$ do not exist as solids.)	Most ($Ca(OH)_2$ is slightly soluble.)
Oxides	O^{2-}		

Note: most of the ionic compounds not listed here may be assumed to be insoluble in water.

Tips for remembering soluble and insoluble ions

One easy way of remembering some common soluble and insoluble compounds is to use the SNAPE and CHOPS initials.

Compounds containing these ions are soluble	
Sodium	Na^+
Nitrate	NO_3^-
Ammonium	NH_4^+
Potassium	K^+
Ethanoate	CH_3COO^-

Compounds containing these ions are usually insoluble unless combined with SNAPE ions	
Carbonate	CO_3^{2-}
Hydroxide	OH^-
Oxide	O^{2-}
Phosphate	PO_4^{3-}
Sulfide	S^{2-}

Another mnemonic to remember the solubility rules uses the following barbecue-themed phrases: ***A Great SNAG***, ***HAPpy***, ***Streaky BaCan*** (bacon) and ***CHOPS***.

(Almost) ALWAYS SOLUBLE — A Great SNAG

Compounds with these are always soluble, but group 17 and sulfates have some exceptions.

Mnemonic	Component	Exceptions
A	**A**mmonium	
Great	**G**roup 17	**1**
S	**S**ulfates	**1**, **2**
N	**N**itrates	
A	**A**cetates (ethanoates)	
G	**G**roup 1	

Exception 1: HAPpy

Mnemonic	Component
H	**H**g: Mercury
A	**A**g: Silver
P	**P**b: Lead

Exception 2: Streaky BaCan

Mnemonic	Component
Streaky	**St**rontium
Ba	**Ba**rium
Can	**Ca**lcium

(Almost) ALWAYS INSOLUBLE — CHOPS

See the CHOPS table provided. Compounds with these are always insoluble, unless combined with A Great SNAG.

tlvd-3066

SAMPLE PROBLEM 6 Determining the solubility of compounds

Use table 4.5 to determine whether $PbSO_4$ is soluble.

THINK	WRITE
1. Identify the ions present.	Pb^{2+} and SO_4^{2-}
2. Refer to table 4.5, looking at the symbols column, and determine if either ion is listed.	SO_4^{2-} is listed.
3. Identify solubility and check the table for exceptions.	Compounds containing SO_4^{2-} are mostly soluble. $PbSO_4$ is an exception.
4. Answer the question.	$PbSO_4$ is insoluble.

PRACTICE PROBLEM 6

Use table 4.5 to determine whether $MgCl_2$ is soluble.

on Resources

Interactivity Common precipitation reactions (int-6358)

4.4.4 Full equations for precipitation reactions

As described in section 4.4.1, when a salt is dissolved in water it breaks up, or dissociates, into its constituent ions, which then become hydrated. For example:

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

The symbol (aq) indicates that each ion is separated and has become surrounded by water molecules. Silver nitrate is another salt that is soluble in water.

$$AgNO_3(s) \xrightarrow{H_2O} Ag^+(aq) + NO_3^-(aq)$$

If these two solutions are combined, an insoluble compound, a precipitate, is formed. By examining the solubility table, it becomes apparent that the precipitate observed must be silver chloride. The **full equation** can, therefore, be written:

$$NaCl(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)$$

That is, the Ag^+ ions combine with the Cl^- ions to form solid AgCl while the Na^+ and NO_3^- ions remain in solution. When equations for precipitation reactions are written, the substance that forms the precipitate is signified by using the solid symbol of state (s). Na^+ and NO_3^- ions are referred to as spectator ions because they do not take part in the reaction, as shown in figure 4.23.

full equation equation that includes all of the reactants and products that are present

FIGURE 4.23 A silver chloride precipitate formed from mixing sodium chloride and silver nitrate solutions

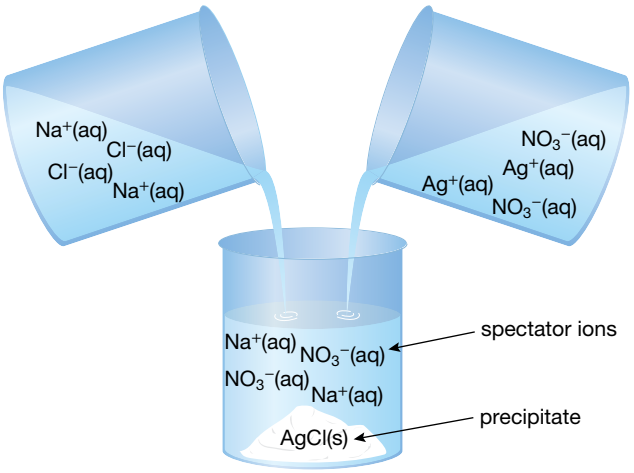

FIGURE 4.24 Silver chloride forms a white precipitate

Steps for writing balanced precipitation full equations

1. Write out the reactants.
2. Write out the products (swap the anions), ensuring the formulas are correct.
3. Write a skeleton equation (just the reactants and products).
4. Balance the equation.
5. Use the solubility table to find out if one of the products is insoluble. This is the precipitate (sometimes a precipitate may not form and on rare occasions two precipitates are formed).
6. Add states, remembering that the insoluble compound, the precipitate, must be solid(s).

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SAMPLE PROBLEM 7 Predicting precipitation reactions

Predict the products and write a balanced equation for the reaction occurring between sodium carbonate solution and copper(II) sulfate solution.

THINK	WRITE
1. Write the reactants.	$Na_2CO_3 + CuSO_4$
2. Predict the products formed, ensuring that you have written the correct formulas.	$Na_2SO_4 + CuCO_3$
3. Write the skeleton equation for the reaction.	$Na_2CO_3 + CuSO_4 \rightarrow Na_2SO_4 + CuCO_3$
4. Balance the equation.	$Na_2CO_3 + CuSO_4 \rightarrow Na_2SO_4 + CuCO_3$ (No balancing is required.)
5. Refer to the solubility table to determine if either of the compounds is insoluble.	Sodium sulfate is aqueous (since all group I compounds are soluble), but copper(II) carbonate is an insoluble compound.
6. Include symbols of state; the insoluble compound has (s) for the symbol of state.	$Na_2CO_3(aq) + CuSO_4(aq) \rightarrow Na_2SO_4(aq) + CuCO_3(s)$

PRACTICE PROBLEM 7

Predict the products and write balanced equations for the reaction $KOH(aq) + Ca(NO_3)_2(aq)$.

4.4.5 Ionic equations

In all the preceding equations, the elements and compounds have been written in their molecular or formula unit forms. However, in aqueous solutions, the reactions are best represented by simpler equations called **ionic equations**. Ionic equations are equations that show only the species that are formed or changed in a reaction. Spectator ions are not included in an ionic equation because they remain in solution unchanged in the reaction.

ionic equation equation that only includes the species that have reacted, without the spectator ions

Steps for writing ionic equations:

1. Write the balanced full chemical equation as described previously.
2. Expand the equation by dissociating all the soluble compounds into their free ions. Check for any molecular substances such as acids and certain bases that react with water to produce ions. Replace the formulas of these substances by the ions that they form.
3. Cancel all free ions that are unchanged on both sides of the equation (the spectator ions).
4. Write out the ions and formulas remaining. This is the ionic equation.
5. Make sure the equation is balanced in charge as well as in the number of atoms and add symbols of state.

Writing ionic equations

Spectator ions are omitted in ionic equations.

tlvd-0562

SAMPLE PROBLEM 8 Writing balanced ionic equations

Write a balanced ionic equation for the reaction occurring between sodium carbonate solution and copper(II) sulfate solution. State the spectator ions.

THINK	WRITE
1. Write out the full equation using the steps already outlined.	$Na_2CO_3(aq) + CuSO_4(aq) \rightarrow Na_2SO_4(aq) + CuCO_3(s)$
2. Expand the equation by dissociating or ionising all of the soluble compounds.	$2Na^+(aq) + CO_3^{2-}(aq) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow$ $2Na^+(aq) + SO_4^{2-}(aq) + CuCO_3(s)$
3. Cancel all free ions that are unchanged on both sides of the equation (these are the spectator ions).	$\cancel{2Na^+}(aq) + CO_3^{2-}(aq) + Cu^{2+}(aq) + \cancel{SO_4^{2-}}(aq) \rightarrow$ $\cancel{2Na^+}(aq) + \cancel{SO_4^{2-}}(aq) + CuCO_3(s)$
4. Write out the remaining ions and formula; this is the ionic equation.	$Cu^{2+} + CO_3^{2-} \rightarrow CuCO_3$
5. Check that the equation is balanced and includes states.	$Cu^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CuCO_3(s)$
6. Spectator ions are those that have not reacted.	Na^+, SO_4^{2-}

PRACTICE PROBLEM 8

Write a balanced ionic equation for the reaction between $MgCl_2$ and NaOH. State the spectator ions.

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EXPERIMENT 4.3 online only

Testing for solubility

Aim

To investigate the formation of precipitates in aqueous solution and gain skill in writing ionic equations to represent them

4.4 Activities

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4.4 Quick quiz on | **4.4 Exercise** | **4.4 Exam questions**

4.4 Exercise

1. Explain the difference between the terms ‘solute’ and ‘solvent’. Give an example of each.
2. Describe how the structure of a water molecule makes it a good solvent for ionic and polar molecular compounds.
3. In order to dissolve the salt lithium bromide, LiBr, in water, what types of bonds need to be broken and what interactions are formed?
4. **a.** What is a precipitation reaction?
   b. Describe what you would observe by mixing two solutions which then form a precipitate.
5. Use table 4.5 to determine whether each of the following ionic compounds is soluble.
   a. $BaSO_4$ **b.** $Pb(NO_3)_2$ **c.** AgI **d.** Sodium sulfide
6. Explain which of the following reactions would be classified as precipitation reactions.
   a. $2HCl(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)$
   b. $Ca(OH)_2(s) \rightarrow CaO(s) + H_2O(g)$
   c. $K_2CO_3(aq) + CaCl_2(aq) \rightarrow CaCO_3(s) + 2KCl(aq)$
7. Predict which of the following pairs of solutions will form a precipitate when mixed.
   a. Iron(III) nitrate and sodium hydroxide **b.** Barium chloride and copper sulfate
   c. Sodium nitrate and ammonium carbonate
8. Predict products and write balanced equations for the following.
   a. $2LiOH(aq) + Ca(NO_3)_2(aq)$ **b.** $Na_2S(aq) + (CH_3COO)_2Pb(aq)$ **c.** $2(NH_4)_3PO_4(aq) + 3CaCl_2(aq)$
9. Write ionic equations for the following reactions and identify the spectator ions.
   a. Barium chloride and magnesium sulfate solutions reacting to form a precipitate of barium sulfate and magnesium chloride solution
   b. Calcium nitrate reacting with sodium carbonate in solution to form solid calcium carbonate and sodium nitrate solution
10. Zahra wishes to test a solution for the presence of chloride ions. Name a soluble salt solution she could add to produce a precipitate containing chloride ions, and write the ionic equation for the reaction.

4.4 Exam questions

Question 1 (1 mark)

With the aid of a solubility table, select the three substances from the following list that are insoluble in water.

$$AgCl,\ NaNO_3,\ NH_4Br,\ BaSO_4,\ MgCO_3,\ K_2CO_3,\ Ca(CH_3COO)_2$$

Question 2 (1 mark)

MC Which of the following ionic equations matches the full equation provided?

$$Pb(NO_3)_2(aq) + 2NH_4I(aq) \rightarrow PbI_2(s) + 2NH_4NO_3(aq)$$

A. $Pb^{2+}(aq) + I^-(aq) \rightarrow PbI_2(s)$
B. $Pb(aq) + 2I(aq) \rightarrow PbI_2(s)$
C. $Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$
D. $NO_3^-(aq) + NH_4^+(aq) \rightarrow NH_4NO_3(s)$

Question 3 (2 marks)

Write the **(a)** full and **(b)** ionic equation for the formation of a gelatinous precipitate when aqueous solutions of cobalt(II) chloride and sodium hydroxide are mixed.

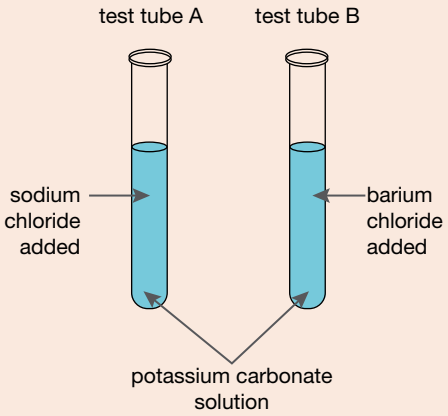
Question 4 (6 marks)

Solutions of sodium chloride and hydrochloric acid are both good conductors of electricity. This means that they both have mobile ions present.

a. Define the term 'dissociation'. **(1 mark)**
b. Define the term 'ionisation'. **(1 mark)**
c. Explain why a solution of sodium chloride can conduct electricity. Use the appropriate term from parts **a** and **b** in your explanation. **(2 marks)**
d. Explain why hydrochloric acid can conduct electricity. Use the appropriate term from parts **a** and **b** in your explanation. **(2 marks)**

Question 5 (8 marks)

Some potassium carbonate solution is poured into two separate test tubes. Some sodium chloride solution is then added to test tube A, and some barium chloride solution is added to test tube B.



a. In which test tube would a precipitate be observed to form? **(1 mark)**
b. Write the formula equation for the formation of the precipitate that you have identified from part **a**. **(2 marks)**
c. Write the ionic equation for this precipitation reaction. **(1 mark)**
d. What are spectator ions? In the example from part **c**, list the spectator ions that are present. **(2 marks)**
e. Suppose that the precipitate is filtered and the remaining liquid (the filtrate) is then heated to evaporate the remaining water. Describe what happens, naming any new substances that might be formed. **(2 marks)**

More exam questions are available in your learnON title.

4.5 A review of bonding

BACKGROUND KNOWLEDGE

- A review of the different types of bonding between atoms, the different models for the structure of materials, and methods of classification using properties and composition

4.5.1 The basis of bonding

We have seen from the previous topics that only the noble gases exist as separate atoms due to their stable outer shell containing eight electrons. The atoms of all other elements form chemical bonds with each other in order to attain a more stable outershell configuration.

Atoms can become stable in one of three ways:

1. by giving electrons to another atom
2. by taking electrons from another atom
3. by sharing electrons with another atom.

When atoms combine to achieve more stable structures, three types of bonding are possible. This is illustrated in figure 4.25.

FIGURE 4.25 The three main types of bonding

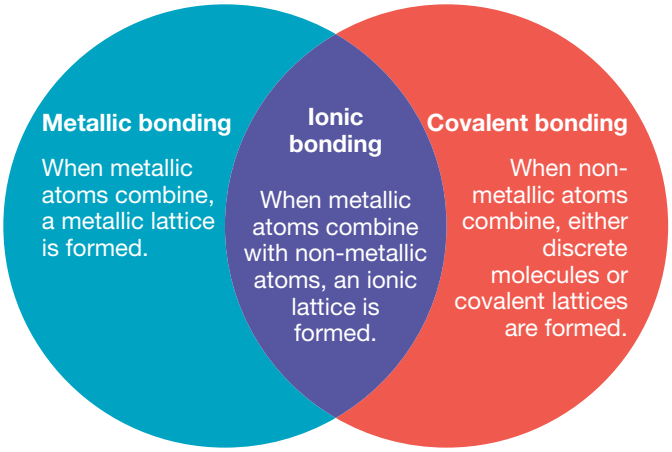
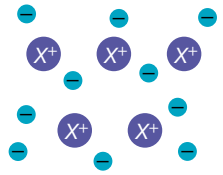
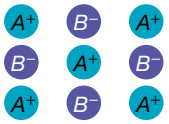
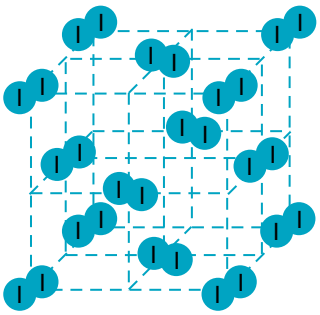
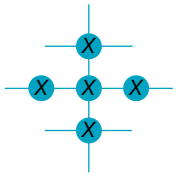
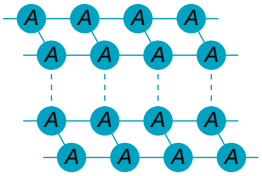
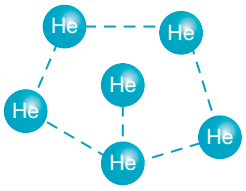


FIGURE 4.26 (a) Rare metallic gold crystals and (b) red ionic crystals of rhodochrosite (manganese carbonate)

4.5.2 Lattice structures

Atoms may bond together to form crystalline solids. The bonding between atoms gives rise to one of six lattice types as seen in table 4.6.

TABLE 4.6 The structures of different crystalline lattices

Type of lattice structure	Type of substance	Particles in the structure	Pictorial representation of the structure	Bonding types present	Examples
Metallic	Element or alloy that contains metals only	Cations and electrons		Metallic	Fe, Ca
Ionic	Compound that contains both a non-metal and a metal	Cations and anions		Ionic	NaCl, KF
Covalent molecular	Elements or compounds that contain non-metals only	Molecules		Covalent, dispersion forces; may also be dipole–dipole and hydrogen bonding	I_2
Covalent network	Elements or compounds that contain non-metals only	Atoms		Covalent	SiO_2, C (diamond)
Covalent layer	Elements or compounds that contain non-metals only	Atoms		Covalent, dispersion forces	C (graphite)
Atomic	Noble gases	Atoms		Dispersion forces	Ne, Ar, Kr

4.5.3 Identifying the bonding in different substances

A material's properties depend on its structure, and its structure is determined by the nature of the constituent particles and the bonds holding them together. We can, therefore, identify the type of bonding in different substances using:

1. the properties of a substance
2. the composition of a substance.

Properties

The properties of different crystalline solids are a consequence of the type of bonding present in the lattice (see table 4.7).

TABLE 4.7 The properties of different crystalline lattices

Type of lattice structure	Melting point	Electrical conductivity			Solubility in	
		Hardness	Solid	Molten or aqueous	Water (polar)	Petrol (non-polar)
Metallic	High	Varies	Good	Good	Insoluble	Insoluble
Ionic	High	Hard	Poor	Good	Most are soluble	Most are insoluble
Covalent molecular	Low	Soft	Poor	Poor	Soluble if polar	Soluble if non-polar
Covalent network	Very high	Hard	Poor	N/A	Insoluble	Insoluble
Covalent layer	Very high	Soft	Good	N/A	Insoluble	Insoluble

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SAMPLE PROBLEM 9 Identifying bonding based on melting point and electrical conductivity

Identify the bonding in substance *A* and substance *B*, given the following information.

a. Substance *A* is hard, melts at 890 °C and conducts electricity in the molten form but not in the solid form.

b. Substance *B* is soft, melts at −183 °C, and is a poor conductor of electricity in solid or liquid form.

THINK	WRITE
a. Electrical conduction of *A* suggests it is ionic, and this is consistent with hardness and melting point.	Substance *A* has an ionic lattice structure in its solid state.
b. The low melting point of *B* suggests it is a molecular lattice, and this is consistent with the softness and electrical conductivity.	Substance *B* has a covalent molecular lattice structure in its solid state.

PRACTICE PROBLEM 9

Identify the bonding in substance *C* and substance *D*, given the following information.

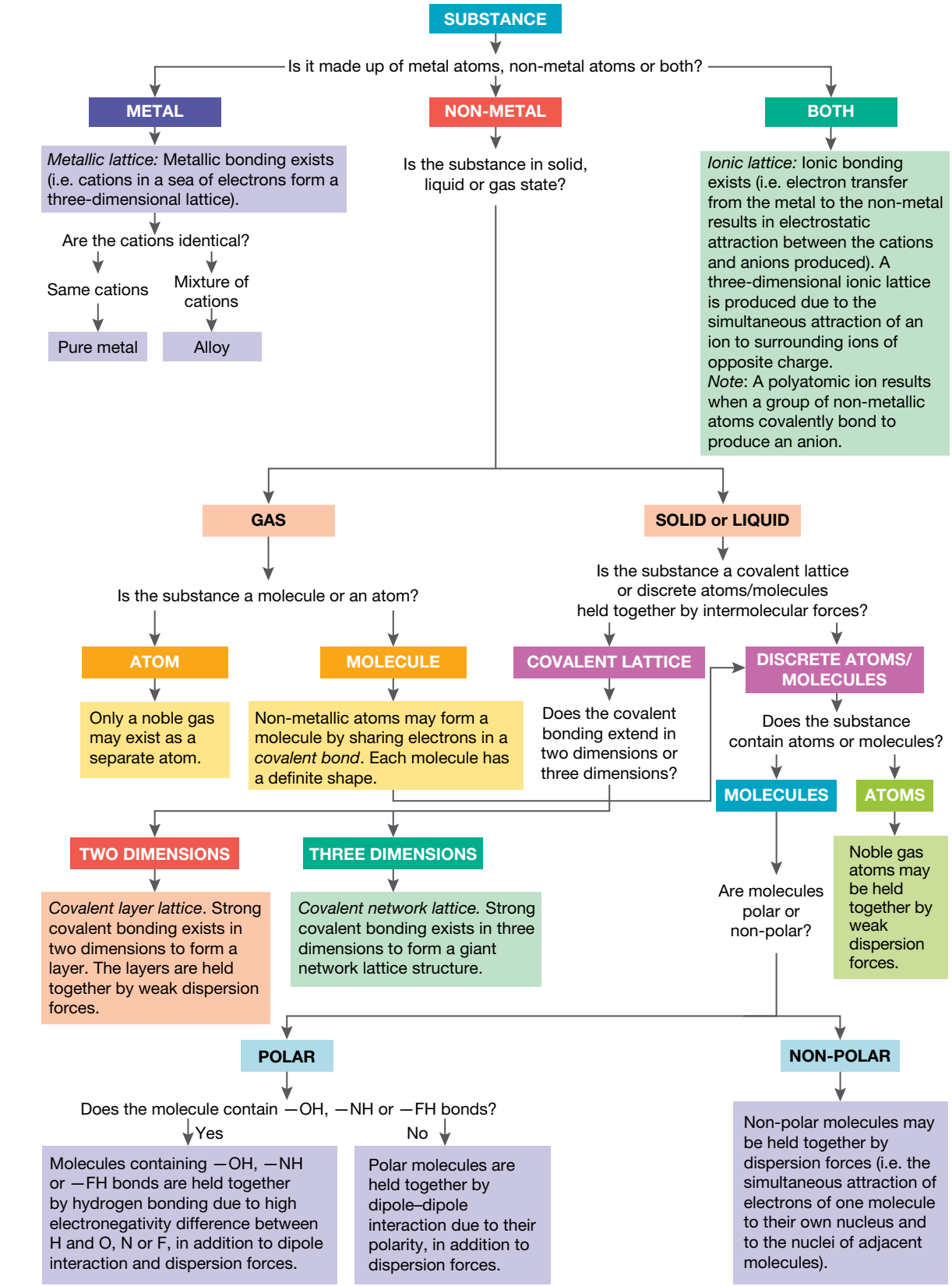
a. Substance *C* is hard, melts at 650 °C and conducts electricity in both the solid and liquid states.

b. Substance *D* is soft and is a good conductor of electricity in the solid state. It melts at 1700 °C.

Composition

Figure 4.27 shows how substances can be identified according to their composition.

FIGURE 4.27 Classification of substances according to composition



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SAMPLE PROBLEM 10 Identifying the type of bonding

Using the flow chart in figure 4.27, identify the bonding in:
a. magnesium (a solid metal)
b. magnesium chloride (a solid salt)
c. chlorine gas.

THINK	WRITE
a. Using the classification of substances flow chart, magnesium is a pure metal, so the bonding is metallic.	Magnesium displays metallic bonding.
b. Magnesium chloride is made up of both metal and non-metal elements; according to the flow chart, the bonding is ionic.	Magnesium chloride displays ionic bonding.
c. Chlorine gas is composed of diatomic molecules of non-metallic atoms. According to the flow chart, the bonding is covalent within the chlorine molecule. Because it is non-polar, dispersion forces also exist between the chlorine molecules.	Chlorine displays covalent bonding within its molecules and dispersion forces between them.

PRACTICE PROBLEM 10

Using the flow chart in figure 4.27, identify the bonding in:
a. solid copper
b. copper(II) oxide
c. oxygen gas.

4.5 Activities

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4.5 Quick quiz on | 4.5 Exercise

4.5 Exercise

1. Identify the bonding in substance *E* and substance *F*, given the following information: substance *E* is soluble in water and has a melting point of –25 °C. Substance *F* is hard and is a poor conductor of electricity. It melts at 2500 °C.
2. Classify each of the following as covalent molecular, ionic or metallic. Justify your responses.
   a. A yellow solid that melts at 105 °C to form a clear yellow liquid. Both solid and liquid are poor conductors of electricity.
   b. A solid that melts at 99 °C to form a silvery liquid. Both solid and liquid are good conductors of electricity.

c. A dark shiny solid that sublimes to form a vapour. It is a poor conductor of electricity and heat.
d. A white solid that melts at 872 °C to form a colourless liquid. The solid does not conduct electricity, whereas the liquid does.

3. Consider the substances in the following table.

Substance	Melting point (°C)	Electrical conductivity		
		Solid	Molten	Dissolved in water
L	–115	—	—	—
M	1260	—	Conducts	Conducts
N	3600	—	—	Insoluble
O	2468	Conducts	Conducts	Insoluble

a. Which substance is metallic?
b. Which substance is ionic?
c. Which substance has a covalent molecular structure?
d. Which substance has a covalent network structure?

4. Use figure 4.27 to identify the bonding in the following substances.
a. Copper
b. Copper(II) chloride
c. Argon
d. Carbon tetrachloride liquid
e. Sulfur dioxide

5. Use figure 4.27 to identify the bonding in the following substances.
a. Steel
b. Diamond
c. Graphite
d. Ammonia gas
e. Iron(II) sulfate

6. Classify the type of lattice structure found in the following solids.
a. Sulfur
b. Silicon dioxide
c. Aluminium fluoride
d. Bronze
e. Tungsten
f. Dry ice

7. a. Use the flow chart shown in figure 4.27 to classify the bonding within and between the particles that make up the following compounds.
b. Draw the shape diagram for each.
i. CO_2
ii. NH_3
iii. CH_4
iv. CH_3OH

8. Which one of the following bonds has the strongest dipole? Justify your response.
A. C—H
B. O—O
C. N—Cl
D. H—F.

9. a. Discuss the bonding in sodium chloride (table salt).
b. Why are salt crystals brittle?
c. What property of sodium chloride makes it useful as a flavouring? Relate this property to its structure and bonding.

10. Explain why, in terms of structure and bonding and their relationship to properties, the following materials have the uses given.
a. Gold is used to make jewellery.
b. Graphite is used as a lubricant in locks.

4.6 Review

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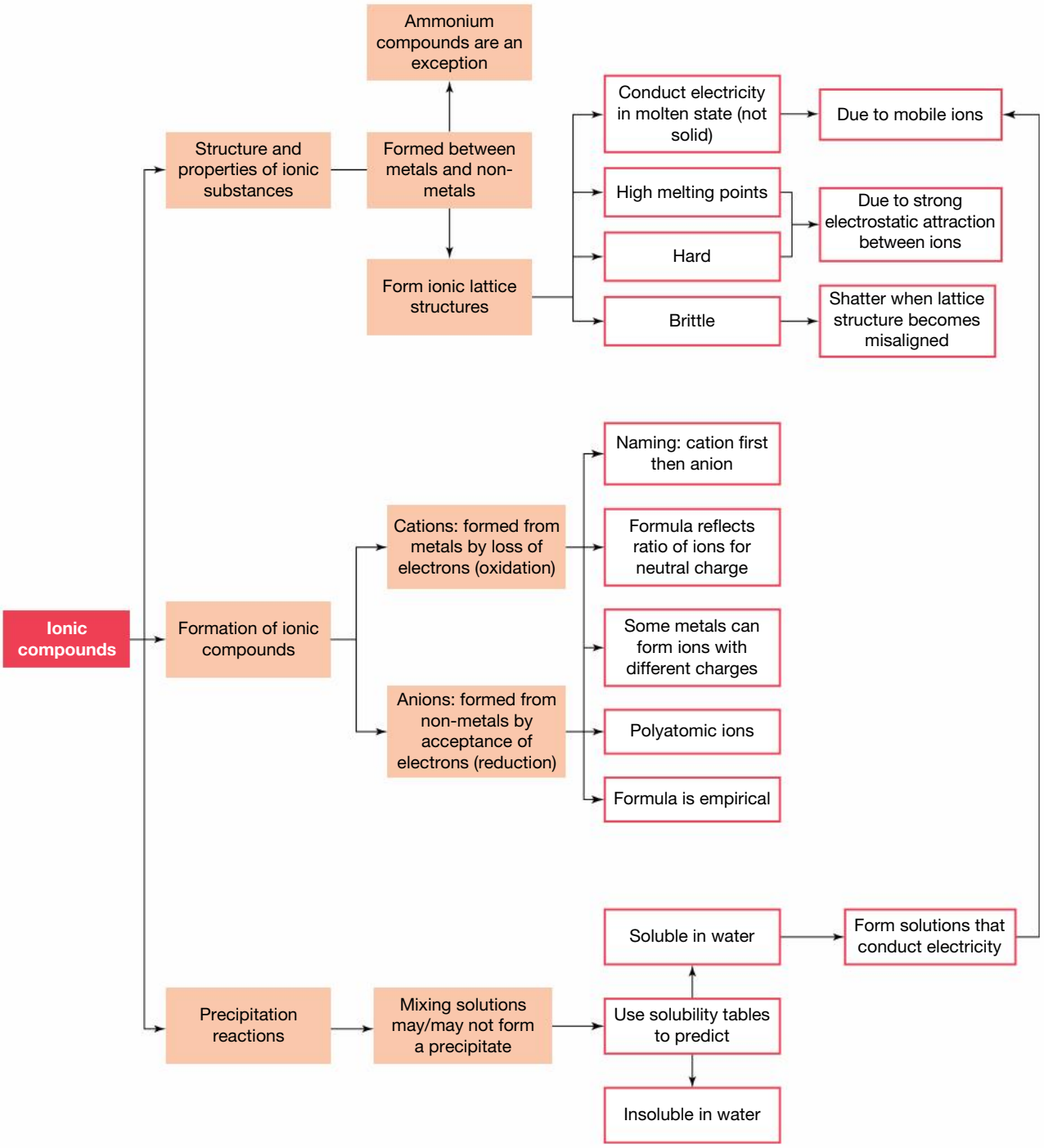
4.6.1 Topic summary



4.6.2 Key ideas summary

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4.6.3 Key terms glossary

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on Resources

Solutions	Solutions — Topic 4 (sol-0803)
Practical investigation eLogbook	Practical investigation logbook — Topic 4 (elog-1609)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 4 (doc-37073) Key idea summary — Topic 4 (doc-37074)
Exam question booklet	Exam question booklet — Topic 4 (eqb-0085)

4.6 Activities

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4.6 Review questions

1. In a sample of potassium chloride (an important electrolyte in the human body), the number of K^+ ions and Cl^- ions are equal. In a sample of potassium oxide, however, the numbers of each ion are not equal. State the ratio of each ion in the potassium oxide and explain why they are not equal.
2. Write simple equations and electron shell diagrams to show what happens when the following pairs of atoms bond.
   a. K and F **b.** Al and O **c.** Be and Cl
3. Explain why ionic salts conduct electricity in the molten and aqueous states but not in the solid state.
4. Draw diagrams to illustrate the brittle nature of salt.
5. A compound made up of *X* and *Y* atoms has the following properties.
   - It has a high melting point and boiling point.
   - It is very soluble in water but not in kerosene.
   - It conducts electricity in the liquid or aqueous state but not in the solid state.
   - The electron configuration of an atom of *X* is 2, 8, 2, while that of *Y* is 2, 6.

   a. What is the formula for the compound?
   b. What type of bonding must it have?
   c. Write a simple equation showing how the bonding is attained.
6. Using table 4.3 (in section 4.3.5), write the formulas for the following compounds.
   a. Sodium fluoride **b.** Barium nitrate
   c. Iron(III) hydroxide **d.** Ammonium hydroxide
   e. Chromium(III) oxide **f.** Sodium hydrogen phosphate
7. Give the names of the following compounds.
   a. $ZnCl_2$ **b.** $Al_2(CO_3)_3$ **c.** Na_2SO_4 **d.** $AgNO_3$ **e.** NaOH
8. Predict whether a precipitate forms when solutions of the following are mixed.
   a. $AgNO_3$ and KCl **b.** H_2SO_4 and $BaCl_2$ **c.** $NaNO_3$ and KCl
   d. $(NH_4)_3PO_4$ and $CaCl_2$ **e.** $(NH_4)_2S$ and $Pb(NO_3)_2$ **f.** $(NH_4)_2S$ and $NaNO_3$
   g. $CaCl_2$ and NaBr

9. Write ionic equations for the reactions that form precipitates in question 8.
10. Write a full equation and an ionic equation for each of the following reactions (if no reaction occurs, write 'no reaction').
    a. Solutions of potassium sulfate and calcium nitrate are mixed.
    b. Lead (II) nitrate solution is added to potassium iodide solution.
    c. Magnesium chloride is added to sodium phosphate solution.
    d. Solutions of potassium sulfide and zinc chloride are mixed.
    e. Solutions of sodium sulfate and potassium bromide are mixed.

4.6 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC Which of the following best describes the formation of an ionic compound?

A. One or more electrons are transferred from a metal to a non-metal, resulting in oppositely charged ions.
B. One or more electrons are transferred from a non-metal to a metal, resulting in oppositely charged ions.
C. One or more electrons are shared between a non-metal and a metal held by electrostatic attraction.
D. Electrostatic attraction between anions, cations and delocalised electrons hold a lattice together.

Question 2

MC Which of the following lists contains only ionic compounds?

A. CH_4, HBr, $KMnO_4$
B. Na_2SO_3, H_2SO_3, SO_3
C. NH_4F, LiCl, KBr
D. MgO, Na_2O, CO_2

Question 3

MC Which of the following ionic compounds has twice as many anions as cations in its crystal lattice?

A. $CaSO_4$ **B.** Li_2SO_4 **C.** $Al(NO_3)_3$ **D.** CaF_2

Question 4

MC Which of the following statements best explains why ionic solids are *not* able to conduct electricity?

A. Ionic solids do not contain any charged particles.
B. Ions are not free to move in ionic solids.
C. No delocalised electrons are present in ionic solids.
D. Ionic solids create ions only when they are dissolved or molten.

Question 5

MC Which of the following would produce a solution with higher conductivity than the others? (Assume equal concentrations.)

A. $Al(NO_3)_3$ **B.** $CuSO_4$ **C.** $NaNO_3$ **D.** $(NH_4)_2SO_4$

Question 6

MC An ionic compound has the formula $M_3(PO_4)_2$ for an ion of an unknown metal, *M*. Which other compounds could *not* be formed by this metal ion?

A. MCO_3
B. MCl_2
C. $M(HCO_3)_2$
D. MNO_3

Question 7

MC Which one of the following mixtures of solutions could result in the formation of a precipitate?

Mixture I: $AgNO_3(aq) + KNO_3(aq)$

Mixture II: $AgNO_3(aq) + KCl(aq)$

Mixture III: $CuSO_4(aq) + KCl(aq)$

Mixture IV: $CuCl_2(aq) + NH_4NO_3(aq)$

A. Mixtures I and III only
B. Only mixture II
C. Mixtures II and III only
D. Mixtures II, III and IV only

Question 8

MC What are the spectator ions in the precipitation reaction that occurs when potassium hydroxide is mixed with cobalt(II) nitrate?

A. Potassium ions and nitrate ions
B. Cobalt(II) and hydroxide ions
C. Nitrate ions only
D. Potassium ions only

Question 9

MC A student mixed together pairs of aqueous solutions of the following reactants to investigate the solubility of compounds.

The table of results is shown.

Reaction	Reactant solutions	Observation
1	Sodium chromate + copper(II) nitrate	Precipitate formed
2	Potassium nitrate + copper(II) nitrate	No precipitate formed
3	Sodium chromate + potassium nitrate	No precipitate formed

Based on these results, what can be inferred?

A. Sodium chromate is insoluble in water
B. Copper(II) chromate is insoluble in water
C. Copper(II) nitrate is insoluble in water
D. Both copper(II) chromate and sodium nitrate are insoluble in water

Question 10

MC Which of the following equations shows the formation of a precipitate when iron(III) chloride and sodium hydroxide solutions are mixed?

A. $FeCl_3(aq) + NaOH(aq) \rightarrow Fe(OH)_3(s) + NaCl(aq)$
B. $FeCl_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(aq) + 3NaCl(s)$
C. $FeCl_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(s) + NaCl(aq)$
D. $FeCl_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(s) + 3NaCl(aq)$

Section B — Short answer questions

Question 11 (2 marks)

Explain, by referring to the ionic and metallic bonding models, why ionic substances are brittle but metals are malleable.

Question 12 (3 marks)

Write the formula for each of the following ionic compounds.

a. Magnesium hydrogen carbonate **(1 mark)**
b. Ammonium sulfate **(1 mark)**
c. Calcium hydroxide **(1 mark)**

Question 13 (2 marks)

Which of sodium chloride, potassium chloride or rubidium chloride would be expected to have the highest melting point? Give a reason for your answer.

Question 14 (2 marks)

Given that silver halides and hydroxide compounds of transition metals are not soluble in water, write the full equations (including symbols of state) for the following precipitation reactions.

a. Silver nitrate solution mixed with ammonium iodide solution **(1 mark)**
b. Zinc nitrate solution mixed with potassium hydroxide solution **(1 mark)**

Question 15 (2 marks)

Write the **(a)** full and **(b)** ionic equations for the formation of solid calcium carbonate when aqueous solutions of sodium carbonate and calcium iodide are mixed.

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5 Separation and identification of the components of mixtures

KEY KNOWLEDGE

In this topic you will investigate:

Separation and identification of the components of mixtures

- polar and non-polar character with reference to the solubility of polar solutes dissolving in polar solvents and non-polar solutes dissolving in non-polar solvents
- experimental application of chromatography as a technique to determine the composition and purity of different types of substances, including calculation of R_f values.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of VCE Chemistry. Experiments and investigations, supported by a practical investigation eLogbook and teacher-led videos, are included in this topic to provide opportunities to undertake investigations and communicate findings.

EXAM PREPARATION

Access exam-style questions and their video solutions in every lesson, to ensure you are ready.

5.1 Overview

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5.1.1 Introduction

Have you ever noticed the rainbow pattern formed by oil as it washes off the road after rain? You may also have noticed this if you have tried to clean an oily garage floor or barbeque hot plate. This is caused by light reflecting from a thin film of oil that floats on the water, rather than mixing with it and dissolving. The reason that oil does not dissolve in water is due to differences in polarity between the oil molecules and the water molecules. Although water is a good solvent, in many situations its use is not desirable and other solvents need to be considered. It, therefore, becomes critical that the role played by molecular polarity in the dissolving process is understood.

FIGURE 5.1 An oil slick on a wet road forms due to differences in polarity between the oil and water molecules.

Another area in which polarity plays an important role is chromatography. This is a simple yet powerful technique that can be used to determine the composition and purity of various substances. As you will see, it can involve very simple equipment, or it can involve quite expensive and sophisticated equipment. Chromatography relies on weak bonds that form between a stationary surface and a mobile mixture that moves across it. The strength of these weak bonds is very sensitive to the relative polarities involved. This is exactly the same phenomenon you may have observed if you have ever spilt methylated spirits on a page of handwriting and noticed the ink separating into different colours.

This topic will focus on the role played by polarity in forming solutions, and on chromatography as a useful technique to identify substances in a mixture. You will learn how to make predictions about the dissolving process and how to perform calculations to assist in identifying substances using chromatography.

LEARNING SEQUENCE

5.1 Overview 188
5.2 Solutions, solvents and chromatography 189
5.3 Review 207

on Resources

Solutions	Solutions — Topic 5 (sol-0804)
Practical investigation eLogbook	Practical investigation logbook — Topic 5 (elog-1610)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 5 (doc-37075) Key idea summary — Topic 5 (doc-37076)
Exam question booklet	Exam question booklet — Topic 5 (eqb-0086)

5.2 Solutions, solvents and chromatography

KEY KNOWLEDGE

- Polar and non-polar character with reference to the solubility of polar solutes dissolving in polar solvents and non-polar solutes dissolving in non-polar solvents
- Experimental application of chromatography as a technique to determine the composition and purity of different types of substances, including calculation of R_f values

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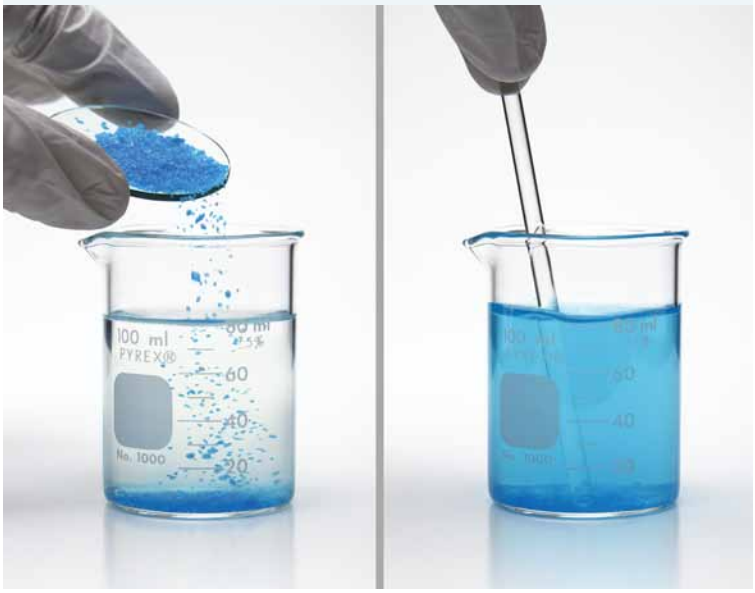
5.2.1 Polar and non-polar solutions

Solutions of various types play an important role in chemistry. A **solution** is made up of two parts:

- The **solute**: the substance that gets dissolved. Solutes may be solids, liquids or gases.
- The **solvent**: the substance that does the dissolving. Solvents are nearly always liquids.

FIGURE 5.2 The initially solid solute copper sulfate pentahydrate, $CuSo_4 \cdot 5H_2O$, is dissolved in the solvent water, H_2O, to create a solution.



In chemical processes and reactions, substances may be present not only as solids, liquids and gases, but also as substances that are dissolved to form solutions. The form in which a substance is present is referred to as its **phase**. Common phases include solid (s), liquid (l), gas (g) and aqueous solution (aq). Aqueous solutions are very common and refer to solutions where water is the solvent. However, if a different solvent is used to dissolve something, the solution that results will represent another type of phase. For example, a solution of iodine (solute) dissolved in ethanol (solvent) would represent a non-aqueous solution (non-aq) and could be referred to as a non-aqueous phase.

You will recall from topic 4 that water is particularly good at dissolving many ionic substances due to the **polarity** of its molecules. (Polarity was discussed in topic 2.) This is due to the formation of ion–dipole bonds that are able to disrupt the ionic lattice and allow dissolving to take place. Furthermore, in topic 10 you will see that this polarity is important in explaining how certain molecular substances such as sugar and various acids also dissolve readily in water. Although it is the most common solvent, many other liquids besides water can also function as solvents. Common examples include ethanol, propanone (also known as acetone), cyclohexane and xylene. Different solvents often dissolve different solutes. Furthermore, solutes soluble in one particular solvent may not be soluble in another or may be soluble to a lesser extent.

solution solute dissolved in a solvent

solute dissolved substance in a solution

solvent substance in which the solute is dissolved

phase the physical form in which a substance is present; common phases include solid, liquid, gas and aqueous solution, but other phases exist as well

polarity localised imbalances in electric charges within a molecule resulting in a negatively charged end and positively charged end

EXTENSION: New solvents and green chemistry

Industrial processes use non-aqueous solvents in huge quantities for processes such as separation and purification, and even as catalysts. They are also used to form solutions in which reactions occur. Often these solvents are organic compounds (compounds containing carbon) that, while suitable for their purpose, have other undesirable properties from a medical and an environmental viewpoint. One of the main such undesirable properties is their tendency to evaporate readily, thus contaminating the factory atmosphere at a local level and even the global atmosphere on a larger scale. Over the past few decades, considerable interest and growth has been seen in a new area of solvent research — ionic liquids.

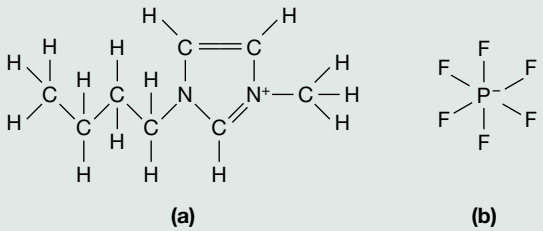
At the simplest level, an ionic liquid is simply an ionic compound that has melted. Sodium chloride, for example, melts at 801 °C and could, therefore, potentially be used as an ionic solvent above this temperature. While this might seem impractical due to the temperature involved, another ionic substance is already used extensively in this capacity as an ionic solvent — cryolite, Na_3AlF_6, has a melting point of 1012 °C and is used in molten form to dissolve alumina, Al_2O_3, during the production of aluminium metal.

The increased interest in ionic liquids in recent years, however, has focused on substances that have much lower melting points, with the term 'ionic liquid' now often being used to refer to ionic substances that are liquids at temperatures below 100 °C. From your work in topic 4, you will recall that ionic substances have high melting points and this, therefore, appears to be a contradiction. The reason is that these liquids usually consist of a large, bulky cation and a relatively bulky polyatomic anion. It is the bulkiness of the ions involved that contributes to the lower melting points. In the ionic compounds studied in topic 4, the ions involved (even the polyatomic ones!) were relatively small, meaning that they were able to pack closely together in the crystal lattice structure, thus increasing the electrostatic attraction between them and the subsequent melting points. The bulky ions possessed by ionic liquids are still attracted to each other due to their charges. However, they cannot pack together as closely as before. Because they are forced to be further apart, the strength of the electrostatic attraction between them is far less and their melting points are consequently lower.

A whole range of these solvents exist, and they can even be 'tailored' to suit a particular application through careful choice of the cation or anion involved. Figure 5.3 shows the ionic structure of a common ionic liquid.

Ionic liquids are increasingly being described as green solvents. They are finding application in roles where they perform the same functions as traditional solvents. They have much lower vapour pressures, meaning that they do not evaporate as easily and, therefore, have the potential to improve air quality at both local and global levels. A huge variety of potential uses for ionic liquids exist, including as lubricants for wind turbines, hydraulic fluids, thermal fluids, biomedical sensors and in dissolving biomass to produce polymers (see topic 8).

FIGURE 5.3 Structures of the (a) cation and (b) anion in a common ionic solvent. Note the bulky appearance of both ions.



Resources

Weblink Ionic liquids

5.2.2 The like dissolves like rule

It has long been observed that *polar* solutes tend to dissolve in *polar* solvents, whereas non-polar solutes do not. The reverse is also observed to be true. That is, *non-polar* solutes tend to dissolve in *non-polar* solvents whereas polar solutes do not. This observation provides a very useful rule for chemists — often referred to as the **like dissolves like rule**.

like dissolves like rule polar solutes are more likely to dissolve in polar solvents, while non-polar solutes are more likely to dissolve in non-polar solvents

Like dissolves like

The like dissolves like rule states that *polar* solutes are more soluble in *polar* solvents, whereas *non-polar* solutes are more soluble in *non-polar* solvents.

This rule provides a useful general rule to chemists when predicting whether a solution will form between a particular solute–solvent combination. From earlier work in topic 4, you will remember that sodium chloride, being ionic in nature, is soluble in water. It is, therefore, expected that it would be virtually insoluble in the non-polar solvents hexane or toluene. On the other hand, common experience tells us that paraffin wax (which is a mixture of large hydrocarbon molecules) is *not* soluble in water. The like dissolves like rule explains this. Paraffin wax is non-polar in nature while water is polar. This means the wax will not dissolve. However, if a non-polar solvent such as xylene or hexane is chosen, the wax will dissolve.

FIGURE 5.4 Oil and water do not mix. The non-polar oil forms a layer on top of the polar water.

This phenomenon can be seen in figure 5.5. Non-polar, covalently bonded iodine, I_2, dissolves in non-polar cyclohexane, C_6H_{12} (top; cyclohexane is lighter than water), but does not dissolve in water.

FIGURE 5.5 Non-polar I_2 is soluble in cyclohexane (top) but insoluble in water (bottom).

Causes of molecular polarity

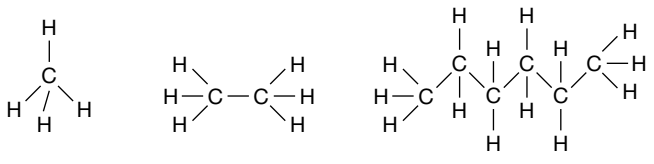
The causes of molecular polarity and non-polarity were discussed in topic 2. You will remember that for a substance to be polar:

- its molecules must contain polar bonds
- its shape must be asymmetrical.

If either or both of these conditions are not present, the substance will be non-polar.

Of particular note in any discussion of solvents and polarity are organic compounds that contain large proportions of carbon and hydrogen atoms bonded together. In topic 2, the molecular structures of methane and ethane are discussed. The carbon–hydrogen bonds involved have little polarity and, more importantly, these molecules are symmetrical. These substances are, therefore, non-polar. This observation enables us to make a very useful generalisation: *the presence of hydrocarbon groups in molecules of a substance will tend to make that substance non-polar.*

FIGURE 5.6 The structures of methane, ethane and hexane. The large number of carbon–hydrogen bonds in hexane leads to the expectation that it will be non-polar.



Degree of polarity

Not all polar solvents are as polar as each other. Comparative polarities may be measured in a number of ways, and sometimes these can even vary from one circumstance to another. Table 5.1 shows the polarities of a number of solvents relative to water. In this table, values close to zero (that is, hexane, xylene and toluene) represent non-polar solvents while the other values indicate how polar a solvent is compared to water.

TABLE 5.1 Relative polarities* of some selected solvents

Solvent	Relative polarity
Hexane	0.00
Xylene	0.07
Toluene	0.09
Propanone (acetone)	0.36
Propan-2-ol	0.55
Ethanoic acid	0.65
Ethanol	0.70
Methanol	0.76
Glycerol	0.81
Water	1.00
Hydrogen fluoride	1.01

*Relative polarity is an average value compared to the polarity of water. The nature and determination of these values is outside the scope of VCE Chemistry.

Although the like dissolves like rule explains whether or not a substance might dissolve, it does not explain the *extent* to which dissolving may take place. Table sugar (sucrose), as expected, is readily soluble in water. This is due to its polarity, which is caused by eight oxygen–hydrogen bonds in each molecule. However, it is less soluble in both methanol and ethanol. This is readily explained by the relative polarities involved. Both methanol and ethanol exhibit a degree of polarity, but it is less than water. This means that sucrose should still dissolve in these solvents, but to a lesser extent than in water. Also, as ethanol is slightly less polar than methanol, the solubility would be expected to be lower in this solvent. Table 5.2 demonstrates this.

FIGURE 5.7 Health authorities are worried that some products contain too much dissolved sugar.

TABLE 5.2 Solubility of sucrose in selected solvents

Solvent	Relative polarity	Solubility at room temperature (g/100 mL)
Water	1.00	20.1
Methanol	0.76	1.0
Ethanol	0.70	0.6
Hexane	0.00	0.0

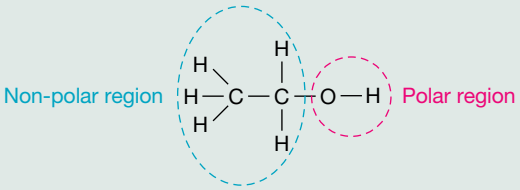
An area in which solvent choice based on molecular polarity is important is chromatography. This very useful identification and separation technique is discussed later in this topic.

EXTENSION: Molecules that exhibit both polar and non-polar character

The polarity figures in table 5.1 represent an overall polarity for the whole molecule. Often, however, many molecules may possess *regions* that are polar and other regions that are non-polar. This becomes more pronounced as certain types of molecules get larger. In cases such as these, whether they are acting as a solute or a solvent, both types of properties may be observed. A good example of this is E10 petrol. This is a mixture that contains 10 per cent ethanol to 90 per cent hydrocarbons. Based on polarity alone, it might be expected that the polar ethanol molecules will not dissolve in the non-polar hydrocarbons that make up the bulk of this petrol. Layers would be expected to form, and the mixture would be described as **immiscible**. However, this is not the case because the ethanol dissolves to form a solution in the petrol. A more detailed examination of the structure of the ethanol molecule explains why this is so. As seen in figure 5.9, although ethanol is polar because of the presence of the OH group, it also contains a C_2H_5 section that is non-polar. It is the attraction of this section to the hydrocarbons in the petrol that allows the ethanol to dissolve.

FIGURE 5.8 Many service stations now sell a blend of petrol and ethanol.

FIGURE 5.9 The regionalised polarity in the ethanol molecule



Another common example of molecules showing both types of behaviour are found in soaps and detergents. The polar sections of these molecules readily dissolve in water, while the non-polar sections dissolve in oil and grease. This is precisely the reason that these substances are so good at cleaning dishes, clothes and skin, compared to water alone.

immiscible when two substances are unable to form a homologous solution

tlvd-3489

SAMPLE PROBLEM 1 Identifying the polarity of solvents

Glucose is a carbohydrate molecule that dissolves easily in water.
Name one other solvent from table 5.1 in which glucose would be expected to dissolve.

THINK	WRITE
1. If glucose dissolves in water, glucose must be polar.	
2. The like dissolves like rule predicts that another polar solvent should, therefore, be able to dissolve glucose.	
3. Look at table 5.1, and pick another polar solvent.	Glycerol

PRACTICE PROBLEM 1

Name one solvent from table 5.1 that glucose would *not* dissolve in.

elog-1860

EXPERIMENT 5.1

online only

Polarity of liquids in a mixture

Aim

To investigate the mixing of water and propan-2-ol (isopropanol)

5.2.3 Introduction to chromatography

Chromatography is a group of techniques that utilise the weak bonding between molecules in a mixture and a surface over which they move. Polarity (or lack of polarity) plays an important role in this process. Chromatography is one the most useful techniques for determining the composition and purity of a mixture.

Chromatographic methods range from those that use simple and relatively cheap apparatus that is simple to use, to those that use expensive, sophisticated and sensitive equipment. Table 5.3 lists the major types of chromatography.

TABLE 5.3 Major types of chromatography

Type of chromatography	Comments
Paper	• Cheap • Moderate accuracy
Thin-layer chromatography (TLC)	• Similar to paper chromatography but more accurate • Offers wider scope for applications
Gas-liquid chromatography (GLC or GC)	• Very sophisticated and sensitive • Capable of quantitative determination
High-performance liquid chromatography (HPLC)	• As per GLC • Also capable of quantitative determination. • Discussed in Unit 4

Some of the situations in which chromatography is used include pharmaceutical research and production, determination of banned substances (for example, in urine samples from athletes), forensic work (for example, the detection and identification of explosives) and checking for the adulteration of consumer products.

The chromatography process

Chromatography is the process where a **mobile phase** (which contains the mixture to be analysed) moves over a **stationary phase** (which contains a material that has a high surface area). As this happens, substances in the mobile phase continually form weak temporary bonds with the stationary surface through **adsorption**. This is followed by the bonds breaking and allowing the substances to move on. This is known as **desorption**.

In paper chromatography, for example, the mobile phase consists of the mixture being separated along with a solvent, while the stationary phase consists of the cellulose fibres from the paper that are naturally coated with a thin layer of water.

This process of adsorption–desorption occurs at different rates due to subtle differences in the weak bonds involved. These are similar to the weak bonds that form between molecules. As you may recall from topic 2, these bonds consist of dispersion forces and are thus influenced by molecular size. They are also influenced by dipole–dipole attractions that are determined by the polarity of the substances involved. Because differences exist in how strongly this adsorption and desorption occurs, substances move over or through the stationary phase at different speeds and are separated from one another as a result.

chromatography a laboratory technique for the separation of a mixture by passing it in solution (mobile phase) over a stationary phase; separation occurs as components move at different rates

mobile phase the mixture to be analysed in the chromatography process

stationary phase a material with a high surface area, used in the chromatography process

adsorption the adhesion of atoms, ions or molecules from a gas, liquid or dissolved solid to a surface

desorption the opposite of adsorption, where molecules are released from the surface of other particles

How chromatography can be used to detect impurities

You may have noticed how certain food dyes, if mopped up with a damp piece of paper towel, spread out into different colours. You may have done simple experiments where a spot of ink is placed onto a piece of filter paper and water allowed to soak up and move the spot, eventually spreading it out into bands of different colours.

FIGURE 5.10 Paper chromatography is a simple and inexpensive form of chromatography.

Figure 5.10 shows such an experiment. The green ink has separated into its blue and yellow components. The red ink has one single component, and the black ink has many. The colours that move furthest are those that adsorb and desorb most weakly, while the colours with stronger adsorption and desorption do not move as far. It is thus immediately obvious if a sample is pure or impure. The presence of an extra band or bands beyond those expected would represent impurities in the sample.

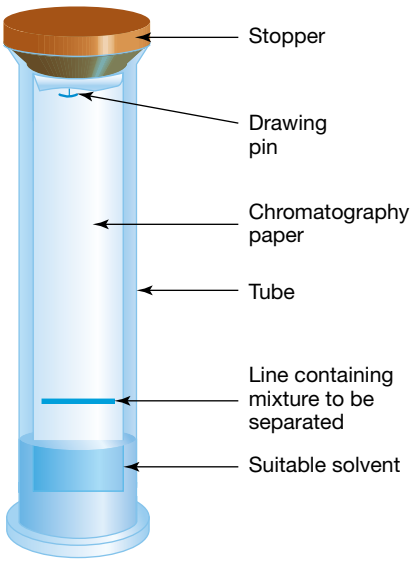
To enable identification of the components of a mixture, the chromatographic procedure would simply be repeated using known samples under identical conditions, and the results compared to those from the mixture.

Chromatography can be used to separate mixtures that have very similar properties and when only tiny amounts of samples are available.

5.2.4 Simple forms of chromatography

Although the following examples use simple equipment, they can still give quite good results as far as identification and purity ascertainment are concerned. Also, it should be remembered that chromatography does not necessarily have to involve coloured solutions. Colourless mixtures can also be analysed, usually with an extra step such as spraying the paper or slide with a chemical to make the bands visible. This is called developing.

FIGURE 5.11 A typical set-up for paper chromatography



Paper chromatography

In paper chromatography, a strip of paper is placed inside a container so that the surrounding vapour becomes saturated from the liquid at the bottom. Different solvents may be used to 'sweep' the mixture up the paper strip. The stationary phase consists of the solvent, usually water, which is adsorbed onto the cellulose fibres in the paper.

A typical experiment would consist of 'running' the sample to be analysed, using a particular solvent, under a given set of conditions. A number of known substances would then be run under identical conditions. The chemicals in the analysed sample would then simply be identified by comparing their behaviour to that of the known substances.

It is important to note in this method that the mixture to be separated is placed on the paper as either a spot or a line. This position should be marked in pencil, never in ink. If ink is used, the dyes in the ink may contaminate the sample being analysed. This is the starting point from which all movement takes place and is called the **origin**.

origin the position on either chromatography paper or a TLC slide that the sample is applied

Thin-layer chromatography

Thin-layer chromatography (TLC) is similar to paper chromatography. A finely divided adsorbent material is coated onto either a glass slide or aluminium foil. The procedures are essentially the same as for paper chromatography. Compared to paper chromatography, however, TLC is quicker and more precise, and can be used with solvent mixtures of a much more corrosive nature. Additionally, a much wider range of materials can be used as the stationary phase.

tlvd-3490

SAMPLE PROBLEM 2 Identifying bonding strength in TLC

A substance was analysed using TLC and was known to be a mixture of four compounds: *A*, *B*, *C* and *D*.
At the conclusion of the process, the order of bands produced, reading from the bottom to the top of the slide, was *C*, *B*, *A*, *D*.
Which of the components in the mixture experienced the strongest bonding to the material in the stationary phase?

THINK	WRITE
1. The stronger the (temporary) bonds are, the slower a component will move along the TLC slide.	
2. The component that moves the shortest distance will be the slowest component.	
3. Component *C* has moved the shortest distance and is, therefore, the component that experiences the strongest forces. (*Note:* This is all comparative. The forces involved are still weak forces in absolute terms.)	Component *C* is the component that experiences the strongest bonding forces.

PRACTICE PROBLEM 2

In the procedure described in sample problem 2, which component would experience the weakest bonding to the material in the stationary phase?

5.2.5 The role of polarity in chromatography

The choice of solvent used in chromatography can have a significant effect on the results obtained. Capillary action moves the mixture over the stationary phase. However, for this to occur, the mixture (and all its components) must dissolve in the solvent. Matching the polar/non-polar properties of the solvent to those in the mixture to be analysed is, therefore, critical. However, even when this is done, insufficient separation of the components may take place. It then becomes necessary to try another solvent, often with a different polarity.

For example, if unsatisfactory results are obtained using water in paper chromatography, ethanol might be used. If this is not satisfactory, other solvents or even a water–ethanol *mixture* might be utilised. This choice is often a matter of trial and error, although a knowledge of polarity and even experience certainly helps. In this respect, TLC offers more scope than paper chromatography. The latter not only (as mentioned) affords a wider range of solvent choice, but also offers a range of options for the stationary phase. For example, if a mixture undergoing TLC moves too quickly over the slide and does not experience adequate separation, a stationary phase with different polarity might be tried to increase the efficiency with which separation takes place.

5.2.6 Using R_f values to identify substances

R_f values provide a convenient way by which the bands obtained during chromatography may be compared.

R_f in chromatography, the retardation factor; the ratio of the distance travelled by the solute to the distance travelled by the solvent

In both forms of chromatography discussed in section 5.2.4, the solvent front moves quicker along the stationary phase than any component in the mixture. It is thus possible to define a fraction called the R_f (retardation factor) value as follows:

$$R_f = \frac{\text{distance moved by solute}}{\text{distance moved by solvent}}$$

These values provide a convenient means for comparison when separations are carried out under identical conditions. Take care to measure from the origin, not the bottom of the paper, to avoid getting an incorrect R_f value.

For example, imagine a pharmaceutical product *C*, that commonly contains two impurities (*B* and *A*) at various stages of its manufacture. To cheaply identify these prior to using more sophisticated forms of chromatography, TLC was used. The following results were obtained:

- Band 1: distance travelled = 3.1 cm.
- Band 2: distance travelled = 4.5 cm.
- Band 3: distance travelled = 7.3 cm.
- Solvent front: distance travelled = 9.0 cm.

A suspected impurity (*A*) was then subjected to a chromatographic determination under identical conditions to yield the following results:

- Band from *A*: distance travelled = 5.2 cm.
- Solvent front: distance travelled = 10.1 cm.

From these results, R_f values may be calculated as follows:

- R_f (band 1) = 3.1 ÷ 9.0 = 0.34.
- R_f (band 2) = 4.5 ÷ 9.0 = 0.50.
- R_f (band 3) = 7.3 ÷ 9.0 = 0.81.
- R_f (band *A*) = 5.2 ÷ 10.1 = 0.51.

From these values, it is now obvious that the band labelled number 2 is most likely to be substance *A*.

FIGURE 5.12 Another simple chromatographic technique is column chromatography, which can be used to separate differently coloured dyes. The dye mixture is poured into the top of the column, which is packed with a fine solid. Opening the tap at the bottom lets the solution flow through so the dyes can be separated.

elog-1861

EXPERIMENT 5.2 online only

Separation of food dyes using paper chromatography

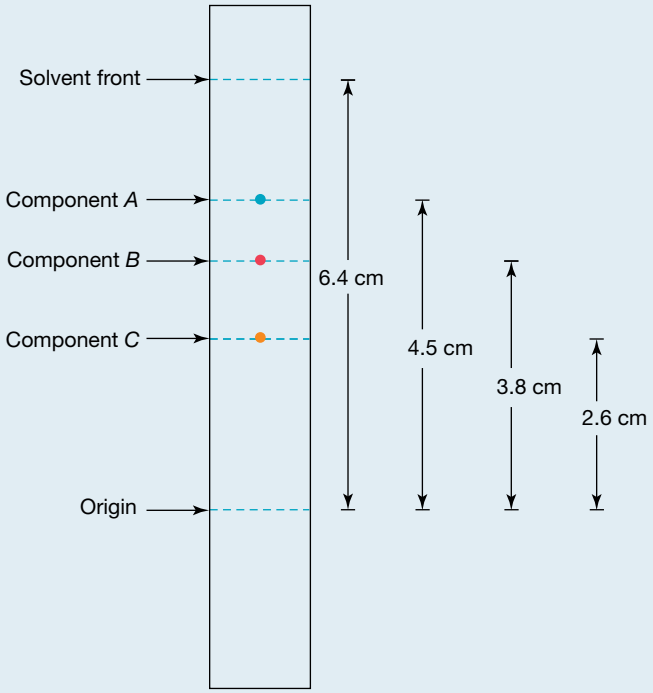
Aim

To examine the separation by paper chromatography of a number of commercially available food dyes

tlvd-3491

SAMPLE PROBLEM 3 Calculating an R_f value

Paper chromatography was used to analyse a mixture consisting of three components as shown in the provided figure.
Calculate the R_f value of component *A*.



THINK

1. Identify the distance moved by component *A*.
2. Identify the distance moved by the solvent front.
3. The R_f value is calculated using the formula:
$$R_f = \frac{\text{distance moved by solute}}{\text{distance moved by solvent}}$$

WRITE

4.5 cm

6.4 cm

$$R_f = \frac{\text{distance moved by solute}}{\text{distance moved by solvent}}$$
$$= \frac{4.5}{6.4}$$
$$= 0.70$$

PRACTICE PROBLEM 3

Calculate the R_f values for component *B* and component *C* using the figure in sample problem 3.

5.2.7 Extension: Instrumental analysis using chromatography

A tour of a modern analytical laboratory would reveal many thousands of dollars worth of instruments. Instruments undreamed of many years ago today routinely analyse materials to a higher level of accuracy than ever before. With current trends in computer design and miniaturisation, many of these instruments are expected to become even cheaper, more powerful and, hence, more widely used than they are now.

From time to time, we hear reports in the media about such accurate methods of analysis. Examples include pesticide residues in food, the swapping of beef for kangaroo meat in exports, or athletes and horses testing positive for banned drugs. Much public outcry can erupt about declining standards and increased abuse. The forgotten fact, however, is that with each new generation of analytical instruments, chemicals are being detected at very low levels — much lower than were previously detectable.

Instrumental methods of analysis have a number of advantages, including:

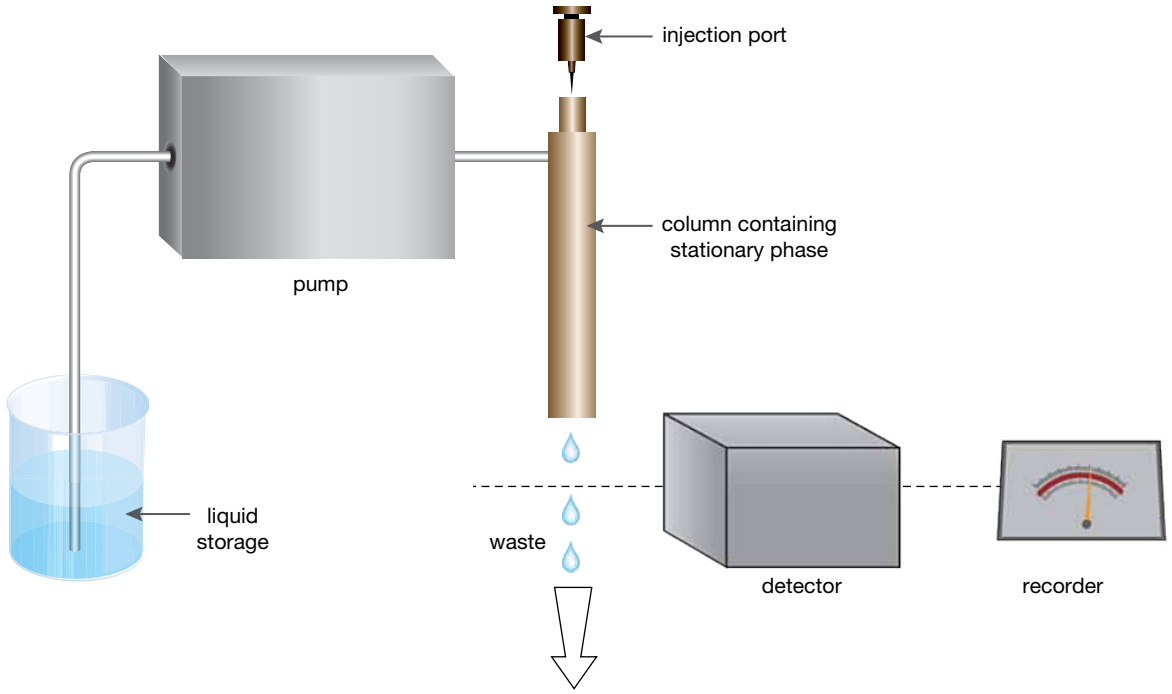
- speed (therefore allowing for testing of more samples)
- increased level of accuracy
- less room for human error
- sensitivity (many instruments require only small samples)
- some are non-destructive
- amounts in a sample may be accurately measured (referred to as quantitative determination).

High-performance liquid chromatography

Along with gas chromatography (see green case study box in this section), high-performance liquid chromatography (HPLC) is a sensitive and sophisticated application of chromatography. (HPLC is sometimes called high-pressure liquid chromatography.)

In HPLC, the mixture is 'swept' through a column that is densely packed with the stationary phase. This is done using a high-pressure liquid. The column is open ended and a suitable detector is placed at the end to detect each component of the mixture as it exits the column. For the same reasons as in paper or thin-layer chromatography, the components in the mixture move through the column at different rates and, therefore, exit the column at different times.

FIGURE 5.13 Schematic of high-performance liquid chromatography



After passing through the column, the components are detected by a suitable device and recorded on a chart as a series of peaks (this chart is called a chromatogram). In many modern instruments, this information can also be presented in tabular form, showing the **retention time (Rt)** and area of each peak. The data can then be fed into programs for graph drawing or for further mathematical evaluation.

retention time (Rt) time taken for a substance to travel through a chromatography column and exit at the end

Two important features of a chromatogram are:

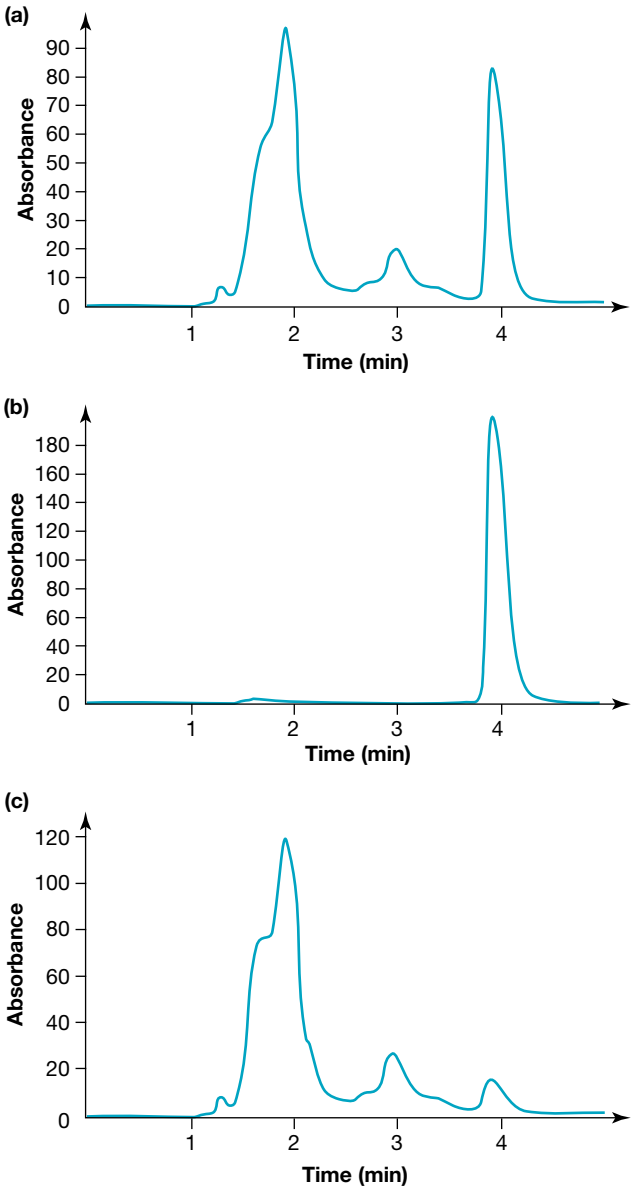
1. The positions of its peaks: the time taken for each component of the sample to travel from the injection port to the end of the column where it is detected is referred to as its retention time, Rt. This corresponds to the position of the peak on the chromatogram. Rt can be used to identify a component — the retention time for an unknown substance is compared to retention times for known substances under the same operating conditions. In this way, Rt values are the instrumental equivalent of the R_f values discussed earlier.
2. The area underneath each peak: the greater the amount of a component, the greater the area under the peak that it causes. This means the measurement can be used to quantitatively identify a substance. The quantitative use of both GC and HPLC is discussed later in this chapter. It should be noted that, when the peaks produced are narrow, the area measurement can be replaced by a measurement of the peak height.

As implied by its name, HPLC is especially suited to the analysis of liquid mixtures. In many cases, these are organic compounds (compounds containing high proportions of carbon and hydrogen) with relative molecular masses of up to 1000.

Figure 5.14 shows some results obtained from testing a brand of decaffeinated coffee. Note that a caffeine standard has been run through the instrument so that the caffeine peak can be identified. The reduction in the height of the peak due to caffeine hence becomes obvious when normal and decaffeinated coffee results are compared.

For a quantitative measurement of the amount of remaining caffeine, the height (or the area under the peak) could be measured and compared to a set of caffeine standards of known concentrations. Quantitative use is discussed in more detail in the next section.

FIGURE 5.14 HPLC chromatograms for (a) normal coffee, (b) caffeine and (c) decaffeinated coffee



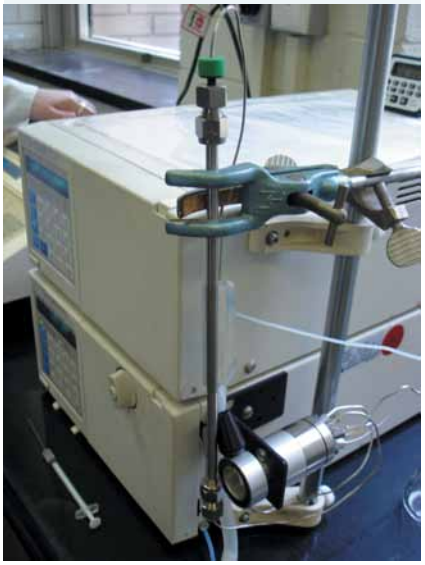
on Resources

Video eLesson High-performance liquid chromatography (HPLC) (eles-0347)

Quantitative analysis using HPLC

Instruments must be accurately calibrated, using a set of standards, before any meaningful quantitative results can be obtained. In each case, this involves running a set of standards of known concentration through the instrument and noting the reading. From these results, a graph of reading versus concentration can be drawn. It is then simply a matter of running the sample to be tested through the instrument, and noting the reading. The calibration graph can then be used to obtain its concentration.

FIGURE 5.15 A high-performance liquid chromatograph



Most modern instruments are programmed to do this automatically. They use statistical algorithms to determine the equation for the line of best fit using the calibration data. The test result for the unknown is then simply fed into this equation and the result displayed.

The following scenario illustrates how this might be done.

Imagine a new brand of throat lozenges called 'Throat Eze' makes the claim that each lozenge contains 1.2 mg of dichlorobenzyl alcohol (DCBA). To test this claim, a government analyst dissolved the lozenge in 500 mL of a solvent made from water and ethanol. A small sample was then injected into a HPLC. A chromatogram containing a large number of peaks was obtained.

The operator then ran a series of DCBA standards of known concentration through the instrument. Chromatograms for each standard were obtained, as well as a measure of the area under each of the reference peaks. The results obtained are shown in table 5.4.

TABLE 5.4 Results from HPLC analysis of dichlorobenzyl alcohol

Concentration of standard (mg L^{-1})	Area under peak (arbitrary units)
1.0	83
2.0	160
3.0	241
4.0	315
Lozenge extract	193

The first step would be to identify which peak on the chromatogram of the sample is due to DCBA. To do this, you would look at the chromatogram of any one (or all) of the standards. As all of these contain *only* DBCA, the retention time, Rt, can be noted. It is then only necessary to look at the samples chromatogram and note the peak that has the same Rt. It is this peak that is then 'measured' to obtain the tabulated value in table 5.4.

The next step is to graph the concentration values against the reading obtained for each peak. This graph is called a calibration curve. It should not be confused with the graph that forms the chromatogram, which is a graph of retention time versus amount (instrument reading).

Figure 5.16 displays the calibration curve obtained from the results shown in table 5.4.

The third step is simply to read the graph 'backwards'. From the lozenge extract reading of 193, simply read across to the line of the calibration curve, and then down. As seen in figure 5.16, a concentration value of 2.4 mg L^{-1} is obtained.

Remembering that the method involved one lozenge being dissolved in 500 mL of solvent, this value would be equivalent to 1.2 mg/500 mL, meaning that the claim made for each lozenge is correct.

FIGURE 5.16 The calibration curve showing the concentrations of DCBA.

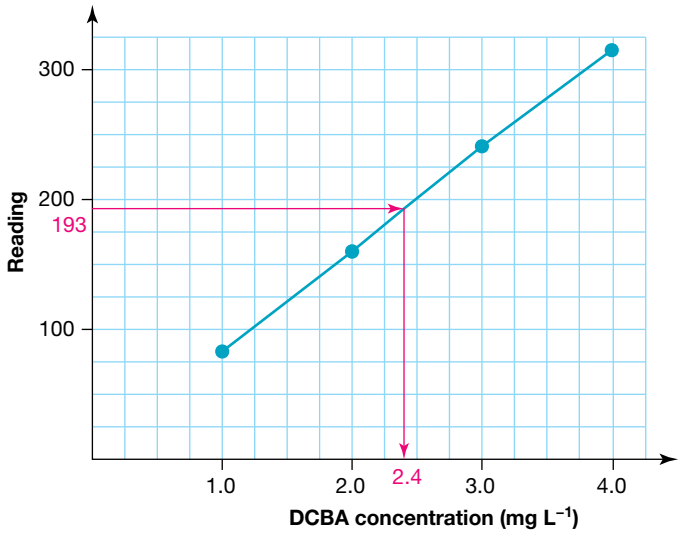
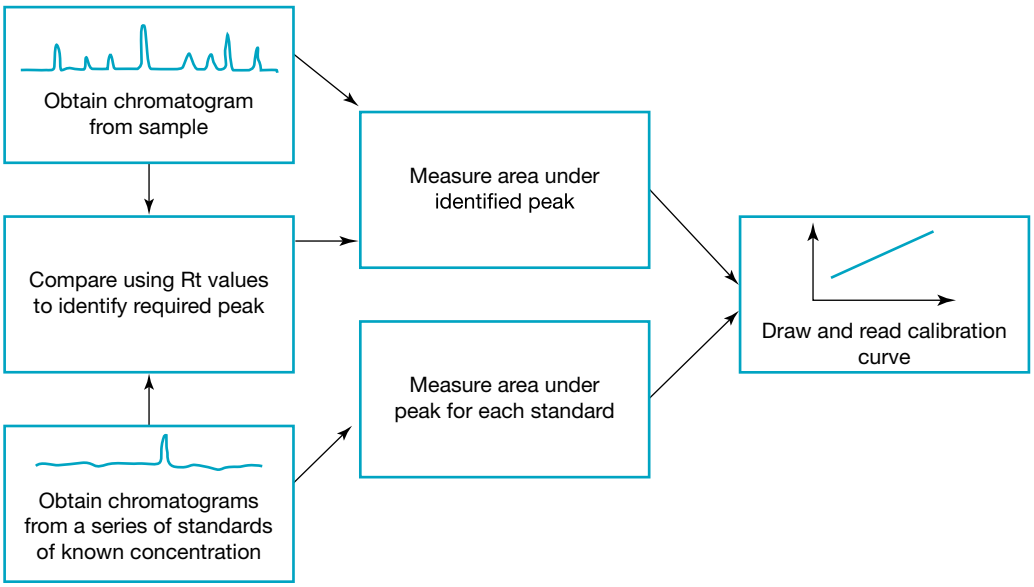


Figure 5.17 summarises this procedure.

FIGURE 5.17 Steps in using a HPLC for qualitative and quantitative analysis



elog-1862

tlvd-4543

EXPERIMENT 5.3

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Separating mixtures using column chromatography

Aim

To separate a mixture of dyes using column chromatography

Resources

Video eLesson Calibration curves (eles-0348)

EXTENSION: Gas chromatography

Gas chromatography (GC) is very similar to HPLC, except that the mixture being analysed is first vaporised and the ensuring separation then takes place in the gas phase. In a similar fashion to HPLC, GC uses a column but in this case it is placed inside an oven so the mixture remains as a gas.

Of all the forms of chromatography, GC is the most sensitive, being able to detect quantities as small as 10^{-12} g from a microlitre (10^{-6} L) sample. In operation, a stream of inert gas such as nitrogen or helium carries the vaporised sample (mobile phase) through the column. The column contains a liquid with a high boiling point, coated onto fine granules of an inert solid. This is why this technique is also sometimes called gas-liquid chromatography (GLC). This column is the stationary phase and performs the same function as in any other type of chromatography — that is, it sorts out a mixture into its components.

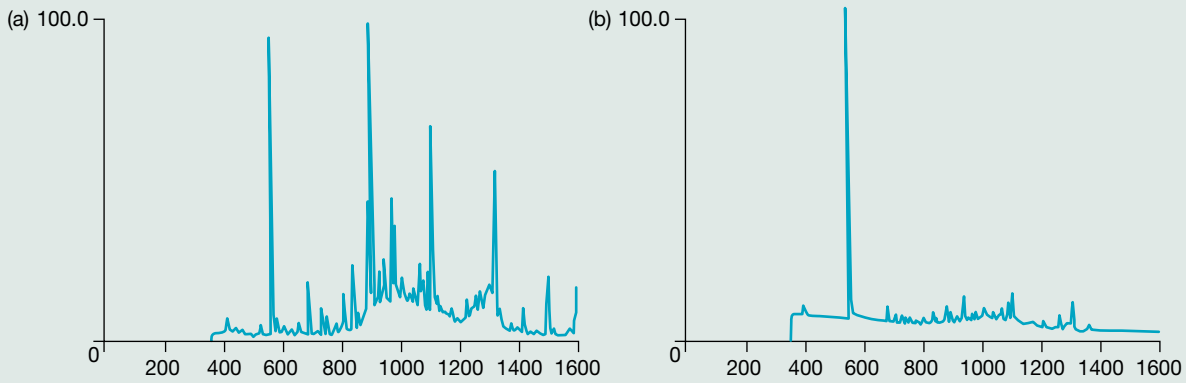
After passing through the column, the components are detected by a suitable device and recorded as a chromatogram. The interpretation and analysis of this chromatogram is exactly the same as for HPLC.

Gas chromatography is ideally suited to analysing liquids that are easily vaporised. Substances harder to vaporise present a problem, and if their boiling temperatures are higher than the coated liquid forming the stationary phase, this method may not be suitable.

Many of the substances that are analysed using GC are organic (that is, compounds of carbon). They typically have relative molecular masses that are less than 300.

Two examples of gas chromatograms are shown in figure 5.18. Figure (a) shows a urine sample from a racing greyhound that was suspected of having been administered amphetamine, while figure (b) is from an amphetamine standard under the same operating conditions. A comparison of the two reveals clearly that the dog has tested positive.

FIGURE 5.18 Gas chromatograms from (a) a racing greyhound and (b) an amphetamine standard



5.2 Activities

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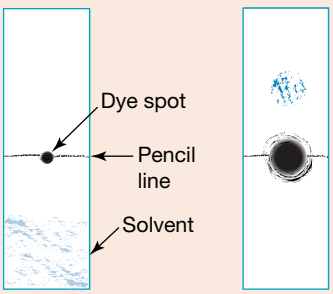
5.2 Quick quiz on | 5.2 Exercise | 5.2 Exam questions

5.2 Exercise

1. MC Waterproofing of certain types of raincoats is carried out by applying a solution of paraffin wax in a propriety solvent called Soluwax.
   Paraffin wax is a solid and consists of large hydrocarbon molecules.
   Why is this process effective?
   A. Water is polar and Soluwax is non-polar.
   B. Water is polar and Soluwax is polar.
   C. Water is non-polar and Soluwax is polar.
   D. Water is non-polar and Soluwax is non-polar.
2. MC The amounts of sucrose (table sugar) able to be dissolved in 1 L of solvent was measured in an experiment.
   The results are shown in the provided table.

Solvent	Amount of sucrose (g) dissolved in 1 L of solvent
W	150.0
X	0.1
Y	5.9
Z	11.8

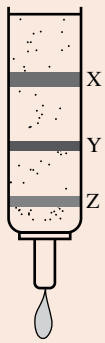
   What is the order of polarity in these solvents, from most to least?
   A. *W*, *X*, *Y*, *Z*
   B. *W*, *Z*, *Y*, *X*
   C. *Z*, *Y*, *X*, *W*
   D. *W*, *Y*, *Z*, *X*
3. Why is it not possible to have an R_f value greater than one?
4. Define the following.
   a. Mobile phase
   b. Stationary phase
   c. Solvent
5. In chromatography, why is it important that the stationary phase has a large surface area?
6. In an experiment, a student used paper chromatography in an effort to identify the different dyes in a sample of pen ink. After a while, the student removed the strip of paper to reveal the result shown in the diagram.
   a. Why is the line near the bottom of the strip drawn in pencil?
   b. Why does the solvent need to be below the pencil line?
   c. As shown in the figure, some dye remained at the original spot. Suggest a modification to the experiment that would establish whether this remaining spot was actually a single dye or a mixture of more dyes.



5.2 Exam questions

Question 1 (1 mark)

MC In column chromatography, a column is packed with the chosen stationary phase and the mixture to be separated is then placed on this at the top of the column. The chosen solvent is then added continually at the top of the column so that the mixture drains down through the column.



A mixture was passed down through a column containing a polar stationary phase and washed through using a less polar mobile phase. As it passed through the column, the mixture separated into three components: *X*, *Y* and *Z*. Based on the diagram, what statement is correct?

A. Component *Z* was more strongly adsorbed to the stationary phase than the other two.
B. Component *Y* moved through the column faster than *Z* but more slowly than *X*.
C. Component *X* had a stronger attraction to the mobile phase than *Y* and *Z*.
D. Components *Y* and *Z* are less polar than *X*.

Question 2 (5 marks)

A student analysed the red dye extracted from a number of lollies using the technique of TLC. The conditions for each experiment were identical, except for the time that each chromatogram was allowed to run. The provided table shows the obtained chromatography results for the lollies.

	Height of red dye front (mm)	Height of solvent front (mm)
Lolly A	23	70
Lolly B	45	65
Lolly C	33	55
Lolly D	36	60

a. Calculate the R_f value for each red dye extract. **(4 marks)**
b. Which lollies contain the same sort of dye? **(1 mark)**

Question 3 (9 marks)

A pharmaceutical company used TLC to test the purity of a drug at a particular stage of its manufacture. This drug was known to contain a mixture of two substances at this stage of the process. It was also expected that two possible impurities might be present.

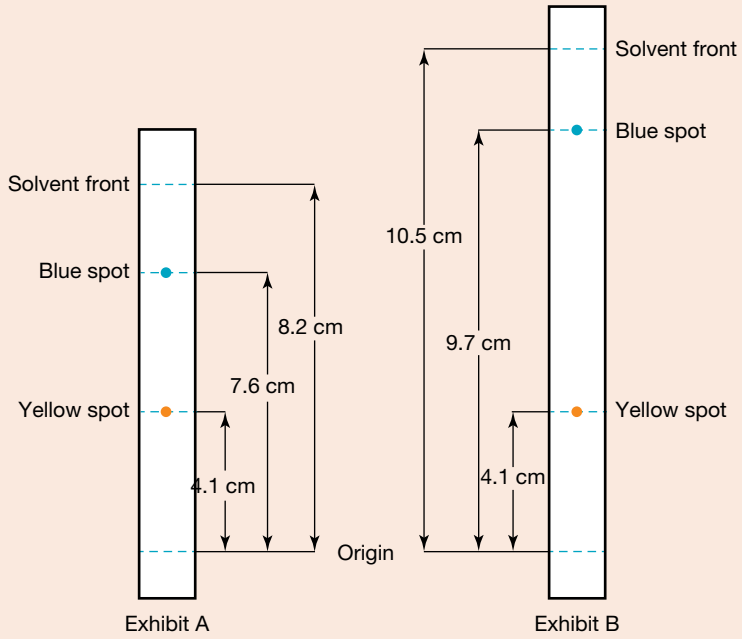
The procedure involved using TLC on a sample of the drug, and then using TLC under exactly the same conditions on separate samples of each of the possible impurities. The results are shown in the table provided.

Chromatography band	Distance travelled (cm)
Band 1	3.1
Band 2	4.5
Band 3	7.3
Solvent front (mixture)	9.0
Impurity *A*	5.2
Solvent front (impurity *A*)	10.1
Impurity *B*	6.0
Solvent front (impurity *B*)	9.5

a. Explain how the results from the sample alone indicate that the sample is not pure. **(1 mark)**
b. Explain how the results may be used to identify which impurity is present. **(2 marks)**
c. Using the method you outlined in part b and showing your working, identify which impurity is present. **(6 marks)**

Question 4 (7 marks)

As part of a forensic science investigation, whether samples of food dye taken from two different parts of a crime scene are the same needs to be determined. Samples from each location were collected and labelled *exhibit A* and *exhibit B*. Back in the laboratory, an analyst used the technique of paper chromatography to try to solve this question. The diagram provided shows the strips of paper produced after the procedure was carried out.



a. Explain why the colours produced travel up the paper at different rates. **(1 mark)**

b. On the basis of these results, what conclusions can be drawn about the following?

i. The yellow component of each exhibit **(2 marks)**

ii. The blue component of each exhibit **(2 marks)**

iii. The likelihood that *exhibit A* and *exhibit B* are the same dye **(2 marks)**

Question 5 (5 marks)

A student is designing an experiment using paper chromatography. From previous results, they know that the R_f values for two dyes they wish to investigate are 0.72 and 0.80.

It is proposed that the dyes be run on the same piece of paper and, for various reasons, that the spots should have a separation of at least 1 cm at the end of the procedure.

What is the minimum height of the chromatography paper the student should use in this experiment?

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5.3 Review

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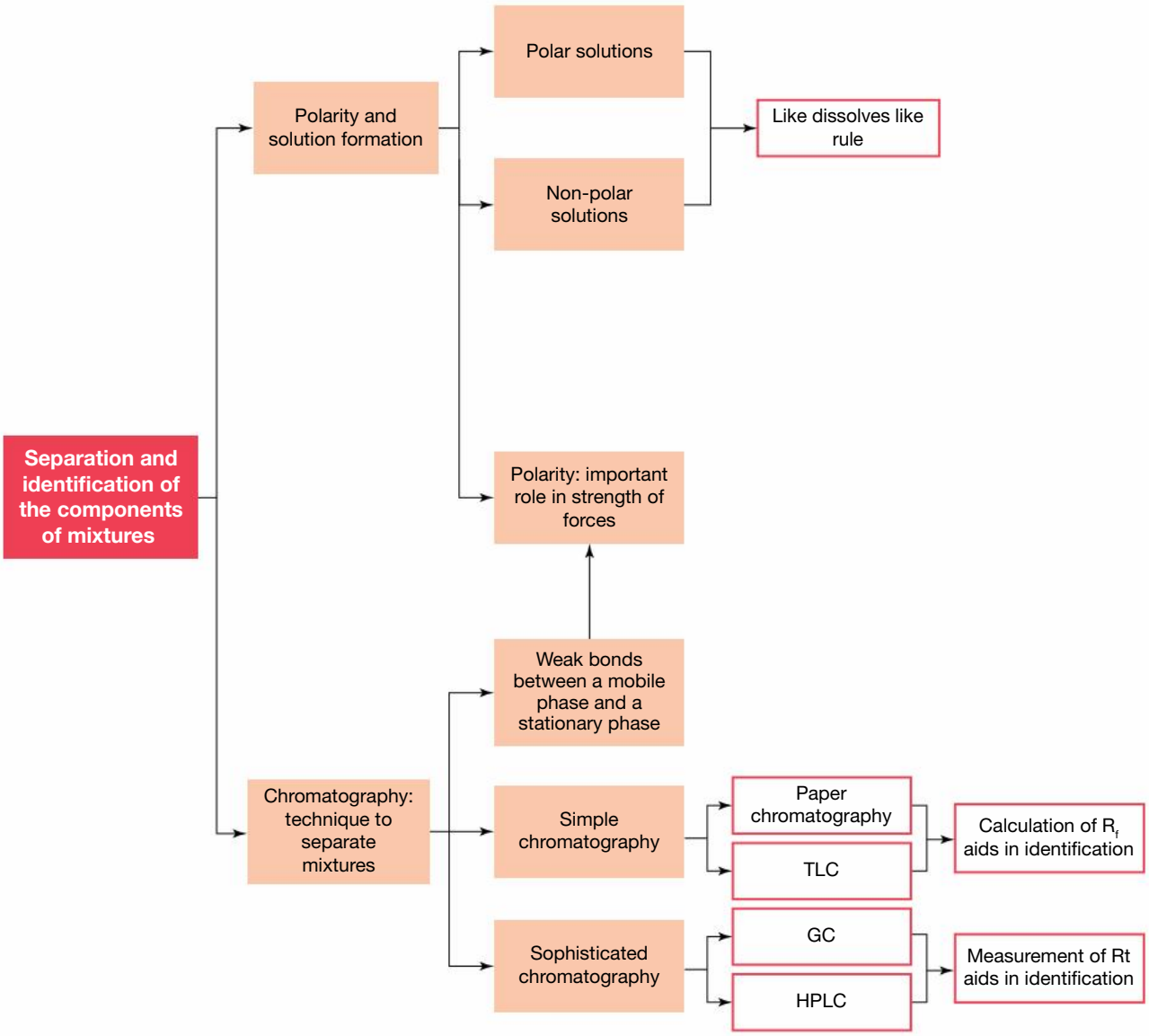
5.3.1 Topic summary



5.3.2 Key ideas summary

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5.3.3 Key terms glossary

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Solutions	Solutions — Topic 5 (sol-0804)
Practical investigation eLogbook	Practical investigation logbook — Topic 5 (elog-1610)
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Exam question booklet	Exam question booklet — Topic 5 (eqb-0086)

5.3 Activities

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5.3 Review questions

1. Many organic chemicals are soluble in water while many others are insoluble. What features in their molecular structures account for this difference?
2. An examination of the formula of a complicated organic molecule reveals that it contains carbon, hydrogen, iodine and sulfur. Is this molecule likely to be water soluble? Explain.
3. A student used paper chromatography to produce a chromatogram from black ink, as shown in the figure. The ink was added to the paper near the bottom, and water was used as the mobile phase.
   a. List the colours shown in order from most strongly adsorbed to least strongly adsorbed.
   b. Predict what would happen if a longer piece of paper was used.

4. Compare and contrast the techniques of paper chromatography and TLC by listing advantages and disadvantages of each technique.
5. The level of glucose, $C_6H_{12}O_6$, in a sample of human blood was analysed by HPLC. After suitable extraction, the sample was analysed, along with a set of known standards. The results are shown in the provided table.

   a. Draw the calibration curve using the data shown.
   b. What is the concentration of glucose in the blood sample in g L^{-1}?
   The normal range for blood glucose is between 0.8 g/L and 1/1 g/L.
   c. Is the glucose within the normal range?

Glucose concentration (g L^{-1})	Peak area
0.4	1650
0.8	3280
1.0	4100
1.2	4920
1.6	6550
Sample	3840

5.3 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC A student has accidentally written on a whiteboard using a permanent marker. After unsuccessfully trying to remove the writing using water, he notes that the permanent marker contains xylene as its solvent.

From these observations, which of the following statements would represent an *invalid* conclusion?

A. The ink in the permanent marker is non-polar.
B. Xylene is a non-polar solvent.
C. The writing might be removed using a less polar solvent than water.
D. A more polar solvent than water should be used to remove the writing.

Question 2

MC An artist has accidentally spilt two different liquids on a recently completed oil painting that is still wet. One liquid was water while the other was hydrocarbon spirit.

Which liquid would cause the most damage to the painting?

A. Water
B. Hydrocarbon spirit
C. Both would cause equal amounts of damage
D. Neither liquid would cause any damage

Question 3

MC A water-soluble compound (labelled *A*) and another compound (labelled *B*) are tested for how easily they dissolve in three different solvents (*X*, *Y* and *Z*).

The results are shown in the table provided. The numbers represent the relative ease of dissolving, with 1 representing the best and 3 representing the worst.

	X	*Y*	*Z*
A	3	1	2
B	1	3	2

The order of polarity in the solvents (from most to least) is:

A. *Y*, *Z*, *X*, and *B* is non-polar.
B. *X*, *Z*, *Y*, and *B* is non-polar.
C. *Y*, *Z*, *X*, and *B* is polar.
D. *X*, *Z*, *Y*, and *B* is polar.

Question 4

MC In carrying out an experiment on a mixture using paper chromatography, a student obtained a chromatogram that showed only partial separation of some of the components near the top of the paper.

Which of the following features of the method would be the best to change in an effort to overcome this problem?

A. Allow the chromatogram to develop for a longer time
B. Use a wider piece of chromatography paper
C. Use a more intense spot of the original mixture at the origin
D. Use a different solvent or solvent mixture

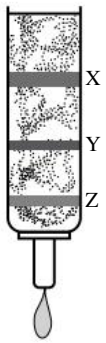
Question 5

MC To accurately identify a substance using paper chromatography, what must happen?

A. A reference chromatogram of the suspected substance must be produced using identical conditions.
B. Significant separation of each component in the mixture must occur so as to get accurate R_f values.
C. The solvent front must move fast enough to stay ahead of the component that is to be identified.
D. A colourless solvent must be used.

Question 6

MC Column chromatography is similar to paper and TLC and is performed in a column as shown.



The origin is at the top of the column where the mixture is placed on the stationary phase previously packed into the column.

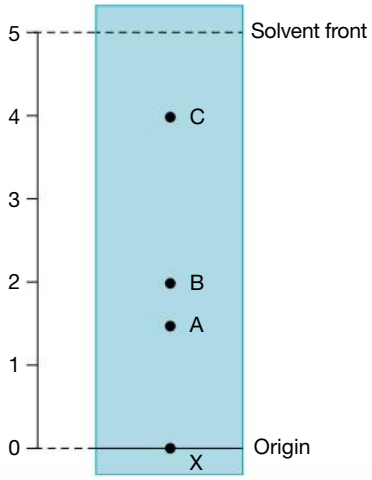
A mixture of components *X*, *Y* and *Z* was passed through a chromatography column containing a non-polar stationary phase and a polar mobile phase. Based on the diagram shown, which of the following statements are correct?

- Statement I: *Z* is the most polar of the three components.
- Statement II: The first component to pass through the column is component *Z*.
- Statement III: *Y* is more strongly adsorbed than *X* but less strongly adsorbed than *Z*.

A. Statements I and II only
B. Statements I and III only
C. Statements II and III only
D. All three statements are correct.

Use the following information to answer questions 7 and 8.

The provided diagram shows a chromatogram obtained using paper chromatography. The sample being tested was originally 'spotted' at position X.



Question 7

MC Which of the following statements is true?

A. Component A is the least strongly adsorbed.
B. Component B would have a lower R_f value than component A.
C. Component C has the highest R_f value.
D. Position X must be below the solvent level to achieve adequate separation.

Question 8

MC What are the R_f values of components A, B and C respectively?

A. 3.3, 2.5, 1.25
B. 0.3, 0.4, 0.8
C. 1.5, 2.0, 4.0
D. 0.3, 0.5, 0.7

Question 9

MC In an experiment using paper chromatography, it was noted that a chemical with a known R_f value of 0.75 had travelled a distance of 4.5 cm from the origin. What would be the distance of the solvent front from the origin (in cm)?

A. 0.2
B. 3.4
C. 6.0
D. 8.0

Question 10

MC In the quantitative use of a HPLC, which of the following would not be done?

A. Measuring the area of the peaks produced
B. Drawing a graph of concentration versus time
C. Drawing a graph of concentration versus peak area
D. Standardising the instrument's readings against known concentrations

Section B — Short answer questions

Question 11 (1 mark)

In a paper chromatography exercise, a mixture of dyes was spotted onto a suspended piece of paper that was just touching a container of water. The water was allowed to soak up through the paper, separating the dye mixture. In this case, what was the mobile phase?

Question 12 (4 marks)

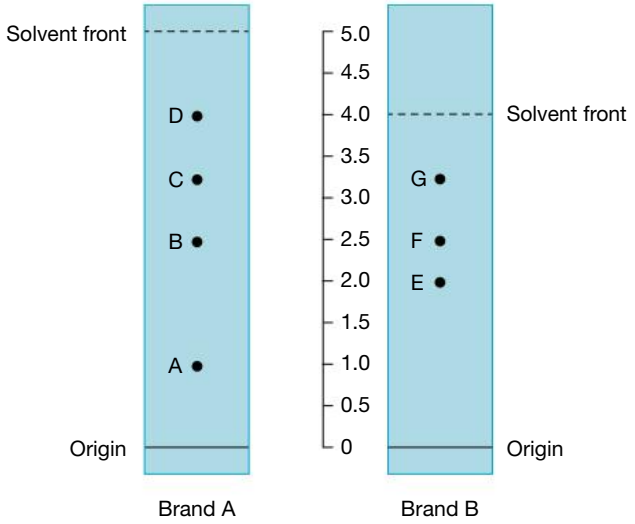
The dyes present in red and purple jellybeans were extracted and then separated using paper chromatography. The following results were obtained:

- Red jellybeans contained erythrosine (R_f = 0.48) and green S (R_f = 0.11).
- Purple jellybeans contained carmoisine (R_f = 0.08) and crocein orange (R_f = 0.58).

Draw the chromatogram for each jellybean.

Question 13 (12 marks)

A student was curious as to whether two different brands of writing ink contained the same dyes. In order to investigate this, she obtained two chromatograms under identical conditions, using paper chromatography.



a. Calculate the R_f values for each of the seven components shown. **(7 marks)**

b. Which ink component was the most strongly adsorbed? **(1 mark)**

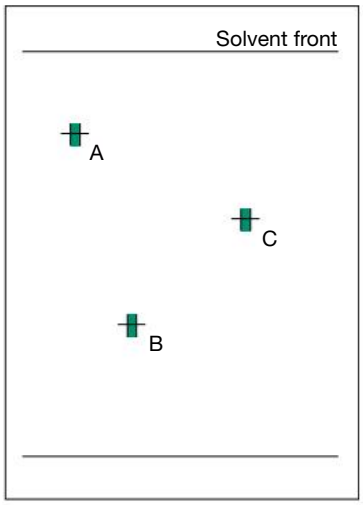
c. Which ink component was the least strongly adsorbed? **(1 mark)**

d. Which components are common to both brands of ink? **(2 marks)**

e. What is the advantage of using R_f values in this experiment? **(1 mark)**

Question 14 (3 marks)

A mixture of amino acids may be separated by chromatography. In one such experiment, the chromatogram shown was obtained by using TLC. (At the end of the experiment, the plate was sprayed with a 2 per cent ninhydrin solution to make the amino acids visible.) The appearance of the plate is shown.



The following table gives the R_f values of some amino acids under the conditions of this experiment.

Amino acid	R_f value
Alanine	0.56
Arginine	0.61
Isoleucine	0.82
Proline	0.89
Taurine	0.34
Valine	0.75

Which amino acids are present in the mixture analysed?

Question 15 (20 marks)

In autumn, the leaves of deciduous trees change colour as they prepare to lose their leaves. These changes in colour occur because the various coloured pigments in the leaves alter their amounts relative to each other.

In order to investigate this phenomenon further, a VCE student obtained some leaves from a liquidamber tree during March that were still green, as well as some that had changed to red. The coloured material was extracted from each leaf and immediately analysed using paper chromatography. The results are shown in the table provided.

Leaf colour	Colour of separated pigment	Distance of pigment from origin (cm)	Distance of solvent front from origin (cm)	R_f value
Green	Pale yellow	5.1	5.3	
	Dark yellow	4.1	5.2	
	Light green	3.2	5.3	
	Dark green	2.6	5.4	
Red	Pale yellow	7.2	7.5	
	Dark yellow	5.8	7.4	
	Red-blue	3.4	7.4	

Samples of known pigments were then also subjected to paper chromatography, under the same conditions as the extracts, and their R_f values calculated. The results obtained are shown.

Pigment	Colour of pigment	R_f value
Carotene	Yellow	0.96
Phaeophytin	Yellow-grey	0.82
Xanthophyll	Yellow-brown	0.79
Chlorophyll b	Yellow-green	0.61
Chlorophyll a	Blue-green	0.48
Anthocyanin	Red	0.46

a. Complete the first table by calculating the R_f values for each component that was separated from the leaves. **(7 marks)**

b. Identify the pigment for each component in the first table, and summarise the changes that have occurred in one or two sentences. **(9 marks)**

c. On the basis of the R_f values quoted, it appears that two pairs of pigments might be difficult to separate and identify. List these two pairs. **(2 marks)**

d. Using other information from the table, one of these pairs should be easy to distinguish. Which pair is this and why? **(2 marks)**

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AREA OF STUDY 1 How do the chemical structures of materials explain their properties and reactions?

OUTCOME 1

Explain how elements form carbon compounds, metallic lattices and ionic compounds, experimentally investigate and model the properties of different materials, and use chromatography to separate the components of mixtures.

PRACTICE EXAMINATION

STRUCTURE OF PRACTICE EXAMINATION		
Section	**Number of questions**	**Number of marks**
A	20	20
B	7	30
	Total	**50**

Duration: 50 minutes

Information:

- This practice examination consists of two parts; you must answer all question sections.
- Pens, pencils, highlighters, erasers, rulers and a scientific calculator are permitted.
- You may use the VCE Chemistry Data Book for this task.

on Resources

Weblink VCE Chemistry Data Book

SECTION A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

1. Which of the following statements about the two particles $^{59}_{27}X$ and $^{59}_{28}Z^{+}$ is correct?

A. They are a pair of isotopes of the same element.

B. They have the same number of neutrons.

C. Particle *Z* is significantly heavier than particle *X*.

D. They contain the same number of electrons.

2. What could the electron configuration $1s^2 2s^2 2p^6 3s^1 3p^6$ represent?

A. A ground state atom of argon, Ar

B. A ground state chloride ion, Cl^-

C. An excited atom of chlorine, Cl

D. An excited potassium ion, K^+

3. Which of the following statements about trends in the periodic table are true?
   - **I.** Atomic radius increases going across a period in the periodic table.
   - **II.** Both ionisation energy and non-metallic character increase across a period in the periodic table.
   - **III.** Metallic character increases going down a group and decreases going across the periodic table.
   - **IV.** Metallic character decreases going down a group and increases going across the periodic table.

   A. Only statements (II) and (III) are correct.
   B. Only statements (I) and (II) are correct.
   C. Only statements (I) and (III) are correct.
   D. Only statements (I) and (IV) are correct.
4. Where would an element with a relatively low first ionisation energy and a relatively high reactivity be found?
   A. *s* block
   B. *p* block
   C. *d* block
   D. *f* block
5. When element Q from group 16 of the periodic table reacts with element J from group 1, what compound is formed?
   A. Ionic with the empirical formula Q_2J
   B. A metal alloy with the empirical formula JQ
   C. Ionic with the empirical formula J_2Q
   D. Ionic with the empirical formula JQ
6. Which of the following best describes a molecular compound?
   A. Ions held together by the electrostatic attraction between positive and negative ions
   B. Atoms containing elements that differ greatly in electronegativity
   C. Atoms in which the elements share electrons
   D. A lattice of non-metallic atoms held together by shared electrons
7. Both bonding and non-bonding electrons are present in the valence structures of covalent compounds. How many non-bonding electrons are present in the structure of methane, CH_4?
   A. 0
   B. 2
   C. 4
   D. 6
8. The following table shows the electronegativity values of five different elements.

Element	*W*	*X*	*Y*	*Z*
Electronegativity	0.8	2.2	3.5	4.0

   Based on the information in the table, what would be expected?
   A. A bond formed between *Y* and *Z* would be strongly polar due to the high electronegativity values.
   B. The compound formed from elements *Z* and *W* would be ionic in nature.
   C. A bond formed between *X* and *Z* would be polar with the more positive end found at *Z*.
   D. The bond formed between *W* and *X* would be more polar than the one formed between *X* and *Z*.
9. The shape of a molecule can have a marked effect on its polarity. Which of the following molecules contains polar bonds but is an overall non-polar molecule?
   A. The linear molecule HF
   B. The tetrahedral molecule $SiCl_4$
   C. The angular molecule H_2S
   D. The pyramidal molecule PBr_3

10. Covalent solids can be classified as covalent network (3D) lattices, covalent layer (2D) lattices or molecular lattices. Which of the following has the substances correctly assigned?

	Covalent network lattice	Covalent layer lattice	Molecular lattice
A	Diamond, C	Graphite, C	Carbon dioxide, CO_2
B	Graphite, C	Ice, H_2O	Carbon dioxide, CO_2
C	Graphite, C	Diamond, C	Carbon dioxide, CO_2
D	Carbon dioxide, CO_2	Graphite, C	Ice, H_2O

A. Option A
B. Option B
C. Option C
D. Option D

11. A student researches the properties of three substances: magnesium sulfate, $MgSO_4$; sulfur dioxide, SO_2; and magnesium metal, Mg. Which of the following statements regarding these substances is correct?
A. Both Mg and $MgSO_4$ have high melting points but only $MgSO_4$ conducts electricity in its molten form.
B. Neither $MgSO_4$ nor SO_2 would conduct electricity in the solid state.
C. Both SO_2 and $MgSO_4$ have high melting points and conduct electricity in their liquid state.
D. Both SO_2 and Mg can conduct electricity in either the liquid or solid state.

12. Fluoromethane, CH_3F, can be solidified if cooled to –138 °C. Apart from dispersion forces, what bond types would be present in a sample of frozen fluoromethane?
A. Both covalent bonds and hydrogen bonding
B. Hydrogen bonding only
C. Covalent bonding and dipole–dipole interactions
D. Covalent bonding only

13. Consider the following qualities of metals *L*, *M* and *Q*:
- Metal *L* tarnishes quickly in air and bubbles in cold water, producing hydrogen gas.
- The lustrous quality and inertness of metal *M* makes it typical for use in jewellery.
- Although metal *Q* does not react with water, it will react in dilute acid.

What is the order of the metals in order of reactivity from the most reactive to the least reactive?
A. $M > Q > L$
B. $Q > L > M$
C. $L > M > Q$
D. $L > Q > M$

14. When compared to a linear economy, a circular economy
A. does not involve mining materials from the earth.
B. does not finish with the disposal of an object or product.
C. does not involve the distribution of an object or product.
D. always makes new objects or products of the same type as the original.

15. Which of the following lists contains only ionic compounds?
A. CH_4, HCl, NaCl
B. H_2O, $CuSO_4 \cdot 5H_2O$, $MgSO_4 \cdot 7H_2O$
C. BaF_2, AlF_3, OF_2
D. $BaCl_2 \cdot 2H_2O$, FeS, $AgNO_3$

16. A student tested an ionic compound for electrical conductivity and found that it did not conduct as a solid, but did when dissolved in distilled water. What is the best reason to explain this?
A. When dissolved, the delocalised electrons are free to move.
B. Dissolving allows the conducting metal cations to separate from the non-conducting non-metal anions.
C. Cations and anions held in the solid ionic lattice can move freely once the lattice is dissolved.
D. The ionic compounds gain electrical conductivity from the surrounding water molecules.

17. When a solution of sodium hydroxide is added to a solution of copper(II) nitrate, a blue gelatinous precipitate is formed. What happened in this reaction?

A. Copper(II) sulfate forms the precipitate, and both sodium and nitrate ions are spectator ions.

B. Sodium nitrate forms the precipitate, and hydroxide ions are spectator ions.

C. Copper(II) hydroxide forms the precipitate, and sodium ions are one of the spectator ions.

D. Copper metal is deposited, and sodium and nitrate ions are spectator ions.

18. When solutions of lead(II) nitrate and potassium iodide are mixed, a precipitate forms. What is the balanced ionic equation for this reaction?

A. $Pb^{+}(aq) + I^{-}(aq) \rightarrow PbI(s)$

B. $Pb^{2+}(aq) + 2I^{-}(aq) \rightarrow PbI_2(s)$

C. $KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + KNO_3(s)$

D. $K^{+}(aq) + NO_3^{-}(aq) \rightarrow KNO_3(s)$

19. Which of the following would be expected to be the *least* soluble in water?

A. Ammonia, NH_3

B. Methane, CH_4

C. Potassium nitrate, KNO_3

D. Glucose, $C_6H_{12}O_6$

20. A TLC experiment was conducted where a polar material was coated onto a glass slide to serve as the stationary phase. Subsequently, four components of roughly equal molar mass were separated and their R_f values calculated as follows:
Component *W* $R_f = 0.72$
Component *X* $R_f = 0.45$
Component *Y* $R_f = 0.36$
Component *Z* $R_f = 0.95$
Which of the components was the most polar?

A. *W*

B. *X*

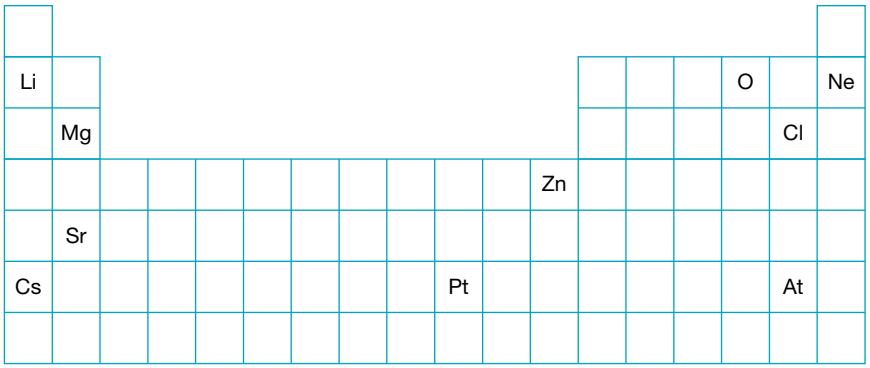
C. *Y*

D. *Z*

SECTION B — Short answer questions

Question 21 (4 marks)

Use only the elements marked in the following periodic table to match the descriptions in the questions. Each element can be selected only once.



a. What is the most electronegative element? **(1 mark)**
b. What is the halogen with the most metallic character? **(1 mark)**
c. Which element is filling its 5*s* subshell? **(1 mark)**
d. Which element is closest in chemical properties to fluorine? **(1 mark)**

Question 22 (3 marks)

Electronegativity is a property that shows distinct trends in the periodic table.

a. Define what is meant by the term 'electronegativity'. **(1 mark)**
b. Electronegativity is found to increase going across a period in the periodic table. Explain this trend. **(1 mark)**
c. Elements with a large atomic radius tend to have lower electronegativities than smaller ones. Explain why this occurs. **(1 mark)**

Question 23 (4 marks)

Complete the following.

a. Draw a diagram representing a metal lattice, and describe the components of the lattice. **(1 mark)**
b. Explain why solid metals are malleable whereas ionic solids are not. **(2 marks)**
c. Both metals and ionic compounds usually have high melting points. Explain why this is the case. **(1 mark)**

Question 24 (6 marks)

The following table lists two molecules.

Complete the following for each molecule.

a. Draw the structure giving its shape according to VSEPR theory and including any non-bonding electrons. **(2 marks)**
b. Classify each as polar or non-polar and mark in the δ– and δ+ regions on opposite ends of any of the polar structures. Leave blank if non-polar. **(2 marks)**
c. Classify the shape of the molecule. **(2 marks)**

Molecule	Nitrogen tri-iodide, NI_3	Sulfur dichloride, SCl_2
Structure of molecule showing shape and non-bonding electrons		
Polar or non-polar		
Shape		

Question 25 (2 marks)

From the following list, identify the process that breaks the type of bond given in the table.

Hint: not all processes will have these bond types broken.

Process:
- Burning methane gas
- Boiling decane
- Evaporating liquid nitrogen
- Melting tin
- Evaporating ethanol
- Melting magnesium *sulfate*

Bond type	Process
Covalent	
Hydrogen bonding	

Question 26 (6 marks)

Nearly all solutions of transition metal compounds form precipitates when mixed with aqueous potassium carbonate.

a. i. Write the full and the ionic equations for the reaction between solutions of iron(II) nitrate and potassium carbonate. Include symbols of state. **(2 marks)**

ii. List the formulas of the spectator ions in this reaction. **(1 mark)**

b. i. Write the full and the ionic equations for the reaction between solutions of iron(III) chloride and potassium carbonate. Include symbols of state. **(2 marks)**

ii. List the formulas of the spectator ions in this reaction. **(1 mark)**

Question 27 (5 marks)

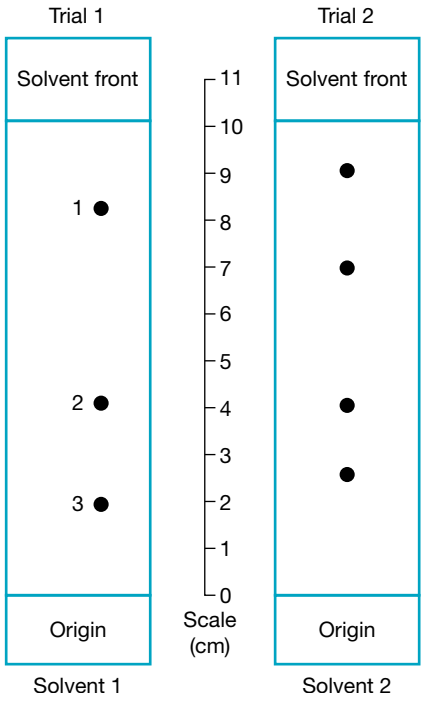
A dye was believed to contain some of the substances listed in the provided table.

To establish which substances were present, two trials using TLC were performed. The second trial used a different stationary phase and a different solvent from the first.

The provided table lists the R_f values established separately for each possible component under the conditions of trial 1 and trial 2.

Substance	R_f value under trial 1 conditions	R_f value under trial 2 conditions
U	0.05	0.80
V	0.60	0.50
W	0.81	0.42
X	0.20	0.70
Y	0.79	0.25
Z	0.41	0.90

The results of each trial are shown in the provided figure.



(a) Why are the R_f values different for solvent 1 and solvent 2? **(1 mark)**

(b) What conclusion can be drawn about spot 1 in trial 1? **(1 mark)**

(c) What are the identities of spots 2 and 3 in trial 1? **(2 marks)**

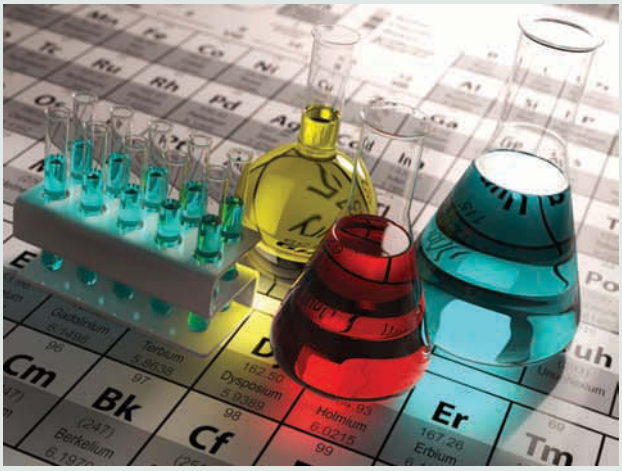
(d) After trial 2, what conclusion can now be drawn about spot 1 in trial 1? **(1 mark)**

PRACTICE SCHOOL-ASSESSED COURSEWORK

ASSESSMENT TASK — DATA ANALYSIS

In this task, you will focus on the trends of the periodic table and their relationship with the atomic model. You will be asked to predict periodic trends and graph these. Additionally, you will apply your practical knowledge to devise experimentation for placing unknown elements in the periodic table.

- This practice SAC comprises three questions; you must complete ALL question sections.
- Pens, pencils, highlighters, erasers, rulers and a scientific calculator are permitted.
- You may use the VCE Chemistry Data Book for this task.



Total time: 55 minutes (5 minutes reading, 50 minutes writing)

Total marks: 50 marks

TRENDS IN THE PERIODIC TABLE

Ionisation energies of the first 20 elements, measured in kJ mol^{-1}

Number	Element	1st	2nd	3rd	4th	5th	6th	7th
1	Hydrogen	1312						
2	Helium	2372	5251					
3	Lithium	520	7298	11 815				
4	Beryllium	900	1757	14 849	21 007			
5	Boron	801	2427	3660	25 026	32 827		
6	Carbon	1087	2353	4621	6223	37 831	47 277	
7	Nitrogen	1402	2856	4578	7475	9445	53 267	64 360
8	Oxygen	1314	3388	5301	7469	10 990	13 327	71 330
9	Fluorine	1681	3374	6050	8408	11 023	15 164	17 868
10	Neon	2081	3952	6122	9371	12 177	15 239	19 999
11	Sodium	496	4562	6910	9543	13 354	16 613	20 117
12	Magnesium	738	1451	7733	10 543	13 630	18 020	23 326
13	Aluminium	578	1817	2745	11 577	14 842	18 379	23 780
14	Silicon	787	1577	3232	4356	16 091	19 805	23 780
15	Phosphorus	1012	1907	2914	4964	6274	21 267	25 431
16	Sulfur	1000	2252	3357	4556	7004	8496	27 107
17	Chlorine	1251	2298	3822	5159	6542	9362	11 018
18	Argon	1521	2666	3931	5771	7238	8781	11 995
19	Potassium	419	3052	4420	5877	7975	9590	11 343
20	Calcium	590	1145	4912	6491	8153	10 496	12 270

Question 1 (23 marks)

a. i. Draw a graph of the number of electrons versus ionisation energy of the approximate values for the first six ionisation energies of aluminium. The first ionisation energy of aluminium is 584 kJ mol^{-1}. **(5 marks)**

ii. Explain your reasoning in drawing your graph. **(3 marks)**

b. Complete the following table by writing the first three ionisation energy equations for aluminium and the resultant electronic configuration. **(6 marks)**

Ionisation energy	Equation	Electronic configuration
First		
Second		
Third		

c. Does Al^{2+} or Al^{3+} have the largest ionisation energy? Use an electronic model to show electron configuration and explain your answer. **(3 marks)**

d. There is a decrease in period 3 ionisation energies from magnesium to aluminium of 736 to 577 kJ mol^{-1} respectively. This is contradictory to the general trend, where ionisation energy increases across a period. Explain why this decrease occurs, referring to the electronic structure of these elements. **(2 marks)**

e. What factors affect the size of the ionisation energy? **(4 marks)**

Question 2 (12 marks)

a. Draw, label and explain graphs to predict the general trend for the *second* period for the following.

i. Electronegativity **(2 marks)**

ii. Ionisation energy **(2 marks)**

iii. Atomic radii **(2 marks)**

b. Draw, label and explain graphs to predict the general trend on *descending* group II for the following.

i. Electronegativity **(2 marks)**

ii. Ionisation energy **(2 marks)**

iii. Atomic radii **(2 marks)**

Question 3 (15 marks)

You are given three unknown elements, *X*, *Y* and *Z*, two of which you are told are in the same group. You have been given access to all the equipment of the school laboratory, but do not have any internet access.

What experimentation would you do to predict where *X*, *Y* and *Z* should be placed in the periodic table and how would you determine which two elements are in the same group?

Write down all experimentations you would perform and why you have chosen these.

on Resources

Digital document U1AOS1 School-assessed coursework (doc-38063)

6 Quantifying atoms and compounds

KEY KNOWLEDGE

In this topic you will investigate:

Quantifying atoms and compounds

- the relative isotopic masses of isotopes of elements and their values on the scale in which the relative isotopic mass of the carbon-12 isotope is assigned a value of 12 exactly
- determination of the relative atomic mass of an element using mass spectrometry (details of instrument not required)
- Avogadro's constant as the number 6.02×10^{23} indicating the number of atoms or molecules in a mole of any substance; determination of the amount in moles of atoms (or molecules) in a pure sample of known mass
- determination of the molar mass of compounds, the percentage composition by mass of covalent compounds and the empirical and molecular formula of a compound from its percentage composition by mass.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of VCE Chemistry. Experiments and investigations, supported by a **practical investigation eLogbook** and **teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

EXAM PREPARATION

Access exam-style questions and their video solutions in every lesson, to ensure you are ready.

6.1 Overview

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6.1.1 Introduction

Atoms are tiny — so tiny that their size is difficult to comprehend. Scientists estimate that a grain of sand contains as many atoms as stars in the known universe. But how do we measure atoms? The scale of atomic size means that chemists rarely deal with atoms one at a time. What is needed is a convenient way to compare and measure masses and amounts of particles in elements and compounds. The mole concept, which is central to most chemical calculations, helps us to work with the vast numbers of atoms that are present in different types of substances.

FIGURE 6.1 International Mole Day occurs from 6:02 am to 6:02 pm on 23 October, because the time and date represent the number for the mole, or Avogadro's constant (6.02×10^{23}).

Once understood, the mole concept is one of the most important concepts in chemistry. It allows a whole range of calculations to be performed. Among these are percentage compositions and empirical formulas, used in important everyday commercial and industrial applications.

In society and industry, developing chemical processes that minimise harm to the environment is becoming increasingly important. To this end, the 12 principles of green chemistry are important. Many new processes will need to be designed to produce new chemicals and improve existing ones. Central to such planning is the ability to predict amounts of raw materials required and desired, and waste product formed. The mole concept plays a central role in how this is done.

LEARNING SEQUENCE

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on Resources

Solutions	Solutions — Topic 6 (sol-0805)
Practical investigation eLogbook	Practical investigation eLogbook — Topic 6 (elog-1611)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 6 (doc-37077) Key idea summary — Topic 6 (doc-37078)
Exam question booklet	Exam question booklet — Topic 6 (eqb-0087)

6.2 Relative isotopic mass and the carbon-12 scale

KEY KNOWLEDGE

- The relative isotopic masses of isotopes of elements and their values on the scale in which the relative isotopic mass of the carbon-12 isotope is assigned a value of 12 exactly
- Determination of the relative atomic mass of an element using mass spectrometry (details of instrument not required)

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6.2.1 Review of atomic number and mass number

BACKGROUND KNOWLEDGE: Atomic number and mass number

You may recall that atoms are composed of a nucleus containing both protons (positive charge) and neutrons (no charge), which are significantly larger than the electrons (negative charge) that occur in shells around the nucleus.

- The atomic number (Z) is the number of protons in the nucleus. All atoms of the same element have the same number of protons in their nucleus.
- The **mass number** (A) of an element is the sum of the protons and neutrons in the nucleus. Protons have approximately the same mass as neutrons and, because the mass of electrons is so small, the mass number only considers protons and neutrons.

Hence, an element can be represented by the following:

$$\begin{matrix}\text{mass number} \rightarrow A \\ \text{atomic number} \rightarrow Z\end{matrix} E \leftarrow \text{symbol for element}$$

6.2.2 Relative isotopic mass

The mass of one hydrogen atom is 1.66×10^{-24} g. This incredibly small size is difficult to manipulate, so a scale based on a multiple of the mass of a single atom of a carbon-12 (^{12}C) isotope was devised. By international agreement, the mass of an atom of carbon-12 is given as 12 **unified atomic mass units (u)**. Therefore, 1 u is $\frac{1}{12}$th the mass of an atom of carbon-12. The carbon-12 isotope is very stable and abundant on Earth.

mass number the total number of protons and neutrons in the nucleus of a particular isotope of an element

unified atomic mass units (u) standard unit of mass based on the mass of the C-12 isotope, where mass of an atom of C-12 equals 12 u; it is approximately equal to one nucleon (proton or neutron)

For example, an atom of hydrogen has a relative mass of 1.0 because the mass of a hydrogen atom ($1.660539040 \times 10^{-24}$ g) is 12 times less (11.999998187) than the mass of a single atom of ^{12}C ($1.992646547 \times 10^{-23}$ g). Mass calculations relative to carbon-12 are demonstrated in table 6.1.

TABLE 6.1 Masses of atoms of three elements

Element	Protons	Neutrons	Protons + neutrons	Mass relative to carbon-12
$^{1}_{1}H$	1	0	1	$\frac{1}{12} \times 12 = 1$
$^{12}_{6}C$	6	6	12	$\frac{12}{12} \times 12 = 12$
$^{59}_{27}Co$*	27	32	59	$\frac{59}{12} \times 12 = 59$

*While cobalt, Co, has 22 isotopes, only one isotype is naturally occurring, ^{59}Co.

Isotopes are atoms of the same element that contain the same number of protons but different numbers of neutrons. Individual isotopes each have a **relative isotopic mass**. A few elements have no variation in the number of neutrons and so have only one relative isotopic mass.

Two equivalent definitions can be used to define relative isotopic mass.

Relative isotopic mass

Relative isotopic mass can be defined as:

- the mass of an atom of an isotope relative to the mass of an atom of the carbon-12 isotope, which has a mass of exactly 12 unified atomic mass units (symbol u)

or

- the mass of an atom of a single isotope relative to $\frac{1}{12}$ of the mass of an atom of the carbon-12 isotope.

Therefore, using the provided definition:

$$\text{Relative isotopic mass} = \frac{\text{mass of an atom of the isotope}}{\text{mass of an atom carbon} - 12} \times 12$$

FIGURE 6.2 Isotopes that have a large difference between the number of protons and neutrons may be unstable and radioactive. Promethium-147 has 61 protons to 86 neutrons, and Thallium-204 has 81 protons to 123 neutrons.

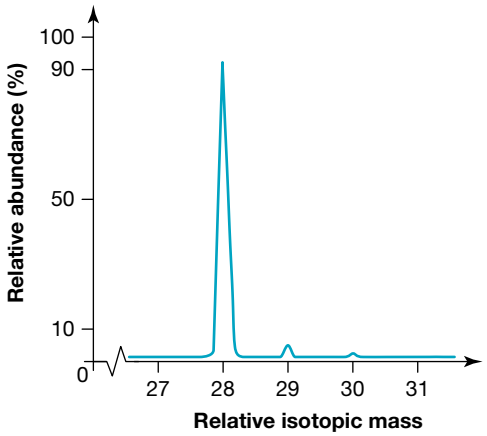
Measuring atomic masses accurately

An instrument called a **mass spectrometer** can be used to measure the relative masses of atoms with great precision. It can, therefore, be used to obtain very accurate values for relative isotopic masses. When a naturally occurring sample of an element is analysed in this instrument, three important pieces of information can be obtained:

- the number of isotopes in the element that was analysed
- accurate values for each isotope's relative isotopic mass
- the relative proportions of each isotope in the sample.

The output from a mass spectrometer is a mass spectrum and provides the preceding information. Figure 6.3 shows a typical mass spectrum for silicon, which indicates that it consists of three isotopes. The abundance of one of these is considerably higher than the other two.

FIGURE 6.3 The mass spectrum of silicon



isotopes forms of an element with the same number of protons but different numbers of neutrons in the nucleus, resulting in elements with different masses

relative isotopic mass mass of an individual isotope of an element on the relative atomic mass scale, on which the masses of particles are compared with the mass of the carbon-12 isotope

mass spectrometer analytical instrument that determines the relative isotopic masses of the different isotopes of an element and abundance

Points to note about a mass spectrum include the following:

- The number of peaks corresponds to the number of isotopes. In this case, silicon has three isotopes.
- The horizontal scale is usually interpreted as relative isotopic mass. However, for technical reasons, it is more correctly a mass : charge ratio.
- The peak heights are a measure of the relative abundances of the isotopes.

EXTENSION: Finding a reference for the mass of atoms

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

This reference standard was later changed from hydrogen to oxygen, but this led to a disagreement between chemists, who used natural oxygen as the standard, and physicists, who used the oxygen-16 isotope only. Having two slightly different lists of atomic masses caused many problems. An agreement was reached in 1961 to change the standard to carbon-12, and this settled the dispute; this change had the added advantage that carbon-12 can be measured very accurately because it is stable and abundant. The units for these measurements are 'unified atomic mass units', although this is often shortened to just 'atomic mass units'. In 1993, the IUPAC proposed the shorter name 'dalton' (with symbol 'Da') for the unified atomic mass unit. The development of an instrument called the mass spectrometer allowed scientists to accurately compare the masses of all atoms.

FIGURE 6.4 Dalton's elemental symbols and atomic weights

The difference between mass number and relative isotopic mass

It is important to understand that mass number and relative isotopic mass *are not the same thing*. By definition:

- Mass number is always a whole number (protons + neutrons in an atom).
- Relative isotopic mass is an accurate measurement made by comparing an atom's mass to the mass of a carbon-12 atom. Relative isotopic masses are rarely whole numbers but they are always very close to an isotope's mass number. The reasons for this are beyond the scope of Unit 1 Chemistry.

A comparison of mass number and relative isotopic mass is shown in table 6.2.

TABLE 6.2 Comparison between mass number and relative isotopic masses for selected isotopes

Isotope	Element	Mass number (protons + neutrons)	Relative isotopic mass (u)
$^{1}_{1}H$	Hydrogen	1	1.007825
$^{2}_{1}H$	Hydrogen	2	2.014102
$^{12}_{6}C$	Carbon	12	12 exactly (by definition)
$^{13}_{6}C$	Carbon	13	13.003355
$^{16}_{8}O$	Oxygen	16	15.994915
$^{35}_{17}Cl$	Chlorine	35	34.968853
$^{37}_{17}Cl$	Chlorine	37	36.965903

Note: due to the closeness of these values, relative isotopic masses are often approximated to the mass number when accuracy beyond a few significant figures is not required.

EXTENSION: How mass spectrometers work

The mass spectrometer is one of the most widely used analytical instruments today. It was developed early in the 20th century based on the work of many preceding scientists. However, most of the credit for its invention goes to the discoverer of the electron, J. J. Thompson, and his protégé, Francis Aston.

Originally developed for the analysis of elements and their isotopes, applications for the mass spectrometer have expanded considerably. Today, it is widely used for the identification of molecules in analytical laboratories and is often part of very sophisticated analytical chains involving other types of instruments.

To learn more about how a mass spectrometer works, explore the online resources listed here.

FIGURE 6.5 A mass spectrometer

on Resources

Digital document EXTENSION: How mass spectrometers work (doc-37627)

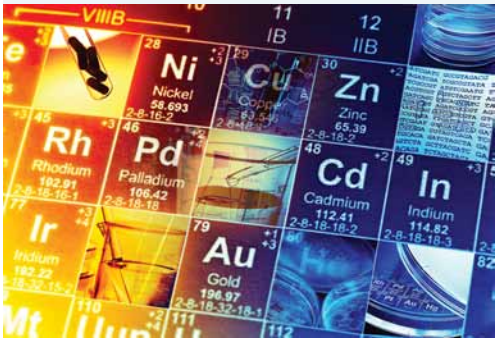
Weblink Mass spectrometry

6.2.3 Relative atomic mass

The **relative atomic mass (A_r)** of an element is defined as the *weighted average of all the relative isotopic masses for the isotopes that make up the naturally occurring isotopic mixture of an element.*

A weighted average takes into account the differing proportions of what it is that is being averaged. In this case, it represents the *average mass of the atoms that are naturally present in an element.* Because the relative atomic mass is obtained from masses that are measured on the relative isotopic mass (RIM) scale (relative to carbon-12), it too represents a comparison to carbon-12. Relative atomic masses, along with atomic numbers, are often shown on periodic tables.

FIGURE 6.6 A section of the periodic table showing relative atomic masses



Relative atomic mass

The relative atomic mass of an element can be calculated from the following formula:

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

relative atomic mass (A_r) mean mass of the naturally occurring mixture of the isotopes of an element on the relative atomic mass scale, on which the masses of particles are compared with the mass of the carbon-12 isotope

For example, consider the mass spectrum of silicon shown in figure 6.3. A careful analysis of this yields the following results:

Isotope ^{28}Si, relative isotopic mass = 27.977, abundance = 92.23%

Isotope ^{29}Si, relative isotopic mass = 28.976, abundance = 4.67%

Isotope ^{30}Si, relative isotopic mass = 29.973, abundance = 3.1%

Using these values, the relative atomic mass of silicon can be calculated as follows:

$$A_r\,(\text{Si}) = \frac{(27.977 \times 92.23) + (28.976 \times 4.67) + (29.973 \times 3.1)}{100}$$
$$= 28.1 \text{ (no units because this is a relative value)}$$

elog-1758

EXPERIMENT 6.1 online only

Determination of relative isotopic mass — a simulation

Aim

To simulate the determination of relative atomic masses using some common everyday objects

tlvd-0522

SAMPLE PROBLEM 1 Calculating the relative atomic mass using isotopes

Lithium consists of two isotopes. One isotope, ^{6}Li, has a relative isotopic mass of 6.02 and an abundance of 7.42 per cent. The other isotope, ^{7}Li, has a relative isotopic mass of 7.02 and an abundance of 92.58 per cent.

a. Is the relative atomic mass closer to six or seven? Explain your answer.

b. Calculate the relative atomic mass of lithium.

THINK	WRITE
a. Nearly 93 per cent of all the isotopes of lithium have a relative isotopic mass of close to seven, so when the masses are averaged it will be much nearer to seven than six.	Closer to seven
b. 1. Recall the relative atomic mass formula and substitute known values.	$A_r(\text{Li}) = \frac{(7.59 \times 6.02) + (92.41 \times 7.02)}{100}$ $= 6.944$
2. Round your answer to three significant figures because this is the fewest number of significant figures used in the question. No units are required because this is a relative value.	$A_r(\text{Li}) = 6.94$

PRACTICE PROBLEM 1

Copper has two isotopes: ^{63}Cu, which has the relative isotopic mass of 62.93, and ^{65}Cu, which has the relative isotopic mass of 64.92. The proportions of each isotope are 69.15 per cent and 30.85 per cent respectively. Calculate the relative atomic mass, A_r, of copper.

Using significant figures

- All digits in numbers expressed in standard form are significant; for example, 4.320×10^{-6} has four significant figures.
- All non-zero numbers are significant; for example, 42.3 has three significant figures.
- Zeros between two non-zero numbers are significant; for example, 4.302 has four significant figures.
- Leading zeros are not significant; for example, 0.0043 has two significant figures.
- Trailing zeros to the right of a decimal point are significant; for example, 42.00 has four significant figures.
- For numbers less than one, similar rules apply; for example, 0.4 has one significant figure and 0.04 also has one significant figure, whereas 0.40 has two significant figures and 0.400 has three significant figures.
- Whole numbers written without a decimal point will have the same number of significant figures as the number of digits, with the assumption that the decimal point occurs at the end of the number; for example, 400 has three significant figures. Therefore, a stated volume of '400 mL' will be considered as having three significant figures.

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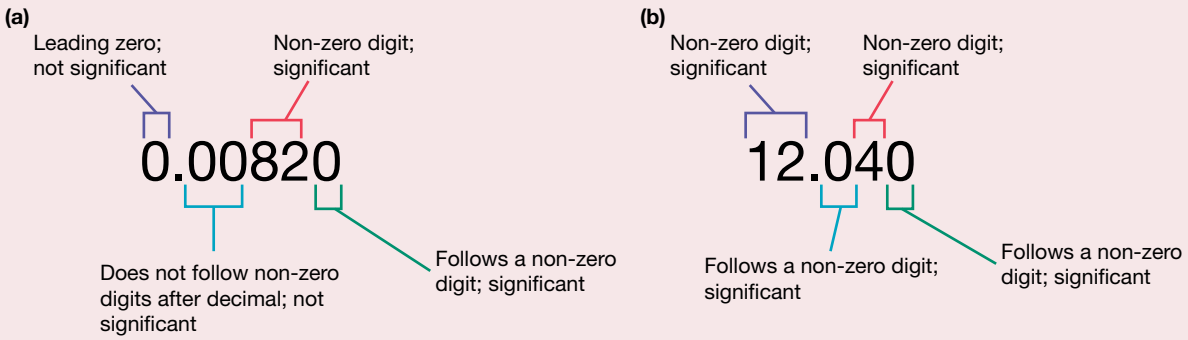
FIGURE 6.7 Examples of significant figures: (a) three significant figures and (b) five significant figures



Adding and subtracting significant figures

When adding and subtracting with decimal places, you should look to the right of the decimal point only. Round your answer to the fewest significant digits to the right of the decimal; for example:

266.38	+	743.394	−	0.000357	=	1009.77
2 decimal places		3 decimal places		6 decimal places		2 decimal places

Multiplying and dividing significant figures

When multiplying and dividing, you should retain as many significant figures as in the number with the fewest significant digits; for example:

8950.23	×	39.8	÷	11.238	=	3.17×10^4
6 sig.figs		3 sig.figs		5 sig.figs		3 sig.figs

The first non-zero digit, reading from left to right in a number, is the first significant figure.

Relative masses, M_r, for molecules and ionic compounds

Applying the relative scale to compounds is also possible, and in this case it is called either the relative formula mass or the relative molecular mass.

Relative formula mass (M_r) uses the same symbol as **relative molecular mass (M_r)** and is equal to the sum of the relative atomic masses of the atoms in the formula of the compound. For covalent compounds, the relative molecular mass is used; for ionic compounds, the relative formula mass is used.

relative formula mass (M_r) the sum of the relative atomic masses of the atoms in the formula of an ionic compound

relative molecular mass (M_r) the sum of the relative atomic masses of the atoms in a molecule of a covalent molecular compound

tlvd-0521

SAMPLE PROBLEM 2 Calculating relative molecular and formula masses

Calculate the M_r of the following compounds.

a. SO_2 **b.** $Mg(OH)_2$ **c.** NH_4NO_3

THINK

a. Look up the relative atomic mass (A_r) of the elements that make up SO_2 using the periodic table.
SO_2: 1 × S and 2 × O
Given that this is a covalent molecule, the term *relative molecular mass* would be used. No units are required.

WRITE

$1 \times S$ and $2 \times O$
$M_r = 32.0 + (2 \times 16.0) = 64.0$
$M_r = 64.0$

b. Look up the relative atomic mass of the elements that make up $Mg(OH)_2$.
$Mg(OH)_2$: 1 × Mg and 2 × O and 2 × H.
Remember that $(OH)_2$ is the same as 2 × OH.
Given that this is an ionic compound, the term *relative formula mass* would be used. No units are required.

$1 \times Mg$ and $2 \times O$ and $2 \times H$
$M_r = 24.3 + (2 \times 16.0) + (2 \times 1.0)$
$M_r = 58.3$

c. Look up the relative atomic mass of the elements that make up NH_4NO_3.
NH_4NO_3: 2 × N and 4 × H and 3 × O
The answer is given to one decimal place because this is the fewest number of decimal places in the data used.
Given that this is an ionic compound, the term *relative formula mass* would be used.

$2 \times N$ and $4 \times H$ and $3 \times O$
$M_r = (2 \times 14.0) + (4 \times 1.0) + (3 \times 16.0)$
$M_r = 80.0$

PRACTICE PROBLEM 2

Calculate the M_r of the following compounds.

a. $Ca_3(PO_4)_2$ **b.** WC **c.** $Na_2Cr_2O_7$

on Resources

Interactivity Relative atomic mass (int-6353)

6.2 Activities

learnon

Students, these questions are even better in jacPLUS

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- **Access additional questions**
- **Track your results and progress**

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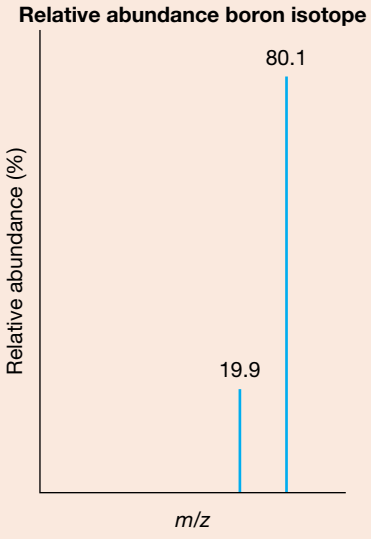
6.2 Quick quiz on | **6.2 Exercise** | **6.2 Exam questions**

6.2 Exercise

1. Complete the table provided for each of the three isotopes indicated.

Isotope	Element	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
$^{16}_{8}O$						
$^{17}_{8}O$						
$^{56}_{26}Fe$						

2. The two isotopes of indium are 9.408 and 9.575 times heavier than the ^{12}C isotope.
   a. Calculate the relative isotopic mass of each isotope.
   b. Use isotopic notation to denote each of these isotopes.
3. Naturally occurring silver consists of two isotopes: ^{107}Ag and ^{109}Ag. Their abundances are 51.8% and 48.2% respectively.
   Estimate the relative atomic mass of silver.
4. Three isotopes of magnesium and their relative abundances are $^{24}_{12}Mg$ (78.8%), $^{25}_{12}Mg$ (10.2%) and $^{26}_{12}Mg$ (11.0%). The mass of each isotope was measured to three significant figures: 24.0, 25.0 and 26.0 respectively.
   a. Sketch a graph of the mass spectrum for magnesium.
   b. Calculate the relative atomic mass of magnesium.
5. Calculate M_r of each of the following compounds.
   a. CO_2 b. NaCl c. H_2O_2 d. H_2SO_4 e. $C_6H_{12}O_6$
6. Gallium has two isotopes. One isotope is $^{69}_{31}Ga$ and has a relative abundance of 60.50%. The relative atomic mass of gallium is 69.70. Find the relative isotopic mass of the other isotope.
7. Thien has just begun VCE Chemistry. After learning about how the relative isotopic mass scale is based on the ^{12}C isotope, she notices that an internet source quotes the value for the relative atomic mass of carbon as 12.01.
   Explain why this value is not exactly 12, as per the definition.
8. Archie has just calculated the relative atomic mass of copper using 63 for the ^{63}Cu isotope and 65 for the ^{65}Cu isotope. On the wall in his classroom, he notices a poster that quotes the relative isotopic masses for these isotopes as 62.9296 and 64.9278 respectively.
   Obviously, these figures can be rounded to whole numbers when appropriate, but they are not the same.
   Explain the difference between the 63 as used in ^{63}Cu and 63 when it is used as an approximation to 62.9296.
9. The green colour of fireworks can be produced using the element boron. Boron's two isotopes are $^{10}_{5}B$ and $^{11}_{5}B$. Use the mass spectrum provided in the figure to calculate the relative atomic mass of boron.



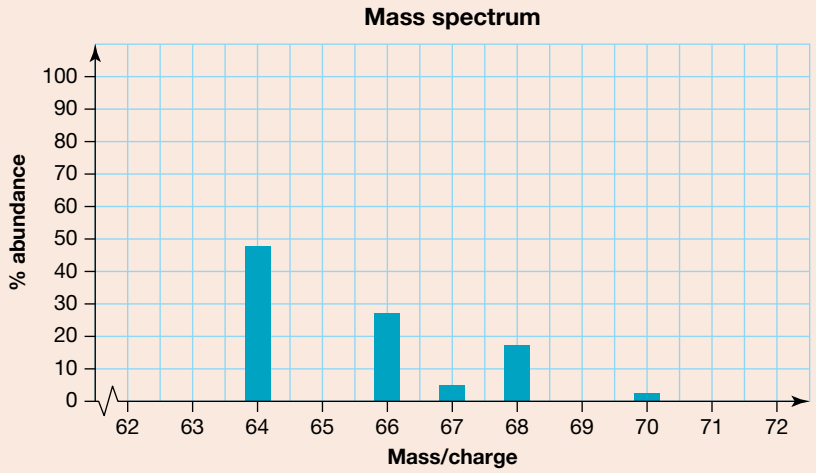
6.2 Exam questions

Question 1 (2 marks)

Explain why relative atomic masses are typically not whole numbers yet mass numbers are always integers.

Question 2 (1 mark)

MC The mass spectrum of is a mystery element is given.

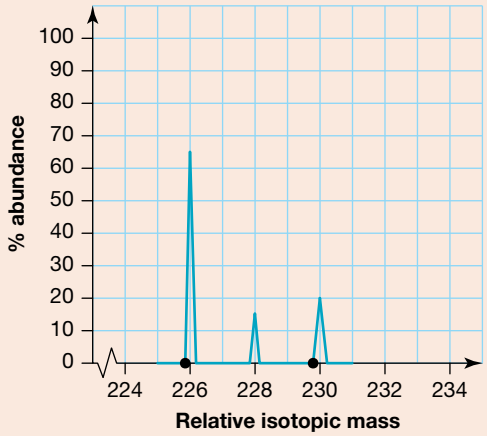


Based on the data provided, what is the element most likely to be?

A. Copper **B.** Zinc **C.** Gallium **D.** Neodymium

Question 3 (2 marks)

The mass spectrum of an element is shown.



a. State the percentage abundances of each of the isotopes present in the mass spectrum. **(1 mark)**

b. Calculate the relative atomic mass of the element. **(1 mark)**

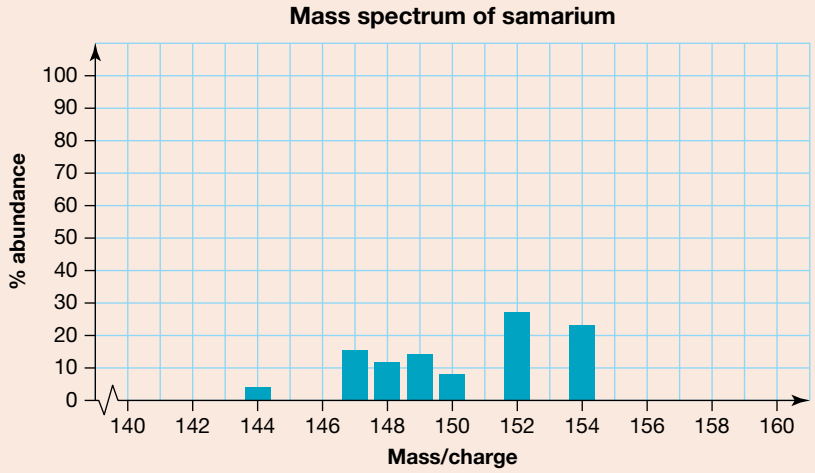
Question 4 (2 marks)

Antimony has a relative atomic mass of 121.8 and exists as two stable isotopes. The lighter isotopic has a relative isotopic mass of 120.90 and the heavier isotope has a relative isotopic mass of 122.90.

Calculate the percentage abundance of the lighter isotope.

Question 5 (4 marks)

Consider the following mass spectrum of samarium.



a. How many isotopes are present in the mass spectrum? **(1 mark)**

b. Give the isotopic symbol ${}^{A}_{Z}X$ and % of the most abundant isotope. **(2 marks)**

c. What is the relative isotopic mass of the lightest isotope? **(1 mark)**

More exam questions are available in your learnON title.

6.3 Avogadro's constant and the mole

KEY KNOWLEDGE

- Avogadro's constant as the number 6.02×10^{23} indicating the number of atoms or molecules in a mole of any substance; determination of the amount in moles of atoms (or molecules) in a pure sample of known mass

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6.3.1 The mole concept

If one drop of water, H_2O, has a mass of 0.0300 g, then this contains 1.00×10^{21} molecules of water. Dealing with such large values would prove overly complicated, and ratios of masses of different substances would be difficult to see. Clearly, in chemistry, we need a measurement unit that can manage these large numbers of atoms and molecules. The **mole** concept helps us count atoms and is fundamental to chemical calculations. The term 'mole' simply represents a number. It is a unit of measurement, just as 'dozen' refers to 12 and 'kilo' represents 1000.

FIGURE 6.8 One drop of water contains approximately 1.00×10^{21} molecules of water.

mole 6.02×10^{23} particles of a substance

The Avogadro constant

The mole, abbreviation mol, symbol n, is the SI unit of amount of substance. A mole contains exactly $6.022\,140\,76 \times 10^{23}$ elementary entities. This number is the fixed numerical value of the **Avogadro constant** (N_A or L), when expressed in mol^{-1}. Therefore, the number of atoms in one mole of a pure element is 6.02×10^{23} mol^{-1}.

The mole

One mole contains exactly $6.022\,140\,76 \times 10^{23}$ elementary entities.

It is typically rounded to three significant figures.

Avogadro's constant: N_A or $L = 6.02 \times 10^{23}$ mol^{-1}

This definition of the mole was only recently adopted. Prior to 20 May 2019, the mole was defined as 'the amount of substance that contains as many particles (atoms, ions or molecules) as there are atoms in exactly 12 g of the ^{12}C isotope'. The number of carbon atoms in 12 g of ^{12}C isotope has been experimentally calculated to be 602 140 760 000 000 000 000 000, or 6.02×10^{23} mol^{-1}. However, in 2018 the General Conference on Weights and Measures determined that all SI (Système international (d'unités)) will not be defined in terms of any physical objects but rather defined by constants (due to their exact nature).

While the definition of the mole has changed, from the number of atoms in exactly 12 g of the ^{12}C isotope to a constant of the same value, the new definition is fundamentally based on the ^{12}C isotope. You may have noticed that the definition of relative isotopic mass is also based on the ^{12}C isotope. A good reason exists for this. The relative isotopic scale allows us to deal with atoms 'one at a time'. The mole concept allows us to deal with much larger groups 'one mole at a time'. The common reference to ^{12}C means that to find out the mass of one mole of any element, all you have to do is add 'g' to the relative atomic mass.

Avogadro constant the number of elementary particles (atoms) in exactly 12 g of carbon-12; this number is now defined as 6.02×10^{23}

FIGURE 6.9 One mole each of mercury, zinc, silicon, aluminium, sulfur and bromine. The quantity of each of the substances shown contains 6.02×10^{23} particles.

To find the mass of one mole of an element

To find the mass of one mole of an element, simply add 'g' to the relative atomic mass of that element.

FIGURE 6.10 Pure carbon is also used in air, gold and water purification, decaffeination, metal extraction, poison and sewage treatment, and in air filters in gas masks.

EXTENSION: Amedeo Avogadro

Amedeo Avogadro is one of Italy's most famous modern scientists. In the commemorative stamp shown in figure 6.11, celebrating 100 years since his death, the text translates to 'Equal volumes of gas under the same temperature and pressure conditions contain the same number of molecules'. This statement is known as Avogadro's law. In tribute to his work on molecular theory, Avogadro's constant, N_A or L, relates the molar mass of a substance (molecule, atom or ion) to the mass of a sample. One mole of any sample will contain exactly $6.022\ 140\ 76 \times 10^{23}$ particles. Avogadro's constant is a recognised SI unit.

To find out how Avogadro's number was originally derived, download the digital document listed here.

FIGURE 6.11 Avogadro was recognised for his work on molecular theory.

on Resources

Digital document EXTENSION: Determining Avogadro's number (doc-37628)

6.3.2 Molar mass, *M*, of an element

The **molar mass** (M) of an element is defined as the mass of one mol of the element; that is, 6.02×10^{23} atoms of the element. The unit is grams per mole (g mol^{-1}). Note that when dealing with molecular elements, ambiguity may arise. For this reason, it is always a good idea to specify whether you are referring to the atoms or molecules of the element. For example:

molar mass mass of one mole of a compound expressed in grams per mol

Molar mass of carbon *atoms*

Molar mass of carbon *atoms* = mass of 1 mol of C *atoms*

$A_r(C) = 12.0$

$\therefore$ mass of 1 mol of C *atoms* = 12.0 g mol^{-1}

That is,

$M(C) = 12.0$ g mol^{-1}

$\therefore$ 12.0 g of carbon contains 6.02×10^{23} *atoms*

Molar mass of oxygen *atoms*

Molar mass of oxygen *atoms* = mass of 1 mol of O *atoms*

$A_r(O) = 16.0$

$\therefore$ mass of 1 mol of O *atoms* $= 16.0\ g\ mol^{-1}$

That is,

$M(O) = 16.0\ g\ mol^{-1}$

$\therefore$ 16.0 g of oxygen contains 6.02×10^{23} *atoms*

Molar mass of oxygen *molecules*

Molar mass of oxygen *molecules* = mass of 1 mol of O_2 *molecules*

$M_r(O_2) = 2 \times 16.0 = 32.0$

$\therefore$ mass of 1 mol of O_2 molecules $= 32.0\ g\ mol^{-1}$

That is,

$M(O_2) = 32.0\ g\ mol^{-1}$

$\therefore$ 32.0 g of oxygen contains 6.02×10^{23} *molecules* of oxygen

$\therefore$ 32.0 g of oxygen contains $2 \times 6.02 \times 10^{23} = 1.20 \times 10^{24}$ *atoms* of oxygen

6.3.3 Calculating molar mass, *M*, of covalent (molecular) and ionic compounds

The molar mass of a compound is defined as the mass of one mol of the compound expressed in grams per mole ($g\ mol^{-1}$).

Covalent compounds

For compounds consisting of molecules, the molar mass (M) is numerically equal to the relative molecular mass (M_r) expressed in grams per mole ($g\ mol^{-1}$). For example:

Molar mass of water *molecules*

Molar mass of water molecules = mass of 1 mol of H_2O molecules

$$M_r(H_2O) = (2 \times 1.0) + (16.0)$$
$$= 18.0$$

$\therefore$ mass of 1 mol of H_2O molecules $= 18.0\ g$

$\therefore M(H_2O) = 18.0\ g\ mol^{-1}$

$\therefore$ 18.0 g of water contains 6.02×10^{23} molecules of water

Molar mass of chlorine *molecules*

Molar mass of chlorine gas = mass of 1 mol of Cl_2 molecules

$$M_r(Cl_2) = 2 \times 35.5$$
$$= 71.0$$

$\therefore$ mass of 1 mol of Cl_2 molecules $= 71.0\ g$

$\therefore M(Cl_2) = 71.0\ g\ mol^{-1}$

$\therefore$ 71.0 g of chlorine gas contains 6.02×10^{23} molecules of chlorine

Ionic compounds

The M_r of an ionic compound is found by adding the A_r of each atom in the formula of the compound. For example:

$$\begin{aligned} M_r(CuSO_4) &= A_r(Cu) + A_r(Cu) + (4 \times A_r(O)) \\ &= 63.5 + 32.1 + (4 \times 16.0) \\ &= 159.6 \end{aligned}$$

The molar mass (M) of $CuSO_4$ is 159.6 g mol^{-1}.

That is,

$$M(CuSO_4) = 159.6 \text{g mol}^{-1}$$

tlvd-0523

SAMPLE PROBLEM 3 Determining the molar mass of compounds

What is the molar mass, M, of calcium nitrate, $Ca(NO_3)_2$?

THINK

1. The M_r of an ionic compound is found by adding the A_r of each atom in the formula of the compound.
   Brackets indicate the number of multiples; e.g. $(NO_3)_2$ is the same as $NO_3 + NO_3$.

WRITE

Atom	Number		A_r		Total A_r
Ca	1	×	40.1	=	40.1
N	2	×	14.0	=	28.0
O	6	×	16.0	=	96.0

2. Add the total A_r to determine the M_r.
   TIP: Molar mass must have units of g mol^{-1}, unlike relative masses which have no units.

$M_r = 40.1 + 28.0 + 96.0$
$= 164.1$
$M(Ca(NO_3)_2 = 164.1$ g mol^{-1}

PRACTICE PROBLEM 3

What is the molar mass, M, of calcium perchlorate, $Ca(ClO_4)_2$?

6.3.4 Calculating moles of atoms and molecules in a known mass

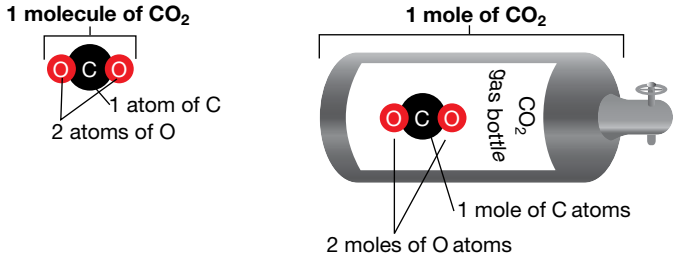
We have discussed a relationship between the number of particles and the mass of the substance. This means that, when we measure the mass of a substance equal to its molar mass (for example, 18.0 g of H_2O), we are also indirectly counting out 6.02×10^{23} particles of that substance (for example, water molecules). If the molecule or formula unit contains more than one kind of atom, we can also calculate the number of different atoms present in one mole of the substance.

For example, 1.00 mol of carbon dioxide molecules contains:

- 1.00 mol of CO_2 molecules
- 6.02×10^{23} CO_2 molecules
- 1.00 mol of C atoms
- 6.02×10^{23} C atoms
- 2.00 mol of O atoms
- 1.20×10^{24} O atoms.

The molar mass of CO_2 is 44.0 g mol^{-1}.

FIGURE 6.12 Whether a substance is made up of atoms, molecules or ions, the same principle applies: one mole of any substance always contains 6.02×10^{23} particles.



These observations suggest the following:

$$\text{mass} \underset{\text{multiply by molar mass}}{\overset{\text{divide by molar mass}}{\rightleftharpoons}} \text{moles}$$

To calculate moles from a known mass

$$n = \frac{m}{M}$$

where:

n = number of moles

m = mass

M = molar mass.

$$\text{number of particles } (N) \underset{\text{multiply by } 6.02 \times 10^{23}}{\overset{\text{divide by } 6.02 \times 10^{23}}{\rightleftharpoons}} \text{moles}$$

To calculate moles from a known number of particles

$$n = \frac{\text{number of particles}}{6.02 \times 10^{23}} = \frac{N}{N_A}$$

where:

n = number of moles

N = number of particles

N_A = Avogadro's constant.

elog-1759

EXPERIMENT 6.2 online only

Estimating the number of objects

Aim

To investigate the number of objects in a sample by weighing

The formula $n = \frac{m}{M}$ is one of the most useful formulas in chemistry. You will use it on numerous occasions to calculate the number of moles in a pure sample when you are given its mass. You will also use it in conjunction with the second formula $\left(n = \frac{N}{N_A}\right)$ in situations where you want to know the actual number of atoms, molecules or ions in a sample.

For example, 1 litre of water has a mass of 1000 g. From this we can calculate the number of mol of water as well as the numbers of water molecules, hydrogen atoms and oxygen atoms that it contains.

$$\begin{aligned} M(H_2O) &= 18.0\ \text{g mol}^{-1} \\ n(H_2O) &= \frac{m}{M} \\ &= \frac{1000}{18.0} \\ &= 55.6\ \text{mol} \\ N(H_2O)\ \text{molecules} &= n \times N_A \\ &= 55.6 \times 6.02 \times 10^{23} \\ &= 3.34 \times 10^{25}\ \text{molecules} \end{aligned}$$

Because one molecule of water (H_2O) contains two atoms of H and one atom of O:

$$\begin{aligned} N(\text{H atoms}) &= N(H_2O\ \text{molecules}) \times 2 \\ &= 2 \times 3.34 \times 10^{25} \\ &= 6.68 \times 10^{25}\ \text{atoms} \end{aligned}$$

and

$$\begin{aligned} N(\text{O atoms}) &= N(H_2O\ \text{molecules}) \\ &= 3.34 \times 10^{25}\ \text{atoms} \end{aligned}$$

tlvd-0524

SAMPLE PROBLEM 4 Calculating the number of atoms in a given mass

A gas balloon contains 22 g of helium atoms. How many helium atoms are present in the balloon?

THINK

1. To determine the number of helium atoms, first use the formula $n = \frac{m}{M}$ to determine the number of mole of helium atoms.

2. To determine the number of helium atoms in the calculated number of moles of helium, rearrange the formula $n = \frac{N}{N_A}$.
   Round your answer to two significant figures because this is the number of significant figures used in the question.
   TIP Avogadro's constant, N_A, can be found in the VCE Chemistry Data Book.

WRITE

$$\begin{aligned} n(\text{He}) &= \frac{m}{M} \\ &= \frac{22}{4.0} \\ &= 5.5\ \text{mol} \end{aligned}$$

$$\begin{aligned} n &= \frac{N}{N_A} \\ N(\text{He}) &= n \times N_A \\ &= 5.5 \times 6.02 \times 10^{23} \\ &= 3.3 \times 10^{24} \end{aligned}$$

3.3×10^{24} atoms of helium are in 22 g of helium.

PRACTICE PROBLEM 4

How many atoms of oxygen are in 66.05 g of ammonium sulfate, $(NH_4)_2SO_4$?

on Resources

Interactivity The mole concept (int-6354)

Weblinks What is mole day?
How big is a mole?

6.3 Activities

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6.3 Quick quiz on | **6.3 Exercise** | **6.3 Exam questions**

6.3 Exercise

1. Calculate the molar masses of the following.
   - **a.** Ozone, O_3
   - **b.** Iodine, I_2
   - **c.** Calcium oxide, CaO
   - **d.** Hydrogen chloride, HCl
2. MC A glass of water contains 250 g of water. How many water molecules are present in this mass?
   - **A.** 13.9
   - **B.** 250
   - **C.** 8.36×10^{24}
   - **D.** 1.51×10^{26}
3. Calculate the number of moles present in each of the following.
   - **a.** 1.064 g of palladium
   - **b.** 17.524 g of strontium
   - **c.** 18.385 g of tungsten
   - **d.** 0.1401 g of nitrogen gas molecules
4. Calculate the number of moles present in each of the following samples.
   - **a.** 3.01×10^{23} atoms of molybdenum
   - **b.** 6.02×10^{21} atoms of zirconium
   - **c.** 9.03×10^{22} atoms of uranium
   - **d.** 1.00×10^{23} molecules of carbon dioxide
5. MC What is the mass of 0.06750 mol of sodium carbonate?
   - **A.** 5.603 g
   - **B.** 7.155 g
   - **C.** 106.0 g
   - **D.** 1570 g

6. **MC** Which of the following contains the largest number of nitrogen atoms?
   - **A.** 2 mol of ammonium nitrate (NH_4NO_3)
   - **B.** 1 mol of iron(III) nitrate, $Fe(NO_3)_3$
   - **C.** 3.5 mol of nitrogen dioxide, NO_2
   - **D.** 60 g of nitrogen gas, N_2
7. Calculate the mass of each of the following.
   - **a.** 5.25×10^{24} molecules of glucose, $C_6H_{12}O_6$
   - **b.** 1.83×10^{21} molecules of nitrogen dioxide, NO_2
   - **c.** 3.56×10^{14} molecules of carbon dioxide, CO_2
   - **d.** 4.13×10^{28} molecules of carbon disulfide, CS_2
   - **e.** 3.62×10^{24} molecules of dinitrogen tetroxide, N_2O_4
8. **MC** Which of the following substances has the greatest mass?
   - **A.** 200 g of magnesium
   - **B.** 5.00 mol of sulfur
   - **C.** 1.2×10^{24} atoms of helium
   - **D.** 3.5×10^{22} molecules of alanine, $C_3H_7O_2N$
9. The relative atomic mass of hydrogen is 1.008.
   - **a.** What is the molar mass of hydrogen atoms?
   - **b.** Use your answer from (a) to calculate the average mass, in grams, of a hydrogen atom.
   - **c.** Explain why your answer from part (b) represents an *average* mass.
10. Use a spreadsheet program to produce a spreadsheet that converts amounts of substances to moles. Set up your spreadsheet according to the following template, and then complete the table. Use a formula and then the 'fill' function.

Conversion of amounts of substances to moles							
	A	B	C	D	E	F	G
	Substance	Amount (g)	Molar mass (*M*)	Number of atoms in the molecule	Number of moles (*n*) of substance	Number of molecules	Total number of atoms
1	Water, H_2O	3.2					
2	Methane, CH_4	2.7					
3	Ammonia, NH_3	0.056					
4	Ethanoic acid, CH_3COOH				27.3		
5	Benzene, C_6H_6				0.56		
6	Octane, C_8H_{18}				2.34		
7	Ethanol, CH_3CH_2OH					6.0×10^{24}	
8	Ozone, O_3					1.27×10^{21}	
9	Sulfuric acid, H_2SO_4						3.0×10^{26}
10	Carbon dioxide, CO_2						7.5×10^{25}

6.3 Exam questions

Question 1 (1 mark)

MC Which of the following contains the most atoms?

A. 1.25 mol of silver nitrate, $AgNO_3$
B. 5.00 mol of copper, Cu
C. 2.50 mol of silver, AgCl
D. 1.25 mol of copper(II) chloride, $CuCl_2$

Question 2 (1 mark)

MC What is the amount (in mol) of iron atoms and oxygen atoms in 3.33 mol of iron(III) oxide, Fe_2O_3?

A. 2 mol Fe and 3 mol O
B. 6.66 mol Fe and 9.99 mol O
C. 1.20×10^{24} mol Fe and 1.80×10^{24} mol O
D. 2.00×10^{24} for both Fe and O

Question 3 (1 mark)

MC 1.365 mol of a metal carbonate, MCO_3, has a mass of 206.4 g. What is the metal involved?

A. Zinc **B.** Molybdenum **C.** Nickel **D.** Zirconium

Question 4 (2 marks)

Answer the following.

a. What is the number of molecules in 4.50×10^{-3} mol of carbon dioxide? **(1 mark)**
b. How many carbon atoms are in this amount? **(1 mark)**

Question 5 (1 mark)

MC Consider a 0.0500 mol sample of pure sucrose, $C_{12}H_{22}O_{11}$. Which of the following is correct?

Statement I: The sample has a mass of 17.1 g.

Statement II: The sample contains 3.01×10^{22} sucrose molecules.

Statement III: The sample contains 1.35×10^{24} carbon atoms.

Statement IV: The sample contains twice as many hydrogen atoms as oxygen atoms.

A. All statements are correct.
B. Only statements I, II and IV are correct.
C. Only statements I and III are correct.
D. None of the statements are correct.

More exam questions are available in your learnON title.

6.4 Using the mole concept

KEY KNOWLEDGE

- Determination of the molar mass of compounds, the percentage composition by mass of covalent compounds and the empirical and molecular formula of a compound from its percentage composition by mass

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6.4.1 The mole as a tool for identification

In the previous subtopic, you saw how to determine the molar mass of a compound from a known formula. Determining the molar mass experimentally is also possible. As you progress through your study of chemistry, you will discover that the number of moles of a substance may be calculated in a number of ways.

These methods can be used to aid in the identification of an unknown substance, especially where other clues are also available. Consider the following example.

A sample of gas is collected from burning sulfur in air. The sample has a mass of 0.259 g. Through alternative methods, the number of moles in the sample was determined to be 4.03×10^{-3} mol.

From this information, the molar mass of the gas may be calculated. By rearranging the formula for calculating moles, we get:

$$M = \frac{m}{n}$$
$$\therefore M = \frac{0.259}{0.0403}$$
$$M = 64.3 \text{ g mol}^{-1}$$

Given the clue that this compound is an oxide of sulfur, a formula of SO_2 easily explains the molar mass that is obtained.

6.4.2 Calculating percentage composition

The chemical formula of a compound tells us about the relative numbers of atoms of its constituent elements. However, the composition of a compound is often expressed in terms of the percentage that each element contributes to its mass. This is called **percentage composition**. For instance, although water, H_2O, and hydrogen peroxide, H_2O_2, both contain the same number of hydrogen atoms, the percentage that hydrogen contributes to the mass of a molecule of water is higher than its contribution to a molecule of hydrogen peroxide.

Calculating percentage composition

$$\% \text{ of } x \text{ in compound} = \frac{\text{mass of } x \text{ in 1 mol of compound}}{\text{molar mass of compound}} \times \frac{100}{1}$$

In the preceding case, the percentage of H in water is:

$$\% \text{ of H in } H_2O = \frac{2 \times 1.0}{2 \times 1.0 + 16.0} \times \frac{100}{1}$$
$$= \frac{2}{18} \times \frac{100}{1}$$
$$= 11\%$$

In hydrogen peroxide, the percentage is:

$$\% \text{ of H in } H_2O_2 = \frac{2 \times 1.0}{2 \times 1.0 + 2 \times 16.0} \times \frac{100}{1}$$
$$= \frac{2}{34} \times \frac{100}{1}$$
$$= 5.9\%$$

percentage composition the percentage contribution that each element makes to the mass of a compound

tlvd-3114

SAMPLE PROBLEM 5 Calculating percentage compositions of elements in compounds

The amino acid proline is showing promise as a green chemistry catalyst. The formula of proline is $C_5H_9NO_2$.
Calculate the percentage of carbon in proline.

THINK	WRITE
1. Calculate M_r for proline.	$M_r\,(C_5H_9NO_2) = (5 \times 12.0) + (9 \times 1.0) + (14.0) + (2 \times 16.0)$ $= 60.0 + 9.0 + 14.0 + 32.0$ $= 115.0$
2. Think in terms of molar amounts and masses. Focus on the contribution from carbon.	115.0 g of $C_5H_9NO_2$ must contain 60.0 g of C.
3. Apply the formula to determine the percentage of carbon by mass. Round your answer to three significant figures because this is the fewest number of significant figures in the data used in the division.	$\%\,(C) = \frac{60.0}{115.0} \times \frac{100}{1}$ $= 52.0\%$ Proline contains 52.0% carbon.

PRACTICE PROBLEM 5

Calculate the percentage composition by mass of oxygen in proline.

6.4.3 Empirical formula of a compound

The **empirical formula** of a compound gives the simplest whole number ratio of the atoms or ions present in the compound. Formulas for ionic compounds are already in the simplest ratio of elements (because it is the ratio of elements in the lattice structure) and so the formula for an ionic compound is its empirical formula.

Empirical formula

The empirical formula of a compound is the simplest whole number ratio of atoms present in the compound.

Empirical formulas can only be found experimentally by determining the mass of each element present in the total mass of the compound. This *ratio of elements by mass* must be converted to a *ratio of elements by numbers*. This is done by calculating the number of moles of each element.

The steps involved in finding an empirical formula are as follows:

1. Write down the *symbols* of the elements present.
2. If percentages are given, assume that the *mass* of the sample is 100 g and all percentages become grams. If the actual masses are given, simply write them down.
3. Convert masses to *moles*.
4. Find the simplest whole number *ratio* of the atoms by dividing all numbers of moles by the smallest number of moles.
5. If necessary, multiply by a *factor* to convert all numbers to whole numbers.

empirical formula formula that shows the simplest numerical ratio in which atoms are combined

tlvd-0525

SAMPLE PROBLEM 6 Finding the empirical formula of a compound

A compound of sulfur contains 2.4 per cent hydrogen, 39.0 per cent sulfur and 58.6 per cent oxygen. Find the empirical formula of the compound.

THINK

1. Follow the steps outlined. Remember to assume that the *mass* of the sample is 100 g and all percentages become grams.
   TIP Always organise the grid of elements in the order that they appear in the question and if you get a ratio of, say, exactly 2.5 to 3, double it so both are whole numbers; e.g. 2.5 to 3 becomes 5 to 6.
2. The empirical formula can be written directly from the ratios in the table — in this case, two hydrogen, one sulfur and three oxygen.

WRITE

Step 1	Symbols	H	S	O
Step 2	Masses	2.4 g	39.0 g	58.6 g
Step 3	Moles	$\frac{2.4}{1.0} = 2.4$	$\frac{39.0}{32.1} = 1.2$	$\frac{58.6}{16.0} = 3.7$
Step 4	Ratio/divide by smallest	$\frac{2.4}{1.2} = 2.0$	$\frac{1.2}{1.2} = 1.0$	$\frac{3.7}{1.2} = 3.1$

H_2SO_3

PRACTICE PROBLEM 6

A sample of potassium carbonate yielded 6.85 g of potassium, 1.06 g of carbon and 4.20 g of oxygen upon analysis. Calculate its empirical formula.

6.4.4 Molecular formula of a covalent compound

The formula of an ionic substance is simply the ratio of ions that it is made from. This means that the formula of most ionic substances is already their empirical formula. Ionic substances do not contain molecules. (A few exceptions to this exist. For example, sodium peroxide is made up of sodium ions (Na^+) and peroxide ions (O_2^{2-}), and, therefore, has the formula Na_2O_2.)

The **molecular formula** of a compound, however, represents the actual composition of a compound that is made up of *molecules*. These substances feature covalent bonding, rather than ionic bonding. A molecular formula is either the same as its experimentally determined empirical formula, such as in ionic compounds, or is a whole-number multiple of it.

For example, acetylene is a gas that is used in welder's torches, whereas benzene is a highly flammable liquid that is a recognised carcinogen. Although both compounds have the same empirical formula, CH, the molecular formula of acetylene is C_2H_2, and the molecular formula of benzene is C_6H_6. Similarly, consider the empirical formula CH_2O. Methanal, also known as formaldehyde, has the molecular formula CH_2O, with a boiling point of –21 °C. It is commonly used as a preservative of biological samples and in the manufacture of pressed wood products. A multiple of two of the same empirical formula gives ethanoic acid, $C_2H_4O_2$, also known as acetic acid, but commonly recognised as vinegar. It has a boiling point of 117 °C. A multiple of six of the same empirical formula, $C_6H_{12}O_6$, gives a variety of compounds, including many of simple sugars, such as glucose.

molecular formula actual number of atoms of each element in a molecule of a substance

FIGURE 6.13 Methanal (formaldehyde, CH_2O), ethanoic acid (vinegar, $C_2H_4O_2$) and simple sugars (glucose, $C_6H_{12}O$) all have the empirical formula of CH_2O.

Molecular formulas are calculated in the same manner as empirical formulas, but with the addition of one extra step and piece of information at the end. This is the molar mass. The predicted molar mass from the empirical formula is then simply compared to the actual molar mass.

In the acetylene/benzene example just discussed, the predicted molar mass for CH is 13 g mol^{-1}.

Acetylene has an actual molar mass of 26 g mol^{-1}, so:

$$\text{scaling factor} = \frac{\text{actual molar mass}}{\text{predicted molar mass}}$$
$$= \frac{26}{13}$$
$$= 2$$

$$\therefore \text{Acetylene molecular formula} = (CH) \times 2 = C_2H_2$$

Benzene has an actual molar mass of 78 g mol^{-1}, so:

$$\text{scaling factor} = \frac{\text{actual molar mass}}{\text{predicted molar mass}}$$
$$= \frac{78}{13}$$
$$= 6$$

$$\therefore \text{Benzene molecular formula} = (CH) \times 6 = C_6H_6$$

tlvd-3115

SAMPLE PROBLEM 7 Determining the empirical and molecular formulas of a compound

The production of ethanol as a sustainable fuel is growing rapidly across the world. Ethanol produced in this fashion is termed *bioethanol*. Analysis of a bioethanol sample revealed that it consisted of 52.2 per cent carbon, 13.0 per cent hydrogen and 34.8 per cent oxygen by mass.

a. Calculate the empirical formula of bioethanol.

b. In a separate determination, the molar mass of bioethanol was found to be 46 g mol^{-1}. Calculate the molecular formula of bioethanol.

THINK

a. Assume that the mass of the sample is 100 g and all percentages become grams.
Calculate moles and hence the empirical formula.

WRITE

Symbols	C	H	O
Masses	52.2 g	13.0 g	34.8 g
Moles	$\frac{52.2}{12.0} = 4.35$	$\frac{13.0}{1.0} = 13.0$	$\frac{34.8}{16.0} = 2.18$
Ratio/divide by smallest	$\frac{4.35}{2.18} = 2$	$\frac{13.0}{2.18} = 6$	$\frac{2.18}{2.18} = 1$
Write empirical formula	C_2H_6O		

b. 1. Calculate the predicted molar mass using the derived empirical formula.

$$\text{Predicted molar mass} = (2 \times 12.0) + (6 \times 1.0) + (1 \times 16.0)$$
$$= 46.0 \text{ g mol}^{-1}$$

2. Determine if the empirical formula needs to be scaled up, and by how much.

$$\text{Scaling factor} = \frac{\text{actual molar mass}}{\text{predicted molar mass}}$$
$$= \frac{46}{46.0}$$
$$= 1$$

i.e. no scaling up necessary

3. Adjust the empirical formula (if necessary) as required to obtain the molecular formula.

C_2H_6O

PRACTICE PROBLEM 7

Another sustainable fuel is biodiesel. One of the compounds that it can be made from is stearic acid. Analysis of stearic acid determined that it consisted of 76 per cent carbon, 12.7 per cent hydrogen and 11.3 per cent oxygen. Its molar mass was found to be 284 g mol^{-1}.

a. Calculate the empirical formula of stearic acid.

b. Calculate the molecular formula of stearic acid.

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tlvd-4564

EXPERIMENT 6.3 online only

Determining the empirical formula of magnesium oxide

Aim

To deduce the empirical formula of magnesium oxide

elog-1781

EXPERIMENT 6.4 online only

Stoichiometry of a reaction

Aim

To investigate quantitatively the reaction between iron and copper(II) sulfate pentahydrate

6.4 Activities

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6.4 Quick quiz on | **6.4 Exercise** | **6.4 Exam questions**

6.4 Exercise

1. Calculate the percentage composition of the element shown in brackets in each of the given compounds.
   a. C_8H_{18} (C) **b.** CH_3Cl (Cl) **c.** H_2SO_4 (S) **d.** SiH_4 (Si)
2. 7.870 g of a compound containing only vanadium and chlorine was found to contain 2.548 g of vanadium. Calculate its empirical formula.
3. An empirical formula of a compound is calculated by determining the ratio of moles for each element that it contains.
   a. Explain why it is a ratio of moles that is used and not a ratio of masses.
   b. Explain why the ratio of moles is the same as the ratio of actual atoms.
   c. Which of the formulas in the following list is *not* an empirical formula? $MgSO_4$, C_3H_8, NH_3, NaCl, C_4H_{10}
4. Hydroquinone is a liquid that is used as a photographic developer. Determine the empirical formula of hydroquinone given that it contains 65.4% carbon and 29.1% oxygen, with the remainder being hydrogen.
5. As part of a chemistry project, Xavier was wondering which of the following compounds would contribute the greatest mass of carbon into the atmosphere, assuming the compound was completely burnt and equal masses were compared. The compounds under consideration are C_8H_{18}, CH_3CH_2OH, CH_4, C_3H_8. Show by calculation which one Xavier should choose.
6. Borax is a naturally occurring compound that is used in the manufacture of optical glasses. Analysis of borax shows that it is made up of three elements: sodium (22.8%), boron (21.5%) and oxygen.
   Determine the empirical formula of borax.
7. MC An ionic compound containing iron, sulfur and oxygen was analysed and found to have 27.9% iron and 24.1% sulfur. What is the empirical formula of the compound?
   A. $Fe_2(SO_4)_3$ **B.** $FeSO_4$ **C.** FeS **D.** $FeSO_3$

8. A gardener wishes to add 50 g of nitrogen to a watering can to achieve a desired concentration for some fertiliser. Calculate the mass of the following nitrogen-containing compounds they would need to add to achieve the desired concentration.
   a. Ammonium sulfate; $(NH_4)_2SO_4$
   b. Ammonium nitrate; NH_4NO_3
   c. Ammonium phosphate; $(NH_4)_3PO_4$
   d. Urea; $CO(NH_2)_2$
9. Methanol is an alternative fuel to petrol that has been used in experimental cars. Determine the empirical formula of methanol, given that it is composed of 49.9% oxygen, 37.5% carbon and 12.6% hydrogen.
10. Calculate the molecular formulas for the compounds labelled a. to g. from the information in the following table.

Compound	Empirical formula	Molar mass (g mol^{-1})
a.	C_5H_{12}	72.0
b.	C_3H_5	82.0
c.	C_3H_4	80.0
d.	CH_2	84.0
e.	CH_2	140.0
f.	CH_2O	180.0
g.	CH_2Cl	99.0

6.4 Exam questions

Question 1 (1 mark)

MC Which of the following has the greatest percentage by mass of oxygen?
A. Magnesium oxide
B. Methanoic acid, HCOOH
C. Copper(II) oxide
D. Water

Question 2 (3 marks)

A compound of gadolinium (element 64) and sulfur contains 23.44% sulfur. What is the empirical formula of the compound?

Question 3 (2 marks)

A hydroxide of iron was found to contain 62.14% Fe and 35.63% O and 2.23% H. What is the empirical formula of the iron hydroxide?

Question 4 (2 marks)

The odour of rancid butter is caused by butyric acid, which contains 54.5% carbon and 9.1% hydrogen, the rest being oxygen. Given that the M_r of butyric acid is 88.0, calculate the molecular formula of the acid.

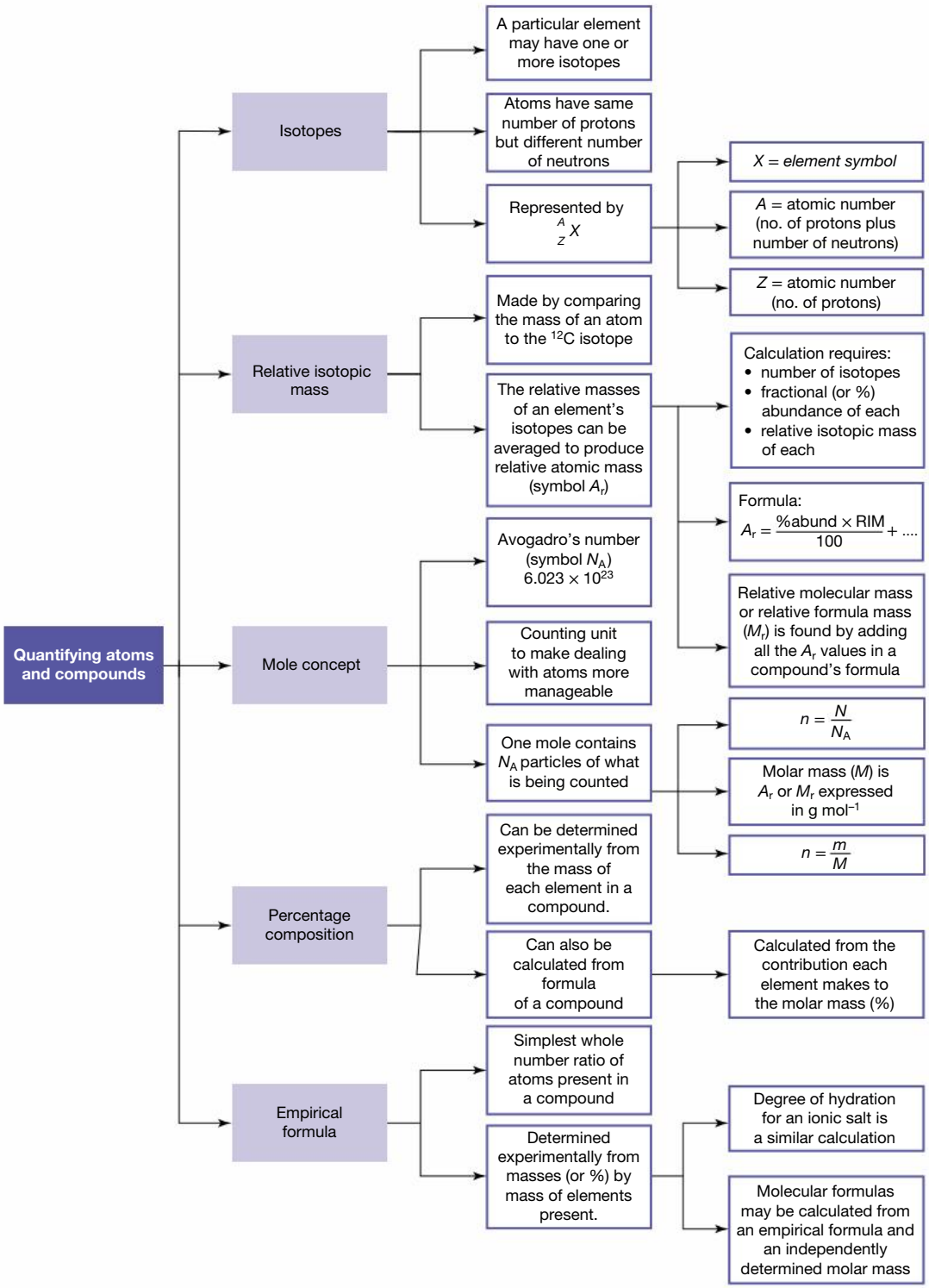
Question 5 (1 mark)

Galena, PbS, is the most common ore mined for lead. A sample of an impure galena bearing ore was found to have lead composition of 78.4%. What mass (in grams) of this ore would be required to produce 1 tonne of lead?

More exam questions are available in your learnON title.

6.5 Review

6.5.1 Topic summary



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6.5.2 Key ideas summary

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6.5.3 Key terms glossary

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6.5 Activities

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6.5 Review questions

1. Silicon-containing ores have three isotopes: 92% silicon-28, 5% silicon-29 and 3% silicon-30.
   a. Draw the mass spectrum for silicon.
   b. What is the atomic number of silicon?
   c. What are the relative isotopic masses of the three silicon isotopes?
   d. What is the relative atomic mass of silicon?
2. In 2.0 moles of $(NH_4)_3PO_4$ (fertiliser) are
   a. _____ moles of nitrogen atoms.
   b. _____ moles of hydrogen atoms.
   c. _____ moles of oxygen atoms.
   d. _____ moles of phosphorus atoms.
   e. _____ (total number) nitrogen atoms.
   f. _____ grams of nitrogen atoms.
3. Complete the following table.

Name	Formula	Molar mass (*M*)	Mole (*n*)	Mass (*m*)
Sodium hydroxide	NaOH			3.41 g
	CCl_4		1.40 mol	
Sodium carbonate		106 g mol^{-1}	1.00 mol	
	KCl		0.25 mol	
Ammonium phosphate				8.46 g

4. Calculate the mass of each of the following.
   a. 160 mol Fe
   b. 0.075 mol silicon dioxide
   c. 4.23 mol NO_2

5. Sodium fluoride is thought to reduce tooth decay, especially in children. It is, therefore, added to some brands of toothpaste. If a tube of toothpaste contains 0.013 g of sodium fluoride, answer the following.
   a. How many moles of sodium fluoride does this represent?
   b. How many fluoride ions does this represent?
6. Diamond is a naturally occurring form of pure carbon. The mass of a diamond is measured in a unit called a carat, where 1.00 carat = 0.200 g. How many atoms of carbon are in the following?
   a. A 1.00-carat diamond b. A 3.15-carat diamond
7. Aspirin is a drug that has been widely used as a pain killer and is still used in some heart medications. It has the formula $C_9H_8O_2$.
   Calculate the percentage composition by mass of each element in aspirin.
8. A pure sample of an oxide of uranium is analysed and found to contain 84.8% uranium.
   Calculate its empirical formula.
9. Silver is an unreactive metal that is used to produce jewellery. Its relative atomic mass is 107.9, and it consists of two naturally occurring isotopes: silver-107 and silver-109. Calculate the abundance of the lighter isotope.
10. Caffeine is a stimulant that is found in coffee, tea and chocolate. Analysis of caffeine shows that it contains 49.5% carbon, 28.9% nitrogen, 16.5% oxygen and 5.1% hydrogen by mass. Determine the molecular formula of caffeine given that its molar mass is 194.2 g mol^{-1}.

6.5 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC Rubidium has two stable isotopes. Rubidium-85 has an abundance of 72% and rubidium-87 has an abundance of 28%. What does the mass spectrum of rubidium have?

A. Two peaks: a smaller peak at relative isotopic mass 85 and a larger peak at relative isotopic mass 87
B. Two peaks: a smaller peak at relative isotopic mass 87 and a larger peak at relative isotopic mass 85
C. One peak at about relative isotopic mass 85.5
D. One peak at about relative isotopic mass 86

Question 2

MC A certain element has two isotopes of relative isotopic mass 203 and 205. Given that its A_r value is 204.4, which of the following is the most likely percentage abundance of the heavier isotope?

A. 7% **B.** 30% **C.** 50% **D.** 70%

Question 3

MC Element 77, iridium, has two stable isotopes: ^{191}Ir and ^{193}Ir. Its relative atomic mass is 192.2.

Based on the value of the relative atomic mass, what can be surmised?

A. There is a 50 : 50 ratio of the two isotopes.
B. About $\frac{3}{4}$ of the isotopes are ^{193}Ir.
C. There is a slight majority of ^{191}Ir isotope.
D. There is a slight majority of ^{193}Ir isotope.

Question 4

MC A student weighs out 1.463 g of anhydrous sodium carbonate. (Anhydrous means the compound contains no water.) What is the amount that this represents?

A. 0.013 80 mol **B.** 0.017 63 mol **C.** 1.463 mol **D.** 72.45 mol

Question 5

MC Which of the following contains the most hydrogen atoms?

A. 26.0 g of hydrogen gas, H_2
B. 180 g of water, H_2O
C. 96 g of methane, CH_4
D. 360 g of glucose, $C_6H_{12}O_6$

Question 6

MC Which of the following contains the fewest atoms?

A. 1.50 mol of zinc, Zn
B. 2.50 mol of iodine, I_2
C. 1.00 mol of water
D. 0.10 mol of $CuSO_4$

Question 7

MC Consider a 0.0100 mol sample of pure caffeine, $C_8H_{10}N_4O_2$. Which of the following statements is/are correct?

- Statement I: The sample has a mass of 1.94 g.
- Statement II: The sample contains 6.02×10^{22} caffeine molecules.
- Statement III: The sample contains 6.02×10^{22} hydrogen atoms.
- Statement IV: The sample contains twice as many carbon atoms as nitrogen atoms.

A. Statements I and IV only
B. Statements II and IV only
C. Statements I, III and IV only
D. All statements are correct.

Question 8

MC Some students carried out an experiment to determine the empirical formula of an oxide of magnesium by heating a strip of pure magnesium ribbon strongly in a mostly covered crucible until it formed a white solid.

If they did not cover their crucible enough and allowed a large quantity of white 'smoke' oxide to escape the crucible, what would their calculation of the ratio Mg/O be?

A. Lower, since a lower magnesium oxide mass gives the impression that less oxygen was combined
B. Higher, since a lower magnesium oxide mass gives the impression that less oxygen was combined
C. Higher, since a greater amount of oxygen would be able to enter the crucible and combine with magnesium
D. Lower, since a greater amount of oxygen would be able to enter the crucible and combine with magnesium

Question 9

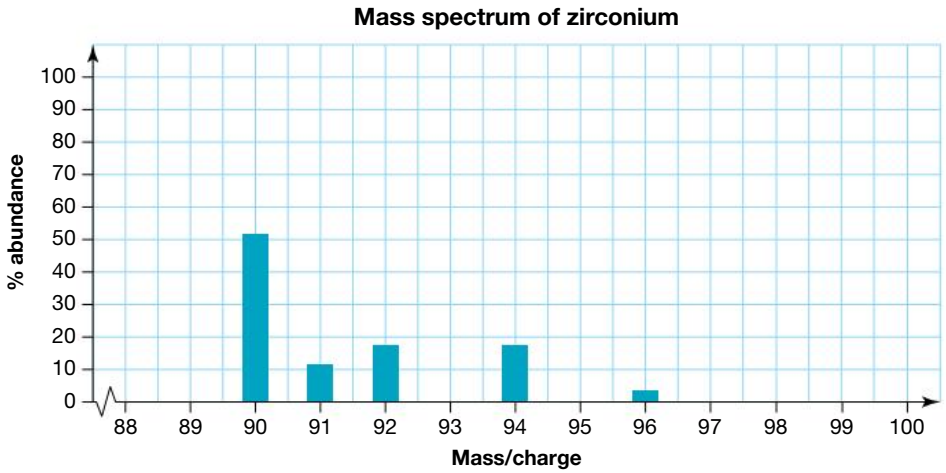
MC A particular amino acid has the following composition by mass: 32.0% carbon, 6.67% hydrogen, 18.67% nitrogen and 42.67% oxygen. The molar mass of the compound is 75 g mol^{-1}.

What is the compound's molecular formula?

A. The same as its empirical formula
B. Twice its empirical formula
C. Three times its empirical formula
D. Not a multiple of its empirical formula

Question 10

MC Consider the following mass spectrum for the stable isotopes zirconium.



Which of the following statements is *incorrect*?

A. The abundance of ^{92}Zr is approximately equal to the abundance of ^{94}Zr.
B. The most abundant isotope of zirconium has 50 neutrons.
C. 53 neutrons would present an unstable nuclear configuration for zirconium.
D. Zirconium does not form stable isotopes with an odd number of neutrons.

Section B — Short answer questions

Question 11 (3 marks)

To prevent a gum disease called scurvy, the minimum daily requirement of vitamin C, $C_6H_8O_6$, is 60 mg.

a. How many moles of vitamin C is this? **(1 mark)**
b. How many molecules is this? **(1 mark)**
c. If 10 g of spinach is found to contain 1.2×10^{-5} g of vitamin C, how much spinach must be eaten to attain the minimum daily requirement? **(1 mark)**

Question 12 (3 marks)

Answer the following.

a. Oxalic acid is a crystalline solid with the molecular formula $H_2C_2O_4$. A sample of pure oxalic acid was found to have a mass of 6.53 g.

i. Find the amount (in mol) of oxalic acid. **(1 mark)**

ii. Find the number of oxygen atoms in this mass of oxalic acid. **(1 mark)**

b. 0.0150 mol of a mystery element has the mass of 1.827 g. What is the element? (Show your working.) **(1 mark)**

Question 13 (2 marks)

Insects of a particular species can identify their mate by using special chemicals called pheromones that transmit chemical messages. The pheromone that serves as a sex attractant for gypsy moths is called disparlure and contains the elements C, H and O. Analysis of disparlure shows that 0.282 g contains 16.00×10^{-3} g of O atoms and 0.228 g of C atoms. The M_r of disparlure is 282 g mol^{-1}. Determine the empirical formula of disparlure.

Question 14 (3 marks)

An oxide of copper is heated in a stream of hydrogen until only the copper remains, according to the equation:

$$Cu_xO\,(s) + H_2\,(g) \rightarrow xCu\,(s) + H_2O\,(l)$$

The data for the experiment is given in the following table. Calculate the empirical formula of the oxide of copper.

Item	Mass (g)
Crucible	27.002
Crucible plus contents before heating	27.128
Crucible plus contents after heating	27.114

Question 15 (2 marks)

A particular organic compound with the non-systematic name of '12-Crown-4' has a molar mass of 128 g mol^{-1} and consists of 37.5% C and 12.5% H, with the remainder being oxygen.

a. Find the empirical formula of this compound. **(1 mark)**

b. Find the molecular formula of this compound. **(1 mark)**

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7 Families of organic compounds

KEY KNOWLEDGE

In this topic you will investigate:

Families of organic compounds

- the grouping of hydrocarbon compounds into families (alkanes, haloalkanes, alkenes, alcohols, carboxylic acids) based upon similarities in their physical and chemical properties including general formulas and general uses based on their properties
- representations of organic compounds (structural formulas, semi-structural formulas) and naming according to the International Union of Pure and Applied Chemistry (IUPAC) systematic nomenclature (limited to non-cyclic compounds up to C8, and structural isomers up to C5)
- plant-based biomass as an alternative renewable source of organic chemicals (for example, solvents, pharmaceuticals, adhesives, dyes and paints) traditionally derived from fossil fuels
- materials and products used in everyday life that are made from organic compounds (for example, synthetic fabrics, foods, natural medicines, pesticides, cosmetics, organic solvents, car parts, artificial hearts), and an assessment of the benefits of those products for society, as well as the health and/or environmental hazards they pose.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of VCE Chemistry. Experiments and investigations, supported by a **practical investigation eLogbook** and **teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

EXAM PREPARATION

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7.1 Overview

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7.1.1 Introduction

Carbon is an exceptional element for many reasons. It forms the basic building blocks of life; it is present in every cell in the body. It is in the food that you eat, the clothes that you wear, and in medications and fuels.

More compounds of carbon exist than compounds of all the other elements combined. Organic chemistry is the study of these compounds. Historically, scientists thought that these carbon compounds could only be made by living organisms — hence the term *organic*. Today, however, we know that this is not the case.

FIGURE 7.1 Carbon is present in the food, fuel and people at this barbeque.

Transport, electricity and heating all rely heavily on hydrocarbons such as coal, petrol and diesel, which are among the simplest of all organic compounds and are derived from crude oil. However, increasing awareness of the devastating impacts of climate change has forced us to consider our uses of carbon, and hydrocarbons in particular.

This topic introduces how chemists sort this vast catalogue of compounds into meaningful groups or families. You will learn about the different types of hydrocarbons and the systematic way of naming compounds based on their structure. You will also explore the environmental impacts of organic compounds and the new ways scientists are using plant-based biomass as a renewable source of the chemicals and products we all use.

LEARNING SEQUENCE

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Solutions	Solutions — Topic 7 (sol-0806)
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Exam question booklet	Exam question booklet — Topic 7 (eqb-0088)

7.2 Hydrocarbon families

KEY KNOWLEDGE

- The grouping of hydrocarbon compounds into families (alkanes, haloalkanes, alkenes, alcohols, carboxylic acids) based upon similarities in their physical and chemical properties including general formulas and general uses based on their properties
- Representations of organic compounds (structural formulas, semi-structural formulas)

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7.2.1 What is organic chemistry?

Organic chemistry is the systematic study of the compounds of carbon, excluding oxides of carbon, carbonates, carbides and cyanides. These exceptions plus the compounds of the remaining elements are considered to be inorganic. Carbon compounds comprise over 90 per cent of all known chemicals and carbon is the only element with properties that make the development of all living systems possible.

The main reason for carbon's unique ability to form a wide range of chemicals is that carbon–carbon bonds are strong. Carbon normally forms four covalent bonds. These strong bonds can form with other non-metals. Carbon can also chemically bond with itself using single, double or triple bonds to form long chain-like structures and even rings.

FIGURE 7.2 Every cell in our body requires the essential organic compound glucose, $C_6H_{12}O_6$. This molecule is used to produce energy for several body tissues, particularly the brain.

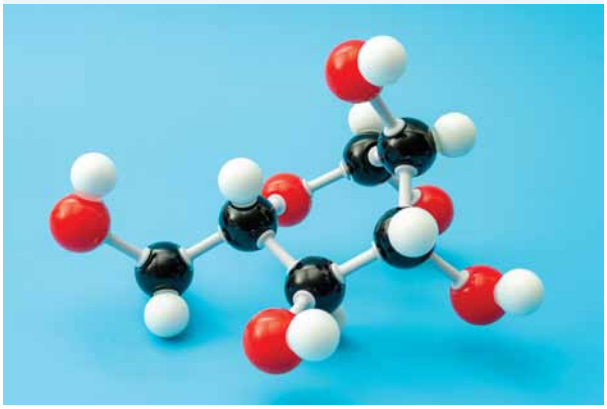
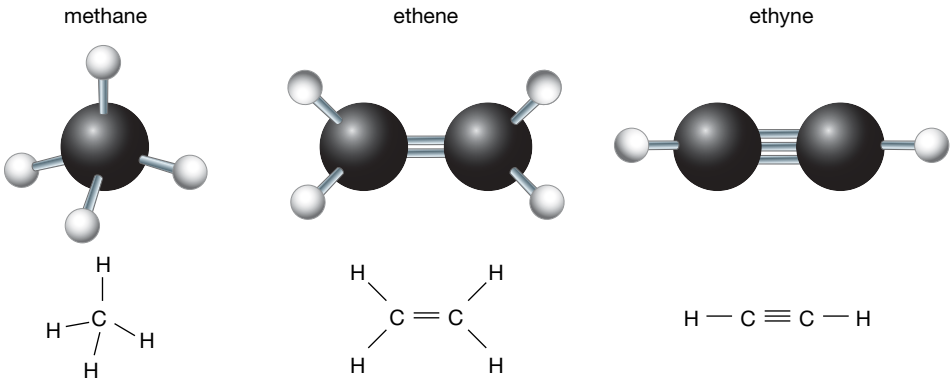


FIGURE 7.3 Ball and stick models and structural formulas of methane, ethene and ethyne



7.2.2 Introduction to hydrocarbons

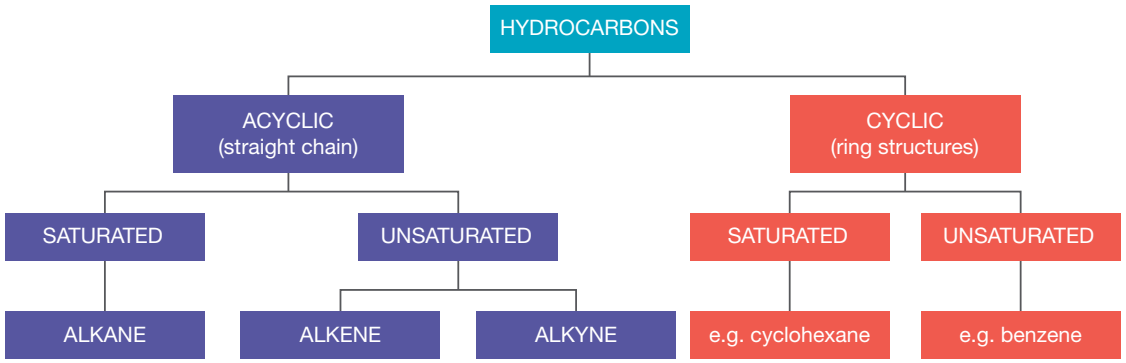
A good place to begin a study of organic chemistry is with **hydrocarbons**. Hydrocarbons are molecules made up of only hydrogen and carbon. These are the simplest organic compounds but are among the most useful. Hydrocarbons are found in **crude oil**, and processing the oil yields many useful products. This process and the products that are formed are covered in more detail later in this topic. The different classes of hydrocarbons are based on the carbon bonding involved. These are indicated in figure 7.4.

organic chemistry the study of carbon-containing compounds and their properties

hydrocarbon compound containing only carbon and hydrogen

crude oil naturally occurring mixture of compounds that are mainly hydrocarbons; also known as petroleum

FIGURE 7.4 The classes of hydrocarbons



Acyclic hydrocarbons have the carbon atoms arranged in a **straight chain** and include **branched** molecules. The following are acyclic hydrocarbons:

- **Alkanes**, which have single carbon-to-carbon bonds. They are examples of **saturated hydrocarbons**.
- **Alkenes**, which contain one carbon-to-carbon double bond.
- **Alkynes**, which contain one carbon-to-carbon triple bond.

Alkenes and alkynes are **unsaturated hydrocarbons** and contain double or triple carbon–carbon bonds.

Cyclic compounds have the carbon atoms arranged in a ring.

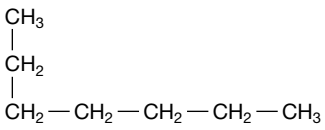
The simplest hydrocarbon, methane, is an alkane. It is a single carbon atom bonded to four hydrogen atoms. The carbon atom forms four equivalent bonds that are 109.5° apart, forming a tetrahedral shape. Figure 7.6 demonstrates different representations of methane. Of these, structural formulas (and semi-structural formulas where appropriate) are the most used representations in organic chemistry.

FIGURE 7.5 Acyclic molecules can be straight chain or branched.

(a)

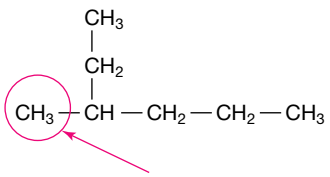
$CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-CH_3$

(b)



These molecules are straight chains even though there is a bend in molecule **b**.

(c)



This molecule is branched.

acyclic hydrocarbons based on straight and branched chains of carbon atoms

straight chain carbon atoms connected in one continuous line without any branches

branched molecule has alkyl groups attached along the carbon chain

alkane a compound of carbon and hydrogen containing only single bonds, no rings, and having the general formula C_nH_{2n+2}

saturated hydrocarbon compound composed of carbon and hydrogen and containing only single covalent bonds

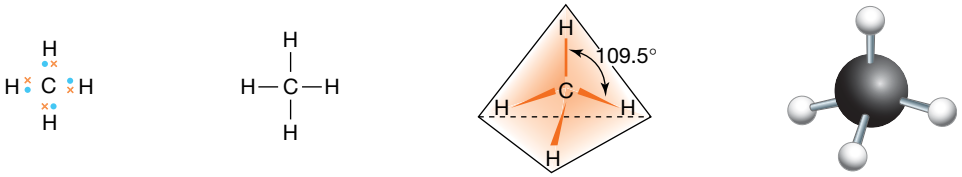
alkene a hydrocarbon that contains one carbon–carbon double bond and the general formula C_nH_{2n}

alkyne a hydrocarbon that contains one carbon–carbon triple bond and the general formula C_nH_{2n-2}

unsaturated hydrocarbon compound containing carbon and hydrogen but possessing one or more double or triple bonds

cyclic compound where the atoms are connected in a ring

FIGURE 7.6 Electron dot structure, structural formula, molecular shape, and ball-and-stick model of methane



General properties of hydrocarbons

In hydrocarbon molecules, the electronegativities of carbon and hydrogen atoms are similar; this means that the C—H bond is nearly non-polar and, because of the symmetry of the molecule, hydrocarbon molecules are non-polar. The bonds formed within a hydrocarbon molecule, called intramolecular bonds, are strong covalent bonds. Because hydrocarbon molecules are non-polar, the only intermolecular forces acting between the molecules are dispersion forces. Other general properties include the following:

- Smaller hydrocarbons have very low melting and boiling points because of the very weak dispersion forces between the molecules. As the length of the chain of carbon atoms increases, so does the melting and boiling point because more dispersion forces are present. Table 7.1 shows this trend in the alkanes. This explains why smaller alkanes are gases, and larger alkanes are solids at room temperature.
- If a molecule is branched instead of a straight chain, this lowers the boiling point. Branches prevent the molecules from coming closer together and so the dispersion forces have less effect. Alkenes have a lower melting point than the corresponding alkanes because the structure of alkene molecules also prevents the molecules becoming close.
- Hydrocarbons are insoluble in water because the water molecules have relatively strong hydrogen bonding between their molecules. To dissolve, the hydrocarbon molecules would need to interact with the water molecules. The hydrocarbon molecules cannot cause the water molecules to separate to form these interactions. Hydrocarbons are more likely to be soluble in non-polar solvents. This is sometimes referred to as the 'like dissolves like' rule.

Intermolecular forces

The stronger the intermolecular forces, the higher the melting and boiling points of compounds.

Uses of hydrocarbons

Hydrocarbons are among the most useful substances in today's society (see table 7.1). For example:

- Alkanes (see section 7.2.3) are widely used as transport fuels. They burn readily and release a large amount of energy in doing so. However, the crude oil from which they are derived is a finite resource. Their use as a fuel also produces carbon dioxide, levels of which are now building up in the atmosphere.
- Alkenes (section 7.2.6) are widely used in the petrochemical industry as pre-cursors for many other useful substances. They are also the raw materials for the manufacture of a wide range of plastics (see topic 8).
- The most useful alkyne is ethyne (known in everyday use as acetylene), a gas used with oxygen for high temperature cutting and welding.

TABLE 7.1 Common straight-chain alkanes

Formula	Name	Phase	Boiling point	Typical use
CH_4	Methane	Gas	–161.5 °C	Natural gas
C_2H_6	Ethane	Gas	–89 °C	Refrigerant
C_3H_8	Propane	Gas	–42 °C	Liquid petroleum gas
C_4H_{10}	Butane	Gas	–1 °C	Manufacture of synthetic rubber
C_5H_{12}	Pentane	Liquid	36.1 °C	Solvent
C_6H_{14}	Hexane	Liquid	68 °C	Filling for thermometers
C_7H_{16}	Heptane	Liquid	98 °C	Testing engine knocking
C_8H_{18}	Octane	Liquid	125.6 °C	Automobile petroleum
C_9H_{20}	Nonane	Liquid	151 °C	Gasoline
$C_{16}H_{34}$	Hexadecane	Liquid	286.8 °C	Lubricating oil
$C_{20}H_{42}$	Icosane	Solid	343.1 °C	Wax candles
$C_{28}H_{58}$	Octacosane	Solid	440 °C	Tar

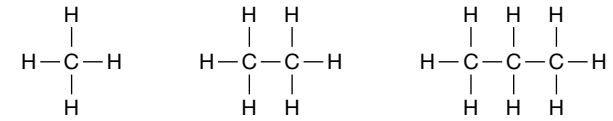
7.2.3 Alkanes

The alkanes are a family of hydrocarbons containing only **single bonds** between the carbon atoms. Notice how the name of each alkane has the same ending: *-ane*. This is how alkanes are identified. The different prefixes, *meth-, eth-, prop-* and so on, show how many carbon atoms are present in each carbon chain (see table 7.2). Table 7.1 summarises some common straight-chain alkanes and their uses.

TABLE 7.2 Standard prefixes for indicating numbers of carbon atoms

Number of carbon atoms	1	2	3	4	5	6	7	8	9	10
Prefix	*meth-*	*eth-*	*prop-*	*but-*	*pent-*	*hex-*	*hept-*	*oct-*	*non-*	*dec-*

FIGURE 7.7 Structural and molecular formulas of the first three alkanes



methane, CH_4 ethane, C_2H_6 propane, C_3H_8

Any series of organic compounds in which each successive member differs by CH_2 from the previous one is called a **homologous series**. Alkanes can be represented by the general formula C_nH_{2n+2}, where n is the number of carbon atoms in the molecule.

Alkanes are also called saturated hydrocarbons. This refers to the fact that all the carbon-to-carbon bonds are single bonds. This means that, for a given number of carbons, the maximum number of hydrogens possible in each molecule.

single bond formed by one pair of electrons shared between two atoms

homologous series series of organic compounds with the same general formula, in which the formula of one molecule differs from the next by a $-CH_2$ group

Homologous series

A homologous series of organic compounds has the same general formula and only differs between members by CH_2.

General formula for alkanes

Alkanes are saturated hydrocarbons with single bonds between the carbon atoms. They have the general formula C_nH_{2n+2}.

Alkanes are valuable as raw materials in the chemical industry and as the main component of petrol and lubricants. Smaller alkanes are used as fuels and for heating and cooking because they are more easily vaporised (more volatile) and flammable. Liquid alkanes are also used in fuels and as solvents.

on Resources

Video eLesson Naming alkanes (eles-2484)

tlvd-0539

SAMPLE PROBLEM 1 Determining the formula of alkanes

Write the formula for the alkane that has 12 carbon atoms per molecule.

THINK

Alkanes have the general formula C_nH_{2n+2}.
If 12 carbons are present, $n = 12$.
The number of hydrogen atoms will be $2n+2 = 2 \times 12 + 2 = 26$.
This molecule is called dodecane.

WRITE

$C_{12}H_{26}$

PRACTICE PROBLEM 1

Write the formula for the alkane that has 16 carbon atoms per molecule.

7.2.4 Representing organic molecules

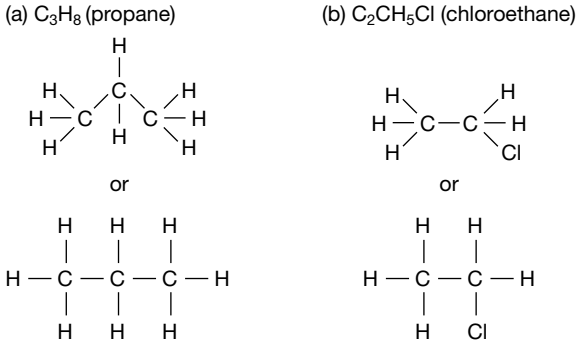
Notice that the equation in figure 7.8 uses structural formulas, which show clearly where atoms are gained and lost in a reaction. Organic molecules can be represented in a variety of ways, as shown in figures 7.8 and 7.9, depending on the context in which they are used. Naming organic compounds will be explained in section 7.3.1. Compounds containing up to eight carbons will be discussed.

Structural formulas

As discussed in topic 3, the **structural formula** of a substance shows the way the atoms in a molecule are connected; that is, it shows all covalent bonds. Although straight-chain alkanes have a zig-zag backbone, they are often drawn with the bonds at right angles to make it easier to see the individual bonds. It is important to remember, however, that the arrangement around the carbon atoms is tetrahedral. Figure 7.8 shows the structural formulas for propane and chloroethane.

structural formula diagrammatic representation of a molecule showing every bond

FIGURE 7.8 Structural formulas for (a) propane and (b) chloroethane

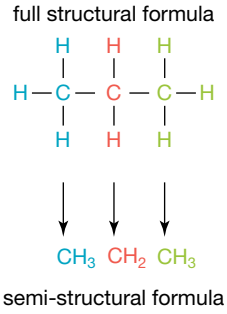


For organic molecules, it is important to show all bonds when structural formulas are drawn.

Semi-structural formulas

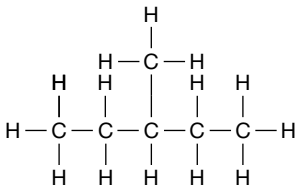
As the name suggests, a **semi-structural formula** (also known as a *condensed formula*), is an abbreviated form of the structural formula. It is frequently used in organic chemistry, especially as molecules become larger. A semi-structural formula avoids the tedious task of showing all bonds and all atoms by grouping sets of atoms together. It shows an abbreviation of the structural formula, on a single line. For example, ethane, C_2H_6, can be represented by the condensed formula CH_3CH_3. The semi-structural formula of propane, C_3H_8, is $CH_3CH_2CH_3$. When abbreviating a structural formula to its condensed formula equivalent, go 'carbon by carbon'.

FIGURE 7.9 Condensing the full structural formula of propane into the semi-structural formula



In a semi-structural formula, brackets are used:

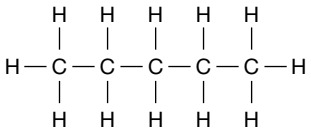
- to indicate side chains, which are written after the carbon to which they are attached — for example:



3-methylpentane

$CH_3CH_2CH(CH_3)CH_2CH_3$

- if there are successive, repeating CH_2 groups — for example:



pentane

$CH_3CH_2CH_2CH_2CH_3$ or $CH_3(CH_2)_3CH_3$

semi-structural formula derived from the structural formula that shows the grouping of atoms in a molecule, carbon by carbon, and is written as a single line of text; also called a condensed formula

empirical formula formula that shows the simplest numerical ratio in which atoms are combined

molecular formula actual number of atoms of each element in a molecule of a substance

skeletal formula structural representation where a molecule is represented by lines and vertices; carbons and hydrogens are not shown but assumed to be at each vertex and at the ends unless otherwise indicated

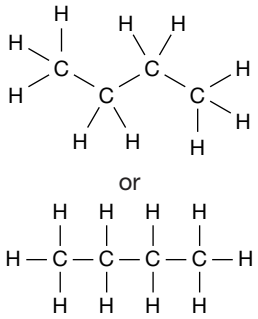
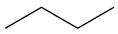
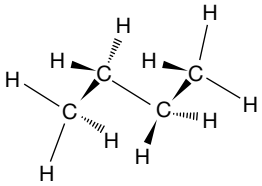
Other formula types

Other commonly used types of formulas are **empirical formulas**, **molecular formulas**, **skeletal formulas** and three-dimensional structural formulas. Examples using butane are shown in table 7.3.

on Resources

Video eLesson Molecular representations of butan-1-ol (eles-2485)

TABLE 7.3 Different ways of representing butane molecules

Formula	Example	Notes
Molecular	C_4H_{10}	This shows the actual number and kinds of atoms in a molecule.
Empirical	C_2H_5	This shows the simplest whole number ratio of atoms in a molecule.
Structural		This shows the actual arrangement of atoms in a molecule. As molecules become longer, the second example of a structural formula tends to be used.
Semi-structural or condensed	$CH_3CH_2CH_2CH_3$ or $CH_3(CH_2)_2CH_3$	These can be written on a single line, with each carbon atom being followed by the atoms that are joined to it. Repeated CH_2 groups can be collected together in brackets with a subscript as shown.
Skeletal		These use lines and vertices to simplify a structural formula. It is assumed that a carbon atom (and enough hydrogens to satisfy carbon's valency) is present at each vertex and at the ends. Double bonds and other different types of atoms are specifically shown. This representation is usually used for more complex molecules and will be encountered in Units 3 and 4.
3D structural or shape diagram		This shows the 3D arrangement of atoms, with the bonds represented as follows: • The continuous line is in the plane of the paper. • The dashed line extends to the back of the paper. • The solid wedge comes out of the plane of the paper.

7.2.5 Reactions of alkanes

Combustion reactions

Alkanes are comparatively unreactive but can burn (combust) in a plentiful supply of oxygen to produce carbon dioxide and water. However, if the supply of oxygen is limited, carbon monoxide or even just carbon may be produced in place of carbon dioxide. The production of carbon monoxide is to be avoided because carbon monoxide is toxic. Chemists use chemical equations to summarise chemical reactions. Using methane as an example, the reactions mentioned may be summarised as follows.

In a plentiful supply of oxygen

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$
$$\text{methane} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water}$$

and in a limited supply of oxygen

$$CH_4(g) + 1\tfrac{1}{2}O_2(g) \rightarrow CO(g) + 2H_2O(g)$$
$$\text{methane} + \text{oxygen} \rightarrow \text{carbon monoxide} + \text{water}$$

or

$$CH_4(g) + O_2(g) \rightarrow C(s) + 2H_2O(g)$$
$$\text{methane} + \text{oxygen} \rightarrow \text{carbon} + \text{water}$$

The second equation can also be doubled, to remove the fractional coefficient:

$$2CH_4(g) + 3O_2(g) \rightarrow 2CO(g) + 4H_2O(g)$$

Both are acceptable equations.

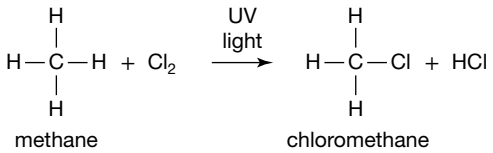
The effect of these reactions can be seen when using a Bunsen burner. When the air hole of a Bunsen burner is open and oxygen is plentiful, methane burns with a blue flame. When the air hole is closed, limiting the supply of oxygen, a yellow flame is the result of the particles of carbon produced (figure 7.10).

FIGURE 7.10 Methane burning with a limited supply of oxygen produces a yellow flame.

Substitution reactions

Alkanes react with chlorine or fluorine in a reaction called a **substitution reaction**. In this type of reaction, one or more of the chlorine or fluorine atoms take the place of one or more of the hydrogen atoms. The products of these reactions are known as **haloalkanes**. For example, in a reaction between methane, CH_4, and chlorine gas, Cl_2, the hydrogen atoms may be substituted in sequence by chlorine to produce a mixture of CH_3Cl, CH_2Cl_2, $CHCl_3$ and CCl_4. The first step in this process is shown in figure 7.11.

FIGURE 7.11 Methane reacts with chlorine gas under UV light to form chloromethane.



7.2.6 Haloalkanes: introduction to functional groups

Haloalkanes are the products of alkane and halogen substitution reactions. The presence of electronegative halogen atoms affects the physical properties of the haloalkanes.

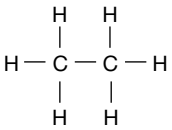
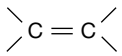
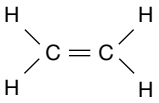
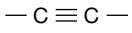
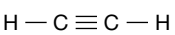
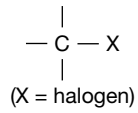
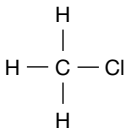
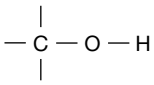
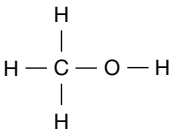
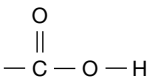
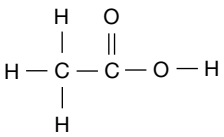
Haloalkanes are not classified as hydrocarbons, because they contain a halogen atom (Cl, F, Br or I). The halogen atom that replaces a hydrogen atom is referred to as a **functional group**. A functional group is an atom or group of atoms that is responsible for characteristic chemical reactions of that homologous series. Examples of some functional groups are listed in table 7.4.

substitution reaction reaction in which an atom (or group of atoms) is removed and replaced by a different atom (or group of atoms)

haloalkane compound with the same structure as an alkane, except that one or more of the hydrogen atoms has been replaced with halogen atoms

functional group an atom or group of atoms attached to or part of a hydrocarbon chain that influence the physical and chemical properties of the molecule

TABLE 7.4 Functional groups

Functional group	Type of compound	Suffix or prefix	Example	Systematic name
	Alkane	*-ane*		Ethane
	Alkene	*-ene*		Ethene
	Alkyne	*-yne*		Ethyne
 (X = halogen)	Haloalkane	*halo-*		Chloromethane
	Alcohol	*-ol*		Methanol
	Carboxylic acid	*-oic acid*		Ethanoic acid

Haloalkanes

The haloalkanes belong to a homologous series comprising of an alkane with one or more hydrogen atoms replaced by a halogen atom.

Properties of haloalkanes

- Haloalkanes have higher melting and boiling points than the corresponding alkanes because their molecules are polar. (The halogen atoms are strongly electronegative.)
- The melting and boiling points of each type of haloalkanes — chloroalkanes, for example — increase with increasing molecular size.
- Haloalkanes are only slightly soluble in water but are more soluble in organic solvents. They are used as organic solvents because they can dissolve non-polar compounds and, apart from fluoroalkanes, they are more reactive than alkanes. They can be reacted to form several other chemicals, including alcohols and amines. (These compounds will be studied in units 3 and 4.)

FIGURE 7.12 Some types of haloalkanes are known as CFCs, which have contributed to the depletion of the ozone layer.

Haloalkanes are used as flame retardants in fire extinguishers, refrigerants, aerosol spray propellants, solvents and pharmaceuticals. They are also used in the synthesis of other organic chemicals. Because of their toxicity and impact on the environment, the use of some haloalkanes has been reduced over time.

7.2.7 Alkenes

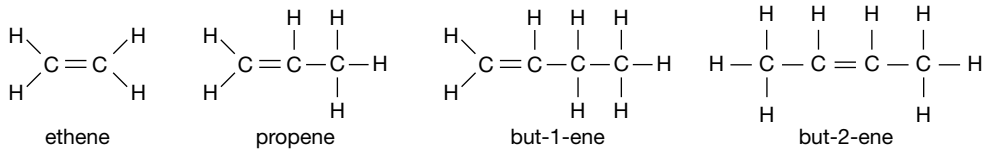
The alkenes make up a family of hydrocarbons that each contains a **double bond** *between two carbon atoms*.

The alkenes have a general formula C_nH_{2n} and are another example of a homologous series. The double bond between the carbon atoms reduces the number of hydrogen atoms in an alkene compared to their corresponding alkane by two.

The structural formulas of the first three alkenes are shown in figure 7.14. Notice that each of the names ends with *-ene*. This is the characteristic ending used for the alkenes.

FIGURE 7.13 This inhaler uses an HFA (1, 1, 1, 2 tetrafluoroethane) propellant. This is a compressed gas that forces the medication out when the inhaler is used.

FIGURE 7.14 Structural formulas of the first three alkenes



General formula for alkenes

Alkenes are hydrocarbons with a double bond between two carbon atoms. They have the general formula C_nH_{2n}.

double bond strong bond between two atoms formed by two pairs of electrons that are shared by the two nuclei

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Video eLesson Homologous series of alkenes (eles-2477)

tlvd-0540

SAMPLE PROBLEM 2 Determining the formula of alkenes

Write the formula for the alkene that has eight carbon atoms per molecule.

THINK

Alkenes have the general formula C_nH_{2n}.
If eight carbons are present, $n = 8$.
The number of hydrogen atoms will be $2n = 2 \times 8 = 16$.
This molecule is called octene.

WRITE

C_8H_{16}

PRACTICE PROBLEM 2

Write the formula for the alkene that has six carbon atoms per molecule.

Reactions of alkenes

Combustion reactions

Alkenes are more reactive than alkanes because of the availability of extra electrons in the double bond to react. They undergo **combustion** reactions with oxygen and, like alkanes, form carbon dioxide and water.

$$\underset{\text{ethene}}{C_2H_4(g)} + 3O_2(g) \rightarrow \underset{\text{carbon dioxide}}{2CO_2(g)} + \underset{\text{water}}{2H_2O(g)}$$

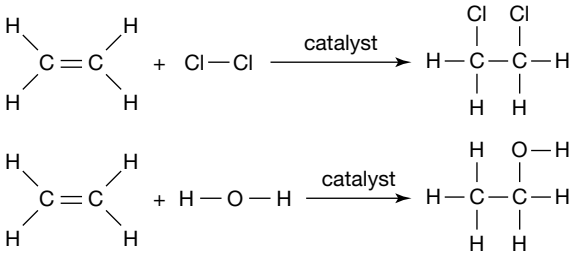
However, this reaction has little commercial value. Small alkenes such as ethene are far more valuable to the chemical industry as raw materials for the manufacture of other chemicals, such as plastics, alcohols and antifreeze. Ethene is also used to ripen fruit.

FIGURE 7.15 Bananas produce ethene as they ripen, which causes other fruit and vegetables to ripen faster.

Addition reactions

Alkenes can undergo **addition reactions**. In these reactions, alkenes react with small molecules such as hydrogen or a halogen. They can also react with water and hydrogen halides. The small molecule attacks and breaks one of the double bonds. In the process, the small molecule is split and is added across the site of the former double bond. With chlorine and water (along with suitable catalysts) the reactions shown in figure 7.16 take place.

FIGURE 7.16 Alkenes break their double bond when they undergo addition reactions.



Alkenes can undergo self-addition in which the alkene molecules join together to make long chains called **polymers**. This will be discussed in topic 8.

Alkenes (and also alkynes) can be referred to as unsaturated hydrocarbons. This is because they contain at least one carbon-to-carbon multiple bond. Consequently, for a given number of carbons, fewer than the maximum number of hydrogen atoms are present per molecule.

combustion reaction between a fuel and oxygen (usually) releasing heat energy

addition reaction a reaction in which one molecule bonds covalently with another molecule without losing any other atoms; only one product is formed

polymer molecule made up of thousands of units (monomers) strung together into very long chains

7.2.8 Alkynes

Alkynes are hydrocarbons that contain a **triple bond** *between two carbon atoms*.

The general formula for the alkynes is C_nH_{2n-2}. The alkynes are another homologous series.

Note that the ending for each of the two alkynes shown in figure 7.17 is *-yne*. This ending shows that a triple bond is present in the molecule.

General formula for alkynes

Alkynes are hydrocarbons with a triple bond between two carbon atoms. They have the general formula C_nH_{2n-2}.

triple bond very strong bond between two atoms formed by three pairs of electrons that are shared by the two nuclei

FIGURE 7.17 Structural formulas of two alkynes

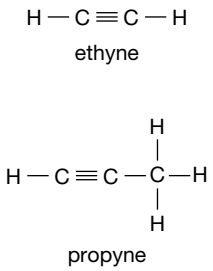


FIGURE 7.18 The combustion of ethyne is used in oxyacetylene welding. The extremely high temperature generated when ethyne is burned with pure oxygen can melt metal and allow surfaces to be welded together.

Reactions of alkynes

Alkynes can undergo combustion (oxidation) and addition reactions. Like all hydrocarbons, the products of the combustion reaction are carbon dioxide and water. The equation for the combustion of ethyne in pure oxygen is

$$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(g).$$

EXTENSION: Cyclic hydrocarbons

The carbon backbone of some hydrocarbons can form a ring. Such compounds are known as cyclic hydrocarbons and have the prefix *cyclo-* before their name.

An important category of hydrocarbons is the unsaturated cyclic compounds making up the aromatic series. The simplest member of this series is benzene, C_6H_6. This molecule has been shown to have two equivalent forms, as shown in figure 7.20 (a) and (b).

Benzene is found naturally in crude oil but is usually synthesised from other hydrocarbons when used in making pharmaceuticals.

FIGURE 7.19 Structural formulas for (a) straight chain hexane and (b) cyclic hexane

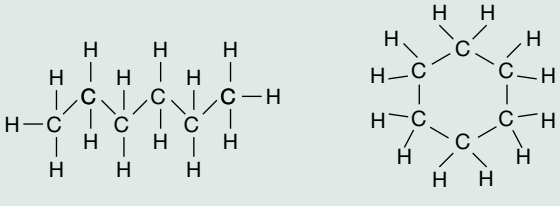
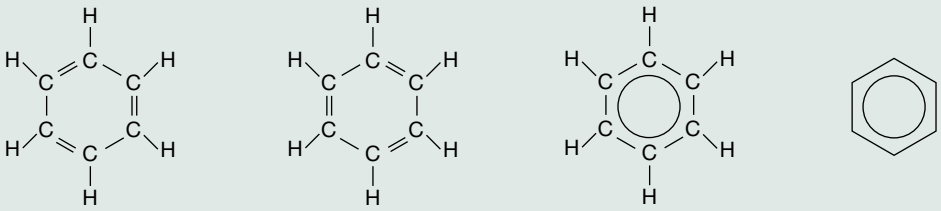


FIGURE 7.20 Four structural representations of benzene

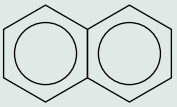
(a) (b) (c) (d)



Scientists have shown experimentally that all carbon–carbon double bonds in benzene are identical and somewhere between single and double bonds in length and bond strength. The equivalence of the bonds is shown by the alternative structure shown in figure 7.20 (c). The circle represents the even distribution of the six electrons around the benzene ring. The six electrons are said to be delocalised. Figure 7.20 (d) shows a shorthand representation of benzene.

Given the information provided for benzene, try to work out the molecular formula for napthalene, represented by the structure in figure 7.21.

FIGURE 7.21 Structural formula of napthalene



7.2 Activities

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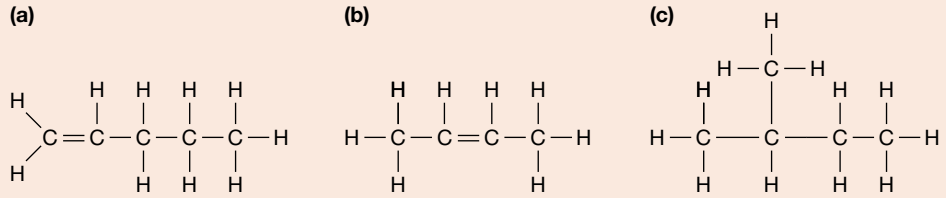
7.2 Quick quiz on | **7.2 Exercise** | **7.2 Exam questions**

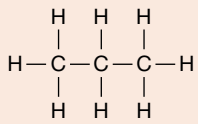
7.2 Exercise

1. MC Which of the following is not classified as an organic molecule?
   - **A.** Theobromine, $C_7H_8N_4O_2$, found in chocolate
   - **B.** Sucrose, $C_{12}H_{22}O_{11}$, also known as sugar
   - **C.** Sulfuric acid, H_2SO_4, found in car batteries
   - **D.** Octane, C_8H_{18}, a component of petrol
2. MC Which of the following lists includes alkanes only?
   - **A.** C_2H_6,C_5H_{10}, C_6H_{12}
   - **B.** C_2H_6, C_4H_{10}, C_6H_{14}
   - **C.** CH_4, C_3H_8, C_7H_{12}
   - **D.** C_3H_6, C_5H_{10}, C_6H_{12}
3. MC What type of reaction is represented by the following equation?
   $CH_4 + Br_2 \rightarrow CH_4Br + HBr$
   - **A.** Substitution
   - **B.** Addition
   - **C.** Saturation
   - **D.** Combustion
4. **a.** Write the formula for the alkane that has 10 carbon atoms per molecule.
   b. Write the formula for the alkene that has 13 carbon atoms per molecule.

5. Write semi-structural formulas for each of the following:

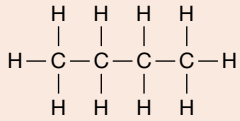


6. What are the products for the combustion of propane in the following circumstances?
   a. A plentiful supply of air
   b. A limited supply of air
7. State two types of reactions that are undergone by the following.
   a. Alkanes
   b. Alkenes
8. As the molecular size of the alkanes increases, a change occurs to the gaseous, liquid and solid phases. Explain why this happens.
9. Which of the following compounds would you expect to have the higher boilling point? Give reasons for your answer.
   a.



propane, C_3H_8

   b.



butane, C_4H_{10}

10. Hexane and 2,3-dimethylbutane both have the formula C_6H_{14}. Their boilling points are –95.3 °C and –128.6 °C. Explain with reference to their structure and bonding why their boilling points are so different.

7.2 Exam questions

Question 1 (1 mark)

MC Which of these statements about alkenes and alkanes is *incorrect*?
A. Alkenes have only double bonds between carbon atoms, whereas alkanes have only single bonds.
B. Alkenes are unsaturated and alkanes are saturated.
C. Alkanes have a higher hydrogen : carbon ratio than alkenes.
D. Both alkanes and alkenes are hydrocarbons.

Question 2 (1 mark)

MC Which of the following lists contains only members of the same homologous series?
A. C_2H_2, C_2H_4, C_2H_6
B. C_3H_8, CH_4, C_5H_{12}
C. CH_3Cl, CH_2Cl_2, $CHCl_3$
D. CH_3COOH, CH_3COOCH_3, C_2H_5OH

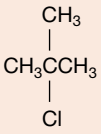
Question 3 (1 mark)

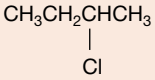
MC Bromine, Br_2, is a reddish-brown liquid used to test for unsaturation because it reacts immediately and the colour disappears. Which of the following statements describes the reaction of bromine with cyclohexene?
A. Bromine is a polar molecule and reacts by adding bromine atoms across the double bond.
B. Bromine is a non-polar molecule and reacts by adding bromine atoms across the double bond.
C. Bromine is a polar molecule and reacts by substituting hydrogen atoms with bromine atoms.
D. Bromine is a non-polar molecule and reacts by substituting hydrogen atoms with bromine atoms.

Question 4 (1 mark)

Consider the following chloroalkanes.

a. $CH_3CH_2CH_2CH_2Cl$

b.


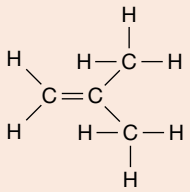
c.


MC Choose the correct order of increasing boiling points.

A. (b) < (a) < (c)
B. (b) < (c) < (a)
C. (c) < (a) < (b)
D. (c) < (b) < (a)

Question 5 (6 marks)

Consider the following molecule.



a. What is its semi-structural formula? **(1 mark)**
b. To what homologous series does it belong? **(1 mark)**
c. How many hydrogen atoms would be in the next member of this homologous series? **(1 mark)**
d. Is it saturated or unsaturated? Explain your answer. **(1 mark)**
e. Explain whether this is a polar or non-polar molecule. **(1 mark)**
f. Name the intermolecular forces operating between molecules of this compound. **(1 mark)**

More exam questions are available in your learnON title.

7.3 Naming organic compounds and isomers

KEY KNOWLEDGE

- Naming of organic compounds according to the International Union of Pure and Applied Chemistry (IUPAC) systematic nomenclature (limited to non-cyclic compounds up to C8, and structural isomers up to C5)

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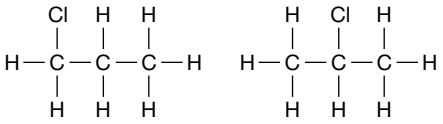
7.3.1 Naming hydrocarbons

Over the years, as organic chemistry grew and more and more compounds were either discovered or synthesised, naming them increasingly became a problem.

The existence of isomers also causes naming issues. Isomers have the same molecular formula but different structural formulas. Only single structures are possible for methane, ethane and propane. However, two positions for the chlorine atom are possible for chloropropane; each requires a name because they will have slightly different properties.

To solve the naming problem, the International Union of Pure and Applied Chemistry (IUPAC) developed a set of rules by which any organic substance could be systematically named depending on the structure of its molecules. This system is in widespread use today, although many common substances still retain their 'old' or 'trivial' names in everyday use. For example, ethyne, mentioned earlier in this topic, is still called acetylene in everyday use.

FIGURE 7.22 Two possible structures for chloropropane, C_3H_7Cl


1-chloropropane

2-chloropropane

Rules for naming hydrocarbons

- Rule 1: Determine the longest chain of carbon atoms. This will form the basis of the name.
- Rule 2: Determine which end is nearest to a branch, a double bond or a triple bond. (A double or triple bond takes precedence over a branch if they are equidistant from either end of the chain.)
- Rule 3: Number the carbon atoms from the end chosen.
- Rule 4: Name any branches first with the ending *-yl* (for example, methane becomes *methyl* and ethane becomes *ethyl*), then the longest chain, and then any single or double bond.
- Rule 5: When two or more branches occur on the same carbon atom, the number of the carbon atom is indicated for each branch, with the names given in alphabetical order (ignoring the prefixes described in rule 6). Place a comma between numbers and a dash between a number and a word.
- Rule 6: When two or more identical branches occur on the carbon chain, the prefixes *di-, tri-* and *tetra-* are used.

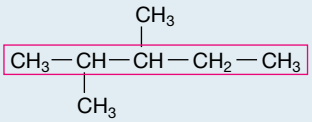
TIP: Avoid the following common errors when naming hydrocarbons:
- not identifying the longest chain possible
- not listing the side branches in alphabetical order
- omitting the prefixes *di-, tri-* and *tetra-* when they are required.

This system uses a set of standard prefixes by which the number of carbons is indicated. These are referred to as alkyl prefixes; for example, methyl, ethyl and propyl. Table 7.2 (section 7.2.3) shows these prefixes for up to ten carbons.

tlvd-0542

SAMPLE PROBLEM 3 Naming organic compounds with single bonds

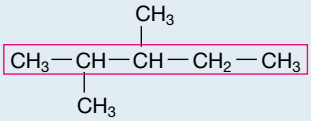
Use the rules for naming organic compounds to systematically name the following compound.

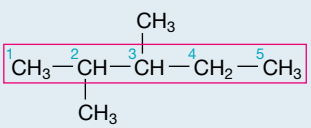


THINK

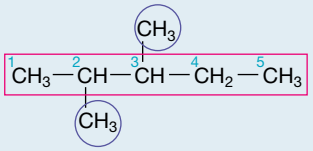
1. Circle or highlight the longest chain.
   The longest unbranched carbon chain in the compound shown contains five carbon atoms separated by single bonds.
2. The carbon atoms of the longest unbranched chain are numbered from left to right, because the branches are closest to the left.

WRITE


pentane



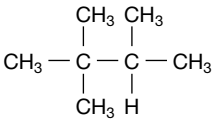
3. The branching chains are —CH_3, so they are methyl groups and attached to carbon atoms 2 and 3.
   The two branches are both —CH_3, so they are named dimethyl.



The name of the molecule is 2,3-dimethylpentane.

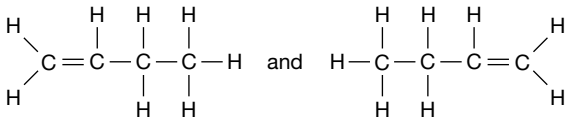
PRACTICE PROBLEM 3

Use the rules for naming organic compounds to systematically name the following compound.



When a double bond is present, it must be included as part of the longest chain and numbering commences at the end closest to the double bond. Figure 7.23 shows two representations of the molecule but-1-ene.

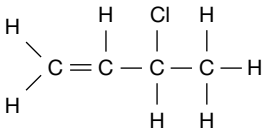
FIGURE 7.23 Both molecules are identical and are named but-1-ene.



When naming alkenes, the double bond must be included as part of the longest chain and given the lowest possible number.

For haloalkanes and haloalkenes, remember to include a number showing where the group is placed. For haloalkenes, the number of the double bond takes priority over the number of the halogen.

For example, 3-chlorobut-1-ene:



tlvd-3163

SAMPLE PROBLEM 4 Drawing organic compounds with double bonds

Draw the structure for 2-methylhex-1-ene.

THINK

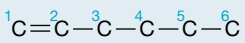
1. The prefix 'hex' indicates the molecule will have six carbons.
2. The 1-ene indicates the molecule is an alkene and a double bond exists between carbons 1 and 2.

WRITE

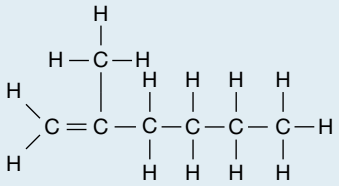
C—C—C—C—C—C

C=C—C—C—C—C

3. Number the carbon chain from the double bond end.



4. 2-methyl indicates a methyl group (—CH_3) on carbon number 2. Ensure sufficient hydrogen atoms have been added to all C atoms so that each carbon atom has four bonds.



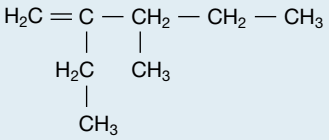
PRACTICE PROBLEM 4

Draw the structure for a molecule of 3-methylpent-1-ene.

tlvd-3164

SAMPLE PROBLEM 5 Naming organic compounds with double bonds and multiple branches

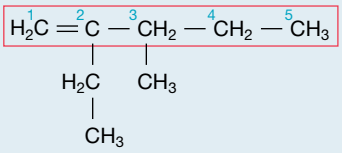
Name the following compound.



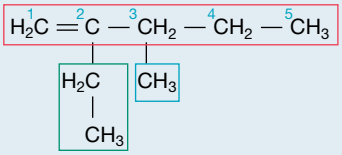
THINK

1. Circle the longest chain, which must include the double bond because it is an alkene. The longest chain in the compound shown has five carbons, so the *pent-* name is given. A double bond is also present so it is pentene. The double bond is nearest to the left end, so number accordingly (starting from the left). The double bond on the carbon atom number 1 is represented as *-1-ene*.

WRITE

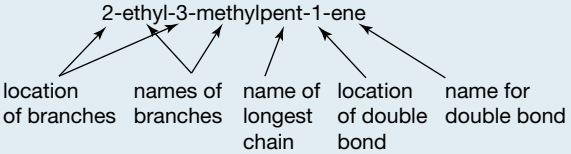


2. An ethyl group is attached to carbon 2 and a methyl group attached to carbon 3. Make sure all parts of the molecule are accounted for.



2-ethyl
3-methyl

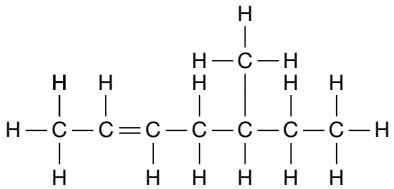
3. Remember to put ethyl and methyl groups in alphabetical order and a dash between the number and a word.



The name of the compound is 2-ethyl-3-methylpent-1-ene.

PRACTICE PROBLEM 5

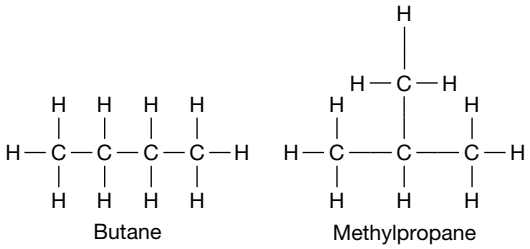
Name the following compound.



7.3.2 Isomers

Isomers are molecules that have the same molecular formula but have differences in the way the atoms are joined together or arranged. Isomerism is very common in organic chemistry and is one of the main reasons that carbon can form such a huge range of compounds. As the number of carbons in a molecule increases, so too does the number of ways that all the atoms can be joined together. In other words, the number of isomers increases. One type of isomerism is structural isomerism; naming these isomers is done using the rules in section 7.3.1.

FIGURE 7.24 The two isomers of C_4H_{10}



Structural isomers have the same molecular formula but different structural formulas. Consider figure 7.24, which shows C_4H_{10} drawn in two ways. Butane is the straight-chain structure, with carbons joined in an unambiguous path from the first carbon atom in the structure to the last. Methylpropane has a branched chain structure.

When considering isomers of hydrocarbons, thinking of carbon atoms attached to the hydrogen atoms as groups is useful. For example:

- A methyl group is one carbon attached to three hydrogens, —CH_3, and they occur at each end of a straight-chain hydrocarbon. Changing the position of a methyl group produces different isomers of the same molecule.
- Similarly, an ethyl group contains two carbon atoms: one attached to two hydrogen atoms and another attached to three hydrogens, —CH_2 —CH_3.

Methyl and ethyl groups are alkyl groups, and they have one fewer hydrogen than the alkane of the same name.

Structural isomers have different physical properties, as well as names, because the molecules have different shapes. In figure 7.24, the only intermolecular forces (between the molecules) each isomer has are dispersion forces. These would be about the same strength because the isomers are of similar size. However, molecules of butane can get closer together than can molecules of methyl propane (because butane is a straight-chain hydrocarbon), meaning that its dispersion forces are slightly more effective. This is reflected in the slightly higher boiling point for butane (−0.5 °C) compared to methylpropane (−11.7 °C).

isomers have the same molecular formula but a different arrangement of the atoms in space

structural isomers compounds that have the same molecular formula but a different semi-structural formula due to differences in the arrangement of the atoms

on Resources

Video eLesson Isomers (eles-2478)

tlvd-0541

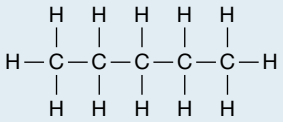
SAMPLE PROBLEM 6 Drawing isomers for alkanes

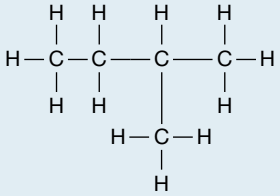
Draw three possible isomers for pentane, C_5H_{12}.

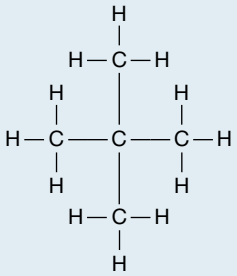
THINK

1. All structures must link five carbons to 12 hydrogens. Each hydrogen must only have one bond. Each carbon must have four bonds. Start with a straight chain of five C atoms, linked to the 12 H atoms.
2. Move one methyl group (one C and three H) to another C in the chain. The numbering of the C atoms in the chain does not matter. The —CH_3 group could be moved to the second C in this case, and the molecule would be the same.
3. Move another —CH_3 group to the same C atom in the chain to create the final isomer of pentane.

WRITE







PRACTICE PROBLEM 6

Draw three possible isomers for hexane, C_6H_{14}.

7.3 Activities

learnon

Students, these questions are even better in jacPLUS

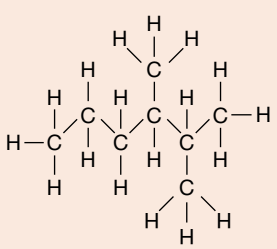
- Receive immediate feedback and access sample responses
- Access additional questions
- Track your results and progress

Find all this and MORE in jacPLUS

7.3 Quick quiz on | 7.3 Exercise | 7.3 Exam questions

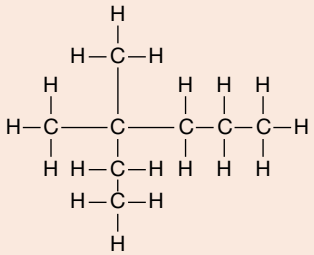
7.3 Exercise

1. MC What is the systematic name of the molecule shown?



A. 2-methyl-3-methylhexane
B. 2,3-methylhexane
C. 4,5-dimethylhexane
D. 2,3-dimethylhexane

2. **MC** What is the systematic name of the compound with the condensed formula $CH_3CH(CH_3)CH_2CH_2CH_3$?
A. Hexane
B. Methylpentane
C. 2-methylpentane
D. 4-methylhexane
3. The compound used in asthma inhalers has the chemical name 1,1,1,2 tetrafluoroethane. Draw the structure of this compound.
4. **MC** How many acyclic structural isomers are there for the molecular formula C_4H_8?
A. 1 **B.** 2 **C.** 3 **D.** 4
5. For each of the following compounds labelled (i), (ii) and (iii), complete the following.
a. How many isomers are there?
b. Draw the acyclic isomers from part **a**.
i. Pentane, C_5H_{12}
ii. Butene, C_4H_8
iii. Propyne, C_3H_4
6. Explain the following.
a. Why are numbers not required to correctly name methylbutane?
b. Why is the numbering in the name propane-3-ol incorrect?
7. What is the systematic name for the following compound?

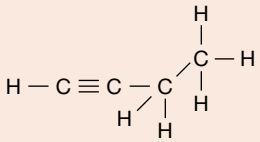


8. Draw the structure for a molecule of 4-ethyl-2,3-dimethylheptane.

7.3 Exam questions

Question 1 (1 mark)

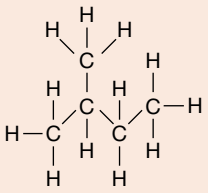
MC What is the systematic name of the molecule shown?



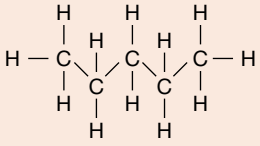
A. But-2-ene **B.** Butyne **C.** But-3-yne **D.** But-1-yne

Question 2 (1 mark)

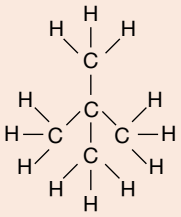
MC Consider the following compound.



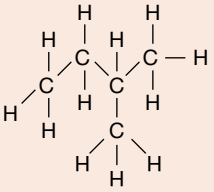
Which of the following are structural isomers of the above compound?



(I)



(II)



(III)

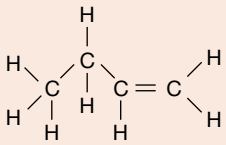
A. All three are structural isomers of the given compound.
B. Only I and II are structural isomers of the given compound.
C. Only II and III are structural isomers of the given compound.
D. Only I is a structural isomer of the given compound.

Question 3 (3 marks)

Name and draw the structural isomers with the molecular formula C_4H_8.

Question 4 (4 marks)

A student incorrectly named the following molecule but-3-ene.



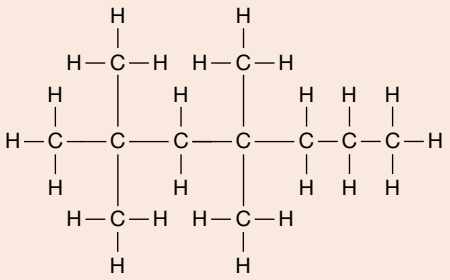
a. What rule was not followed in order for this mistake to occur? **(1 mark)**
b. State the correct name for this molecule. **(1 mark)**
c. Write the semi-structural formula for this compound. **(1 mark)**
d. Draw another isomer of this molecule. **(1 mark)**

Question 5 (13 marks)

a. Draw structural formulas for the following. **(6 marks)**
- **i.** Hex-2-ene
- **ii.** Methylpropene
- **iii.** Pent-2-yne

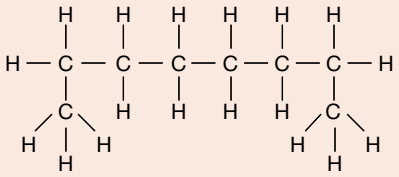
b. Name each of the following.

i.



(1 mark)

ii.



(1 mark)

c. Write semi-structural formulas for each substance from parts **a** and **b**. **(5 marks)**

More exam questions are available in your learnON title.

7.4 Functional groups — alcohols and carboxylic acids

KEY KNOWLEDGE

- The grouping of hydrocarbon compounds into families (alcohols, carboxylic acids) based upon similarities in their physical and chemical properties including general formulas and general uses based on their properties
- Representations of organic compounds (structural formulas, semi-structural formulas) and naming according to the International Union of Pure and Applied Chemistry (IUPAC) systematic nomenclature (limited to non-cyclic compounds up to C8, and structural isomers up to C5)

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7.4.1 Functional groups

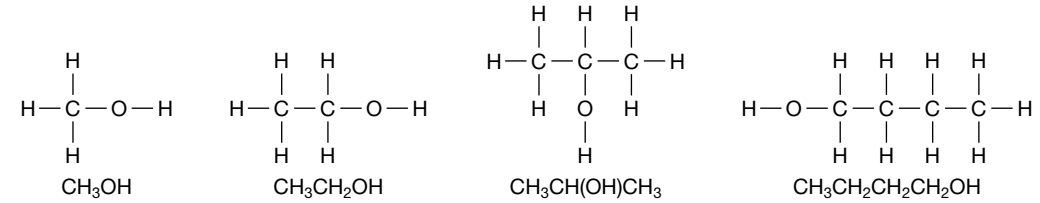
As introduced in section 7.2.6, functional groups are atoms or groups of atoms within a hydrocarbon that influence the properties of a molecule. Chemists use functional groups to simplify and assist with the classification of organic compounds. This allows them to reduce the large number of different compounds into more manageable groups that share common properties. For example, the functional group of the alkenes is C = C, while the functional group of the alcohols is —OH. The presence of the highly electronegative element oxygen, and the oxygen–hydrogen bond in alcohols and carboxylic acids gives them significantly different properties from those of hydrocarbons.

7.4.2 Alcohols

Alcohols are carbon chains containing one or more —OH groups. The —OH group is called a hydroxyl functional group. Names of alcohols have the ending *-ol*.

Alcohols are named on the basis of the alkane that they resemble. The *-e* on the end is dropped and replaced by *–ol*; for example, ethanol is the alcohol that resembles ethane. For instances where more than one location to attach the —OH group is possible, the number of the carbon to which it is attached must be specified. The general formula for an alcohol is $C_nH_{2n+1}OH$.

FIGURE 7.25 The first four alcohols shown as their structural formula and their semi-structural formula. Common uses are also listed.



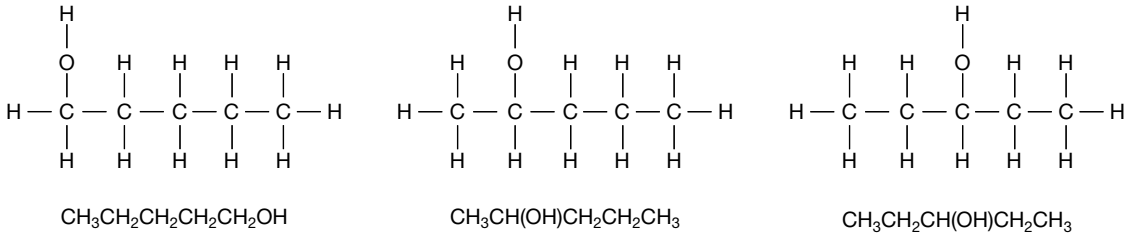
methanol: solvent for varnishes, windscreen washer fluid

ethanol: component of alcoholic beverages

propan-2-ol: rubbing alcohol, solvent

butan-1-ol: solvent, hydraulic fluid

FIGURE 7.26 The structural and semi-structural formulas of the isomers of pentanol, $C_5H_{12}O$



pentan-1-ol

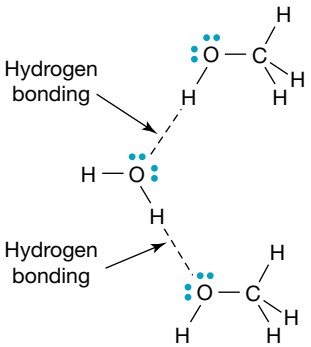
pentan-2-ol

pentan-3-ol

Properties of alcohols

- Alcohols have higher boiling points than the corresponding alkane or alkene, because in similar sized molecules, the hydrogen bonding between the alcohol molecules is stronger than dispersion forces alone that occur between hydrocarbon molecules in alkanes, alkenes and alkynes. More energy is needed to separate methanol molecules than methane molecules.
- The boiling points of alcohols increase with molecular size as the non-polar portion of the molecule increases, because the strength of the dispersion forces increases.
- Similarly, smaller alcohols are soluble in water; hydrogen bonds form between the water molecules and the polar hydroxyl functional groups in the alcohol molecules (figure 7.27). Solubility decreases with increasing size due to the increasing non-polar section of the alcohol molecules (figure 7.28).

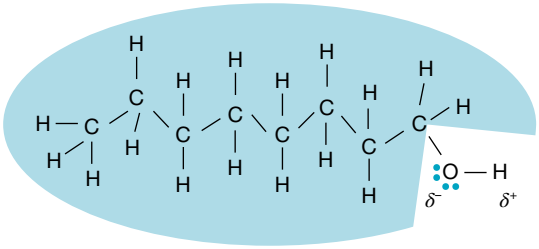
FIGURE 7.27 Smaller alcohols are soluble in water due to hydrogen bonds that form between the alcohol and water molecules.



Alcohols

- Alcohols contain the hydroxyl functional group, —OH.
- They have the general formula $C_nH_{2n+1}OH$.

FIGURE 7.28 Larger alcohols become less and less soluble in water due to the increasing influence of the non-polar hydrocarbon portion of the molecule (shaded). Octan-1-ol is not soluble in water.



Reactions and uses of alcohols

Alcohols undergo combustion reactions in air to produce carbon dioxide and water. Smaller alcohols such as methanol and ethanol are increasingly being used as fuels. Much research is currently occurring into developing efficient ways of producing ethanol from plant materials in a renewable manner (making so-called bioethanol), in order to decrease our dependence on fossil fuels. Many service stations now sell 'E10' — a blend of petrol and up to 10 per cent ethanol.

Using ethanol as an example, the combustion of alcohols may be summarised as follows.

$$\underset{\text{ethanol}}{CH_3CH_2OH(l)} + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$
$$\text{ethanol} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water}$$

Note that ethanol is shown using its semi-structural formula rather than its molecular formula. This is a common practice in organic chemistry.

Ethanol is also used as a solvent (it is widely used and sold as methylated spirits, which is ethanol with a small amount of methanol added) and is present in many products such as perfumes and cosmetics, where a fast rate of evaporation is desirable. It is the type of alcohol in alcoholic drinks. Other commonly used alcohols are methanol which, besides its use as a fuel, is used to make a number of other important chemicals. Also widely used is propan-2-ol (often referred to by its alternative name, isopropanol). This is used as a skin disinfectant and to dissolve oil. It is frequently used to sterilise the skin prior to injections and surgical procedures, and is also used as antiseptic, or 'rubbing alcohol'. Its rapid evaporation rate makes it an ideal choice in these situations.

FIGURE 7.29 Many service stations now sell a blend of petrol and ethanol.

Alcohols also react with carboxylic acids to produce esters. These are discussed in section 7.4.4 (as an extension) and in Units 3 and 4.

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EXPERIMENT 7.1 online only

Properties of alcohols

Aim

To examine properties of a range of alcohols

tlvd-0545

SAMPLE PROBLEM 7 Determining the properties of various alcohols

Consider the following alcohols: ethanol, butan-1-ol and decan-1-ol.

a. Which would have the highest boiling point?

b. Which would be the most soluble in water?

THINK	WRITE
a. All alcohols have hydrogen bonding and dispersion forces between their molecules. The larger alcohols will have stronger dispersion forces between their molecules and, therefore, the higher boiling points. Decan-1-ol is the largest molecule from those given (with a 10-carbon chain).	Decan-1-ol will have the highest boiling point.
b. As the hydrocarbon component of the alcohol increases, alcohols become less soluble in water due to the increasing non-polar nature that results. Hence, the molecule with the smallest hydrocarbon component will be the most soluble. This is ethanol (with a 2-carbon chain).	Ethanol is the most soluble in water.

PRACTICE PROBLEM 7

Consider the following alcohols: propan-1-ol, octan-1-ol and pentan-1-ol.

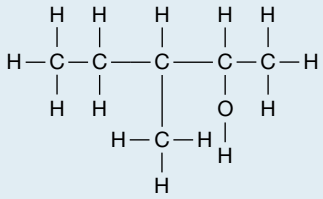
a. Which would have the highest boiling point?

b. Which would be the most soluble in water?

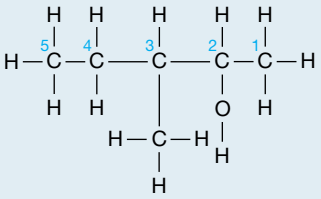
tlvd-0546

SAMPLE PROBLEM 8 Naming alcohols

a. Name the following molecule.



b. Write the semi-structural formula for this molecule.

THINK	WRITE
a. 1. Examine the molecule to find the longest carbon chain and number the chain to give the functional group (—OH) the smallest number. The molecule has five carbons in its longest chain, with the —OH attached to carbon number 2.	 This gives pentan-2-ol.

2. Identify any other hydrocarbon groups on the chain. A methyl group is also attached to carbon number 3, this is denoted by 3-methyl.
3. Write the name of the compound.

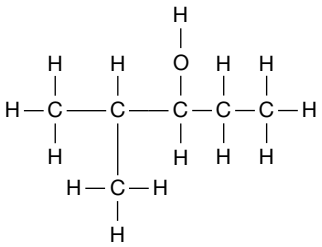
3-methylpentan-2-ol

b. Go carbon-by-carbon along the longest chain, abbreviating the structural formula by grouping the atoms together. Show branches in brackets after the carbon they are attached to.

$CH_3CH(OH)CH(CH_3)CH_2CH_3$

PRACTICE PROBLEM 8

a. Name the following molecule.



b. Write the semi-structural formula for this molecule.

7.4.3 Carboxylic acids

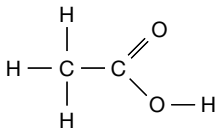
Carboxylic acids are another homologous series. These compounds contain the functional group —COOH at the end of a chain. The —COOH group is called the carboxyl functional group. When naming a carboxylic acid, add the ending *-oic acid* to the alkyl prefix. Remember to count the carbon atom in the —COOH group. For example, methane becomes methanoic acid. Figure 7.31 shows the structure and names of two further carboxylic acids.

Remember: when identifying and numbering the longest chain in a carboxylic acid, always *start* with the —COOH group. The C in this will *always* be number 1 in the sequence.

Carboxylic acids have the general formula $C_nH_{2n+1}COOH$.

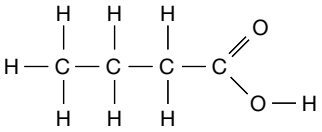
FIGURE 7.30 Painful bee and ant stings contain methanoic acid, HCOOH.

FIGURE 7.31 The structural and semi-structural formulas for two carboxylic acids



CH_3COOH

ethanoic acid (acetic acid)



$CH_3CH_2CH_2COOH$

butanoic acid

carboxylic acids homologous series containing the —COOH functional group

Carboxylic acids

- Carboxylic acids are a homologous series that contain the carboxyl functional group, —COOH.
- They have the general formula $C_nH_{2n+1}COOH$.

Properties of carboxylic acids

Carboxylic acids show trends similar to alcohols with respect to their boiling points and solubility. For example:

- They have relatively high boiling points due to the hydrogen bonding that is present; the hydrogen bonding forms due to both the —OH and the =O components of the —COOH group.
- Short chain carboxylic acids are soluble in water because the functional group can form multiple hydrogen bonds with water molecules. Their solubility decreases as the molecules become larger due to the increasing influence of the non-polar hydrocarbon component.

As their name suggests, carboxylic acids show typical acid properties. They are weak acids and some, such as ascorbic, citric and ethanoic acids, are present in foods.

Reactions and uses of carboxylic acids

- Familiar reactions of carboxylic acids include their reaction with bases to form a salt plus water.
- Carboxylic acids also react with reactive metals (such as magnesium) to produce a salt plus hydrogen gas.
- Carboxylic acids also react with carbonates, making a salt, carbon dioxide and water. Acid–base behaviour is examined in more detail in Unit 2.

Carboxylic acids also react with alcohols to make the important class of compounds called esters.

TABLE 7.5 Sources and uses of some common organic acids

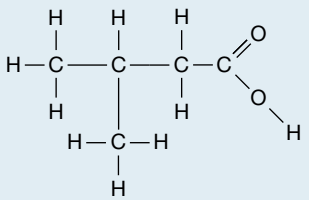
Name	Sources and uses
Methanoic acid (formic acid)	Responsible for sensation caused by nettle, bee and ant stings; used in medicine, food preservation and textile industry
Ethanoic acid (acetic acid)	Main component of vinegar; used as a solvent; salts of the acid used in insecticides and fungicides
Propanoic acid (propionic acid)	Used as an antifungal agent in the baking industry and in ointments (either in salt or acid form)
Butanoic acid (butyric acid)	Odour-causing component of rancid butter, soiled socks and body odour
2-hydroxybenzoic acid (salicylic acid)	Used as a food preservative
2-acetoxybenzoic acid (acetylsalicylic acid, aspirin)	Used for relief of fever, pain and rheumatic conditions

Note: common names are given in brackets

tlvd-0547

SAMPLE PROBLEM 9 Naming carboxylic acids

a. Name the following carboxylic acid.



b. Write the semi-structural formula for this compound.

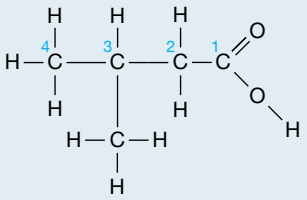
THINK

a. 1. Carboxylic acids are always numbered from the COOH end. Examine the molecule to find the longest carbon chain, numbering from the C in the COOH.
The longest chain is four carbons (counting the C in the COOH).

2. Identify any other hydrocarbon groups on the chain. A methyl is attached to carbon number 3 (this is denoted by 3-methyl).
Write the name of the compound.

b. Go carbon-by-carbon, abbreviating the structural formula by grouping the atoms together. Put side groups in brackets.
The standard abbreviation for the carboxyl group is —COOH.

WRITE

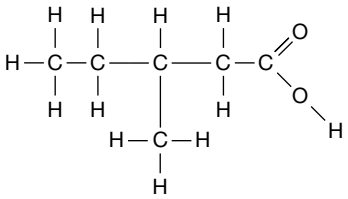


This is a butanoic acid.
3-methylbutanoic acid

$CH_3CH(CH_3)CH_2COOH$

PRACTICE PROBLEM 9

a. Name the following carboxylic acid.



b. Write the semi-structural formula for this compound.

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EXPERIMENT 7.2 online only

Properties of carboxylic acids

Aim

To examine a range of properties of carboxylic acids

7.4.4 EXTENSION: Esters

Although esters are beyond the scope of the Units 1 and 2 study design, knowledge of esters, including how they are produced, is an important component of Units 3 and 4. **Esters** are formed when an alcohol reacts with a carboxylic acid. Esters are a group of compounds that give the pleasant 'fruity' smell to various fruits, and synthetic esters are often used as an additive to create this pleasant smell. Such reactions are called condensation or esterification reactions,

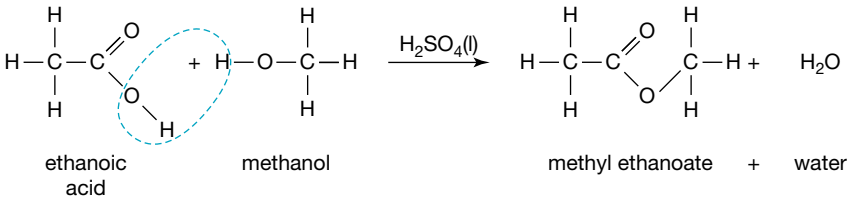
ester an organic compound formed from a condensation reaction between an alcohol and a carboxylic acid

where a small molecule (usually H_2O) is liberated. Concentrated sulfuric acid is used as a catalyst in the reaction that links the alcohol and carboxylic acid together by removing a water molecule.

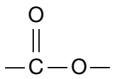
carboxylic acid + alcohol $\xrightarrow{H_2SO_4}$ ester + water

$CH_3COOH(l) + CH_3OH(l) \xrightarrow{H_2SO_4} CH_3COOCH_3(l) + H_2O(l)$

ethanoic acid + methanol $\xrightarrow{H_2SO_4}$ methyl ethanoate + water



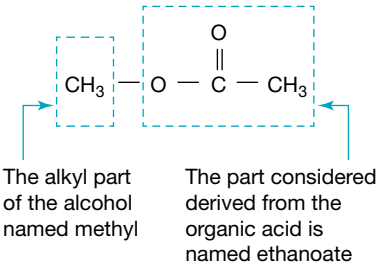
As can be seen from the preceding example, the functional group for an ester can be represented as



This is abbreviated to COO when writing semi-structural formulas.

When naming an ester, the alcohol (alkyl part) name is given first, followed by the acid part. For example, the alcohol methanol becomes *methyl*. The *-oic* ending of the acid is replaced with *-oate*. (For example, the carboxylic acid called ethanoic acid becomes *ethanoate*.) So, methanol reacting with ethanoic acid forms the ester methyl ethanoate.

FIGURE 7.32 Naming methyl ethanoate



Full name of ester: methyl ethanoate

Esters

- Esters are formed when an alcohol reacts with a carboxylic acid.
- They form an homologous series that contains the functional group, —COO.

Esters are an extension concept for the VCE Units 1 and 2 course; however, esters are required for understanding innovations in polymer manufacture using condensation reactions, and the breakdown of polymers using hydrolysis reactions. You will be required to study them in more detail in VCE Units 3 and 4.

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SAMPLE PROBLEM 10 Naming esters produced from alcohols and carboxlyic acids

Name the esters produced from the following alcohols and carboxylic acids and write their molecular formula.

a. Propanoic acid and ethanol
b. Methanoic acid and ethanol
c. Ethanoic acid and methanol
d. Which two esters from those listed are isomers?

THINK	WRITE
a. 1. When naming an ester, state the alcohol part first, and then the carboxylic acid part. Change ending to *-oate.* This means ethanol becomes ethyl and propanoic acid becomes propanoate.	The ester produced is ethyl propanoate.
2. Determine the condensed formula of ethyl propanoate by writing the condensed formulas of the reaction between propanoic acid and ethanol.	$CH_3CH_2COOH + CH_3CH_2OH \rightarrow CH_3CH_2COOCH_2CH_3 + H_2O$
3. Write the molecular formula of ethyl propanoate.	$C_5H_{10}O_2$
b. 1. When naming an ester, state the alcohol part first, and then the carboxylic acid part. Change ending to *-oate.* This means ethanol becomes ethyl and methanoic acid becomes methanoate.	The ester produced is ethyl methanoate.
2. Determine the condensed formula of ethyl methanoate by writing the condensed formulas of the reaction between methanoic acid and ethanol.	$HCOOH + CH_3CH_2OH \rightarrow HCOOCH_2CH_3 + H_2O$
3. Write the molecular formula of ethyl methanoate.	$C_3H_6O_2$
c. 1. When naming an ester, state the alcohol part first, then the carboxylic acid part. Change ending to *-oate.* This means methanol becomes methyl and ethanoic acid becomes ethanoate.	The ester produced is methyl ethanoate.
2. Determine the condensed formula of methyl ethanoate by writing the condensed formulas of the reaction between ethanoic acid and methanol.	$CH_3COOH + CH_3OH \rightarrow CH_3COOCH_3 + H_2O$
3. Write the molecular formula of methyl ethanoate.	$C_3H_6O_2$
d. Isomers have the same molecular formula.	Of those listed, (b) ethyl methanoate and (c) methyl ethanoate are isomers. Molecular formula is $C_3H_6O_2$.

PRACTICE PROBLEM 10

Name the esters produced from the following alcohols and carboxylic acids and write their molecular formula.

a. Propanoic acid and methanol
b. Methanoic acid and propan-1-ol
c. Ethanoic acid and ethanol
d. Which esters from those listed are isomers?

Properties of esters

Esters have relatively low boiling points because the intermolecular attraction is dipole–dipole (and dispersion forces) instead of the hydrogen bonding that occurs in carboxylic acids. Table 7.6 demonstrates this for a similarly sized ester and carboxylic acid.

TABLE 7.6 Ester and carboxylic acid boiling points

Name	Molecular formula	Condensed formula	Type	Boiling point (°C)
Ethyl ethanoate	$C_4H_8O_2$	$CH_3COOCH_2CH_3$	Ester	77.1
Butanoic acid	$C_4H_8O_2$	$CH_3CH_2CH_2COOH$	Carboxylic acid	164

Although esters cannot form hydrogen bonds with themselves, they can with water. This leads to the same solubility trend that is observed for alcohols and carboxylic acids. Smaller esters are moderately soluble but, as the number of carbons increases, they become less and less soluble (see table 7.7).

TABLE 7.7 The trend in ester solubility with increasing molecular size

Name	Formula	Solubility in water (g/100 g)
Ethyl methanoate	$HCOOCH_2CH_3$	10.5
Ethyl ethanoate	$CH_3COOCH_2CH_3$	8.8
Ethyl propanoate	$CH_3CH_2COOCH_2CH_3$	1.7

Reactions and uses of esters

The formation of an ester can be reversed by heating the ester with a dilute acid such as sulfuric acid or a dilute base such as sodium hydroxide. This is called **hydrolysis**. As expected, this regenerates the original alcohol and carboxylic acid.

Esters are important biologically. The important class of biomolecules called lipids (which includes oils and fats) are esters formed between an alcohol called glycerol and long chain carboxylic acids (often referred to as fatty acids). Table 7.8 outlines the uses of some esters.

hydrolysis a reaction in which a molecule is broken into smaller molecules, with water as a reactant

TABLE 7.8 Some esters and their uses

Name	Common name	Uses
Methyl salicylate		Liniment (treatment for sprains and bruises)
Ethyl ethanoate	Ethyl acetate	Nail polish remover, glue
Ethyl butanoate	Ethyl butyrate	Banana flavouring
Propyl ethanoate	Propyl acetate	Pear flavouring
Ethyl hexanoate		Pineapple flavouring
Acetylsalicylic acid	Aspirin	Pain relief, blood thinner

Esters are also becoming popular as a fuel, and are used to make biodiesel. In this production, naturally occurring plant and animal esters (oils and fats) are reacted with methanol where one type of ester is transformed into another to make a fuel that can be used in place of diesel. This is examined in Unit 3.

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EXPERIMENT 7.3 online only

Formation of esters

Aim

To prepare and name a number of different esters

7.4 Activities

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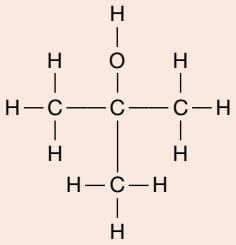
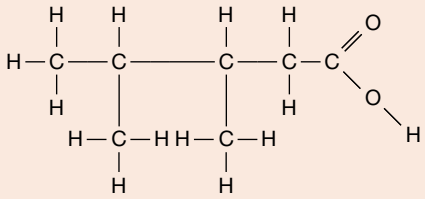
7.4 Quick quiz on | **7.4 Exercise** | **7.4 Exam questions**

7.4 Exercise

1. Write the semi-structural formula for a straight-chain alcohol that contains seven carbon atoms.
2. Write the semi-structural formula for a carboxylic acid that contains six carbon atoms.
3. Give a definition of a functional group. Give two examples.
4. **a.** Name the following molecule.



   b. Write the semi-structural formula for this molecule.
   c. Explain why a number is not required to indicate the CH_3 group in the name from part **a**, but is required to indicate the —OH group.
5. **a.** Name the following carboxylic acid.



   b. Write the semi-structural formula for this compound.
6. **a.** What is a homologous series?
   b. Explain why alcohols are an example of a homologous series. Support your answer with some appropriate formulas.
   c. Explain why carboxylic acids are an example of a homologous series. Support your answer with some appropriate formulas.
   d. Explain why alkanes are an example of a homologous series. Support your answer with some appropriate formulas.

7. Write molecular, structural and condensed formula for the following.
   a. Butan-1-ol b. Butan-2-ol
8. Write molecular, structural and semi-structural formulas for the following.
   a. Ethanoic acid
   b. Propanoic acid
   c. Butanoic acid
9. a. Draw the structural formula for octanoic acid and write its semi-structural formula.
   b. Draw the structural formula for ethyl pentanoate and write its semi-structural formula.
10. Pentanoic acid and ethyl propanoate both have the formula $C_5H_{10}O_2$. Their boiling points are 185 °C and 99 °C respectively.
    With reference to their structures and bonding, explain why this difference occurs.

7.4 Exam questions

Question 1 (1 mark)

MC A saturated carboxylic acid is one that has no double bonds other than the one between carbon and oxygen in the carboxyl functional group. If C_nH_{2n+2} is the general formula for alkanes, what would be the general formula for saturated carboxylic acids?

A. $C_nH_{2n+2}O_2$
B. $C_nH_{2n}O_2$
C. $C_nH_{2n-2}O_2$
D. $C_nH_{2n+2}O_n$

Question 2 (1 mark)

MC What is the systematic name of the compound with the condensed formula CH_3CH_2COOH?

A. Ethanoic acid
B. 1-ethanoic acid
C. Propanoic acid
D. 1-propanoic acid

Question 3 (2 marks)

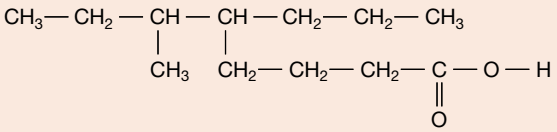
a. How does the carboxyl group affect the polarity of carboxylic acids? Explain. **(1 mark)**
b. How does the carboxyl group affect the water solubility of carboxylic acids? Explain. **(1 mark)**

Question 4 (2 marks)

Use a labelled diagram of a methanol molecule and a water molecule to explain why methanol is soluble in water.

Question 5 (5 marks)

An organic compound has the following formula.



a. What is the name of the homologous series to which this compound belongs? **(1 mark)**
b. Name this compound, using IUPAC nomeclature. **(2 marks)**
c. Draw the structure of one isomer of this compound. **(2 marks)**

More exam questions are available in your learnON title.

7.5 Sources and uses of organic chemicals

KEY KNOWLEDGE

- Plant-based biomass as an alternative renewable source of organic chemicals (for example, solvents, pharmaceuticals, adhesives, dyes and paints) traditionally derived from fossil fuels
- Materials and products used in everyday life that are made from organic compounds (for example, synthetic fabrics, foods, natural medicines, pesticides, cosmetics, organic solvents, car parts, artificial hearts), and an assessment of the benefits of those products for society, as well as the health and/or environmental hazards they pose

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7.5.1 Fossil fuels as a source of organic chemicals

Today, many of the chemicals we use are derived from fossil fuels — including not just fuels but also cosmetics, clothes, tools, plastics, bags, sports equipment, solvents, safety gear, paint, medicines, electronics, toiletries, artificial limbs, construction materials, home furnishings and many more. The list is extensive.

Fossil fuels are **non-renewable,** which means their supply is finite. More importantly, however, fossil fuel use is harmful to the environment for several reasons, including the following:

- The combustion of fossil fuels is the main contributor to greenhouse gas emissions and climate change.
- Many products produced using fossil fuels cannot be recycled. This creates waste, which has a long-term negative impact on the environment.

Manufacturing products using fossil fuels is an example of a linear economy.

One of the most common materials made from fossil fuels are plastics. Plastics are polymers, which have a large molecular structure built up from multiples of small molecules called monomers bonded together. A common polymer, polypropylene, is typically made using fossil fuels (crude oil). An example use of polypropylene is for face masks. These may take up to 450 years to biodegrade. Polymers will be discussed in detail in topic 8.

Given that the main element in these compounds is carbon, and because fossil fuels were formed from long-dead plants and animals, it makes sense to investigate the use of animal and plant materials that are currently available in **biomass**.

FIGURE 7.33 Examples of products made from fossil fuels

FIGURE 7.34 Facemasks are made from polypropylene sourced from fossil fuels.

non-renewable resources that are depleted or consumed faster than they can be replaced

biomass renewable organic matter that can be obtained from living or recently living plants and animals

7.5.2 Plant biomass

Plants have been used directly by humans to make simple products for centuries, including crops for foods and oils, natural remedies and building materials. Conventional biomaterials have also been used for energy and in packaging, clothing and building throughout history. Great potential now exists to be innovative in the production of a greater variety of **feedstock** chemicals from biomass to limit the use of fossil fuels. Some are concerned that using plant biomass in this way would use land that would otherwise grow crops for food, or that deforestation might occur. This argument can be off-set in cases where the non-edible parts of plants (for example, corncobs) or waste material (such as harvest waste or sawdust and wood scraps from the timber industry) is utilised, creating a **sustainable** process. The use of renewable feedstocks is one of the core principles of **green chemistry**, and when products are able to be recycled, they contribute to the growth of a circular economy. Plants commonly used as renewable feedstocks include sugar cane, corn/maize, algae and wastes from the harvest of other crops.

FIGURE 7.35 Maize used a renewable feedstock, with a biogas plant in the background

Chemicals from plant biomass

Plant cells are different from animal cells because they have a cell wall. This cell wall is composed of **cellulose**, which gives strength to the cell walls, and **hemicellulose**, which helps strengthen and connect the cellulose. Both are held together by **lignin**, which is a biopolymer and helps stiffen and protect the two sugar polymers (polysaccharides). Together these are known as **lignocellulose**, and this is a common biomass used as an industrial feedstock.

Each of these three components are natural polymers. Cellulose is a polymer of the monosaccharide, glucose; hemicellulose is a polymer of several monosaccharides (mostly xylose); and lignin is a polymer of phenolic units (see figure 7.36). These chemicals are mainly obtained from woody plants such as trees, shrubs, vegetables and grasses. The difficulty is getting them to dissolve — after all, it would be a problem if plants did dissolve!

Lignocellulose

Lignocellulose is a complex tissue consisting of cellulose, which gives strength to the cell walls, hemicellulose, which helps strengthen and connect the cellulose, and lignin, which binds these compounds together. Lignocellulose must be broken down before it can be used as a feedstock.

feedstock raw material used in industrial processes to produce a product

sustainable development that meets current needs without compromising the needs of future generations

green chemistry the design of new chemical products and processes that decrease or eliminate the use or creation of hazardous wastes and the impact on the environment

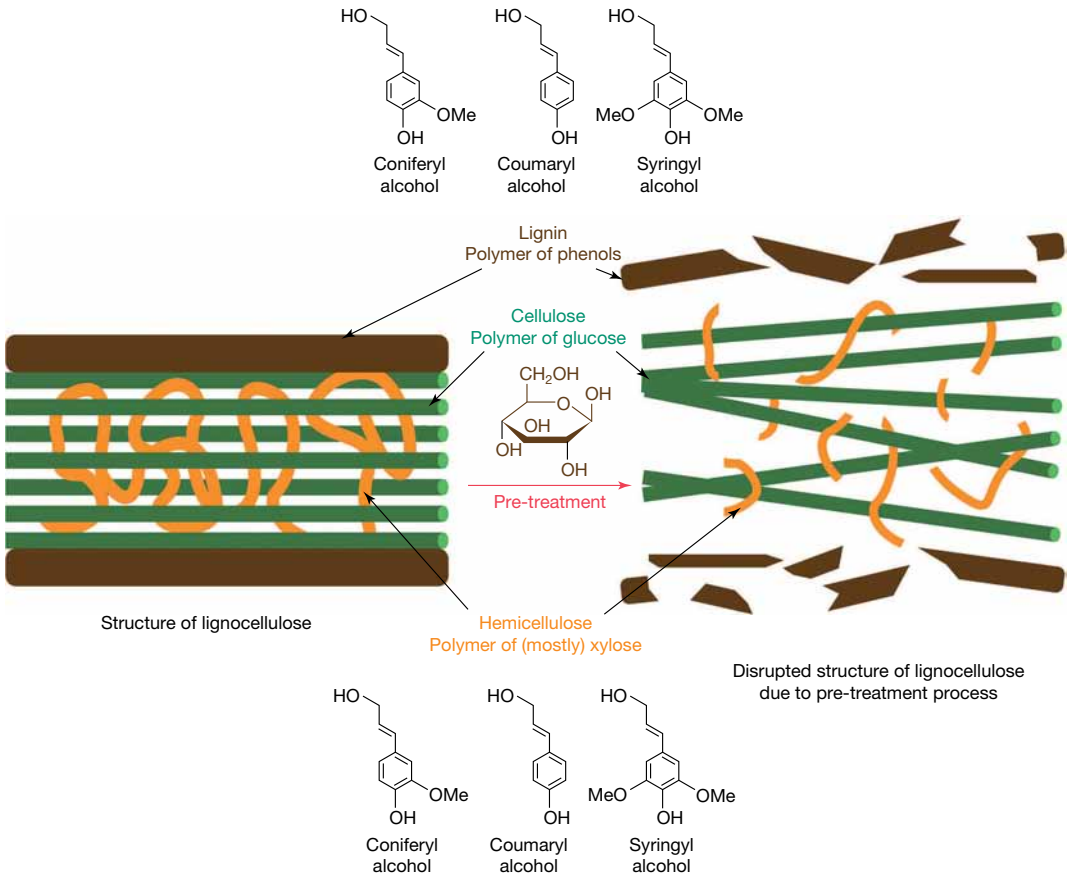
cellulose the most common carbohydrate and a condensation polymer of D-glucose; humans cannot hydrolyse cellulose, so it is not a source of energy

hemicellulose branched polysaccharide found in plant cell walls that have a simpler structure than cellulose

lignin complex organic polymer that provides rigidity to plants

lignocellulose a complex formed from lignin and cellulose

FIGURE 7.36 Lignocellulose must be broken down before it can be used as a feedstock.



Scientists are working to develop an economical and environmentally friendly process to separate these compounds. Examples include:

- The use of ionic liquids, which are salts that are liquid at close to room temperature, to break these large molecules into smaller molecules. Hemicellulose breaks down to lactic acid during fermentation and lactic acid is a useful feedstock to make other chemicals.
- The use of catalysts to reduce the number of steps involved; for example, a rhodium metal catalyst is used with a reactive chemical called phosphinite in another method.
- The use of lignin-degrading bacteria to degrade the mixture.

All these approaches aim to increase yield, reduce the number of steps, decrease the cost of production, minimise waste and find suitable catalysts, as outlined in the principles of green chemistry.

Comparison of fossil fuel chemicals and biomass chemicals

Fossil fuels such as petroleum are mainly straight-chain hydrocarbons with a high proportion of carbon, whereas biomass chemicals often contain ring compounds with less carbon and more oxygen present. As already described, biomass is mainly cellulose containing sugar polymers (carbohydrates). Cellulose and hemicellulose can be broken down by enzyme-catalysed hydrolysis into a range of chemicals that can then be used as feedstocks to produce the required chemicals. This reaction is studied in Unit 4 Chemistry.

Plant biomass is acknowledged as a fuel, but it can also be used in a multitude of products, including as acetone in nail polish remover and palmitic acid in hair products. Pharmaceuticals can be made from biomass through fermentation (decomposition of organic compounds into simpler ones through the action of microorganisms). More effective use of **biotechnology** will result in a wider range of products in the future.

biotechnology the use of biological processes, systems and resources in industrial production

Uses of chemicals from plants

- Lignin can be converted into renewable compounds for the flavouring industry, and into plastics, adhesives, surfactants and lubricants. Lignin and carbon dioxide with a ruthenium-cobalt metal catalyst have been used to produce ethanol, which is a useful solvent, fuel and feedstock for other chemical manufacture. The plastic properties of lignin make it a useful and sustainable product to line coffee cups and replace other plastic items. Polymers manufactured from biomass and are discussed in topic 8.
- Cellulose is a long unbranched polymer made up of glucose monomers. It can be made into paper, textiles, pharmaceuticals, cosmetics, emulsifiers, detergents and explosives.
- Hemicellulose, which forms about 25 per cent of wood biomass, is a smaller branched polymer than cellulose and is made up of several different monomers. It can be used to make furans, which can be used to synthesise lycra. Takeaway coffee cups can be made from hemicellulose as well as environmentally friendly detergents. It can also be used in paints as a stabiliser to replace toxic synthetic oil-based materials, in packaging and as a natural emulsifier in cosmetics. Ethanol, which is a useful feedstock for many other chemicals, can be produced from hemicellulose.

FIGURE 7.37 Plant-based products are continually being researched and developed as alternatives to petroleum-based products, and have a huge range of applications. Hemicellulose is a component of bamboo-fibre products like the coffee cup pictured.

7.5.3 Materials and products made from organic compounds

Organic compounds are in us and all around us. We use them every day in an extraordinary array of materials and products. Organic chemicals are used to manufacture pharmaceuticals, textiles, plastics, fuels, cleaning materials, cosmetics, food additives, natural medicines, vitamins, dyes and paints, fertilisers, pesticides, solvents, furniture and carpets. Some examples of the benefits and hazards of a limited number of products are described in table 7.9.

TABLE 7.9 Examples of materials made from organic compounds

Material	Chemicals present	Benefits	Hazards
Synthetic fabrics; e.g. nylon	Made from polyamides, which are made from petrochemicals	Strong, lightweight, durable, elastic	Can be uncomfortable to wear Most are non-biodegradable (last indefinitely)
		Waterproof	Production uses considerable energy and creates nitrous oxide, a greenhouse gas
		Resistant to sunlight	Uses fossil fuel resources
		Many uses, from activewear to carpets	Source of microfibres, which cause pollution
		Requires less water than natural fibres to produce	Recycling is energy-intensive
		Long-lasting so reusable	Hazardous to workers in production

Material	Chemicals present	Benefits	Hazards
Food additives (chemicals added to the natural organic substances in foods)	Food acids; e.g. citric, tartaric, fumaric, lactic and malic acid	Improve taste and preservatives ensure food lasts longer and is less likely to be wasted	Added food acids can lead to irritation or inflammation
	Organic bases	Improve appearance and regulate pH	Preservative sodium benzoate together with vitamin C can result in the formation of benzene, which is carcinogenic
	Antioxidants; e.g. vitamin C	Increase nutritive value and stop food deterioration due to reaction with oxygen, resulting in longer shelf life	
	Food colouring; e.g. cochineal — red colour obtained from a scaly insect	Improve appearance	Food colourings can cause asthma or hyperactivity in children, and some colourings are suspected of increasing risk of cancer; cochineal can cause allergic reactions
	Monosodium glutamate, MSG	Flavour enhancer	It has been suggested that some people may react to MSG with headaches and numbness
	Esters	Flavour enhancer	
	Sugars added for sweetness	More 'natural' additives can be used (e.g. stevia)	Fructose sweetener is associated with weight gain, diabetes and inflammation of cells; adding extra sugars (identified on labels as substances ending in -ose) can contribute to excess weight gain, tooth decay and increase the risk of heart disease
Car parts, both external and internal	Polypropylene for bumpers, steering wheels and instrument panels Polyurethane for seating Polyamides Polyvinyl chloride for seating	Improve safety through absorbing energy Durable Lightweight Reduce costs through being easy to manufacture Resistant to corrosion Easily moulded Inert Improve flexibility through added phthalates Alternative materials are gradually being substituted	Plastics in cars react with the atmosphere in the presence of heat and release volatile organic chemicals (VOCs) which can irritate eyes, nose and throat, also affect breathing and the central nervous system Phthalates used in PVC affects hormones and reproductive systems Very little is recycled and so waste adds to landfill
Cosmetics	Phthalates in fragrances, hair and nail products and skin lotions	Increase shelf life and help products adhere to the skin and hair	Associated with endometriosis; can affect hormones and reproductive systems; exposures are cumulative
	Parabens in makeup, moisturisers and hair products	Preservative	Found to affect hormones in mice; some types of parabens have been banned in other countries
	Formaldehyde in lotions, shampoos, shower gels	Preservative	Can irritate skin, eyes, nose and respiratory tract; carcinogenic in large amounts
	Triclosan in soap, antiperspirants, bodywash and toothpaste	Anti-bacterial agent and preservative	Animal tests have shown changes in endocrine system; possibility of causing antibiotic resistance and skin cancer

The examples outlined in table 7.9 do not include all the components of the various products currently available. Different brands will have different ingredients and often the potential hazard depends on the amount of the compound present. Many manufacturers have taken measures to eliminate potentially hazardous materials by substituting more natural materials. It is important to read labels on food and personal care products to identify what ingredients are included, and then investigate whether the chemicals are harmful.

FIGURE 7.38 Checking labels for ingredients

7.5 Activities

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7.5 Quick quiz on | **7.5 Exercise** | **7.5 Exam questions**

7.5 Exercise

1. MC Which of the following materials can biomass *not* be obtained from?
   - A. Wood chips
   - B. Wheat stalks
   - C. Agricultural waste
   - D. Recycled plastic
2. MC What does cellulose consists of?
   - A. Very long branched polymer molecules
   - B. Short branched polymer molecules
   - C. Very long unbranched polymer molecules
   - D. Short unbranched polymer molecules
3. MC Identify one of the properties of lignocellulose.
   - A. A compound made of sugars
   - B. Insoluble in water
   - C. A straight-chain hydrocarbon
   - D. Obtained from biowaste from animals
4. Catalysts used in processes involving plant biomass include ruthenium, cobalt and rhodium. State the symbol, atomic number, group and period of these metals.

Metal	Symbol	Atomic number	Group	Period
Ruthenium				
Cobalt				
Rhodium				

5. Describe an environmental advantage of using each of the following, rather than using a traditional organic compound.
   a. Lignin
   b. Hemicellulose
6. a. What is nylon made from?
   b. Identify three benefits in the use of nylon.
   c. Identify three disadvantages in the use of nylon.

7.5 Exam questions

Question 1 (1 mark)

Identify a possible disadvantage of using plant biomass to produce materials.

Question 2 (1 mark)

What feature of plant cells allow them to be used as biomass, where animal cells cannot be used?

Question 3 (6 marks)

a. Describe what lignocellulose consists of. **(2 marks)**
b. Where is lignocellulose obtained? **(2 marks)**
c. Identify two challenges when obtaining the components of lignocellulose. **(2 marks)**

Question 4 (7 marks)

a. State three similarities between the structures of cellulose and hemicellulose. **(3 marks)**
b. State two differences between the structures of cellulose and hemicellulose. **(2 marks)**
c. State two uses of each compound. **(2 marks)**

Question 5 (5 marks)

State the name of each of the chemicals described in the following.

a. Found in both plastics in cars and cosmetics; detrimental to hormones **(1 mark)**
b. Textile made from polyamide **(1 mark)**
c. A red food colouring obtained from an insect that can cause allergic reactions **(1 mark)**
d. Causes the 'new smell' in cars and can irritate eyes, nose and throat **(1 mark)**
e. A flavour enhancer in food that can cause headaches and numbness in some people **(1 mark)**

More exam questions are available in your learnON title.

7.6 Review

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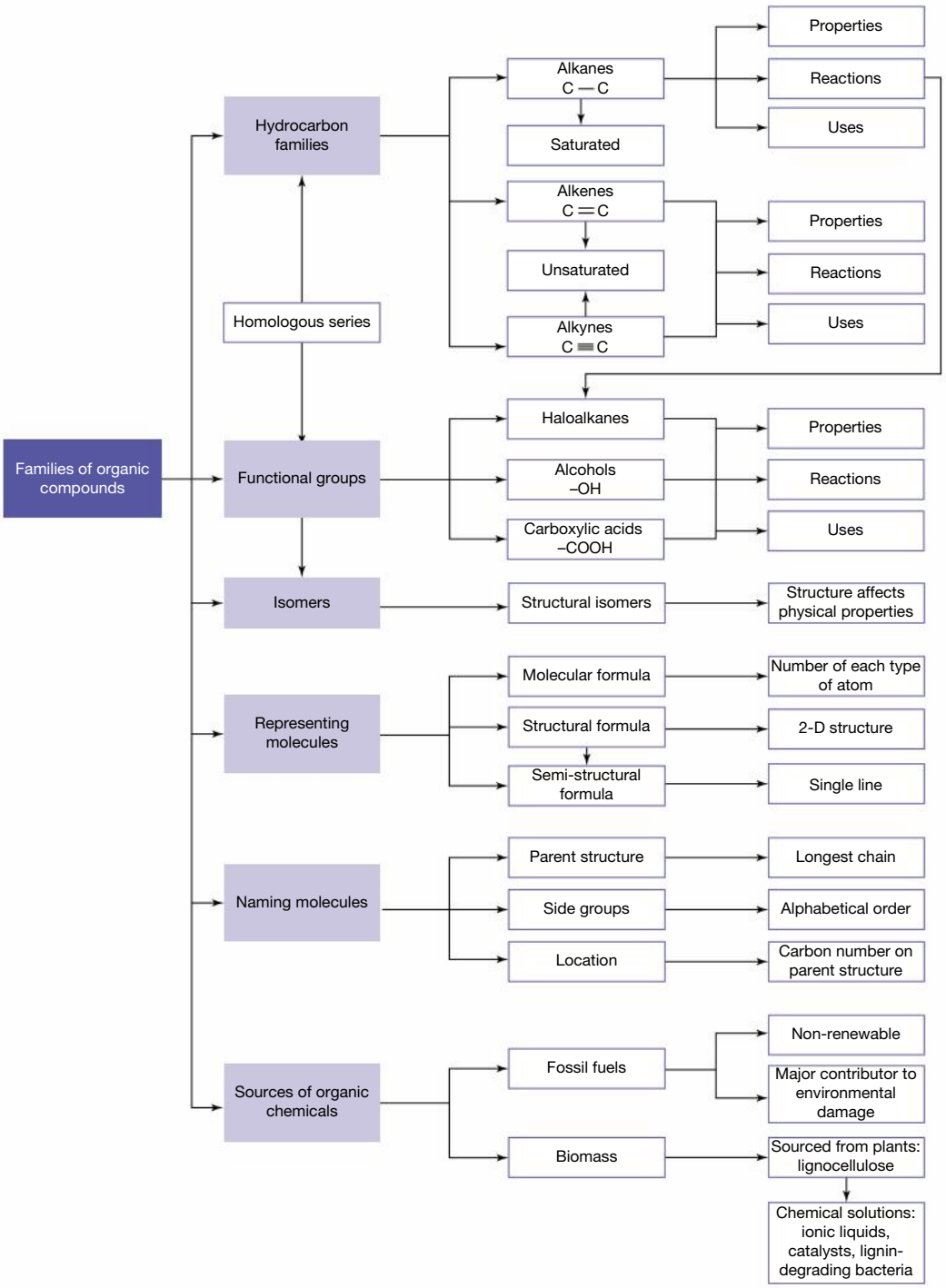
7.6.1 Topic summary



7.6.2 Key ideas summary

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7.6.3 Key terms glossary

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on Resources

Solutions	Solutions — Topic 7 (sol-0806)
Practical investigation eLogbook	Practical investigation eLogbook — Topic 7 (elog-1612)
Digital documents	Key science skills (doc-37066) Key terms glossary— Topic 7 (doc-37079) Key ideas summary — Topic 7 (doc-37080)
Exam question booklet	Exam question booklet — Topic 7 (eqb-0088)

7.6 Activities

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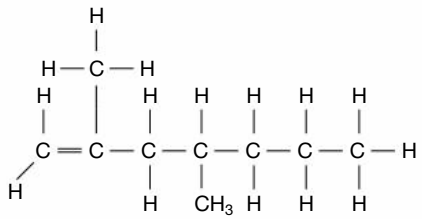
7.6 Review questions

1. Explain why the compounds methene and methyne do not exist.
2. **a.** Distinguish between saturated and unsaturated hydrocarbons.
   b. Which is more reactive?
3. Write a formula for each of the following.
   a. An alkane with 22 carbon atoms
   b. An alkene with 17 carbon atoms
   c. An alkyne with 13 carbon atoms
4. **a.** Complete the following equations, balancing where necessary.
   i. $CH_3CH_3 + Br_2 \rightarrow CH_3CH_2Br +$ _____
   ii. _____ $+ Cl_2 \rightarrow CH_2ClCH_2Cl$
   iii. $CH_3Cl + Cl_2 \rightarrow$ _____ $+ HCl$
   iv. $C_3H_8 + O_2 \rightarrow CO_2 +$ _____
   b. Name each of the types of reaction in part **a**.

5. Write semi-structural (condensed) formulas for each of the following.

a.



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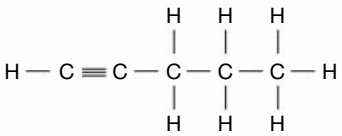

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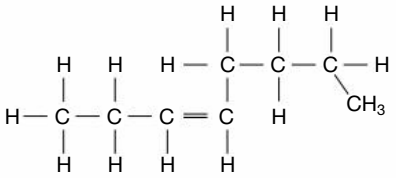

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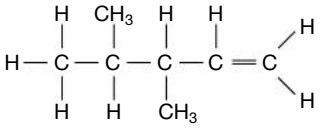

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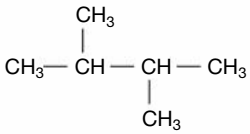
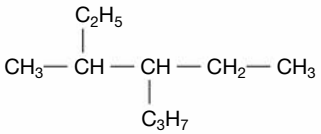


f.



6. Name the following organic compounds.
(*Note:* The longest continuous chain may not be written in a straight line.)

a. $CH_3-CH_2-CH(CH_3)-CH_3$

b. $CH_3-CH(CH_2CH_3)-CH_2-CH_3$

c. $CH_2(CH_3)-CH_2-CH(CH_2CH_3)-CH_2-CH_3$

d. $CH_3-CH_2-C(CH_3)(CH_2CH_2CH_3)-CH_3$

e. $CH_3-CH_2-CH(CH_3)-CH_2-CH(CH_3)-CH_3$

f. $CH_3-C(CH_3)(CH_3)-CH_2-C(CH_3)(CH_3)-CH_3$

7. Draw structural formulas for the following.
   - **a.** 2-methylheptane
   - **b.** 2,3-dimethylbut-2-ene
   - **c.** Pent-2-yne
   - **d.** 2,6-dimethylhept-3-ene
   - **e.** Oct-3-ene
   - **f.** 4-ethy1-3-methylhex-2-ene

8. Which of the following compounds are isomers?
   - **i.** pentane
   - **ii.** 2-methylbutane
   - **iii.** 2,3-dimethylbutane
   - **iv.** 2,2-dimethylpropane

9. Define the term *isomer*. Illustrate your answer with the isomers of pentane and butene.

10. Butane is a hydrocarbon used as fuel for cigarette lighters and portable stoves.
    - **a.** What is a hydrocarbon?
    - **b.** What family of organic compounds does butane belong to?
    - **c.** State the general formula for this family of hydrocarbons.
    - **d.** State the formula of the next compound in this family.
    - **e.** Butane has two isomers. Draw the two isomers and name them.

7.6 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC Which of the following is *not* the formula of an alkane?

A. C_4H_{10} **B.** C_7H_{16} **C.** $C_{11}H_{22}$ **D.** $C_{14}H_{30}$

Question 2

MC Which one of the following hydrocarbons is unsaturated?

A. C_2H_6 **B.** C_7H_{16} **C.** C_5H_{12} **D.** C_4H_8

Question 3

MC Which one of the following is a hydrocarbon that readily undergoes substitution reactions?

A. Ethane **B.** Ethene **C.** Methylpropene **D.** Chloromethane

Question 4

MC Which one or more of the following statement is *not* correct?

A. A functional group remains unchanged in most chemical reactions.
B. A functional group is common to all members of the particular homologous series.
C. A functional group may be regarded as being responsible for a characteristic set of chemical reactions for a given homologous series.
D. Successive members of homologous series differ by $-CH_2$.

Question 5

MC What is an alkyl group?

A. A strongly basic group in an organic molecule
B. A group of organic compounds related to the alkanes
C. A group equivalent to an alkane molecule less one hydrogen atom
D. A group of organic compounds with the general formula C_nH_{2n+2}

Question 6

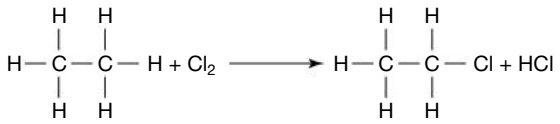
MC Which of the following statements best describes a saturated compound?

A. A saturated compound, when used as a solvent, does not dissolve any more solute at a particular temperature.
B. A saturated compound contains carbon and hydrogen atoms only.
C. A saturated compound has only single bonds between carbon atoms.
D. A saturated compound does not react with chlorine.

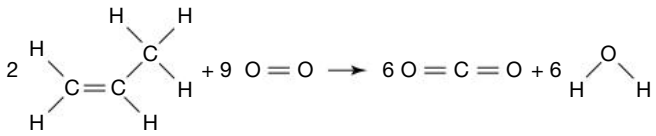
Question 7

MC Which of the following equations represents an addition reaction?

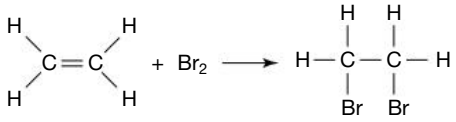
A.



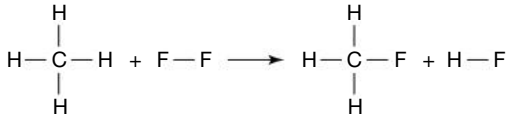
B.



C.



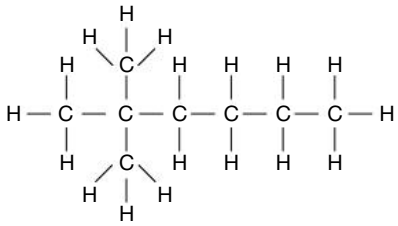
D.



Question 8

MC Which of the following structures is *not* an isomer of C_8H_{18}?

A.



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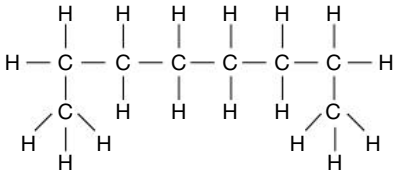

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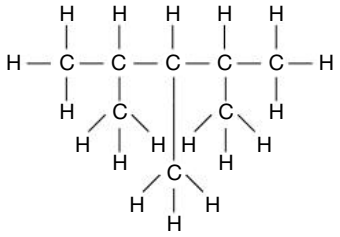

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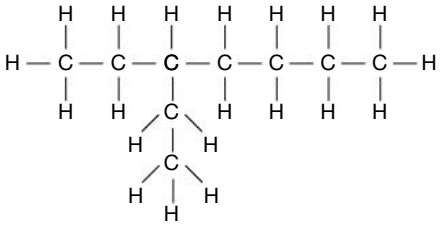
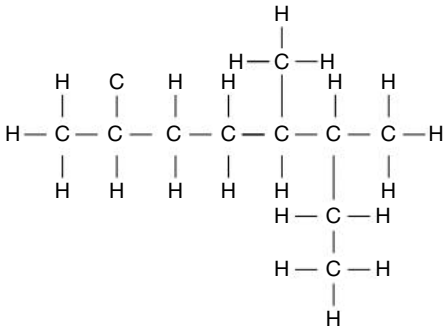


Question 9

MC What is the correct systematic name for the compound shown below?



A. 4-methyl-5-ethylhexane
B. 2-ethyl-3-methylhexane
C. 4,5-dimethylheptane
D. 3,4-dimethylheptane

Question 10

MC Which organic compound in the following group has the highest boiling point?

A. Acetylene (ethyne)
B. Ethane
C. 1,1,1-trichloroethane
D. Ethanol

Section B — Short answer questions

Question 11 (7 marks)

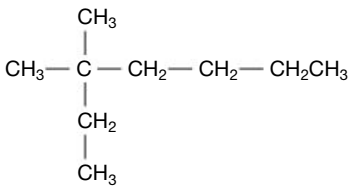
a. In terms of bonding, explain why heavier hydrocarbons present in diesel oil (those with a high number of carbon atoms) have higher boiling temperatures than lighter hydrocarbons such as propane (those with fewer carbon atoms). **(2 marks)**

b. In terms of bonding, explain why lighter hydrocarbons such as petrol are more flammable than heavier ones such as paraffin wax. **(2 marks)**

c. List the following compounds in order of increasing boiling points and give reasons for your answer: chloroethane, CH_3CH_2Cl, ethane, CH_3CH_3, and methane, CH_4. **(3 marks)**

Question 12 (2 marks)

Jock isolated the following compound and named it 2-ethyl-2-methylhexane. Lillian said that the compound was incorrectly named. Explain why Lillian is right and give the correct name for the compound.



Question 13 (8 marks)

You find a container labelled as dimethylpentane. Draw the possible structures for this compound and give each its correct name.

Question 14 (10 marks)

Ethene, C_2H_4, is widely used in the chemical industry in the production of plastics, antifreeze and solvents. It is unsaturated and a member of the alkene homologous series.

a. Explain the meaning of the terms unsaturated and homologous series. **(2 marks)**

b. Ethene reacts with hydrogen to produce a single product, *A*. Compound *A* reacts with chlorine, Cl_2, to produce gaseous compound *B* and hydrogen chloride gas. Identify compounds *A* and *B* and write balanced equations, with states, for each of these reactions. **(4 marks)**

c. Ethene also burns in excess oxygen to form two products. Write a balanced equation, with states, for this reaction. **(2 marks)**

d. Explain how biomass could be used as a renewable feedstock for the production of ethene. **(2 marks)**

Question 15 (9 marks)

Analysis of a certain compound reveals that it contains only carbon, hydrogen and oxygen.

a. A small amount of crushed calcium carbonate is added to a sample of the compound. A gas is observed to form. Name the homologous series to which this compound belongs. **(1 mark)**

b. If you had not performed the test described in part a, name another homologous series that this compound might belong to. **(2 marks)**

c. If the compound has four carbon atoms, name the compound. **(1 mark)**

d. State its molecular formula. **(2 marks)**

e. Calculate the percentage composition of carbon, hydrogen and oxygen present. **(3 marks)**

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8 Polymers and society

KEY KNOWLEDGE

In this topic you will investigate:

Polymers and society

- the differences between addition and condensation reactions as processes for producing natural and manufactured polymers from monomers
- the formation of addition polymers by the polymerisation of alkene monomers
- the distinction between linear (thermoplastic) and cross-linked (thermosetting) addition polymers with reference to structure and properties
- the features of linear addition polymers designed for a particular purpose, including the selection of a suitable monomer (structure and properties), chain length and degree of branching
- the categorisation of different plastics as fossil fuel–based (HDPE, PVC, LDPE, PP, PS) and as bioplastics (PLA, Bio-PE, Bio-PP); plastic recycling (mechanical, chemical, organic), compostability, circularity and renewability of raw ingredients
- innovations in polymer manufacture by condensation reactions and break down by hydrolysis reactions, contributing to the transition from a linear economy towards a circular economy.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of VCE Chemistry. Experiments and investigations, supported by a **practical investigation eLogbook** and **teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

EXAM PREPARATION

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8.1 Overview

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8.1.1 Introduction

Originally, toothbrushes were made from the neck hairs of wild hogs, and these hairs were attached to either a bone or bamboo handle. But with the advent of modern plastics, getting a new toothbrush usually requires only a short trip to the local supermarket. Toothbrushes today are typically constructed with nylon bristles fixed to a polypropene handle. Interestingly, toothbrushes were one of the first commercial uses of nylon. Each of these materials has specific properties that make it suitable for this purpose. Unlike hog's hair, which is a natural polymer, both nylon and polypropene are synthetic polymers. What are polymers, and how are they customised and produced?

FIGURE 8.1 Bamboo-handled toothbrushes are becoming popular as an alternative to traditional plastic toothbrushes.

This topic introduces the polymerisation processes by which complex organic molecules, called polymers, may be formed from simpler molecules, monomers. The different methods of polymer production are described, along with how polymer properties can be altered by different processes and additives. Bamboo toothbrushes are becoming more popular because they are biodegradable. Environmentally aware consumers must take care when purchasing these toothbrushes because often the bristles are made from nylon because it is more durable than natural materials. The advantages and disadvantages of the uses of polymers are discussed in this topic, as well as which plastics can be recycled to form other useful materials. Innovations in polymer manufacture are presented, with a discussion of how we can incorporate our use of plastics into the circular economy.

Prior knowledge of hydrocarbons is built on, especially alkenes, and the effects of molecular bonding — both the intramolecular bonding that exists within molecules and the types and properties of intermolecular bonding that exists between molecules.

LEARNING SEQUENCE

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8.2 Polymers 311
8.3 Linear and cross-linked polymers 320
8.4 Polymer selection 326
8.5 Plastic recycling and innovations in design 337
8.6 Review 348

on Resources

Solutions	Solutions — Topic 8 (sol-0807)
Practical investigation eLogbook	Practical investigation logbook — Topic 8 (elog-1613)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 8 (doc-37081) Key ideas summary — Topic 8 (doc-37082)
Exam question booklet	Exam question booklet — Topic 8 (eqb-0089)

8.2 Polymers

KEY KNOWLEDGE

- The differences between addition and condensation reactions as processes for producing natural and manufactured polymers from monomers
- The formation of addition polymers by the polymerisation of alkene monomers

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8.2.1 Types of polymers

The name **polymer** comes from the Greek *polymers* ('of many parts'). Now that we have grown so used to plastics, imagining what life was like without them is difficult. Many of the items we use every day, from objects as insignificant as pens and zippers to larger items such as furniture, refrigerators and the interiors of cars, are made from plastics. Even our clothes can be made from synthetic fibres. Many modern products are produced from a selection of plastics, each one chosen for its specific properties.

- A polymer molecule is made up of thousands of units strung together into very long chains. The simple molecules that are strung together are called **monomers**.
- Monomers link together to form a polymer chain in a process called **polymerisation**. The monomers used to make the polymer can be the same or different.
- **Copolymers** are formed when two or more different monomers are used.
- Most polymerisation reactions require a catalyst.

Many polymers are named by placing the prefix 'poly' in front of the monomer. In several (but not all) cases, the older or trivial name is used for the monomer instead of the IUPAC name. Polyethylene is the common name for polyethene, for example, and sometimes just the initials PE are used. Polymers can be categorised in different ways — for example, whether they are natural or synthetic, or the type of reaction process used to form them.

Features of **natural polymers**:

- Most natural polymers are made of proteins or cellulose.
- Plants are made of cellulose, and many contain another polymer called starch. Cellulose and starch consist of sugar monomers and are formed by condensation polymerisation. Wool, silk, cotton, linen and rubber are all naturally occurring polymers.
- Animals are made of proteins that are also condensation polymers. Skin, organs, muscles, enzymes, hair and fingernails are all made of protein.

Features of **synthetic polymers**:

- Synthetic polymers are made from raw materials obtained from fossil fuels or, more recently, from **biomass**.
- Many synthetic polymers are called **plastics** in everyday life. The term 'plastic' actually refers to something pliable and/or able to be moulded, either in its final state or at some stage in its production.

Polymers are made by either of two main methods: **addition polymerisation** or **condensation polymerisation**.

- Addition polymerisation: the polymer is produced by a reaction between monomers that have a carbon–carbon double bond. For example, molecules of ethene can react using a suitable catalyst to form the polymer polyethylene (common name) or polyethene (IUPAC name).

polymer molecule made up of thousands of units (monomers) strung together into very long chains

monomer molecule that links together to form a polymer

polymerisation the formation of giant molecules by repeated monomers that are joined by covalent bonds

copolymer polymer formed from the polymerisation of two monomers

natural polymers naturally occurring polymers in plants and animals

synthetic polymers polymers made from fossil fuels or biomass

biomass renewable organic matter that can be obtained from living or recently living plants and animals

plastic polymer that can be moulded when hot and retains its shape when cooled

addition polymerisation the process in which monomers with at least one double carbon–carbon bond react together to form a polymer by addition reactions

condensation polymerisation polymerisation in which two monomers combine and a smaller molecule is eliminated

- Condensation polymerisation: a type of reaction in which two molecules combine to form a larger molecule, producing a small molecule such as HCl or H_2O as a by-product. If the molecule lost is water, it can also be described as a dehydration reaction.

8.2.2 Addition polymers

For addition polymerisation to occur, the monomers must contain a *carbon-to-carbon* double bond. That is, the monomer must be an alkene. In the process of joining the monomers together, one of these bonds is broken and the electrons that are freed up are used to form a new bond to the next monomer. This links the monomers together and forms the long chains. Alkenes, and especially the smaller members such as ethene and propene, are thus ideal candidates to act as monomers.

Synthetic addition polymer: polyethene

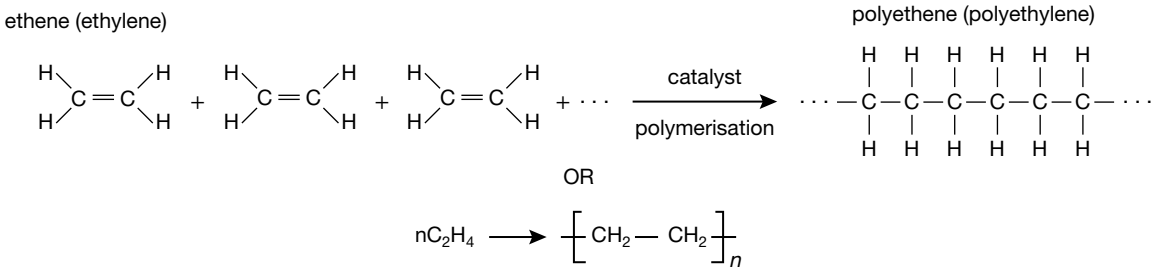
The simplest monomer that can undergo addition polymerisation is ethene, which results in the polymer **polyethene** (see figures 8.2 and 8.3). Ethene has been produced from fossil fuels, but recent research has resulted in the development of processes to produce organic chemicals such as ethene from biomass (as explained in subtopic 7.5.2).

When ethene is subjected to high pressure, it changes from a gas to a liquid. If the liquid ethene, still under pressure, is heated in the presence of a catalyst (a small quantity of oxygen), an **addition reaction** takes place in which the ethene molecules join together and form a long chain (polyethene or polyethylene). The length of these chains can vary from 4000 to 20 000 carbon atoms, depending on the temperature and pressure used.

The second representation of the reaction provided in figure 8.3 is called the condensed or bracketed form. It can be seen that the empirical formula of the polymer is the same as the monomer.

FIGURE 8.2 Milk bottles are made of high-density polyethylene (HDPE).

FIGURE 8.3 The formation of the polymer polyethene. The double bond between the two carbon atoms in the monomer breaks, allowing long chains to form.



Polymers are very large, chain-like molecules. Due to their large size, they will have very high molar masses. These values can commonly be in the region of 500 000 g mol^{-1}.

Many other addition polymers exist — all of which are formed from monomers containing a carbon–carbon double bond. These are dealt with later in this topic.

polyethene polymer consisting of monomers of ethene

addition reaction reaction in which one molecule bonds covalently with another molecule without losing any other atoms; only one product is formed

Drawing polymers

When drawing sections of a polymer it is important to remember to include open bonds at each end to show that the molecule continues.

tlvd-0553

SAMPLE PROBLEM 1 Calculating the molar mass of a polyethene sample

Calculate the approximate molar mass of a polyethene sample, the average chain length of which is 35 000 carbon atoms.

THINK	WRITE
1. Recall that in addition polymerisation, all the atoms in the original monomers are used in the production of the resulting polymer (nothing is lost).	Carbon atoms = 35 000
2. Recall that the molecular formula of ethene is C_2H_4. Hence, two carbon atoms exist per monomer of ethene.	Number of ethene monomers $= \frac{35\,000}{2}$ $= 17\,500$
3. Determine the approximate molar mass of the polyethene sample by multiplying the number of ethene monomers by the molar mass of ethene. Record your answer to 3 significant figures.	Molar mass ethene $= 2 \times C + 4 \times H$ $= 2 \times 12.0 + 4 \times 1.0$ $= 28.0\ g\ mol^{-1}$ Approximate molar mass $= 17\,500 \times 28.0$ $= 4.90 \times 10^5\ g\ mol^{-1}$

PRACTICE PROBLEM 1

Calculate the approximate molar mass of a polyethene sample, the average chain length of which is 5000 carbon atoms.

on Resources

Video eLesson Addition polymerisation (eles-2481)

Natural addition polymer

Natural **rubber** is produced from latex, a milky white addition polymer that can be harvested from rubber trees by making a cut in the bark and collecting the sap as it runs out. A well-managed plant can yield approximately 1.8 kg of dry crude rubber per year.

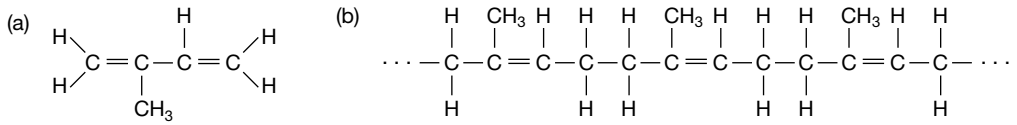
Rubber is a completely **amorphous polymer**. The monomer in natural rubber is **isoprene** or 2-methylbuta-1,3-diene. Isoprene polymerises to form long chains, as shown in figure 8.4, and the molecular formula is written as $(C_5H_8)_n$.

rubber an elastic polymer formed from the latex of tropical rubber trees and plants

amorphous polymer polymers that do not have a crystalline structure; their atoms are held together loosely and their arrangement is not ordered or predictable

isoprene volatile liquid hydrocarbon obtained from petroleum, the molecule of which forms the basic structural unit of natural and synthetic rubbers

FIGURE 8.4 Structural formulas of (a) the monomer isoprene and (b) the polymer for raw (natural) rubber.



This chain is similar to that of polyethene (polyethylene) but with an important difference: rubber still contains double bonds that can be attacked by oxygen and, unlike polyethene, rubber can perish. Another disadvantage of natural rubber is that it is not elastic. When stretched, the long, tangled chains straighten out and remain this way, with little tendency for them to return to their original shape. Natural rubber is also susceptible to temperature changes, becoming very brittle when cold and sticky when hot.

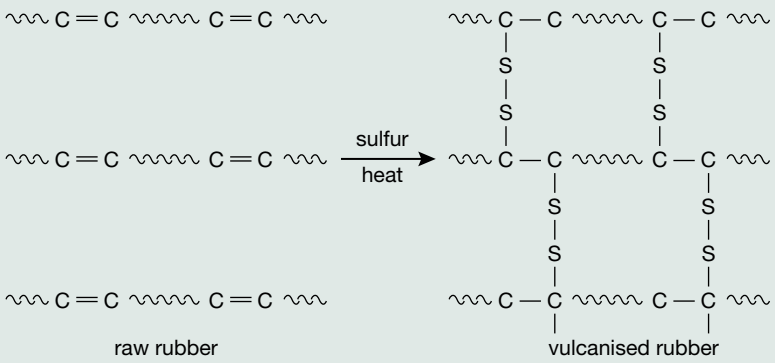
FIGURE 8.5 Harvesting of natural rubber

EXTENSION: Vulcanising rubber

American inventor Charles Goodyear experimented for many years to find a way of countering the tendencies of natural rubber. In 1839, Goodyear accidentally dropped a piece of rubber that had been treated with sulfur onto a hot stove. He later found that this sample had improved elasticity and greater resistance to temperature change. His discovery formed the basis of the process of **vulcanisation**, which is still widely used in the rubber manufacturing industry today to improve the durability and elasticity of natural rubber.

Natural rubber is vulcanised in an industrial process where it is mixed with sulfur and heated. The sulfur atoms form cross-links between chains of rubber molecules, reducing the number of double bonds, as shown in figure 8.6.

FIGURE 8.6 Formation of vulcanised rubber



When vulcanised rubber is stretched, the sulfur linkages stop the chains slipping past one another and the rubber returns to its original shape when the stretching force is removed. Vulcanised rubber is an example of a cross-linked polymer where the sulfur atoms link straight chains together. Rubber is used for tyres, carpet backing, tyre tubing and the soles of sports shoes.

vulcanisation the process in which sulfur is added to rubber and heated to cause cross-linking of polymer chains, increasing the strength of rubber

Addition polymers

Addition polymers are formed from monomers, which contain a carbon–carbon double bond. The bond is broken and a new bond forms with the next monomer, creating very long polymer chains.

8.2.3 Condensation polymers

Condensation polymers form from monomers that have *two functional groups per molecule*. Such molecules are said to be **difunctional**. Polymerisation occurs when one functional group of one monomer reacts with one functional group on another. A small molecule (which is often, but not always, water) is removed in this process. Because each monomer has one more functional group on it (they are difunctional), this process can be repeated on and on to produce the long polymer chains.

Many natural polymers such as proteins and polysaccharides are examples of condensation polymerisation. Common synthetic polymers produced using this method include nylon, polyester and polyethene terephthalate (PET) plastic. As you might expect, many of these reactions show similarities to the esters introduced in topic 7.

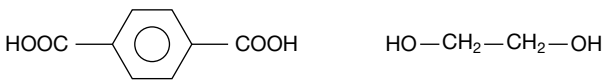
Synthetic condensation polymer: PET (or PETE) plastic

Plastic soft-drink bottles are a good example of an article made from **PET** polymer. Its full name is polyethene terephthalate. It has also been used to make fibres and is often included in carpets designed for high traffic areas. It is an example of a **polyester**.

To make PET, terephthalic acid (also known as benzene-1,4-dicarboxylic acid) and ethylene glycol (also known as ethane diol) are reacted together. The structures of these two compounds are shown in figure 8.8.

FIGURE 8.7 Disposable drink bottles are made of PET plastic.

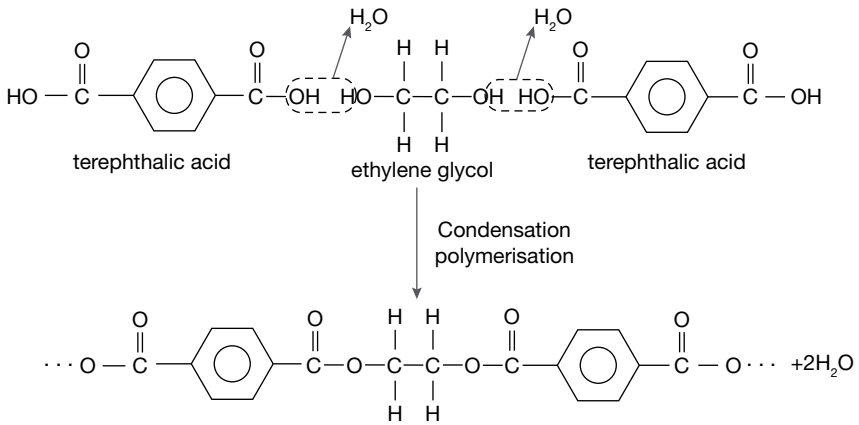
FIGURE 8.8 The structures of terephthalic acid and ethylene glycol



terephthalic acid — ethylene glycol

The reaction is shown in figure 8.9. Note the removal of the small molecule (water), which makes this another example of a condensation polymerisation.

FIGURE 8.9 The reaction between terephthalic acid and ethylene glycol



difunctional a molecule that contains two functional groups

PET thermoplastic polymer produced from the monomer ethylene terephthalate; typically used as synthetic fibres (polyester) and packaging (PET); commonly recycled

polyester synthetic polymer in which the structural units are linked by ester bonds

EXTENSION: Nylon

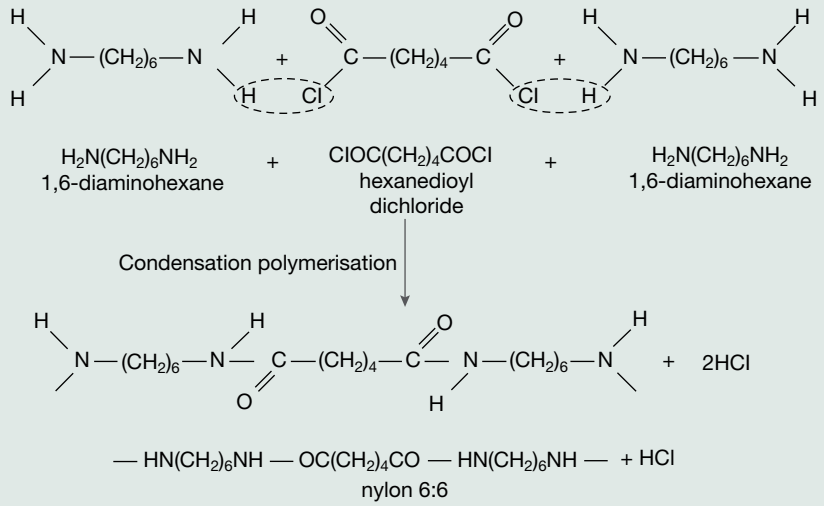
Nylon can be extruded when molten to form fibres or sheets of strong, durable and elastic material. It is used to make fabrics and other products, such as the bristles of brushes and velcro. Its invention had a great impact on the textile and clothing industry.

Nylon 6:6 is one type of copolymer formed by the condensation polymerisation of two monomers: hexanedioyl (adipyl) chloride and 1,6-diaminohexane. An amine group in 1,6-diaminohexane reacts with a chloride group in adipyl chloride, forming a bond between the two molecules and eliminating a molecule of hydrochloric acid. Further molecules add on in the same way to build up a long chain.

Nylon is a linear chain containing up to 100 repeated units. The name 'nylon 6:6' refers to the existence of six carbon atoms on each of the units. You will notice that a repeated linkage pattern exists consisting of —NHOC— units. This is called the amide group and results when the small molecule HCl is eliminated from the two functional groups during the condensation reaction. Nylon is classed as a polyamide fibre. This same linking group is found in the protein polymer chain that makes up wool, although there it is referred to as a peptide link.

FIGURE 8.10 Electron micrograph of velcro, which is made of nylon. Velcro was inspired by burrs on plants that attach themselves to passing objects through thousands of tiny hooks.

FIGURE 8.11 Nylon 6:6

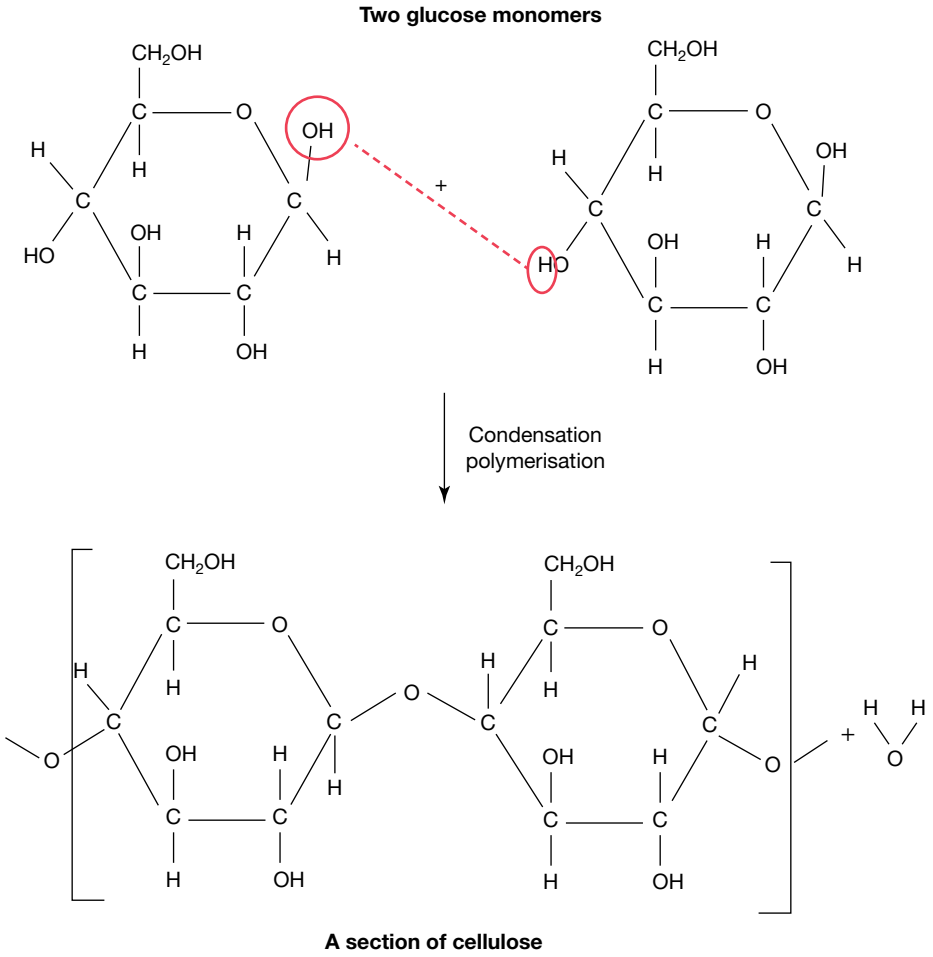


A modification of this manufacturing process uses hexanedicarboxylic acid, $HOOC(CH_2)_4COOH$, in place of hexanedioyl dichloride. This produces exactly the same nylon product, but has the advantage that the small molecule removed is water.

Natural condensation polymer: cellulose

Cellulose is the main component in the walls of plant cells giving them strength and stability. It is a natural polymer found in plant materials such as grains, vegetables, wood, paper, linen and cotton. Cellulose is the most abundant polysaccharide — that is, it is a polymer made up of sugar monomers. It is a linear polysaccharide formed by condensation polymerisation of the glucose monomer releasing a water molecule between each pair of monomers.

FIGURE 8.12 Cellulose is a natural condensation polymer formed by glucose monomers.



Condensation polymers

A condensation polymer is formed in a reaction in which two molecules combine to form a larger molecule, producing a small molecule such as HCl or H_2O as a by-product. If the molecule lost is water, it can also be described as a dehydration reaction.

Common addition and condensation polymers and their uses are summarised in table 8.1.

TABLE 8.1 Common polymers and their uses

Polymer	Type of polymerisation	Use
Polyethene	Addition	Cling film, squeeze bottles, milk crates
Polypropene (polypropylene)	Addition	Mixing bowls, ice-cream containers, moulded chairs
Polyvinyl chloride (PVC)	Addition	Water piping, upholstery covering, toys, flexible laboratory tubing, suitcases, packaged meats, juice bottles, credit cards, shampoo bottles
Nylon	Condensation	Hosiery, carpets, clothing, bearings, fishing lines, tennis racquet strings, toothbrush bristles, mascara brush bristles
Polystyrene	Addition	Foam for insulation, plates, cups, trays, packaging, surfboards, beanbag filling, refrigerator and washing machine components
Polytetrafluoroethylene (PTFE, Teflon®)	Addition	Washers, coating on frying pans, razor blades, skis
Polymethyl methacrylate (Perspex®)	Addition	Advertising signs, telephones, light fittings, aeroplane windows
Polyethene terephthalate (PET)	Condensation	Soft drink bottles, other clear plastic bottles, hard wearing carpets, as a polyester filling (for pillows, weather jackets, etc.)
Urea formaldehyde or phenol formaldehyde (Bakelite®)	Condensation	Backs of television sets, ash trays, vacuum jugs
Epoxy resin (Araldite®)	Condensation	Handles of screwdrivers
Melamine formaldehyde (Formica®)	Condensation	Kitchen cupboards and benches
Polyurethane	Condensation	Foam rubber, pillow filling, packaging

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Weblinks Polymerisation animation
Polymers
Nobel Prize in Chemistry for conductive polymers

8.2 Activities

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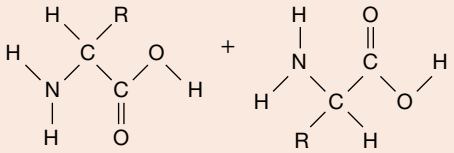
8.2 Quick quiz on | **8.2 Exercise** | **8.2 Exam questions**

8.2 Exercise

1. Explain why alkenes are able to act as monomers for addition polymerisation, but alkanes are not.
2. Give examples of the following.
   a. Three natural polymers
   b. Three synthetic polymers
3. Calculate the approximate molar mass of a polyethene sample, the average chain length of which is 50 000 carbon atoms.

4. Ethene is a gas that can be stored for long periods of time. Explain why it can be stored as a gas without it polymerising into polyethene.
5. A sample of polyethene is found to have an approximate molar mass of 450 000 units.
   a. Calculate the approximate number of monomer units that are joined to make the molecules in this sample.
   b. What is the approximate chain length in terms of number of carbon atoms?
6. Are alkynes able to act as monomers for addition polymerisation? Explain.
7. How is addition polymerisation similar to condensation polymerisation?
8. How does addition polymerisation differ from condensation polymerisation?
9. Proteins are natural condensation polymers made of amino acids monomers. Explain the meaning of the following terms.
   a. Natural polymer b. Condensation polymer c. Monomer
10. Draw a circle around the parts of these two molecules that combine and are lost as a result of a condensation reaction.



8.2 Exam questions

Question 1 (1 mark)

MC Which of the following best describes polymerisation?

A. Small molecules decompose to form a new substance.
B. Two or more chemicals react together.
C. Small molecules react to form very long molecules.
D. Small molecules react to form a thicker substance.

Question 2 (1 mark)

MC Which of the following molecules is NOT able to form an addition polymer?

A. C_2H_3Cl B. $C_2H_2Cl_2$ C. CH_3CHCCl_2 D. CH_3CHCl_2

Question 3 (1 mark)

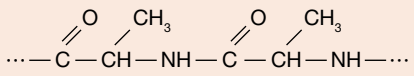
Draw a segment of polypropene including three monomers.

Question 4 (4 marks)

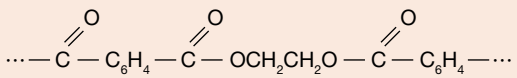
For each of the polymers **a** and **b** provided, answer questions i and ii.

i. Write the structural formula of the monomer(s) from which the polymer is made.
ii. State whether the monomer has undergone addition or condensation polymerisation.

a.

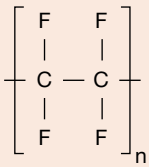


b.



Question 5 (4 marks)

Polytetrafluoroethylene is used as non-stick coatings on cookware. The structure of this polymer can be represented by the diagram provided.



a. Name and draw the structure of the monomer from which this polymer is formed. **(2 marks)**
b. Write the molecular formula of this monomer. **(1 mark)**
c. Name the type of reaction used to form the polymer. **(1 mark)**

More exam questions are available in your learnON title.

8.3 Linear and cross-linked polymers

KEY KNOWLEDGE

- The distinction between linear (thermoplastic) and cross-linked (thermosetting) addition polymers with reference to structure and properties
- The features of linear addition polymers designed for a particular purpose, including the selection of a suitable monomer (structure and properties), chain length and degree of branching

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8.3.1 Thermoplastic and thermosetting polymers

Plastics may be classified in a number of ways. One of these, as we have already seen, is on the mechanism of their manufacture; that is, whether they are addition polymers or condensation polymers. Another method is on the basis of their thermal behaviour. This method provides insight into their uses and applications as well as their ability to be recycled.

- **Thermoplastic** (or thermosoftening) plastics may be repeatedly melted, reshaped and hardened by cooling. An example of a thermoplastic polymer is polystyrene.
- **Thermosetting** plastics do not melt but char when heated. These plastics must be moulded or shaped during their manufacture. Bakelite is an example of a thermosetting plastic material.

Table 8.2 lists a range of common polymers classified as to their thermal behaviour.

TABLE 8.2 The reaction of some common polymers to heat

Polymer	Reaction to heat	Classification
Polyethene	Melts/softens	Thermoplastic
Polypropene	Melts/softens	Thermoplastic
Polyvinyl chloride	Melts/softens	Thermoplastic
Nylon	Melts/softens	Thermoplastic
Polystyrene	Melts/softens	Thermoplastic
Polytetrafluoroethylene (PTFE, Teflon®)	Melts/softens	Thermoplastic
Polymethyl methacrylate (Perspex®)	Melts/softens	Thermoplastic
Polyethene terephthalate (PET)	Melts/softens	Thermoplastic
Urea formaldehyde or phenol formaldehyde (Bakelite®)	Chars	Thermosetting
Epoxy resin (Araldite®)	Chars	Thermosetting
Melamine formaldehyde (Formica®) — condensation	Chars	Thermosetting
Polyurethane	Chars	Thermosetting

Thermoplastic and thermosetting polymers obviously have very different properties. These can be explained if we look closer at the long polymer chains themselves, and the forces and other physical factors that operate between them.

thermoplastic polymers that soften on heating; also known as thermosoftening

thermosetting polymers that do not soften on heating and char if heated strongly

linear polymer polymer that does not have side chains

8.3.2 Linear polymers

Linear polymers are thermoplastic polymers. As their name suggests, linear polymers are polymers that form as long chains. In any sample of a linear polymer, huge numbers of such chains will be present and these, due to their length, will invariably become tangled.

Intermolecular forces and the degree of entanglement have significant effects on linear polymers:

- Intermolecular forces operate between these chains because each chain is, in effect, a very large molecule.
- These forces between the chains, along with their entanglement, will give a degree of rigidity to the overall polymer. The stronger the forces and the higher the entanglement, the harder the polymer will be. The weaker the forces and the lesser the degree of entanglement, the softer and more flexible it will be.
- When heat is applied to such polymers, the weak intermolecular forces between the chains are, at least partially, overcome, and the chains can move relative to one another. This results in the overall polymer becoming softer and more liquid-like. Hence, it is able to be moulded into new shapes, which will once again 'set' upon cooling as the intermolecular forces between the chains once again take effect.
- This process of melting and softening is additionally assisted by the increased thermal vibration of the long-chain polymer molecules with increasing temperature. This also assists them to slide past each other and form new shapes.

Therefore, the observed properties of thermoplastic polymers — that is, their ability to be melted and softened and then moulded into new shapes, and the variety that they display in terms of flexibility and rigidity — can all be explained in terms of their long linear polymer chains and how these chains interact with one another.

Linear (thermoplastic) polymers

- Linear polymers are thermoplastic polymers.
- Thermoplastic polymers may be repeatedly melted and reshaped and are hardened by cooling.
- Heating of thermoplastic polymers overcomes the intermolecular forces between the long polymer chains.
- The length and degree of entanglement of the chains affects their properties.

8.3.3 Cross-linked polymers

Polymers can also be produced where the long chains are formed first and then deliberately connected together by strong covalent bonds. This process is called **cross-linking**. This means that the forces that operate between the chains are now strong covalent forces rather than the weak intermolecular forces seen in linear polymers.

If only a small amount of cross-linking is introduced, an **elastomer** results. As the name suggests, elastomers can be stretched because the chains can still move, but in a restricted manner, past each other. Once the stretching force is removed, however, the cross-links pull the polymer back into its original shape. Examples of elastomers include vulcanised rubber (refer to Extension: Vulcanising rubber), where a small amount of cross-linking using sulfur produces disulfide bonds between the chains. Most elastomers are thermosets but some are thermoplastic. Polyacrylic rubber and silicone rubber are examples of elastomers.

FIGURE 8.13 What property of vulcanised rubber makes it suitable for car tyres and rubber hoses?

cross-linking the bonding between two polymers

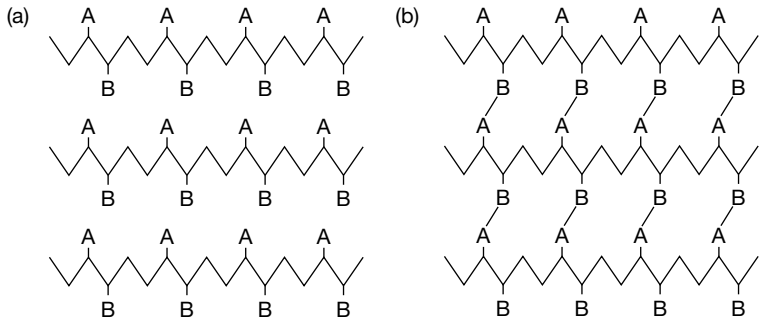
elastomer polymer that can be stretched and return to its original shape

Thermosetting polymers

If the degree of cross-linking within a polymer is large, then, as expected, the entire arrangement will be very rigid and strong due to the covalent bonding that now exists in all dimensions. Such polymers will not soften or melt when heated but can char at high temperatures. These are thermosetting polymers, and can be referred to as thermosets.

To make a thermosetting polymer, long linear chains with functional groups attached to them are produced first. The cross-linking is then brought about either by using heat or by adding a chemical to react between the lateral functional groups linking the chains together.

FIGURE 8.14 (a) A collection of linear polymers with lateral functional groups A and B. (b) When the A and B groups are linked, a rigid cross-linked polymer results.



For example, a particular article may be made by placing the preliminary linear polymer into a mould and then heating it. Cross-links form, producing a rigid thermosetting article in the mould. Note that thermosetting polymers may still be classed as plastics because, in the early stages of their manufacture (before the cross-linking is introduced), they can still be moulded.

A familiar example of the second method of producing cross-links is when two-part glues are mixed. One tube contains the preliminary linear polymer with the lateral functional groups. The other tube contains the 'hardener'. When mixed, this chemical reacts with the lateral functional groups, bonding them together and linking the chains together with strong covalent cross-links.

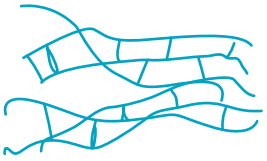
Cross-linked (thermosetting) polymers

- Cross-linked polymers are thermosetting polymers.
- Cross-linked polymers are created by long polymer chains with functional groups subsequently joined by strong covalent bonds.
- Large amounts of cross-linking create rigid, strong compounds due to covalent forces.

8.3.4 Summary of thermoplastics, elastomers and thermosets

Table 8.3 provides a summary of the different properties of thermoplastics, elastomers and thermosets.

TABLE 8.3 Comparing thermoplastics, elastomers and thermosets

Property	Thermoplastics	Elastomers	Thermosets
Effect of heat	Softens/melts	Varies	Does not soften/melt; chars
Number of cross-links	Nearly zero	A small number	A much larger number
Forces between chains	Weak intermolecular forces; often dispersion forces and/or hydrogen bonding	Weak intermolecular forces and some strong covalent bonds	Strong covalent bonds
Hardness	Variable	Varies according to degree of cross-linking	Usually hard
Capacity for recycling	A number are ideally suited	Limited	Very limited
Diagrammatic representation	thermoplastic	elastomer	 thermoset

EXTENSION: The first plastics

The first plastics were developed in the middle of the 19th century, following a competition set up by a manufacturer of billiard balls. In those days, billiard balls were made of ivory and the manufacturer, Phelan and Collender, offered a substantial prize for the discovery of a satisfactory alternative. One of the entrants in the competition was John Wesley Hyatt, who developed a substance called celluloid. He did not win the prize, because celluloid is volatile, and the billiard balls would have exploded on impact. However, celluloid was the first thermoplastic — a substance moulded using heat and pressure that retains its shape once cooled — and celluloid became a widely used polymer. Its applications included dental plates and men's collars, but it is most identified with its role in the early photographic and motion picture industries, where it was used to make film.

FIGURE 8.15 The casings and dials of early telephones, radios and many other appliances were frequently produced from Bakelite.

In 1907 the first completely synthetic polymer, called Bakelite, was produced by Belgian American chemist Leo H Baekeland (1863–1944). Bakelite, unlike celluloid-based plastics, is a thermosetting polymer and cannot be softened by heat once it has set. It is also resistant to common acids and solvents. These properties made it virtually indestructible, and Bakelite was a landmark in the history of plastics. Its discovery started a large plastics industry and the age of plastics had begun.

8.3 Activities

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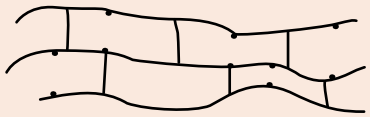
8.3 Quick quiz on | 8.3 Exercise | 8.3 Exam questions

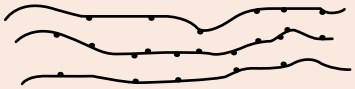
8.3 Exercise

1. MC Which one of the following is a characteristic of thermoplastic?
   - A. Cross-links between chains
   - B. Can only be moulded during manufacture
   - C. Chars when heated
   - D. Has weak dispersion forces between chains
2. MC Which diagram best represents an elastomer?

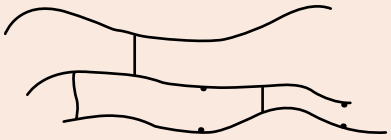
A.



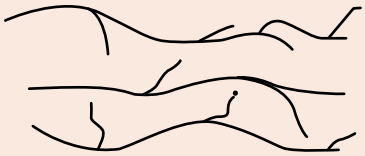
B.



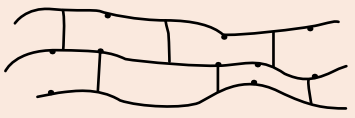
C.



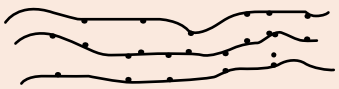
D.



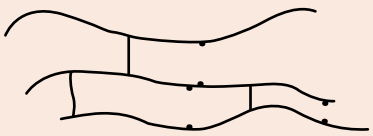
3. Describe the interchain bonding and properties of a thermosetting polymer.
4. MC Plastic used for making which of these materials would be expected to have a large degree of interchain cross-linking?
   - A. Calculator casing
   - B. Water bottles
   - C. Food wrap
   - D. Stretching 'bungee' ropes
5. MC Which of the polymers with segments provided would be expected to melt when heated gently?



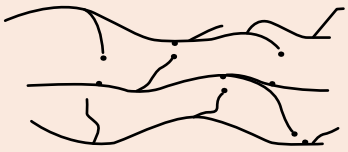
(I)



(II)



(III)



(IV)

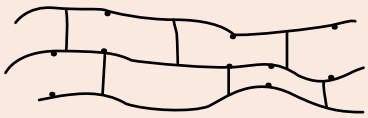
   - A. Only polymers I and III would melt.
   - B. Only polymers II and IV would melt.
   - C. Polymers II, III and IV would melt.
   - D. All polymers would melt.

6. Explain why the bonding between chains is stronger in thermosetting polymers than in thermoplastic polymers.
7. Describe the difference(s) between thermoplastic polymers, elastomers and thermosetting polymers in terms of the following.
   a. Bonding
   b. Properties
8. a. In general terms, describe the steps in the production of an object made from a thermosetting polymer.
   b. Why are thermosetting polymers still described as plastics, even though they cannot be melted and remoulded?
9. A number of common thermoplastics are recycled in large amounts, whereas articles made from thermosetting polymers are rarely recycled. Explain why this is so.
10. Two linear polymers, *A* and *B*, are of roughly equal molar mass. Hydrogen bonding exists between the chains in *A* whereas dispersion forces are the only forces operating between the chains in *B*.
    How will the properties of polymer *A* compare to polymer *B*?

8.3 Exam questions

Question 1 (1 mark)

MC Which diagram best represents an elastomer?

A.


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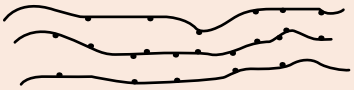
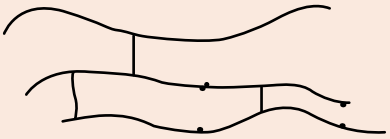
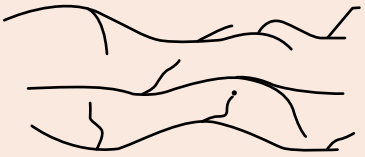
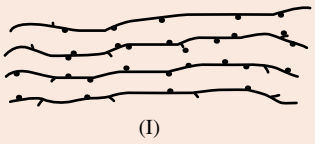


C.


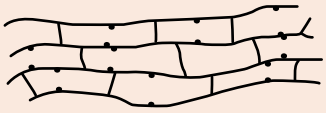
D.


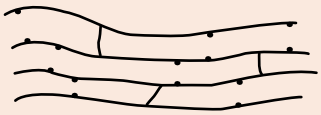
Question 2 (6 marks)

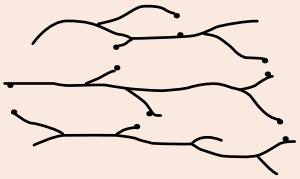
Use the provided diagrams to answer the questions that follow.

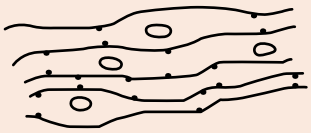

(I)

(II)


(III)


(IV)


(V)


(VI)

Which of the polymers provided:

a. could be used for a rigid mobile phone casing? **(1 mark)**
b. could be used for hard but transparent material? **(1 mark)**
c. could be used as a flexible food wrap? **(1 mark)**
d. has covalent interchain bonding? **(1 mark)**
e. does not melt but chars when heated? **(1 mark)**
f. has an additive that increases flexibility? **(1 mark)**

Question 3 (1 mark)

MC Which of the following statements about polymers is/are true?

Statement I: Polymers are covalently bonded molecules.

Statement II: Polymers are typically electrical insulators.

Statement III: All polymer chains link to other chains by cross-linking.

Statement IV: All polymers are solids at room temperatures due to their covalent bonding.

A. Only statements I and II are true.
B. Only statements I, II and IV are true.
C. Only statements II, III and IV are true.
D. All statements are true.

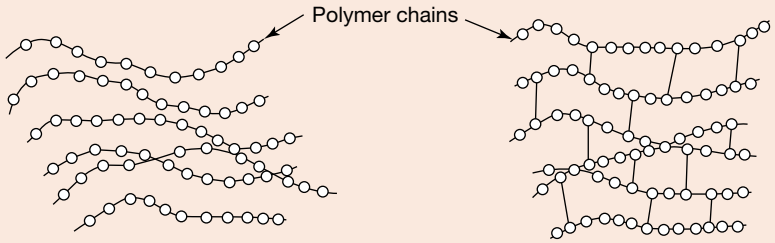
Question 4 (6 marks)

Plastic can be described as thermosetting or thermoplastic.

a. Explain the difference in these two types of plastic. **(2 marks)**
b. Give an example of a thermosetting plastic and describe the property that makes it useful. **(2 marks)**
c. Give an example of a thermoplastic and describe the property that makes it useful. **(2 marks)**

Question 5 (7 marks)

Two polymer chains are shown.



Polymer A Polymer B

a. Describe the type of bonding in and between the polymer chains. **(4 marks)**
b. Explain which polymer you would expect to soften easily when heated. **(2 marks)**
c. Explain whether you would expect either of these polymers to conduct electricity. **(1 mark)**

More exam questions are available in your learnON title.

8.4 Polymer selection

KEY KNOWLEDGE

- The features of linear addition polymers designed for a particular purpose, including the selection of a suitable monomer (structure and properties), chain length and degree of branching

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8.4.1 Features of linear polymers

In 1963, the Nobel Prize–winning chemist Giulio Natta said, 'A chemist setting out to build a giant molecule is in the same position as an architect designing a building. He has a number of building blocks of certain shapes and sizes, and his task is to put them together in a structure to serve a particular purpose.' Today we can take this statement to mean that, due to their versatility, polymers can be produced for almost any imagined purpose.

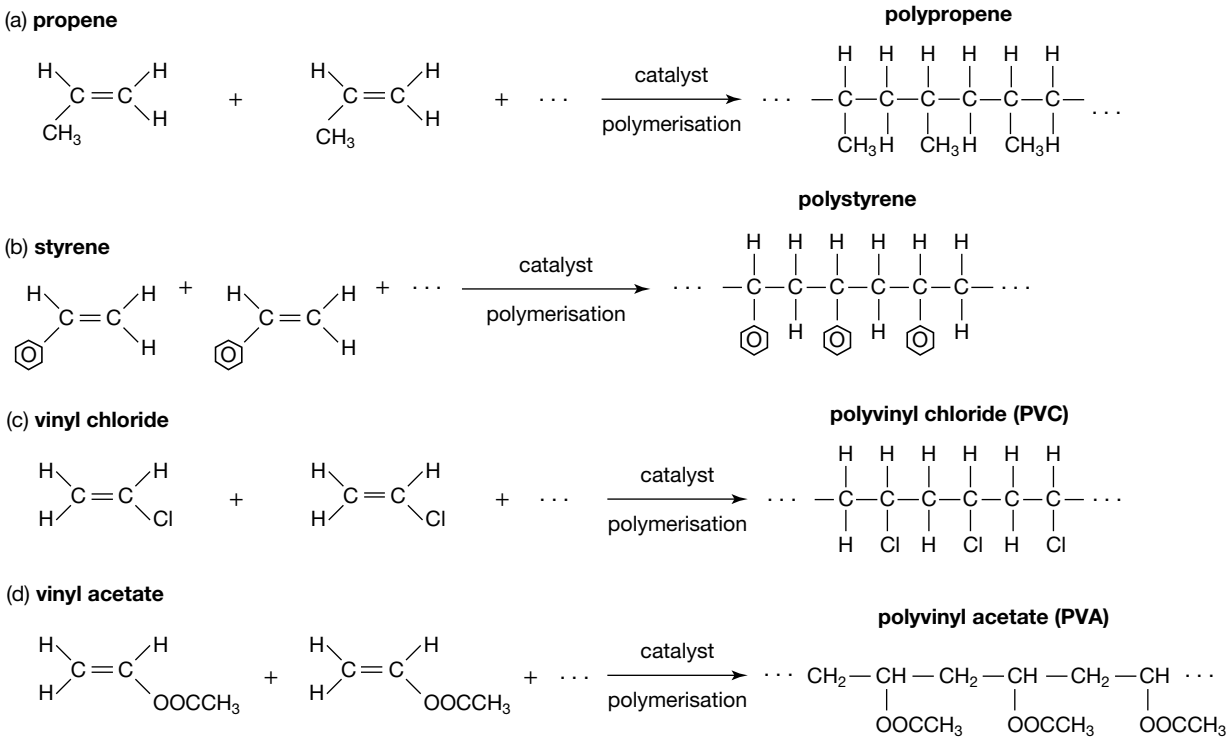
The versatility of polymers is due to their many different properties. They can be hard or soft, flexible or rigid, transparent or coloured, brittle or able to be stretched. These properties of polymers are affected by the:

- choice of monomer
- length of the polymer chains
- degree of branching.

Choice of monomer

As seen earlier in this topic, ethene can be polymerised to produce polyethene. The critical feature that enables this to occur is the presence of the carbon–carbon double bond in the monomer, ethene. By using ethene and substituting other atoms or groups of atoms in place of hydrogen atoms, molecules can be produced in which the double bond is still intact. These molecules will, therefore, be able to act as monomers by which a further range of polymers with different properties may be produced. Figure 8.16 shows this process for a number of commonly used polymers. You will note that, as is the case for ethene, the method of polymerisation is addition polymerisation and that all monomers contain a carbon–carbon double bond.

FIGURE 8.16 The formation of (a) polypropene from propene, (b) polystyrene from styrene, (c) polyvinyl chloride (PVC) from vinyl chloride and (d) polyvinyl acetate (PVA) from vinyl acetate



Choosing monomers

To create an addition polymer, monomers should contain a carbon–carbon double bond.

Structure and properties of monomers

The functional groups that are present in the monomers affect not only the properties of the monomers but also the properties of the polymers that they form. Non-polar polymers — for example, polyethene (see figure 8.17), polypropylene and polytetrafluoroethene — are resistant to water because they are unable to form any interactions with polar water molecules. In addition, polypropylene (see figure 8.18) has a lower density than polyethene because of the methyl group attached to every second carbon atom. (However, the density of polyethene is also affected by the degree of branching.)

FIGURE 8.17 Polyethene is non-polar so only weak dispersion forces exist between the chains of the polymer (the degree of branching also affects the properties).

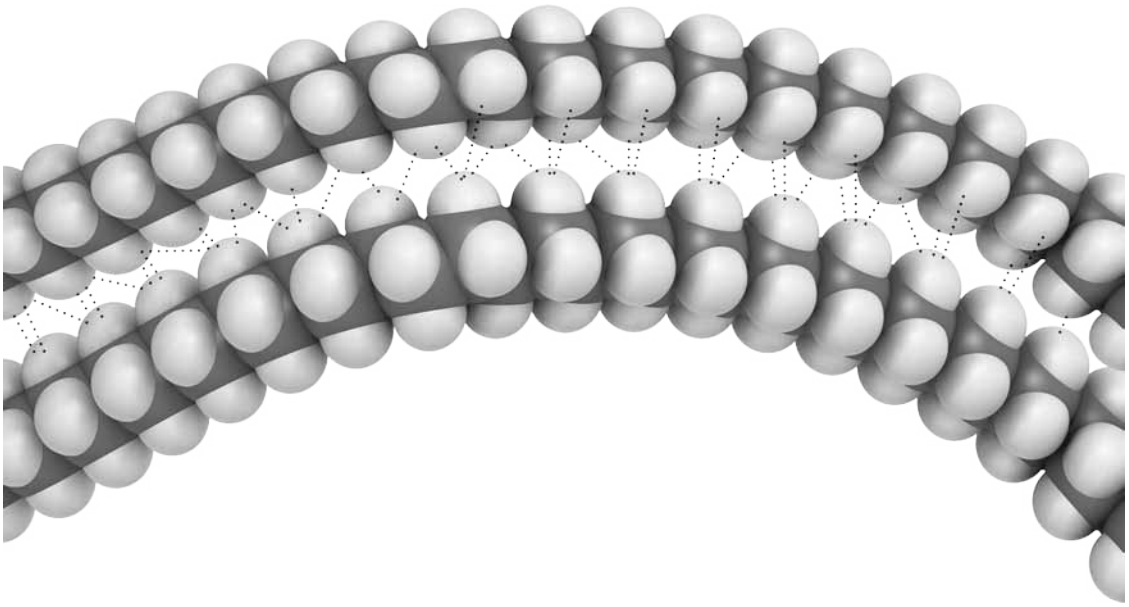
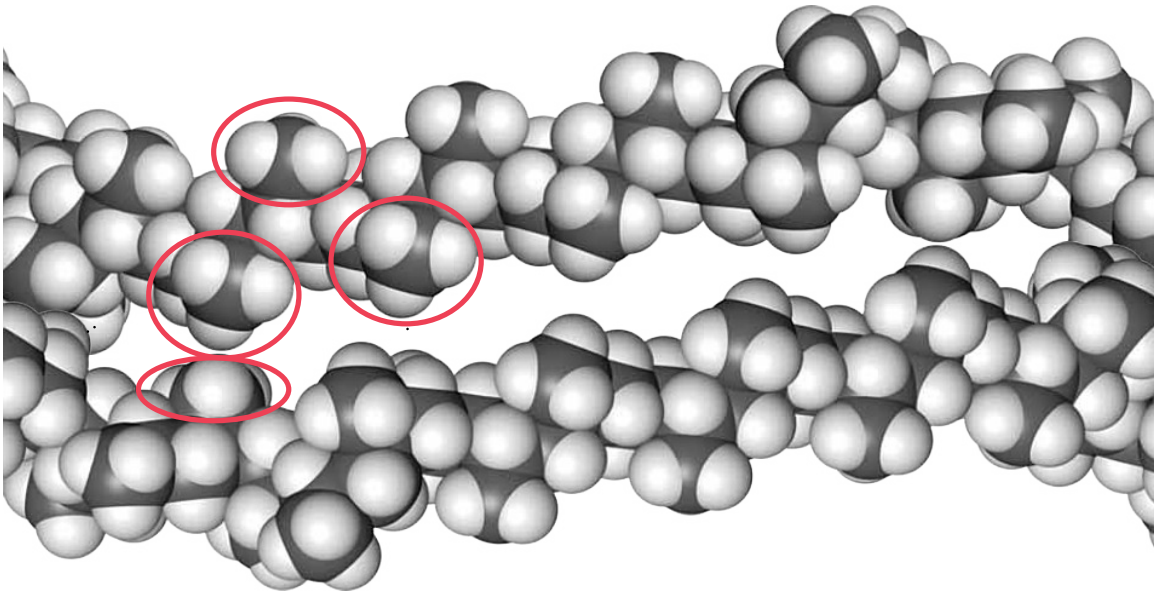


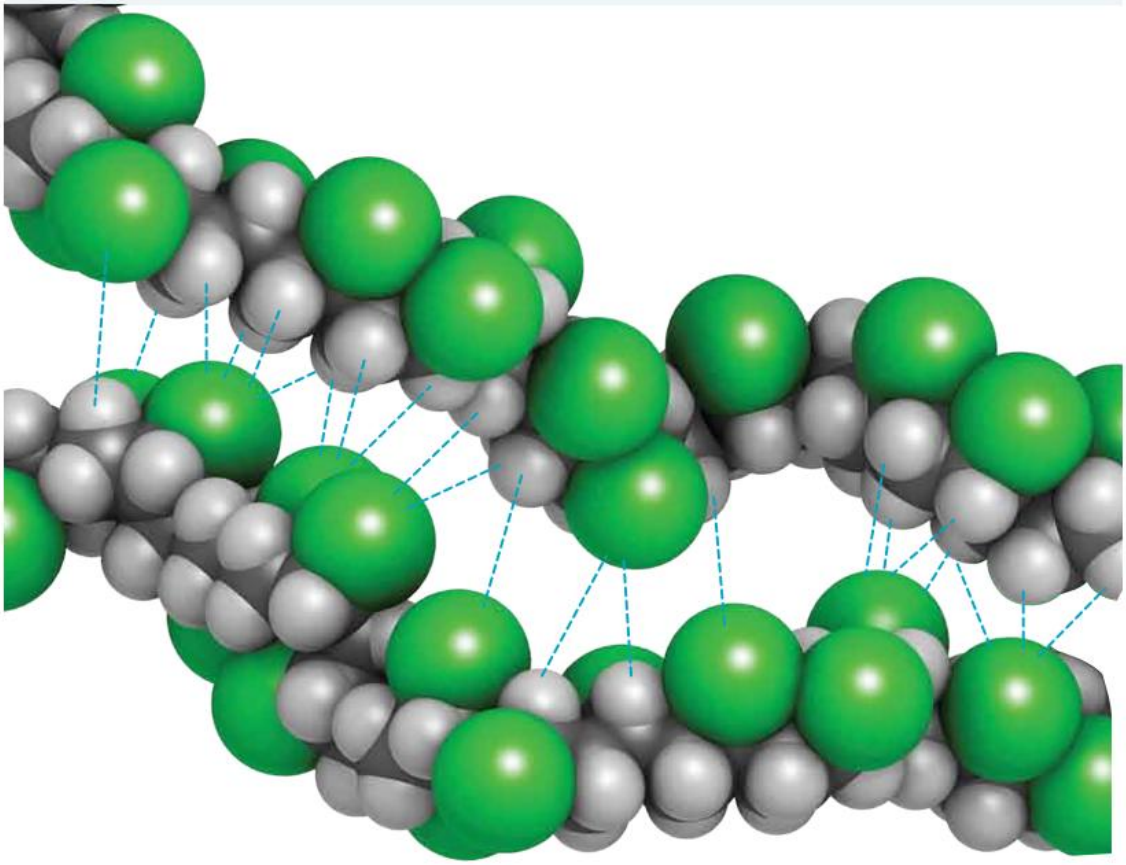
FIGURE 8.18 Polypropylene is less dense than unbranched polyethene, because the methyl groups (circled in red) disrupt the uniform structure and close packing of chains.



Polytetrafluoroethene, or polytetrafluoroethylene, also known as Teflon, has highly electronegative fluorine atoms attached to the carbon atoms. These form strong polar covalent bonds with tightly held electrons. The fluorine atoms will not interact with other atoms and so the polymer is chemically inert — which is why it is used in non-stick pans. It also has applications in the manufacture of waterproof fabric and for coating medical tubing and equipment.

Polyvinyl chloride (see figure 8.19) is a polymer that is very widely used to make window frames, pipes, medical devices and wire insulation. Its structure is similar to polyethene, but with a chlorine atom attached to every second carbon atom. A polar bond is formed, which results in dipole–dipole attractions between the polymer chains. This results in a rigid strong material that is chemical resistant and weatherproof.

FIGURE 8.19 Polyvinyl chloride (PVC) chain showing chlorine atoms (green). Dipole–dipole attractions (blue) can be formed between chlorine atoms (which have slightly negative charges) and hydrogen atoms (which have slightly positive charges) on neighbouring chains.

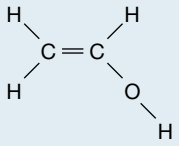


A polymer that is water-soluble is polyvinyl alcohol. It is used in adhesives and emulsifiers and to make other chemicals. Another use is to make plastic film for detergents and disinfectants because it will biodegrade. As might be expected, the solubility is the result of the hydroxyl groups in the polymer that can form hydrogen bonds with water.

tlvd-0554

SAMPLE PROBLEM 2 Drawing a section of a polymer of polyvinyl alcohol

Polyvinyl alcohol has many uses. Play 'slime' and the water-soluble plastic used in laundry and dishwashing pods are just two of these. It is made from vinyl alcohol, the structure of which is shown.

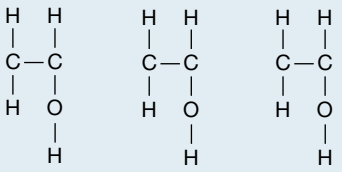


Draw a section of the polymer that is produced when this alcohol is polymerised to make polyvinyl alcohol.

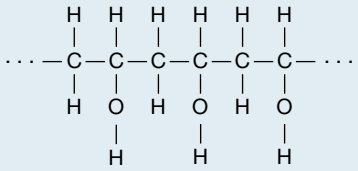
THINK

1. Draw a number of vinyl alcohol molecules next to each other, but with only a single bond between the carbons.

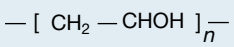
WRITE



2. Place a new bond between the molecules to link them together. Show three dots at each end to signify only a section is shown.

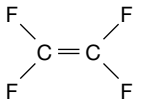


Alternatively, you can show the answer in condensed form.



PRACTICE PROBLEM 2

Teflon® is made from polytetrafluoroethylene (PTFE). The monomer used to make this polymer is shown.

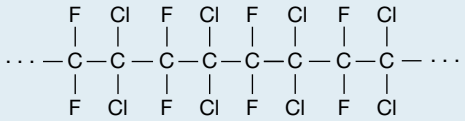


Draw a section of the polymer that is produced when this monomer is polymerised to make PTFE.

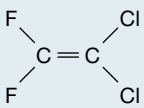
tlvd-0555

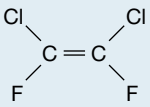
SAMPLE PROBLEM 3 Determining the monomer that makes up a polymer

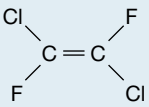
A section of a polymer produced at a research laboratory has the structure shown.



Which of the following three possible monomers is it made from?

(a) 

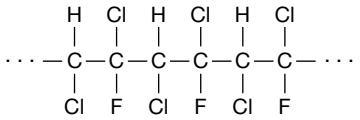
(b) 

(c) 

THINK	WRITE
1. The polymer shows alternating F and Cl atoms above and below the carbon backbone. It, therefore, cannot be (b).	
2. Each F is attached to the same carbon, as is each Cl. It cannot be (c). Structure (a) is the only monomer that will produce the required pattern when polymerised.	Monomer (a) is the only monomer that will produce a polymer with the fluorine and chlorine atoms in the correct pattern.

PRACTICE PROBLEM 3

Draw the monomer that the following polymer is made from.



Tip for drawing monomers and polymers

When drawing the structural formula for monomers and polymers, remember it is important to retain the position of the functional groups above or below the carbon backbone.

elog-1825

tlvd-0626

EXPERIMENT 8.1 online only

Cross-linking an addition polymer to make slime

Aim

To investigate how the properties of a linear polymer may be altered by the introduction of weak cross-linking between its chains

on Resources

Interactivity Making polymers (int-3849)

Length of polymer chain

Longer polymer chains are, in fact, larger molecules. This means that the dispersion forces between chains are stronger. Additionally, the longer polymer chains are, the more they become tangled. They are also less likely to slide over each other than smaller chains. These features mean that it is harder for the chains to move with respect to one another. Consequently, polymers with longer chains tend to be harder/stronger and less flexible, and are able to withstand higher temperatures before they soften or melt.

Length of polymer chains

Polymers with longer chains tend to be harder/stronger and less flexible; they can withstand higher temperatures before they soften or melt.

Degree of branching

In reality, polymerisation does not produce purely linear chains. For various reasons, branching occurs that causes side chains of various lengths to form. A linear polymer may, therefore, be more accurately thought of as branches on a tree. As a general rule, the more extreme the conditions required for polymerisation, the greater the number and length of such side branches. This phenomenon is illustrated by polyethene.

Ethene can be polymerised to produce both low- and high-density polyethene. Polyethene was discovered in 1933. Scientists had been experimenting with the effect of heat and very high pressures on ethene. It was not until 1939 that full commercial production of polyethene began.

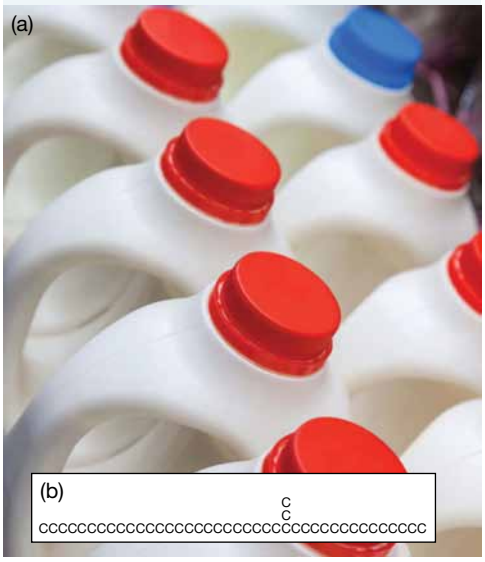
- **Low-density polyethene (LDPE)** was the first polyethene produced. Its polymer chains support many long side branches, producing a low-density substance (see figure 8.20). Since the only forces causing these polymer chains to attract each other are dispersion forces, the effect of the branches is to keep the chains apart. Because the attraction becomes weaker as the chains are further apart, the density of the resultant compound is low, and LDPE is soft, flexible and translucent, with a waxy surface that repels water. Among its uses are plastic bags.
- **High-density polyethene (HDPE)** was developed in the early 1950s, when ethene was polymerised using lower pressure and lower temperatures. The result was a polymer of ethene with very few branches, and any branches that developed were short. Dispersion forces act more effectively on these chains because they can pack more closely together, and this type of polyethene is more rigid, stronger and more opaque than LDPE. It is slightly flexible and also has a waxy surface that repels water. It is commonly used to make bottles (see figure 8.21).

low-density polyethene (LDPE) thermoplastic polymer with a large number of side branches produced from the monomer ethene

high-density polyethene (HDPE) thermoplastic polymer with very few side branches produced from the monomer ethene

FIGURE 8.20 (a) A translucent, flexible LDPE bag and (b) the carbon backbone of LDPE

FIGURE 8.21 (a) Opaque, rigid HDPE milk containers and (b) the carbon backbone of HDPE



LDPE and HDPE

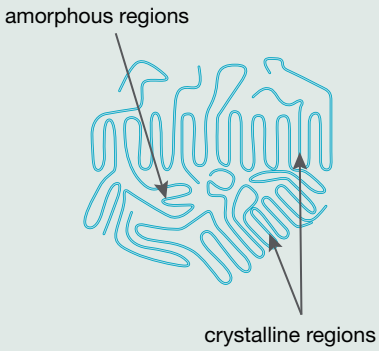
- LDPE has many long side branches and is soft, flexible and translucent, with a waxy surface.
- HDPE has shorter and fewer branches and is more rigid, stronger and more opaque.

EXTENSION: Other factors affecting the properties of polymers

Degree of crystallinity

Polymer chains can be arranged in two ways. They can be crystalline, in that they are regularly organised into lines, or randomly packed with no particular order. Polymers can be partially crystalline or almost totally amorphous. The percentage of regularly ordered chains usually ranges between 10 and 80 per cent. The more crystalline a structure is, the greater its hardness, tensile strength and opacity. This is because the chains are packed closer together in the crystalline regions, resulting in the dispersion forces that operate between then being more effective. This close packing also means that light waves are scattered rather than transmitted, resulting in opacity. Amorphous polymers are more easily deformed and often transparent. The factors that affect crystallinity include chain length, branching and interchain bonding, as seen previously for LDPE and HDPE. HDPE is more likely to conform to a crystalline structure because of its unbranched carbon chains.

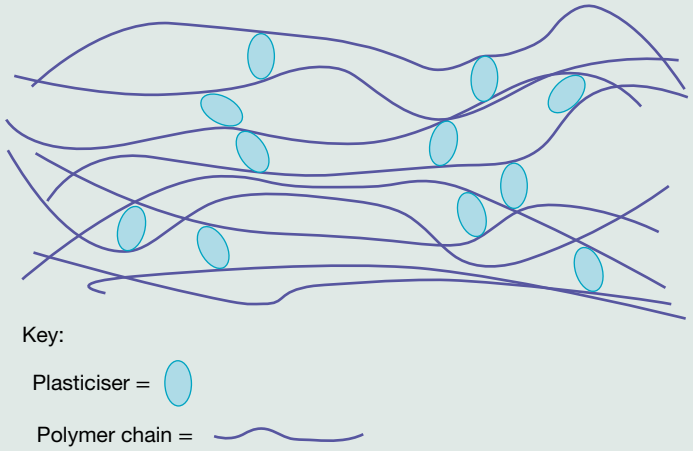
FIGURE 8.22 The strength of a polymer is increased if more crystalline areas are present.



Addition of plasticisers

Plasticisers are major components added to polymers such as PVC during production. They are small molecules that cause the polymer chains to move slightly further apart, resulting in a softer and more flexible polymer. Many different plasticisers are available but the most commonly used are phthalates.

FIGURE 8.23 Plasticisers are added to polymers to increase flexibility.



Other additives

Additives in polymers can have many different uses. Examples include the following:

- UV stabilisers, which absorb UV rays to prevent the polymer breaking down
- flame retardants to reduce the tendency of the polymer to burn
- dyes to add colour or provide patterns.

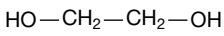
plasticisers small molecules that are added to some polymers to improve their flexibility

8.4.2 Polymer properties and uses

The versatility of polymers results from the factors mentioned in the previous section, and from the degree of cross-linking that may or may not be present. This means a polymer may be chosen (or made) to suit nearly any purpose, based on its properties. Table 8.4 gives just a few examples of how a polymer's properties make it useful for a particular application.

FIGURE 8.24 Many items, such as this in-line skate, are fabricated from a diverse range of plastic components. Each of these is chosen for its particular properties.

TABLE 8.4 Some examples of polymer properties and related uses

Polymer	Monomer	Properties	Use
High-density polyethene (HDPE) $-[CH_2-CH_2]_n-$	Ethylene (ethene) $CH_2=CH_2$	Semi-rigid, translucent, more resistant to heat than LDPE, chemically inert	Milk and juice bottles, water pipes, toys, wheelie bins
Low-density polyethene (LDPE) $-[CH_2-CH_2]_n-$	Ethylene (ethene) $CH_2=CH_2$	Flexible, transparent, water repellent, generally chemically inert, electrical insulator	Plastic shopping bags, food wrap, electrical wire covering
Polyvinyl chloride $-[CH_2-CHCl]_n-$	Vinyl chloride $CH_2=CHCl$	Flexible*, electrical insulator, resists weather (water repellent) and the effects of sunlight	Garden hoses, raincoats, electrical wire covering, spouting, downpipes, non-food containers
Polypropene $-[CH_2-CH(CH_3)]_n-$	Propylene (propene) $CH_2=CHCH_3$	Rigid, water repellent, strong, chemically inert	Plastic furniture, plastic buckets, food containers, car parts
Polystyrene; polystyrene (expanded) $-[CH_2-CH(C_6H_5)]_n-$	Styrene $CH_2=CHC_6H_5$	Tough, rigid, lightweight, soluble in organic solvents, thermal insulator	Plastic cutlery, CD cases, laboratory ware, packaging, insulation, coffee cups
Polytetrafluoroethylene (PTFE) (Teflon®) $-[CF_2-CF_2]_n-$	Tetrafluoroethylene $CF_2=CF_2$	Low coefficient of friction, chemically inert	Non-stick cookware, low-friction bearings, plumbing tape
Polyethene terephthalate (PET/PETE)#	terephthalic acid  ethylene glycol	Tough, flexible, clear, lightweight, shatterproof, inert	Soft drink bottles, water bottles, moulded packaging, carpet, clothing

* Varies and depends on the use of plasticisers; can range from stiff to flexible.
# PET is a condensation polymer. All others shown are addition polymers.

The different structures of thermoplastic polymers and thermosetting polymers explain their different properties. A further consequence of this is their ability to be recycled. As a rule, thermosetting polymers are currently not able to be recycled, whereas some (but not all) thermoplastic polymers can be. The ability of thermoplastic polymers to be softened and melted by heat and then remoulded is ideal for recycling purposes.

8.4 Activities

learnon

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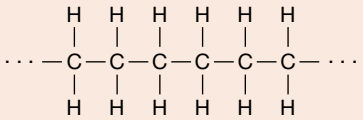
8.4 Quick quiz on | **8.4 Exercise** | **8.4 Exam questions**

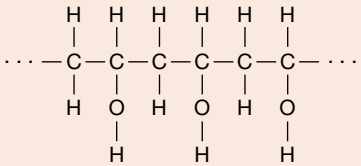
8.4 Exercise

1. Explain how chain length and the degree of branching affect the properties of a linear polymer.
2. Explain why plasticisers can make a polymer softer and more flexible.
3. With reference to the structures of polyethene and polyvinyl alcohol, shown, explain why one is water repellent and the other is water soluble.

polyethene



polyvinyl alcohol

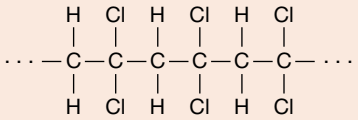


4. Explain why polymers with regularly spaced side groups along their carbon backbones are more likely to show a higher degree of crystallinity than those where the side groups are more randomly orientated.
5. An application requires a polymer that can soften under moderate heat but set again when it cools. When set, it needs to be rigid and water repellent.
   Discuss how you would design a polymer to meet these specifications.
6. Modern cars contain many parts that are made of plastic; for example, bumper bars and front side panels. Give two advantages and one disadvantage of using polymers in this way.
7. Give one advantage for the use of polymers in place of each of the following materials.
   a. Metal
   b. Wood
   c. Glass
   d. Wool
   e. Surface coating
8. PVC is the third most produced polymer after polyethene and polypropene. However, a much greater amount of polyethene and polypropene are recycled than is PVC. Give possible reasons for this discrepancy.

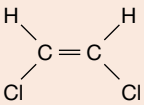
8.4 Exam questions

Question 1 (2 marks)

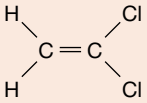
The structure of polyvinylidiene chloride is shown. The structures of two possible monomers from which it might be made are also shown.



Monomer I



Monomer II



a. Which is the correct monomer and why? **(1 mark)**

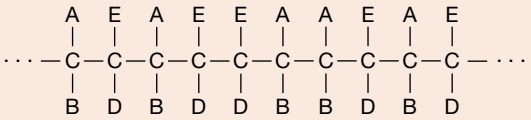
b. Draw the polymer produced if the other monomer is polymerised. **(1 mark)**

Question 2 (2 marks)

Polypropene, polystyrene and polytetrafluoroethene are all water repellent. Explain with reference to their structure and bonding.

Question 3 (1 mark)

The structure of a section of hypothetical polymer is as follows.



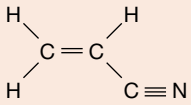
Draw the structure of the monomer from which it is produced.

Question 4 (2 marks)

You have been asked to produce a polymer that is flexible and suitable for sealing left over foods. You have a choice of either ethene or vinyl alcohol as raw materials. Discuss which would be the preferable starting chemical. Refer to the monomer, length of chain and degree of branching in your response.

Question 5 (4 marks)

Acrylonitrile is a monomer produced from propene and ammonia. It is used to manufacture nitrile rubber, which can then be used in the manufacture of seat belts.



a. Draw a section of the polymer formed from three acrylonitrile molecules. **(2 marks)**

b. Show another method of representing the polymer. **(1 mark)**

c. Name the polymer. **(1 mark)**

More exam questions are available in your learnON title.

8.5 Plastic recycling and innovations in design

KEY KNOWLEDGE

- The categorisation of different plastics as fossil fuel–based (HDPE, PVC, LDPE, PP, PS) and as bioplastics (PLA, Bio-PE, Bio-PP); plastic recycling (mechanical, chemical, organic), compostability, circularity and renewability of raw ingredients
- Innovations in polymer manufacture by condensation reactions and break down by hydrolysis reactions, contributing to the transition from a linear economy towards a circular economy

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8.5.1 Categorising and recycling plastics

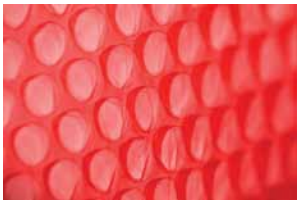
You may have noticed the triangular symbols on the base of many plastic items. These codes identify the type of plastic and help you to make informed decisions about recycling. The presence of these triangles does not mean that the item is recyclable or if the plastic is made from biomass.

TABLE 8.5 Symbols for types of plastics and recycling

Symbol and name	Details	Examples of uses	Recycling information
 PETE Polyethene terephthalate	• Most used consumer plastic • Mainly for single use products • Not heat resistant • Difficult to clean properly, should not be reused	Textiles such as fleece garments, carpets, stuffing for pillows and life jackets; soft drink and water bottles 	Widely recycled
 HDPE High-density polyethene	• Hard wearing, does not break down • Stronger than PET • Reusable • Suitable for freezing	Compost bins, detergent bottles, pipes, plumbing fittings, household bags and irrigation pipes, milk, juice, water and detergent bottles 	Easily recycled

(continued)

TABLE 8.5 Symbols for types of plastics and recycling *(continued)*

Symbol and name	Details	Examples of uses	Recycling information
3 V PVC	• Can leach toxins • Not suitable for food and drinks • Useful for outdoor products because it doesn't break down • Produces toxic chemicals when heated and this limits the ability to be recycled	Juice bottles, detergent bottles, PVC piping, credit cards 	Not easily recycled but can be made into more PVC products such as flooring
4 LDPE Low-density polyethene	• Thin and flexible • Easy and inexpensive to produce • Safe to use with food	Frozen food bags, bin liners, squeezable bottles, flexible container lids, cling film, carry bags and packaging film, bubble wrap 	Soft, scrunchable plastics can be returned to the supermarket for recycling
5 PP Polypropene	• Hard and light weight • Withstands heat • Resistant to grease and chemicals • Safe to reuse	Reusable microwave containers, kitchen ware, nappies, yoghurt containers, straws, disposable cups and plates, landscaping border stripping, battery cases, margarine tubs 	Can be recycled but depends on local council
6 PS Polystyrene	• Lightweight and soft • Flammable • Inexpensive to make • Easily moulded • Can release harmful chemicals, particularly if heated	Packing 'peanuts', disposable cups, plates and trays, insulation, disposable takeaway containers 	Not easily recycled and must be disposed of in general waste; avoid use if possible

Symbol and name	Details	Examples of uses	Recycling information
7 OTHER Other	• Strong and tough • Includes polycarbonates and other plastics • Possible release of hazardous BPA	Beverage bottles, baby milk bottles, electronic casing, lenses for sunglasses and safety goggles 	Generally not recyclable

Bio-based plant compostable plastics are being developed that can go into general waste. They are *not* recyclable.

8.5.2 Bioplastics

Millions of tonnes of plastics are consumed annually in the world. As discussed earlier in this topic, the polymers in these plastics are sourced from non-renewable fossil fuels. They are also usually non-biodegradable, causing severe environmental problems on land and in the sea. Microplastics, the remnants of broken-down plastics, are in our oceans, soil, food and us.

The use of **bioplastics** has been proposed as a solution to this problem because they can be sourced from polymers obtained from biomass — for example, agricultural, cellulose and corn starch waste.

The advantage for the environment is the reduction of the use of raw materials from fossil fuels; however, *even if a plastic is sourced from biomass, it does not mean it is biodegradable*. In addition, the manufacture of bioplastics may involve the use of fossil fuels for energy for production and recycling. The impact on the environment is dependent on the different types of bioplastics that are produced and their end-of-life disposal. Terms used to describe the end-of-life possibilities of all types of plastic include the following:

- **Degradable**: all plastic is degradable, which means it can be broken up into smaller and smaller fragments. However, it may never degrade into simple molecules, and the time taken for degradation can be considerable. Additives can be added to plastics to enable them to degrade more quickly in sunlight; for example, photodegradable plastics.
- **Biodegradable**: biodegradable plastic can be broken down completely into water, carbon dioxide and compost by microorganisms, provided the conditions are optimal. If oxygen is limited, a more damaging greenhouse gas, methane, is obtained. This decomposition can take years. Note that some fossil-fuel plastics are biodegradable, and some bioplastics are not biodegradable.
- **Compostable**: compostable plastic will biodegrade under the right conditions in a compost site within 180 days. Microorganisms break the plastic down into carbon dioxide, water, inorganic compounds and biomass at a similar rate as other organic materials in the compost pile, leaving no toxic residue.
- **Recyclable**: plastic is recyclable when waste materials can be converted into new materials and objects. Recycling usually takes place at a specialised recycling facility.

bioplastic a polymer obtained from renewable biomass; for example, agricultural, cellulose and corn starch waste

degradable can be broken into smaller and smaller fragments or simpler chemical structures

biodegradable can be broken down or degraded by natural means, including by microbes

compostable a product that can be degraded by microorganisms in a moist, warm environment to produce matter that can support plant life in a relatively small amount of time

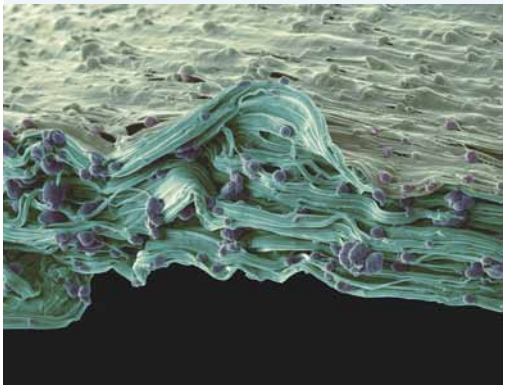
recyclable materials which can be reprocessed into new materials

8.5.3 Types of bioplastic

Polylactic acid (PLA)

Imagine using plastic screws in surgery that gradually degrade into harmless lactic acid while broken bones mend themselves, saving repeat surgery to remove them. **Polylactic acid (PLA)** is such a material. It is a versatile, thermoplastic polymer obtained from plant products such as corn starch, maize or sugar cane. It is biodegradable and cost-efficient to produce. PLA is a polyester with molecular formula $(C_3H_4O_2)_n$, it is soluble in some organic solvents, has a relatively low heat resistance and is easy to melt and manipulate. Applications of PLA include bottles, food packaging, products in the healthcare and medical industry, shrink wrap, plastic bags (see figure 8.25), PPE, disposable cups and cutlery, and 3D printing.

FIGURE 8.25 SEM image of a plastic bag made from PLA. The starch grains can be seen in purple.

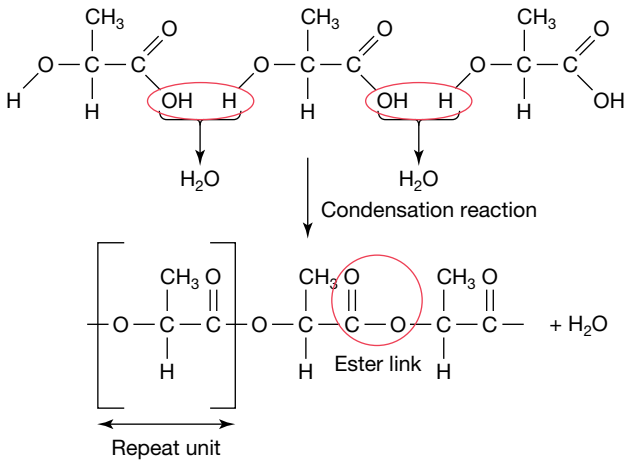


Manufacture of PLA

The production of PLA is sustainable because it comes from plant-based raw materials. Production involves several steps:

1. The raw material, starch, is extracted from the plant material.
2. The sugar dextrose is processed from the starch.
3. Dextrose is converted to lactic acid by fermentation using microorganisms.
4. Either of two methods of condensation polymerisation:
   a. producing lactide as an intermediate
   b. direct condensation polymerisation of lactic acid.
5. Further condensation reactions result in a monomer that polymerises to form PLA. It may need to be strengthened with additives.

FIGURE 8.26 Formation of PLA by condensation reaction



A benefit of PLA is that it breaks down when exposed to the environment. If left in the ocean, it would degrade in 6 to 24 months — whereas conventional plastics will last for hundreds or even thousands of years. PLA is categorised with an identification code of 7, the miscellaneous category that includes all plastics that don't fit into the other categories (refer to table 8.5). PLA begins to break down with a hydrolysis process (by breaking ester links) followed by enzymatic or microorganism-based activity leading to fragmentation of the PLA and, eventually, harmless substances. Exposure to thermal or UV radiation can increase the rate of degradation.

polylactic acid (PLA) a versatile, thermoplastic, biodegradable polymer obtained from plant products such as corn starch, maize or sugar cane

Degradation rates depend on:

- additives to the PLA
- pH — hydrolysis of ester bond is favoured at low pH
- molecular weight — degradation decreases as molecular weight increases
- crystallinity — a more orderly arranged structure is more difficult to degrade than a random arrangement.

Disadvantages of the use of PLA include the following:

- It is unsuitable for high-temperature applications.
- The use of fertilisers is required to grow the plants used in production.
- Land used to grow crops for production may be needed for food. (Using non-edible or waste plant matter is preferable.)
- It must be taken to a specialised facility for composting (not home-compostable).
- Production involves some water usage (although less than for synthetics).
- Additives may not break down.
- If mixed into general plastic waste, it contaminates the recycling process of other plastics.

Biopolyethene (BioPE)

Biopolyethene (common name is biopolyethylene) can be made from various raw materials including sugar cane, sugar beet and wheat grain, so it is a renewable resource. It also reduces greenhouse gas emissions because it removes carbon dioxide from the atmosphere. However, it is not biodegradable and cannot be composted. Because it is a thermoplastic, it can be melted and remoulded into different shapes and recycled (see figure 8.27). Biopolyethene, once manufactured, is chemically identical to polyethene made from fossil fuel–based feedstocks. It is used to manufacture shopping bags, films, bottles and car parts. Polyethene can be recycled and converted into new products using current technology, regardless of the initial source.

FIGURE 8.27 Recycled polyethene can be made into granules that are then melted and remoulded into new materials.

Manufacture of bio-polyethene

Production involves the following steps:

1. Yeast is used to ferment the sugars in the plant biomass into ethanol according to the equation:

$$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(l) + 2CO_2(g)$$

2. The ethanol is distilled and dehydrated to form ethene.
3. The ethene undergoes addition polymerisation to form polyethene.

Ethanol made from sugarcane requires one fewer step than ethanol made from starch because the starch must be converted to sugar first. In addition, the production using sugar cane can be powered using the combustion of bagasse removing the need for external power input. Bagasse is the pulp remaining from sugarcane after the fermentation and distillation process occurs. This production method is more sustainable than using other plant materials because the waste is reduced and one less step is used. Disadvantages of this process include the use of food crops for fuels and fossil fuels being used for harvesting and vehicles.

Bio-based polypropene (bio-PP)

Propene is the next most useful monomer for chemical production after ethene. **Biopropene** is a thermoplastic and has similar properties to **bioethene** but it is harder and more resistant to heat. It is also recycled similarly. This polymer is used in injection moulding, textiles, film, ropes and other applications.

biopropene chemically identical to propene but produced from biomass

bioethene chemically identical to ethene but produced from biomass

Manufacture of bio-PP

Production involves the following steps:

1. Raw materials are sourced from biomass. Ideally non-edible plants are used, but corn, sugar cane, vegetable oil can be substituted. The next steps in the process depend on the feedstock available.
2. In one method, biomass is fermented to produce propan-2-ol (IPA).
3. The propan-2-ol is then dehydrated to obtain propene.
4. The propene undergoes addition polymerisation to produce biopolypropene (polypropylene).

Neither biopolyethene nor biopolypropene is biodegradable; both must be recycled.

8.5.4 Plastics and the circular economy

Incorporating plastics into the **circular economy** is a complex and immense task. Plastics are so common in many aspects of our lives because they are cheap, widely available and solve problems. For example, plastic packaging helps keep food fresh, which prevents food wastage. But a considerable amount of unnecessary packaging is used on many consumer items.

FIGURE 8.28 Plastic helps keep food fresh, but we need to reconsider unnecessary uses of plastic.

Current challenges in incorporating plastics into the circular economy include:

- Most feedstocks are still sourced from fossil fuels, but inroads have been made with biobased plastics.
- Most plastics at the end of life are sent to landfill or incinerated to supply energy. The majority are not yet recycled because of the associated difficulties in recycling.
- Biodegradable plastics in landfill are problematic because they can lead to methane emissions, which contribute to global warming. Both methods of disposal (landfill/burning and biodegradation) are not sustainable and are detrimental to the environment.

However, treatment for plastics that are collected at the end of life are developing and include mechanical, chemical and organic recycling.

In **mechanical recycling**, the plastic waste is collected and sorted. Suitable plastics are then shredded, washed, dried and melted before being converted into pellets. The plastics do not have their structure significantly changed in this process. The pellets can be used in the manufacture of new products. Unfortunately, this process can only be applied a few times because the quality of the plastic is degraded each time as the length of the polymer chains is shortened. To rectify this, new plastic is combined with the recycled plastic before different products are manufactured. Mechanical recycling is energy efficient and decreases the need to consume more raw materials.

Most types of plastics can be converted into their original molecular building blocks in **chemical recycling**, regardless of colour or composition. This means that high-quality raw material is used to produce a wider range of new plastics and other chemicals that are not degraded. These include safe food containers and medical items. The cycle for plastics is, therefore, complete. Catalysts and/or high temperatures assist the process, and the presence of contaminants is not a problem. The process is complex and begins with the collection of the waste. The plastic waste is heated and then undergoes one of three processes:

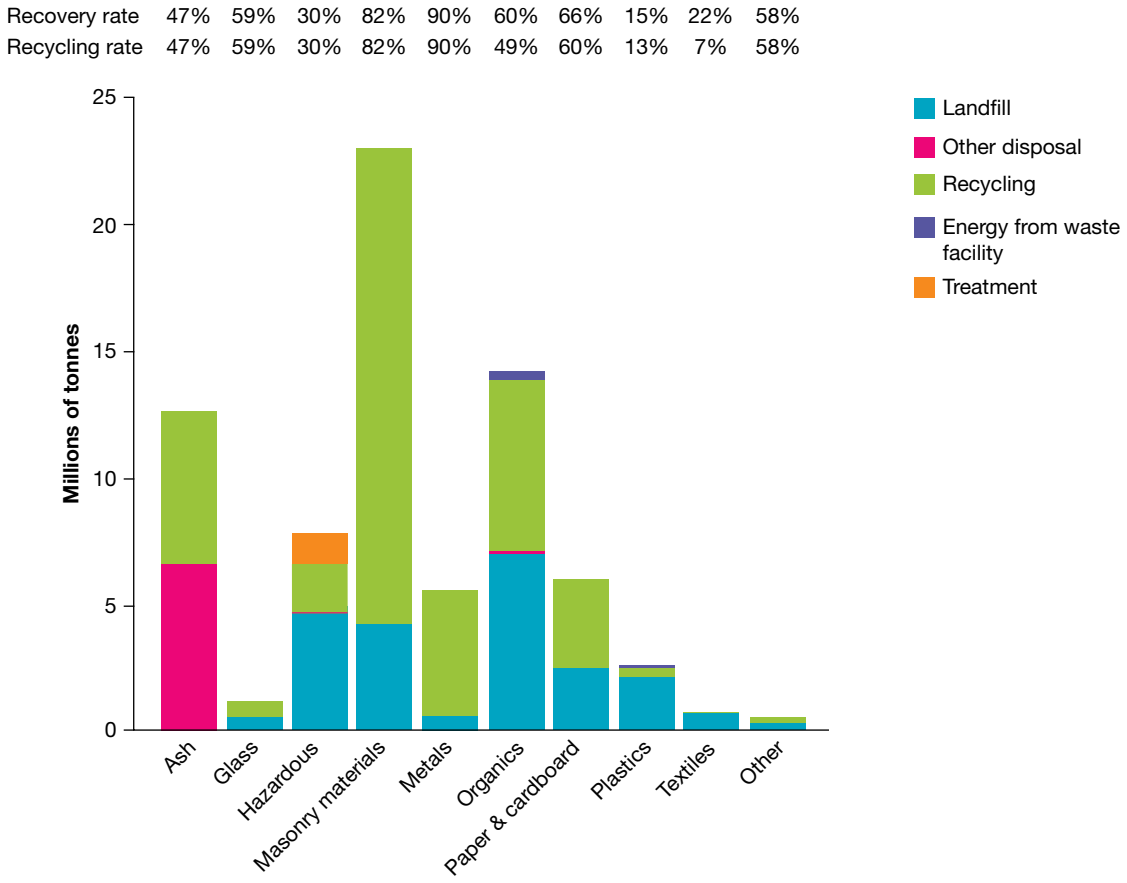
- *Hydrogenation:* the polymer is broken down using hydrogen and heat into products that go to be processed in refineries.
- *Gasification:* the polymer is heated with oxygen to make synthesis gas that can be used to make hydrocarbon feedstocks.
- *Pyrolysis:* the polymer is heated without oxygen to make hydrocarbons for use as fuels or in refineries.

circular economy economy in which waste and pollution is eliminated, products at the end of life are reused, repaired or recycled, and the environment is not harmed

mechanical recycling physical recycling of plastics where the plastic is shredded and melted, and converted into pellets

chemical recycling recycling using heat or chemicals in which the chemical structure of the polymer is converted into monomers that can then be reused in chemical processes

FIGURE 8.29 The proportion of plastics collected is very low and most of what is collected is deposited in landfills. Only 15% of generated plastic waste was collected and only 13% of this waste was recycled with a small proportion (less than 3%) used for energy production.



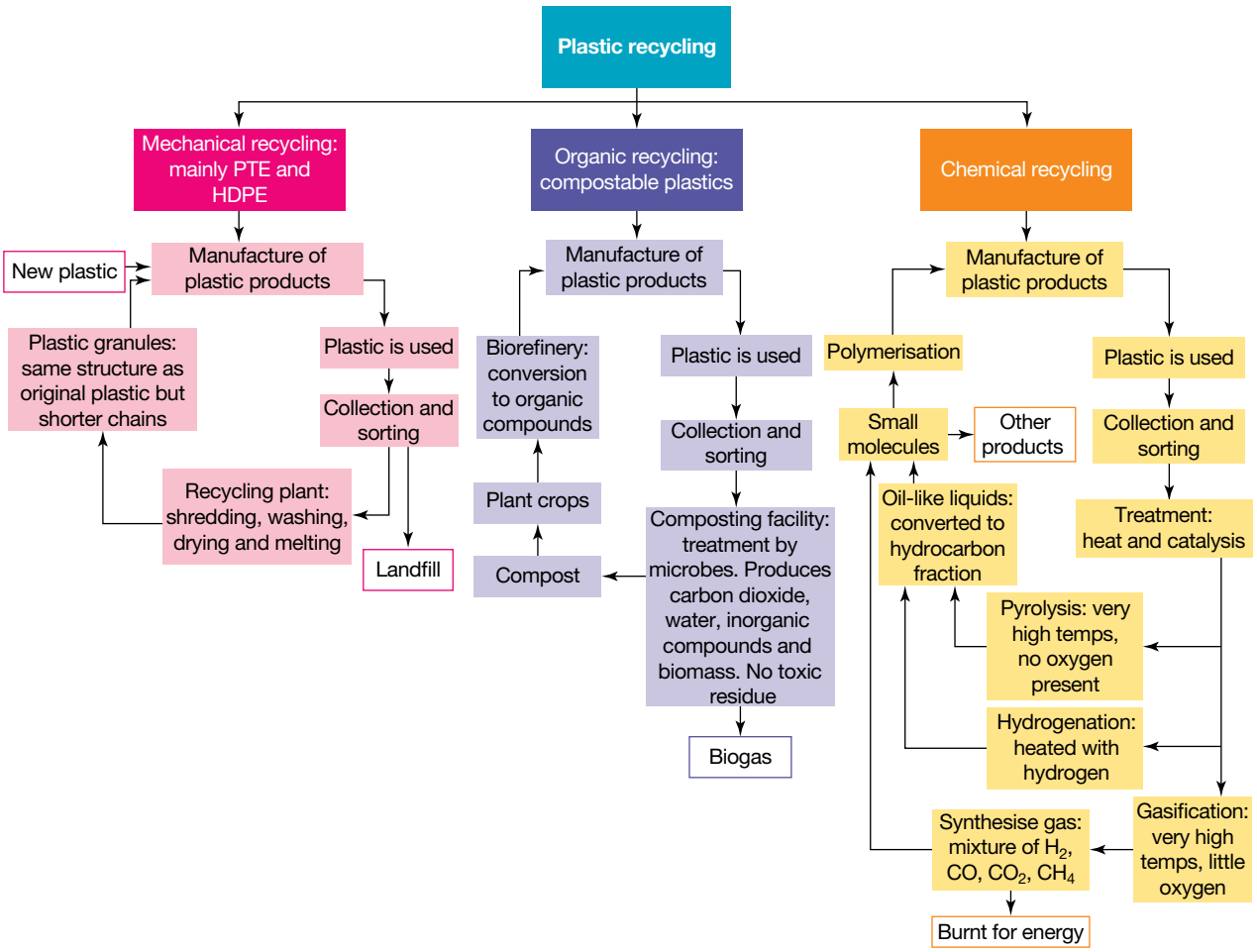
Chemical recycling of PLA is discussed further in section 8.5.5.

Organic recycling (composting) is the aerobic treatment (industrial composting) or anaerobic treatment (bio-gasification) of specifically designed compostable plastic waste using microorganisms in humid, warm surroundings. In aerobic composting, fungi and aerobic bacteria consume the biopolymers for energy, and the by-products are carbon dioxide, water vapour and heat. Part of the biopolymer is converted into biomass as part of the compost. Generally, the composting takes place at a specialised recycling facility that has higher temperatures than home compost heaps. Moreover, mixing compostable plastic with other plastics can make the recycling of the other plastics difficult.

A system using anaerobic digestion to turn some waste plastic into energy is being developed. The anaerobic digestion process varies upon disposal feedstock and is also dependent on temperature. The microbes treat the plastic and the wastewater simultaneously. They then die and leave behind liquid and solid biomass, which can be used as fertiliser and biogas. The biogas can be separated into methane and carbon dioxide. Further research is required to find specific microbes that are evolving to degrade different plastics.

organic recycling waste material is treated using microorganisms in humid, warm conditions

FIGURE 8.30 Incorporating types of recycling into the circular economy



Considerations of plastic product use, reuse and recycling as part of the circular economy are summarised in table 8.6.

TABLE 8.6 Plastic products and the circular economy

Towards a circular economy	Examples
1. Eliminate use where possible	Avoid packaging for items such as clothing, toys and sporting equipment Use cloth shopping bags instead of plastic Consider need to purchase plastic item or packaging
2. Reuse items	Reusable coffee cups, water bottles, etc. Refillable containers for bulk goods
3. Recycle plastics	Make bottles, furniture, shoes and clothing from recycled plastics
4. Make recycling easier	Deposits for containers to facilitate collection Make collection and recycling centres more accessible Collect more plastics so they do not end up in waterways and oceans
5. Use compostable plastics	Clear labels on containers so consumers can make a choice Automated sorting of plastics at waste collection facilities so appropriate treatment occurs

Towards a circular economy	Examples
6. Develop innovative design	Products that last longer and are less likely to be disposed of Effective and economical methods of sorting and recycling Design plastics for easier degradation
7. Use renewable energy	Renewable energy to be used for manufacture, recycling and transport of plastics
8. Use renewable raw materials	Use biomass to produce plastics to replace fossil fuel–based plastics
9. Use less hazardous chemicals	Ensure that chemicals used in the manufacture, use and recycling process are safe for living organisms and the environment

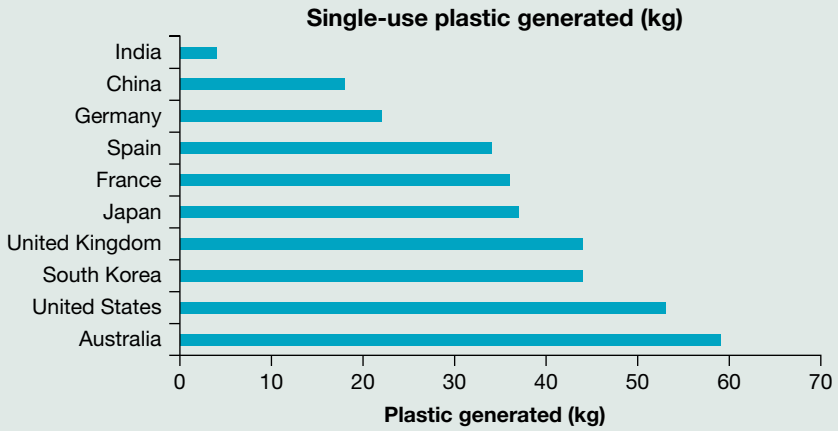
CASE STUDY: Australia's use of plastics

Australians produce more waste plastic per person than people in any other country. Households are the largest contributor, supplying 47 per cent of all plastic waste. Significant changes need to occur within our society to move towards a circular economy.

- Only 9 per cent of the 2.5 million tonnes of plastic waste generated in 2019 was sent for recycling, while 84 per cent was sent to landfills.
- About 130 000 tonnes of plastic ends up in the marine environment each year.

Since 2021, exporting plastic waste has been banned so new strategies are needed. The good news is that the federal government has put into place a National Plastics Plan to address the problem of plastic waste that, among other proposals, aims to have 100 per cent of packaging reusable, recyclable or compostable by 2025.

Figure 8.31 Per capital single use plastic generation



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Weblink Government action on plastics and packaging

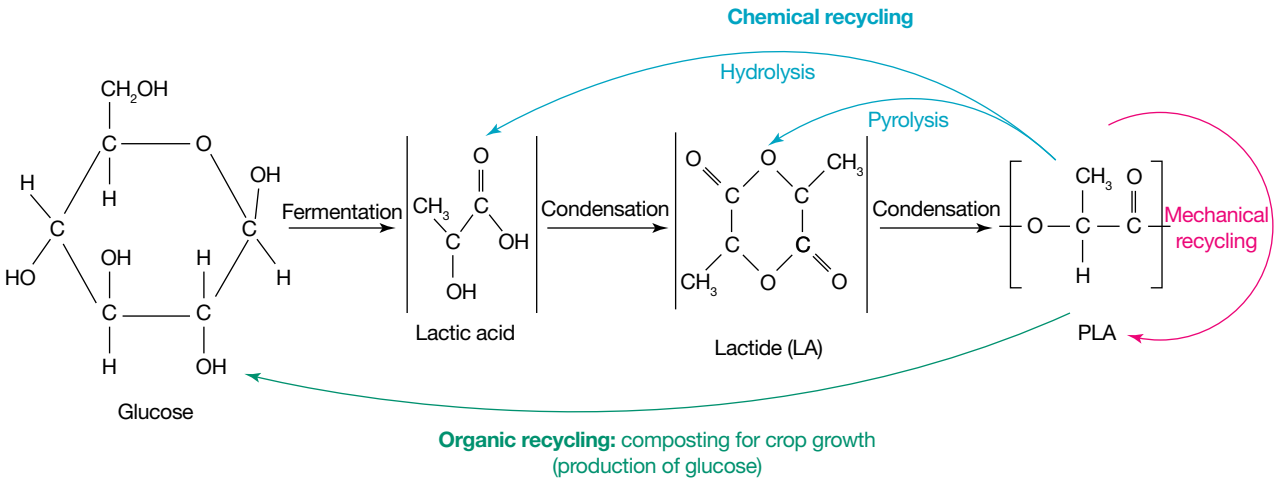
8.5.5 Innovations in polymer manufacture

One innovation in polymer manufacture is the more efficient and sustainable design of chemical recycling using catalysts to reduce energy usage. Another is investigating the use of different bacteria or fungal enzymes to break down plastics such as PET into reusable monomers. Many condensation polymers can be recycled by hydrolysis reactions into their monomers.

The new bioplastic condensation polymer PLA, for example, is widely used and can be recycled in several ways, as shown in figure 8.32. The processes include:

- organic recycling through composting in specialised composting facilities
- mechanical recycling
- chemical recycling using hydrolysis or pyrolysis, undertaken using a zinc catalyst.

FIGURE 8.32 Organic, mechanical and chemical recycling options for PLA



Significant social and economic changes are required to develop more innovative approaches to reuse, repair and recycle plastics, as well as the instigation of effective and sustainable process design. Bioplastics made from raw materials from biomass are only a partial solution to the problem. In addition, the production of plastics must be free from hazardous materials to preserve the health of people and the environment. It seems the sustainable way of minimising plastic waste is to eliminate the use of plastics where possible. New products should be designed with a 'whole-of-life' approach, where product creation and disposal/recycling are considered before a product reaches the market. A move from a linear economy of 'take-make-dispose' to a circular economy is possible but requires a concerted effort of all communities, businesses, industries and governments to collect and effectively treat plastic waste, and to invest in innovative design technologies for new products and recycling.

8.5 Activities

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8.5 Quick quiz on | **8.5 Exercise** | **8.5 Exam questions**

8.5 Exercise

1. **a.** What is a bioplastic?
   b. Give three examples.
2. Explain the difference between degradable, biodegradable and compostable plastics.
3. State three advantages and three disadvantages of using plants to produce plastics.
4. **a.** What is polylactic acid (PLA)?
   b. List the steps in production of PLA.
   c. State three properties of PLA.

5. Compare and contrast the two bioplastics polylactic acid (PLA) and bioethene (bio-PE).
6. Identify four difficulties of applying the circular economy to plastics.
7. **a.** State the three general methods for recycling plastics at the end of their life.
   b. Evaluate if the three methods of treating plastics at the end of life are compatible with the idea of a circular economy.
8. Suggest four approaches to limiting the environmental impact of plastics in society that are consistent with the principles of the circular economy.
9. Suggest one way in which the following can assist in dealing with problems of plastic use.
   a. Households **b.** Business **c.** Industry **d.** Government
10. Describe an experiment to investigate the biodegradability of three types of plastic.

8.5 Exam questions

Question 1 (1 mark)

MC Which of the following polymers can be recycled?

A. LDPE and HDPE
B. Thermosetting polymers
C. Elastomers
D. Cornstarch polymers

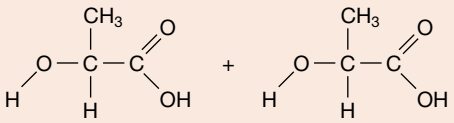
Question 2 (1 mark)

MC Which of the following statements regarding recycling is false?

A. Recycled plastic items, such as water bottles, are melted and then re-produced as the same type of object.
B. Recycling of plastics reduces the need to use declining oil reserves in the manufacture of plastics.
C. Recycling reduces the accumulation of plastics in landfill and in waterways.
D. Making products from recycled plastic reduces the amount of energy required to produce the object.

Question 3 (4 marks)

a. Explain whether PLA is an addition or a condensation polymer. **(2 marks)**
b. Demonstrate your answer using the monomers of lactic acid provided. **(2 marks)**



Question 4 (3 marks)

Victoria has a higher rate of plastic recycling (41 per cent) than other states in Australia. It also has a higher proportion of plastics manufacturers and recycling plants.

Annually, about 1.5×10^5 tonnes of plastic is recycled in Victoria.

a. Calculate the mass (in tonnes) of plastics consumed annually in Victoria. **(1 mark)**
b. If 69 per cent of all recycled plastics comes from packaging and 79 per cent of packaging is recycled, calculate the mass of packaging consumed per year. **(2 marks)**

Question 5 (6 marks)

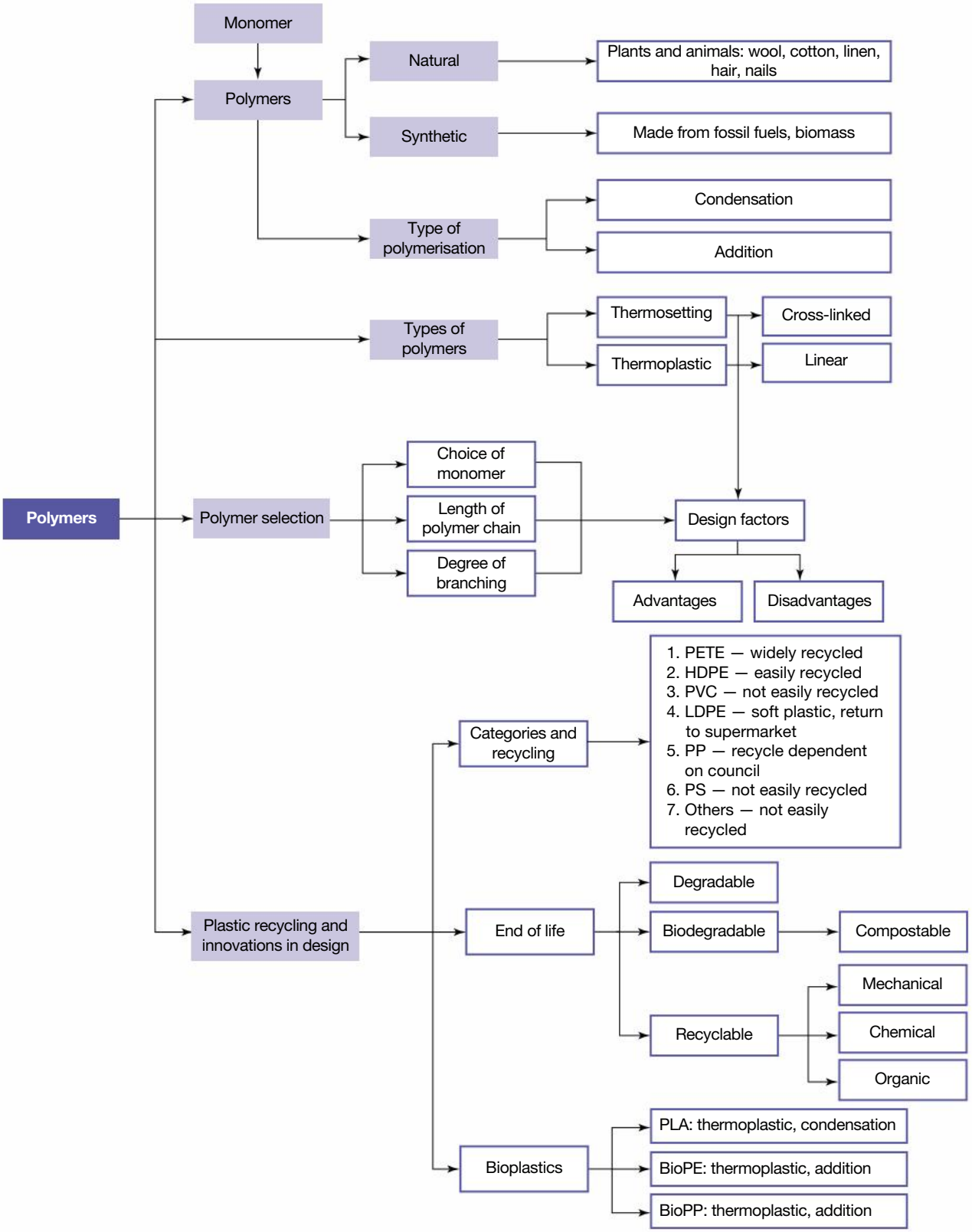
A water-soluble plastic is now available that is biodegradable and gradually breaks down leaving no toxic residue.

a. Suggest three possible uses of water-soluble plastics. **(3 marks)**
b. Describe two advantages of using water-soluble plastics. **(2 marks)**
c. Describe one disadvantage of using water-soluble plastics. **(1 mark)**

More exam questions are available in your learnON title.

8.6 Review

8.6.1 Topic summary



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8.6.2 Key ideas summary

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8.6.3 Key terms glossary

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Solutions	Solutions — Topic 8 (sol-0807)
Practical investigation eLogbook	Practical investigation logbook — Topic 8 (elog-1613)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 8 (doc-37081) Key ideas summary — Topic 8 (doc-37082)
Exam question booklet	Exam question booklet — Topic 8 (eqb-0089)

8.6 Activities

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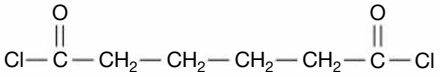
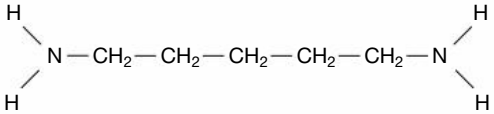
8.6 Review questions

1. List three advantages and three disadvantages of using polymers.
2. Which of the following molecules could possibly act as monomers in a polymerisation reaction? In each case, explain why or why not.
   a. C_4H_{10} **b.** C_3H_6 **c.** C_2H_6 **d.** C_4H_6 **e.** C_2H_4 **f.** C_2H_2 **g.** C_3H_8
3. Distinguish between thermosetting and thermoplastic polymers and give three examples of each.
4. In what ways is an elastomer similar to
   a. a thermoplastic substance?
   b. a thermosetting substance?

5. The two monomers provided form a polymer and another product.

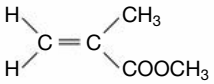


Monomer 1

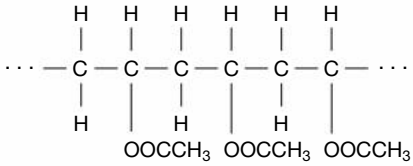


Monomer 2

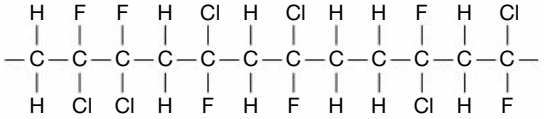
   a. Explain what type of polymer is formed.
   b. Name and write the formula of the other product formed.
6. LDPE is soft and flexible whereas HDPE is stronger and more rigid. Explain.
7. Explain why the recycling of plastics is desirable.
8. Describe why the coding system used for plastics is beneficial.
9. Perspex is an addition polymer that has the appearance of glass. It is made from the monomer provided.



   a. Draw part of the polymer of perspex.
   b. In what situations would the use of perspex be superior to the use of glass?
   c. In what situations would the use of glass be superior to the use of perspex?
10. A polymer used in the manufacture of wood glue is shown.



   a. Draw the structure of the monomer from which the polymer is derived.
   b. Describe the forces that would hold these chains together.
   c. Show the structure in the condensed (bracketed) form.
11. a. Draw the full structural formula of the unsaturated molecule represented by the semi-structural formula CHClCHF.
   b. Draw a section of the polymer produced when the monomer from part **a**. undergoes polymerisation.
   c. A section of an addition polymer is provided.



Draw the structural formula of the monomer that it is produced from.

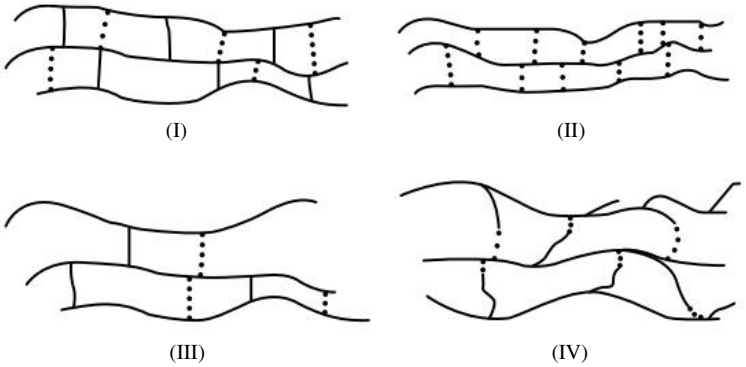
8.6 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC Which of the polymer segments shown are from thermoplastic polymers?

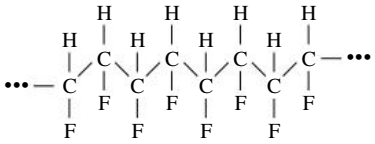


A. Polymer I only
B. Polymer II only
C. Polymers I and III only
D. Polymers II and IV only

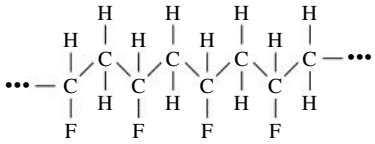
Question 2

MC Tetrafluoroethene is the monomer used to make polytetrafluoroethylene (PTFE). The following polymer segments show four monomers of the polymer. Which segment corresponds to PTFE?

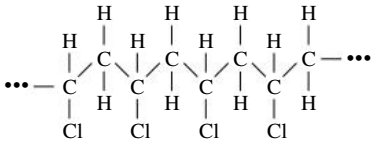
A.



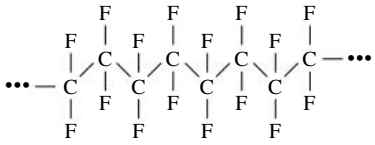
B.



C.



D.



Question 3

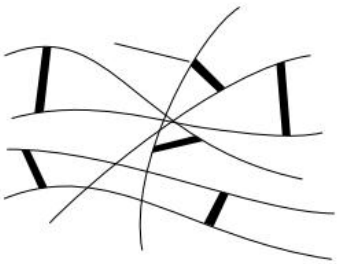
MC An elastomer is a linear polymer with what?

A. A large number of cross-links between its chains
B. A small number of cross-links between its chains
C. No cross-links between its chains
D. N cross-links but with tangled chains

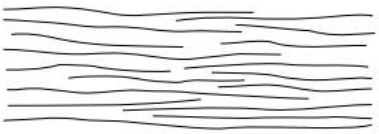
Question 4

MC Which of the following diagrams depicts a thermosetting polymer?

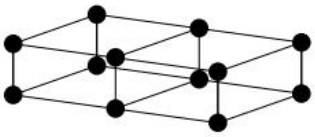
A.



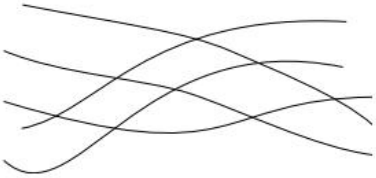
B.



C.



D.



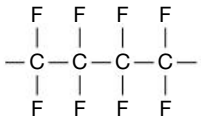
Question 5

MC Why are thermoplastic polymers more suited to recycling than thermosetting polymers?

A. They are lighter and easier to collect in roadside collection services.
B. Most have short-term uses and quickly become waste.
C. They have cross-links between their polymer chains that are easily broken by heat.
D. Their properties mean that they can be re-melted and remoulded into new articles.

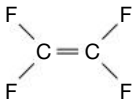
Question 6

MC Teflon is made from the polymer polytetrafluoroethylene (PTFE). It has the following structure.

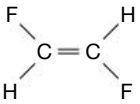


Which of the following is the monomer for this polymer?

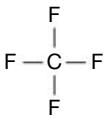
A.



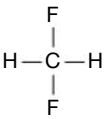
B.



C.

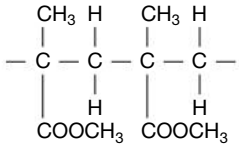


D.

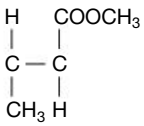


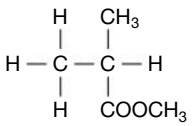
Question 7

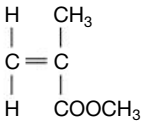
MC Polymerisation of methyl methacrylate produces perspex. The structure of perspex is as follows.

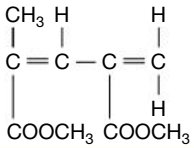


What is the structure of methyl methacrylate?

A. 

B. 

C. 

D. 

Question 8

MC What happens when the chain length of a polymer increases?

A. The melting temperature is lowered
B. The substance becomes more viscous
C. Solubility in non-polar solvents increases
D. Density decreases

Question 9

MC Which type of plastics are accepted at most recycling facilities?

A. 1 PET and 2 HDPE
B. 1 PET and 3 PVC
C. PVC and PLA
D. PLA and 1 PET

Question 10

MC What would applying the principles of the circular economy include?

A. Stopping plastic production
B. Eliminating hazardous materials
C. Ensuring landfills are situated away from towns
D. Making better use of fossil fuels

Section B — Short answer questions

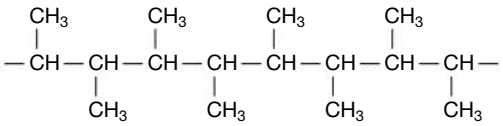
Question 11 (5 marks)

a. Draw the structural formulas of the following molecules.

i. But-2-ene

ii. But-1-ene **(2 marks)**

b. Examine the section of polymer that is shown.



From which of the substances in part **a**. could this polymer have been made? **(1 mark)**

c. Draw a section of the polymer chain that could be made from the other molecule (i.e. the one *not* used for part **b**). **(2 marks)**

Question 12 (8 marks)

High-density polyethene (HDPE) and polyvinyl chloride (PVC) are both formed with chains that have only a small amount of branching. They show a number of similarities, but also a number of differences.

a. Draw the structural formulas of the monomers from which each is made. **(2 marks)**

b. For samples of each in which the length of the polymer chains is about the same, what would be some of the factors that affect the difference in strength and hardness of these two polymers? Explain. **(2 marks)**

c. One of these polymers has plasticisers added to it to increase its versatility. Explain how the addition of plasticisers affects the properties of this polymer. **(2 marks)**

d. Polyethene can also be produced in a different form called low-density polyethene (LDPE). Describe the main difference in molecular structure between LDPE and HDPE and the effect that this has on the polymer's properties. **(2 marks)**

Question 13 (6 marks)

Polyethyne (polyacetylene) is a polymer that is made from the monomer ethyne. In 1977, it was discovered that it could be modified to become an electrically conductive polymer. This led to an explosion in research and possible applications. In 2000, the Nobel Prize in Chemistry was awarded for work in this field.

a. Draw the structural formula for a molecule of ethyne and use this to explain why it can be polymerised. **(2 marks)**

b. Draw a section of polyethyne that shows at least three monomer units. **(2 marks)**

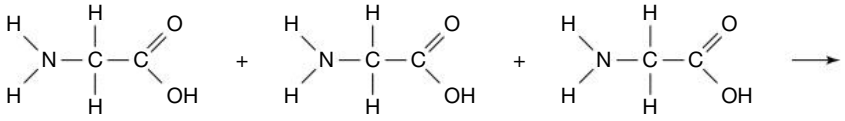
c. Show your answer from part **b**. using condensed (bracketed) notation. **(1 mark)**

d. How would you expect the reactivity of polyethyne to compare to polyethene? **(1 mark)**

Question 14 (4 marks)

Proteins are natural polymers that are an essential part of all living organisms. Glycine is the simplest of the 20 amino acids that the body uses to make proteins.

a. Draw a section of protein formed from the reaction of three glycine molecules. **(2 marks)**



b. State the type of reaction and explain your reason. **(2 marks)**

Question 15 (6 marks)

Draw a diagram showing how the principles of the circular economy might apply to the production, use and disposal of biopolyethene. State any reactions involved.

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AREA OF STUDY 2 How are materials quantified and classified?

OUTCOME 2

Calculate mole quantities, use systematic nomenclature to name organic compounds, explain how polymers can be designed for a purpose, and evaluate the consequences for human health and the environment of the production of organic materials and polymers.

PRACTICE EXAMINATION

STRUCTURE OF PRACTICE EXAMINATION		
Section	**Number of questions**	**Number of marks**
A	20	20
B	8	30
	Total	**50**

Duration: 50 minutes

Information:

- This practice examination consists of two parts; you must answer all question sections.
- Pens, pencils, highlighters, erasers, rulers and a scientific calculator are permitted.
- You may use the VCE Chemistry Data Book for this task.

on Resources

Weblink VCE Chemistry Data Book

SECTION A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

1. Analysis of an element using a mass spectrometer reveals that it contains a number of isotopes. The two most common are 15.331 and 15.498 times heavier than the ^{12}C isotope.
The relative isotopic masses of these two isotopes would be closest to what?

A. 153 and 155
B. 15.3 and 15.5
C. 184 and 186
D. 18.4 and 18.6

2. Element *X* exists in only two stable isotopes, ^{72}X and ^{74}X. If the lighter isotope makes up 70 per cent of all *X* atoms, what will the relative atomic mass of *X* be?

A. Less than 72
B. Greater than 74
C. Between 72 and 73
D. Between 73 and 74

3. What is the molar mass of $CuSO_4 \cdot 5H_2O$?

A. 159.6 g mol^{-1}

B. 249.6 g mol^{-1}

C. 177.6 g mol^{-1}

D. 127 g mol^{-1}

4. One mole of iron(III) nitrate nonahydrate, $Fe(NO_3)_3 \cdot 9H_2O$, contains

A. 3 mol of O atoms.

B. 9 mol of O atoms.

C. 10 mol of O atoms.

D. 18 mol of O atoms.

5. A 27.77 g sample of a mystery element contains 0.1460 mol atoms of the element. What is the element most likely to be?

A. Beryllium, Be **B.** Osmium, Os **C.** Niobium, Nb **D.** Rhenium, Re

6. Which of the following contains the greatest number of hydrogen atoms?

A. 10 g of hydrogen gas, H_2

B. 6.02×10^{23} molecules of glucose, $C_6H_{12}O_6$

C. 6.02×10^{24} methane, CH_4, molecules

D. 5 mol of ammonia, NH_3

7. Which of the following compounds contains the highest percentage of nitrogen by mass?

A. $CO(NH_2)_2$ **B.** CH_3NH_2 **C.** $CH_2(NH_2)_2$ **D.** NH_4NO_3

8. A sample of a compound containing only carbon, hydrogen and chlorine was analysed and found to contain 0.15 mol of carbon, 0.2 mol of hydrogen and 0.1 mol of chlorine. Its molar mass was found to be 222 g mol^{-1}. What is the molecular formula of this compound?

A. $C_{1.5}H_2Cl$ **B.** C_3H_4Cl **C.** $C_3H_4Cl_2$ **D.** $C_6H_8Cl_4$

9. Hydrocarbons can be classified as alkanes, alkenes or alkynes. Which of the following statements is true?

A. Due to the presence of multiple bonds, alkenes and alkynes tend to be more polar than alkanes.

B. Alkenes have a larger proportion of hydrogen than alkanes.

C. Alkynes are the only compounds that have just triple bonds between each carbon.

D. Alkanes, alkenes and alkynes have only dispersion forces for their intermolecular attraction.

10. Which of the following statements about the organic molecule 3, 3-dimethylpent-1-ene are true?

I. It is an unsaturated molecule.

II. It has a total of seven carbon atoms.

III. It has a different molecular formula from 3-ethylpent-1-ene.

A. Only statements (I) and (II) are true.

B. Only statements (I) and (III) are true.

C. Only statement (III) is true.

D. Only statement (I) is true.

11. What is the number of hydrogen atoms present in one molecule of propanoic acid?

A. 4 **B.** 6 **C.** 8 **D.** 10

12. Which of the following is expected to have the *lowest* boiling point?

A. CH_3OH **B.** CH_3CH_3 **C.** $CH_3CH_2CH_3$ **D.** CH_3NH_2

13. What is the correct systematic (IUPAC) name of $CH_3CH_2CH(CH_3)CH_2CH_2Cl$?

A. 1-chloro-3-methylpentane

B. 3-methyl-6-pentachloride

C. 1-chloro-2-methylpentane

D. 1-chloro-4-ethylbutane

14. What is the number of different structural isomers that have the formula C_5H_{12}?

A. 5 **B.** 4 **C.** 3 **D.** 2

15. Which of the following will affect the properties of a polymer?

A. The monomer(s) used
B. The degree of branching in the polymer chains produced
C. The length of the polymer chains produced
D. All of the above

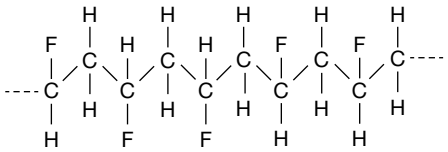
16. Which of the following lists contains only plastics that are based on fossil fuels?

A. HDPE, PP, PS, PVC
B. PVC, PLA, LDPE, PS
C. PLA, HDPE, LDPE, PP
D. PP, PS, PVC, PLA

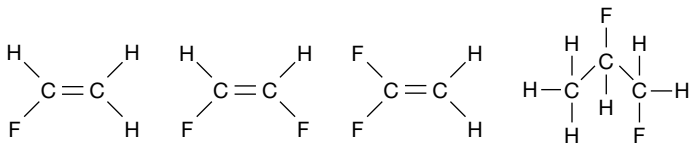
17. A polymer made by the condensation mechanism will always what?

A. Be the only product produced in the reaction
B. Produce a thermoplastic polymer
C. Involve monomers containing carbon–carbon double bonds
D. Produce a by-product in the reaction

18. Consider the following polymer segment.



Which monomer would be used for producing this polymer?



Molecule A Molecule B Molecule C Molecule D

A. Molecule A **B.** Molecule B **C.** Molecule C **D.** Molecule D

19. Which of the following statements about thermosetting and thermoplastic polymers is correctly classified?

	Thermosetting	Thermoplastic
A	Contains extensive cross links between polymer chains	Forms rigid materials that cannot be softened on melting
B	Chars when moderately heated	Only dispersion forces operate between polymer chains
C	Only dispersion forces operate between polymer chains	Polymer chains are kept separated by the addition of plasticisers
D	Forms rigid materials that cannot be softened on melting	Unable to be recycled

A. Option A **B.** Option B **C.** Option C **D.** Option D

20. Which of the following statements about bioplastics is true?

A. Their monomers come from biomass and they are always biodegradable.

B. Their monomers come from fossil fuels and they are sometimes biodegradable.

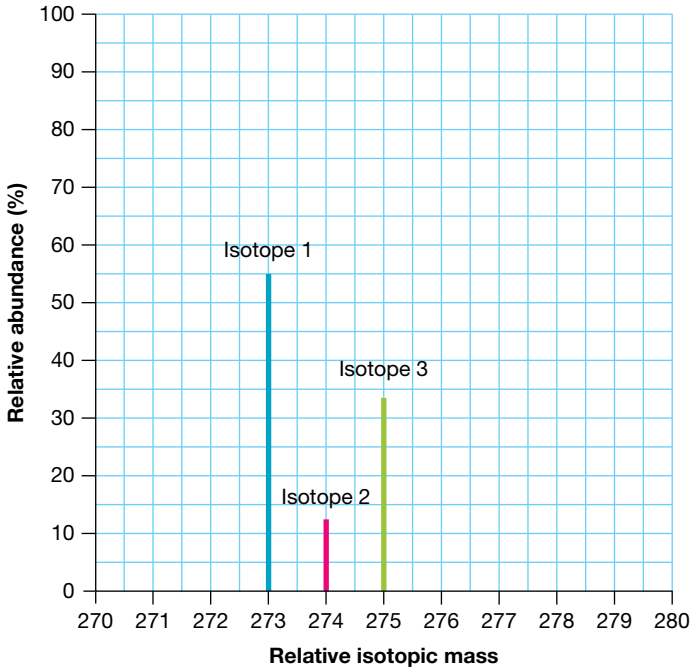
C. Their monomers come from biomass but they may not be biodegradable.

D. Their monomers come from fossil fuels and they are not biodegradable.

SECTION B — Short answer questions

Question 21 (3 marks)

The mass spectrum of a rare element is shown in the following figure.



a. Use the data in the mass spectrum to complete the following table. **(2 marks)**

	Isotope 1	Isotope 2	Isotope 3
Relative isotopic mass			
Relative abundance (%)			

b. Using the data, calculate the relative atomic mass, A_r, of the element. **(1 mark)**

Question 22 (5 marks)

Ethanoic acid, CH_3COOH, is a naturally occurring compound found in vinegar.

a. In 20.6 grams of ethanoic acid, calculate the following.

i. The amount in moles of ethanoic acid **(1 mark)**

ii. The amount of carbon atoms in moles **(1 mark)**

iii. The number of hydrogen atoms **(2 marks)**

b. What mass in mg of ethanoic acid would be present in 5.0×10^{-4} moles of the compound? **(1 mark)**

Question 23 (4 marks)

A compound containing carbon, hydrogen and chlorine is analysed and found to contain 2.640 g of carbon, 0.661 g of hydrogen and 7.809 g of chlorine. Its molar mass is 101 g mol^{-1}.

Calculate its molecular formula.

Question 24 (3 marks)

Write balanced equations for the following reactions. You may use structural formulae, condensed formulae or molecular formulae in your equations. Symbols of state are not required.

a. The reaction of propene with chlorine gas **(1 mark)**

b. The reaction between ethane gas and chlorine gas under UV light **(1 mark)**

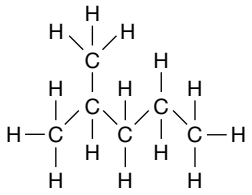
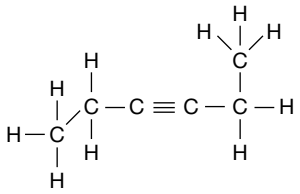
c. The complete combustion of butane **(1 mark)**

Question 25 (4 marks)

Draw and systematically name two *structural isomers* for a pair of alcohols that have the molecular formula C_3H_8O.

Question 26 (4 marks)

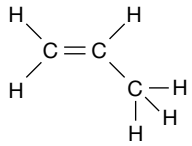
a. Complete the following table, providing either the systematic (IUPAC) name or the structural formula as required for the compounds. **(3 marks)**

Structural formula	Systematic name
	
	
	Methanoic acid

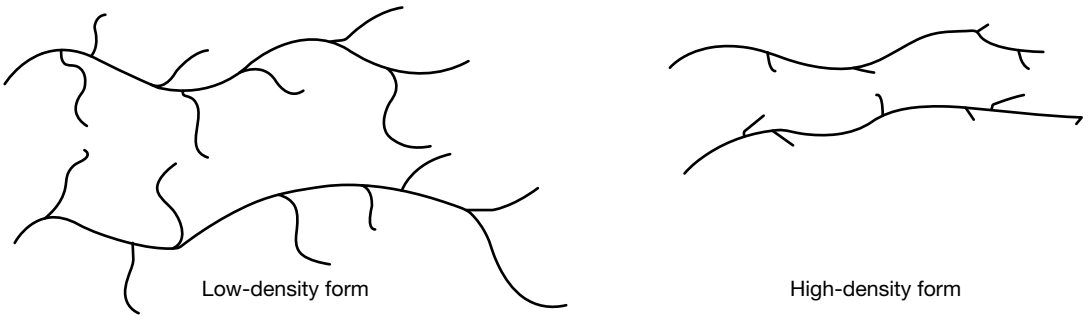
b. Write the semi-structural (condensed) formula for ethylbutanoate. **(1 mark)**

Question 27 (3 marks)

a. Draw a segment of polymer (at least three units) formed by addition polymerisation of the monomer shown. **(1 mark)**



b. Depending on the reaction conditions, this polymer may be formed in a low-density form or a high-density form, as illustrated in the following diagram.



Apart from their different densities, describe two other ways that their physical properties would differ, making reference to the bonding involved. **(2 marks)**

Question 28 (4 marks)

Rayon is a fibre made from reconstituted cellulose. Unfortunately, in the past, producing rayon has required the use of certain chemicals that have serious health and environmental concerns. Because of this, a new method involving less dangerous chemicals has been invented, and promising research has emerged into producing cellulose from sunlight, water and carbon dioxide.

a. Choose two raw materials from the following list from which rayon could be made.
corn, recycled plastic, paper, sawdust, recycled cotton fabric, recycled aluminium **(2 marks)**

b. Give two benefits if the production of cellulose from sunlight, water and carbon dioxide is proven to be both possible and economic. **(2 marks)**

PRACTICE SCHOOL-ASSESSED COURSEWORK

ASSESSMENT TASK — A media analysis/response

In this task, you will have one week to complete the preliminary investigation, create blog posts and complete a timed, written response to an essay question and stimulus.

- Part A is a sequence of preparatory responses or the creation of a blog.
- Part B is the final written response. This will be written on lined paper provided to students. Students are permitted their blog notes, textbook summaries, pens and pencils during this time.

Time allowed for part A: one full week.

Time allowed for part B: one full lesson.

Total marks: 26 marks

Introduction

Most people receive their news about science through the popular media, rather than through academic research papers. You will read scientific articles and consider any social, health, environmental benefits and concerns related to polymers and plastics.

You will demonstrate your understanding of the key knowledge and skills in Outcome 2 and synthesise your blog notes and textbook summaries to write a response to the essay question and stimulus.

Task

Part A: The following table provides five stimulus topics and questions, which will help you with your response to the essay question given in part B. The stimulus material provided is not exclusive, and you may do additional research. You have one full week to investigate each topic and related questions, and complete your preparatory or blog notes.

Topic	Prompting questions and tasks	Resource(s)
Environmental impact of plastics.	• What are the environmental implications of plastic waste? • How can society be made aware of the environmental impact caused by plastic waste disposal?	'Plastic trash in the ocean is a global problem, and the US is the top source — a new report urges action' https://theconversation.com/plastic-trash-in-the-ocean-is-a-global-problem-and-the-us-is-the-top-source-a-new-report-urges-action-172848
Future plastics	• Identify and describe the synthetic future polymers in these articles. • Why is society interested in these synthetic future polymers? • Explain any challenges and issues with their usage.	'Five synthetic materials with the power to change the world' https://theconversation.com/five-synthetic-materials-with-the-power-to-change-the-world-37131 'New lightweight material is stronger than steel' https://news.mit.edu/2022/polymer-lightweight-material-2d-0202
Circular versus linear economy	• What is the difference between a linear economy approach and circular economy approach? • What are the advantages and disadvantages of linear and circular economies?	'There's a genius, sustainable economic system we could all be using. Here's how' https://www.sciencealert.com/can-we-build-a-circular-economy-to-save-the-planet
Bioplastic, PHA	• Why are PHAs beneficial to society? • One of the monomers used to make PHA is 3-hydroxybutanoic acid. Draw and label the functional groups present in 3-hydroxybutanoic acid. • What type of reaction (addition or condensation) describes the formation of PHA from 3-hydroxybutanoic acid? Write a balanced equation and explain your answer. • Referring to the PHA formed from 3-hydroxybutanoic acid, identify the products which form after PHA is hydrolysed.	'PHA: promising, versatile, biodegradable' https://www.biobasedpress.eu/2016/08/pha-promising-versatile-biodegradable/
Waste control technology	• What is carbon capture? Why is carbon capture important to society? • Explain how polymer technology could be used as a carbon capture method. • Explain the benefits of and issues when using waste-eating enzymes.	'New polymer membrane tech improves carbon capture efficiency' https://scitechdaily.com/new-polymer-membrane-tech-improves-carbon-capture-efficiency/ 'Plastic-eating enzyme could eliminate billions of tons of landfill waste' https://news.utexas.edu/2022/04/27/plastic-eating-enzyme-could-eliminate-billions-of-tons-of-landfill-waste/

Part B: Stimulus

Australia uses around 70 billion pieces of soft plastics, including food wrappers, each year.

- *Australians used 3.4 million tonnes of plastics in 2018–2019.*
- *One million tonnes of Australia's annual plastic consumption is single-use plastic; 84% of plastic is sent to landfill and only 13% is recycled.*
- *Our use of plastic is increasing and across the world will double by 2040.*
- *By 2050, it is estimated that plastic in the oceans will outweigh fish*

Source: *National Plastics Plan summary 2021*, Department of Climate Change, Energy, the Environment and Water (https://www.awe.gov.au/environment/protection/waste/publications/national-plastics-plan-summary).

Part B: Essay question

With reference to polyethene and polyhydroxyalkanoates, how can society move away from a linear economy approach and towards a circular economy approach?

In your response to this question, refer to:

- environmental, social and health implications
- chemically balanced equations
- carbon capture technology.

Resources

Digital document U1AOS2 School-assessed coursework (doc-38065)

9 Research investigations

KEY KNOWLEDGE

online only

This investigation has three sections to the key knowledge — scientific evidence, sustainability and scientific communication — each of which you must consider and include in your final report. Each of the points is intertwined with the key scientific skills outlined for the research investigation.

Scientific evidence

- the distinction between primary and secondary data
- the nature of evidence and information: distinction between opinion, anecdote and evidence; and scientific and non-scientific ideas
- the quality of evidence, including validity and authority of data and sources of possible errors or bias
- methods of organising, analysing and evaluating secondary data
- the use of a logbook to authenticate collated data.

Sustainability

- sustainability concepts and principles: green chemistry principles, sustainable development, and the transition from a linear economy towards a circular economy
- identification of sustainability concepts and principles relevant to the selected research question.

Scientific communication

- chemical concepts specific to the investigation: definitions of key terms; and use of appropriate chemical terminology, conventions and representations
- characteristics of effective science communication: accuracy of chemical information; clarity of explanation of chemical concepts, ideas and models; contextual clarity with reference to importance and implications of findings; conciseness and coherence; and appropriateness for purpose and audience
- the use of data representations, models and theories in organising and explaining observed phenomena and chemical concepts, and their limitations
- the influence of social, economic, legal and/or political factors relevant to the selected research question
- conventions for referencing and acknowledging sources of information.

KEY SCIENCE SKILLS

- plan and conduct investigations
- construct evidence-based arguments and draw conclusions
- analyse, evaluate and communicate scientific ideas.

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9.1 Overview

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9.1.1 Introduction

One of the key skills of a scientist is the ability to research — that is, to question and investigate a specific aspect of science, such as 'Which is the cheapest plastic to recycle?' You might already know the answer to this question because you have read about the topic through other scientists' research investigations published in the media, or from communication in textbooks and/or the internet.

FIGURE 9.1 Effective research involves careful analysis of information and evidence.

In Unit 1 and Unit 2 AOS 3, you will drive your own investigation from concept to communication. For Unit 1 AOS 3, you will investigate and explain how chemical knowledge is used to create a more sustainable future in relation to the production or use of a selected material. You will think critically and creatively, and apply science inquiry skills to communicate the relevant chemical concepts associated with your investigation. You will also critically examine the information and data available to answer your research question, and identify the sociocultural, economic, political, legal and ethical implications of the selected investigation in terms of sustainability.

LEARNING SEQUENCE

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9.1 Overview
9.2 Investigating how chemistry can create a more sustainable future
9.3 Scientific evidence, and analysing and evaluating sources
9.4 Models and theories to understand observed phenomena
9.5 Effective science communication
9.6 Review

on Resources

Solutions	Solutions — Topic 9 (sol-0808)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 9 (doc-37083) Key ideas summary — Topic 9 (doc-37084)
Exam question booklet	Exam question booklet — Topic 9 (eqb-0090)

9.2 Investigating how chemistry can create a more sustainable future

KEY KNOWLEDGE

- Sustainability concepts and principles: green chemistry principles, sustainable development, and the transition from a linear economy towards a circular economy
- Identification of sustainability concepts and principles relevant to the selected research question

KEY SCIENCE SKILL

- Plan and undertake investigations

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9.2.1 Sustainability concepts and principles

As discussed in topic 1, sustainability can be defined as satisfying our current needs without risking the needs of the next generations. For example, bringing reusable bags to the supermarket is more sustainable than buying plastic or paper bags every time. The Brundtland Report produced by the United Nations expressed the need for a sustainable society given the finite resources on Earth and the need to protect the environment and ecosystems.

In chemistry, we think of sustainability in terms of three perspectives:

- sustainable development
- green chemistry principles
- the move from a linear economy towards a circular economy.

Sustainable development

Sustainable development involves devising ways of continuing or renewing a material; for example, aluminium cans and foil are completely recyclable using a melting process where the molten aluminium can be reused. In 2015, the UN Sustainable Development Summit of 193 countries agreed to the 2030 Agenda for Sustainable Development, and clearly outlined 17 Sustainable Development Goals (SDGs).

FIGURE 9.2 The United Nation's 17 Sustainable Development Goals

Source: United Nations, 2015.

The SDGs most relevant to VCE Chemistry are:

- Goal 2 — Zero hunger
- Goal 6 — Clean water and sanitation
- Goal 7 — Affordable and clean energy
- Goal 9 — Industry, innovation and infrastructure
- Goal 11 — Sustainable cities and communities
- Goal 12 — Responsible consumption and production
- Goal 13 — Climate action
- Goal 14 — Life below water
- Goal 15 — Life on land.

Chemistry plays a significant role in sustainable development, with current research focusing on innovative processes to develop new materials such as biodegradable plastics and to reduce our use of natural resources, while being energy efficient. The practices of sustainability have been carried out by First Nations peoples of Australia for thousands of years. They exemplify a collective responsibility and respect for the land and sea through their traditional practices; for example, the management of the mangrove habitats in the coastal areas.

Green chemistry principles

Chemical processes are involved in our everyday lives and products that we use. As our natural resources are being depleted over time and society demands safer products that are produced sustainably, a global shift has occurred towards sustainable materials. This change has caused the chemical industry to look at alternative manufacturing processes to balance society's needs with the impact on the environment and human health.

As discussed in topic 1, Paul Anastas and John Warner developed the 12 Principles of Green Chemistry in 1991. Seven of the principles are included as part of your study in unit 1:

- *1 — Waste prevention:* rather than deal with waste, it is favourable to prevent waste.
- *2 — Atom economy:* use all reactants to maximise efficiency of process.
- *4 — Designing safer chemicals:* utilise chemicals that are non-toxic and safe.
- *6 — Design for energy efficiency:* manufacturing occurs with environmental and economic impact significantly limited with energy efficiency optimised.
- *7 — Use of renewable feedstocks:* reduce fossil fuels in favour of renewable (plant-based) materials.
- *9 — Catalysis:* catalysts utilised in pathway that is energy-efficient with reduced waste
- *10 — Design for degradation:* products when no longer required are able to biodegrade without toxicity

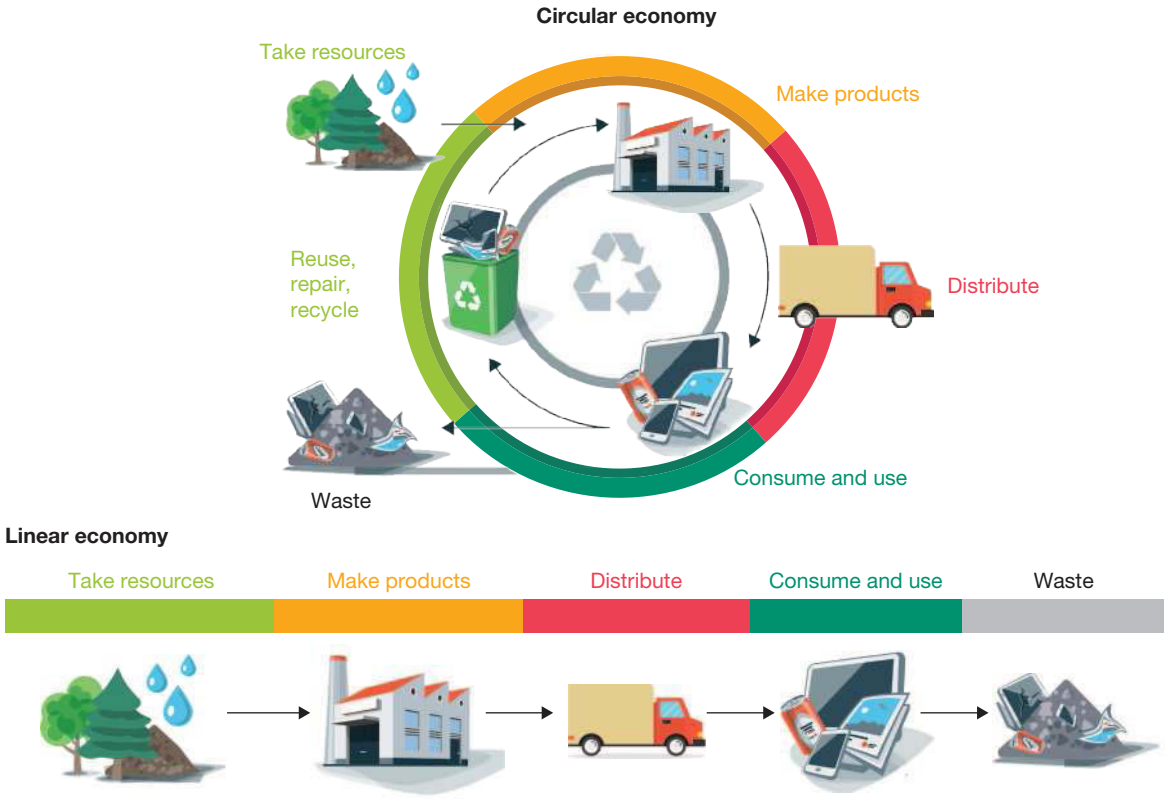
Linear and circular economies

A **linear economy** is the process by which a material is made from start to finish by a manufacturer and where the material is not reclaimed to reuse; for example, non-recyclable thermoset plastics. In the interests of sustainable development for future generations, the **circular economy** is now becoming the adopted model for manufacturing. The circular nature allows for a product to be repaired, repurposed, recycled and remade to be used again, which permits a continuous sustainable cycle and minimises any waste. For example, artificial turf that has been worn down can now be separated into its rubber granules, sand and plastic fibres and then reused or recycled. The linear economy is often referred to as 'take, make, dispose', whereas the circular economy is referred to as 'remake, reuse, recycle'. This is illustrated in figure 9.3.

linear economy raw materials are converted into products that are used and then discarded as waste

circular economy economy in which waste and pollution is eliminated, products at the end of life are reused, repaired or recycled, and the environment is not harmed

FIGURE 9.3 The circular economy versus the linear economy



9.2.2 Why do we conduct research investigations?

A research investigation is an exploration of a scientific question or questions, and the presentation of findings. In Outcome 3, you have the opportunity to use your developing research skills to investigate a recent discovery, innovation, advance, case study, issue or challenge related to sustainability.

A variety of scientific investigation methodologies can be used depending on the aim of the investigation and the question under investigation. For this investigation, you will not be conducting a practical investigation (controlled experiment or fieldwork) to generate your own primary data. You will probably complete a **case study** or a **literature review**. This form of investigation involves considerable planning, and expertise in working scientifically to understand a method, review results, analyse findings and communicate conclusions. Your investigation will involve researching, evaluating and interpreting information, and presenting it in a cohesive, accurate and detailed report.

Researching allows us to:

- gain a comprehensive understanding of a topic and research question
- discover current and prior investigations, including findings regarding a topic
- make informed decisions
- investigate patterns and relationships
- understand differences between strong evidence that has been peer reviewed and weaker evidence that shows bias
- understand ways to conduct research, including applying the research and the development of new scientific understanding.

case study an investigation of a particular event or problem, involving the discussion of knowledge learned from a situation

literature review using and analysing secondary evidence and data from a group of scientists research experiments

9.2.3 Choosing a topic and developing a question

For Unit 1 Outcome 3, the investigation is based on the following overarching question: 'How can chemical principles be applied to create a more sustainable future?'. This is then subdivided into four research investigation topics (see figure 9.4).

FIGURE 9.4 For Unit 1 Outcome 3, you can choose from four topics for your research investigation.

Investigate a topic that interests you. You may wish to consider one of the questions from any investigation topic provided in the options in the VCAA Chemistry study design. Alternatively, in conjunction with your teacher, you may develop your own research question that relates to Unit 1 Area of Study 1 and/or Area of Study 2. Typical questions begin with how, why, what, when, who or where. Record the question(s) that you are going to investigate. These questions may change as you start researching.

9.2.4 Identification of sustainability concepts and principles relevant to the selected research question

When you have selected your research question, you need to consider the related sustainability concept and principles. The following prompts will help you to identify the sustainable concepts and principles of your topic and assist in answering your research question.

Sustainable development: Does your material or process meet our current needs without putting at risk future generations?

Is the process or material you're investigating sustainable? Consider energy usage, source of raw materials, possible reduction in environmental damage, and economically viability or cost savings. For example, the energy to produce 'green steel' comes from renewable electricity and hydrogen instead of coal-dependent electricity.

Green chemistry: Does the process or material include one of the 12 Principles of Green Chemistry?

For example, the ninth principle is prevention. Does the process you're investigating eliminate all waste, or does it produce waste but turn the waste into a reusable product?

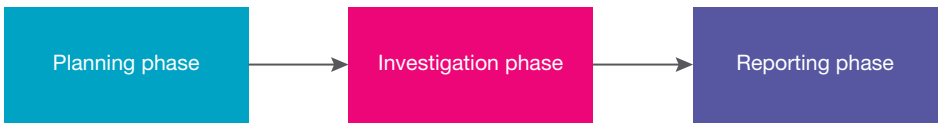
Circular economy: Does the process or material follow a circular or linear economy?

For example, is the product's life cycle from start to finish a 'take, make, dispose' model, or does it follow the sustainable circular economy model of 'remake, reuse recycle'?

9.2.5 Planning and conducting a research investigation

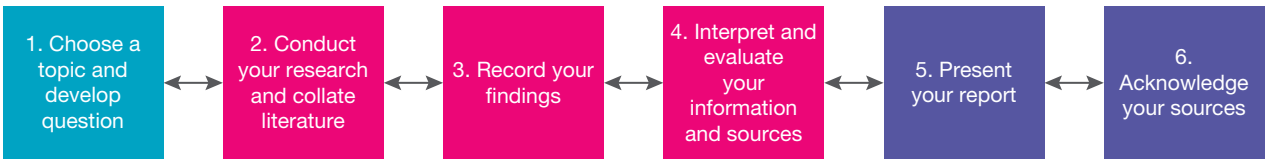
Once your key questions have been decided, it is time to commence your research using the abundance of available sources. When planning and conducting your research investigation, following a structure helps you make sure your research is valuable and appropriate to the questions, and allows you to apply critical thinking, inquiry and communication skills. You will use both your prior knowledge from topics you have covered throughout the year, as well as gather new ideas and information to help draw conclusions. Your investigation will be made up of three main phases, as outlined in figure 9.5.

FIGURE 9.5 Phases of research investigations



These three phases can be further subdivided into six steps that will help you plan your investigation and your report. It is important to note that you may move back and forward between these steps as your research investigation develops.

FIGURE 9.6 Planning and presenting your research investigation



Throughout the planning and presenting of a research investigation, it is important to consider:

- the aim and type of research question
- the target audience results are being communicated to
- the format of communication
- if research is required outside of class time
- time allocated in class
- how work will be checked and authenticated
- referral to your rubric and/or success criteria to check all requirements have been attempted.

Resources

Weblink Circular economy

9.2 Activities

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9.2 Quick quiz on | **9.2 Exercise** | **9.2 Exam questions**

9.2 Exercise

1. Explain, using an example, why the practices of First Nations peoples of Australia are sustainable.
2. How could you increase the sustainability of your school, and monitor the effectiveness of your sustainability program?
3. Sustainability is often referred to as having three interconnecting spheres:
   - social
   - environmental
   - economic.

   Explain how these align with the three perspectives of sustainability in chemistry.
4. Victoria is now transforming to a circular economy, and many sustainability initiatives are emerging from Victorians who have developed their own circular economy.

   Research innovative circular economies (you could start by using the **Circular economy** and **Circular story** weblinks in the Resources panel). Can you think of how you could create your own circular economy?
5. Why is it important to conduct research investigations?

9.2 Exam questions

Question 1 (1 mark)

MC Biopolymers are an example of which of the 12 principles of green chemistry?

A. #5 — Safer solvents
B. #6 — Design for energy efficiency
C. #9 — Use of catalysis
D. #10 — Design for degradation

Question 2 (1 mark)

MC Using bio-based feedstocks is an example of which UN Sustainable Development Goal?

A. Goal 2 — Zero hunger
B. Goal 6 — Affordable and clean energy
C. Goal 12 — Responsible consumption and production
D. Goal 14 — Life below land

Question 3 (1 mark)

In Kenya's Nyanza Province, a safe-water and hygiene program supplied clay pots to 45 public primary schools for safe water storage. The schools were also given a year's supply of water disinfectant, 200-litre plastic water tanks with taps for hand-washing, and soap. Two teachers from each school received educational materials. These covered water treatment, how to safely store water and good classroom hand-washing practices. The teachers formed safe-water clubs with students, who were encouraged to share the information with their parents.

Source: Adapted from: https://www.unicef.org/wash/schools/files/Raising_Even_More_Clean_Hands_Web_17_October_2012(1).pdf

Identify the SDGs being addressed in the above program.

Question 4 (6 marks)

Hydrogen gas can be synthesised via several methods, including reacting methane with steam and by reacting with a type of heated coal (coke) and steam.

The reactions to produce hydrogen gas via these two methods are as follows.

Methane and steam: $CH_4(g) + H_2O(g) \rightarrow 3H_2(g) + CO(g)$

Coke and steam: $C(s) + H_2O(g) \rightarrow H_2(g) + CO(g)$

a. Calculate the atom economy for the production of hydrogen by reacting methane and steam. **(2 marks)**
b. In terms of atom economy, determine which of these processes for the production of hydrogen gas is greener. **(2 marks)**
c. Methane gas can be produced by anaerobic fermentation of organic waste. Based on this information, which is a greener choice of reactant — methane or coal? **(1 mark)**
d. What other principle of green chemistry does this relate to? **(1 mark)**

More exam questions are available in your learnON title.

9.3 Scientific evidence, and analysing and evaluating sources

KEY KNOWLEDGE

- The distinction between primary and secondary data
- The nature of evidence and information: distinction between opinion, anecdote and evidence; and scientific and non-scientific ideas
- The quality of evidence, including validity and authority of data and sources of possible errors or bias
- Methods of organising, analysing and evaluating secondary data
- The influence of social, economic, legal and/or political factors relevant to the selected research question
- The use of a logbook to authenticate collated data

KEY SCIENCE SKILLS

- Plan and undertake investigations
- Construct evidence-based arguments and draw conclusions
- Analyse, evaluate and communicate scientific ideas

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9.3.1 The distinction between primary and secondary data

Data is a set of facts that are collected, observed or generated. Data can be from primary sources or from secondary sources. **Primary data** is data that is collected by a researcher from first-hand sources; examples include surveys, measurements or experiments. **Secondary data** is analysis of data from other people's investigations and is the readily available form of data collected from various sources, such as government censuses and publications, internal records of organisations, reports, books, journal articles and websites.

FIGURE 9.7 Finding appropriate information for your investigation takes time and research.

Secondary data offers several advantages, including that it is easily available, and saves time and cost for the researcher. However, some disadvantages are associated with this; for example, because the data is gathered for purposes other than the problem in mind, its usefulness may be limited in terms of relevance and/or accuracy.

Lots of valuable information is out there, and part of research is deciphering the right information for your investigation and identifying information that may be irrelevant, biased or misleading.

9.3.2 The quality of evidence

Scientific and non-scientific ideas

For your investigation, distinguishing between scientific and non-scientific ideas is important. Think about the components of scientific reporting: clear evidence; acknowledgement of limitations, uncertainties and errors; a concise investigation including data and an analysis of results; and conclusions relating to a research question.

primary data direct or firsthand evidence about some phenomenon

secondary data researchers' comments on or summaries and interpretations of primary data

Many sources, especially on the internet, are not scientific in nature and do not have a high level of validity. They do not follow a scientific method; rather, they are based on intuition or personal experiences and opinions. It is important to gain an understanding of the **reliability** and authority of sources you are using for your secondary data.

reliability the consistency of a measurement across multiple trials

correlation measure of a relationship between two or more variables

causation when one factor or variable directly influences the results of another factor or variable

randomised describes assignment of individuals to an experiment or control group at random, not influenced by external factors

control group a group that is not affected by the independent variable and is used as a baseline for comparison

controlled variables variable that is kept constant across different experimental groups

Strong and weak evidence

In your research investigation, you need to be able to recognise the difference between strong and weak evidence in order to gain the best possible information for your report.

Research and investigations that have strong evidence include the following features:

- a basis in facts derived from studies with high validity and minimal bias
- statistical evidence to support conclusions
- a clear distinction between **correlation** and **causation** — two variables may often have some correlation (they both increase, for example) but no causation (one variable does not cause the change in value of the other)
- data from investigations that have a repeatable and reproducible method — which include those that are **randomised** and have a **control group** or **controlled variables**
- peer-reviewed research formed from scientific ideas.

These features are summarised in figure 9.8.

FIGURE 9.8 Strong evidence involves many facets.

Weak evidence usually cannot be supported by facts and data. It may be more subject to bias and misinterpretations, may not clearly address the question being asked or may draw conclusions that are inappropriate to the research and data obtained. The contrast between strong and weak evidence is summarised in table 9.1.

TABLE 9.1 Contrasting the features of strong and weak evidence

Strong evidence	Weak evidence
• Is based on facts derived from studies with high internal validity • Is related to the topic • Provides support for stated conclusions • Is logically organised • Is free of bias • Comes from a source that is trustworthy and reputable	• Is based on personal opinion or anecdote • Is irrelevant or unrelated to the topic • Provides little support for stated conclusions • Is not logically organised • May contain bias • Comes from a source that is unknown or untested

Distinguishing between opinion, anecdote and evidence

As part of any scientific methodology (such as a case study, controlled experiment, fieldwork or correlational study), it is important that you can construct evidence-based arguments with clarity.

As part of your investigation, you will conduct preliminary research to provide background knowledge for your investigation and to justify the conclusions you are constructing.

When assessing research, it is important to understand whether it is an opinion, an anecdote or evidence.

- Often **opinions** are not based on evidence, but rather on an individual's personal beliefs. The strength of evidence can depend on who is giving an opinion. The opinion of an expert in a field is often stronger than the opinion of an individual unaware of the subject matter. However, opinions are much more likely to be biased and skewed in a certain direction compared to anecdotes and strong evidence. For example, an opinion of an individual may be, 'I don't believe mercury is harmful.'
- **Anecdotes** are often based on personal experience and are a brief account of an incident that occurred to a person. Again, these can be biased, because storytelling can lead to stretching the truth or incorrect recollections. Because they are based on personal experience, they also do not represent a large sample size. For example, an anecdote related to the preceding opinion may be, 'I remember when they used mercury in thermometers.'
- **Evidence** is a much stronger way to help draw conclusions, relying on facts and data as a basis of research. In all investigations, evidence rather than opinion or anecdotes should be used both in how you communicate your findings and in the research you conduct to construct conclusions. For example, evidence related to mercury may be, 'In 2013, governments agreed to the Minamata Convention on Mercury. The convention obliges government parties to take a range of actions, including to address mercury emissions into the air and to phase out certain mercury-containing products due to its toxicity.'

Sources of error and bias

Many sources of error and bias exist, and these need to be taken in to consideration when conducting research. No experiment or investigation is perfect; **human error**, **random error** or **systematic errors** can affect the collection and interpretation of data. Errors may include uncalibrated machines being used (leading to results that are subject to systematic errors). Therefore, it is important to not have all your research based on one piece of information, but rather allow for a broader sample to reduce the effect of errors that may have occurred. Errors are further investigated in subtopic 17.6.

Bias also has a great impact on research, in which results are influenced by external factors. Bias may occur for a variety of reasons. Researchers may have a conflict of interest in the investigation, in which specific results are desired. This may be for financial gain, career aspirations or personal reasons. It is important to ensure that minimal bias exists in investigations. Bias is also further examined in subtopic 17.5.

opinions statements based on an individual's personal belief and experience

anecdote an individual's story based on personal experience rather than strong evidence

evidence reliable and valid data used to support or refute a hypothesis, model or theory

human error mistakes made by individuals conducting an investigation, such as mistakes in entering data or the use of the wrong type of chemical in a reaction

random errors chance variations in measurements; result in a spread of readings

systematic errors errors that affect the accuracy of a measurement that cannot be improved by repeating an experiment; usually due to equipment or system errors

Some examples of bias are outlined in table 9.2.

TABLE 9.2 Types of bias

Type of bias	Outcome	Example
Measurement	Manipulating results and measurements to get a desired outcome	Changing the results of individuals on a placebo versus an experimental drug
Selection	Not randomly assigning individuals	Deliberately assigning family members or those with a milder form of an illness to the test group
Sampling	Subjects not being representative of the population	Selecting only individuals under 25 to test a new drug designed for women of all ages
Response	Only certain individuals choosing to be part of a trial, which may not be representative of everyone	Only individuals in a high socioeconomic area choosing to take part

Reliability and authority of sources

Finding reliable sources with a high level of authority is vital to allow for the most valid conclusions to be drawn in your investigation. The following sources typically provide reliable and authoritative sources of data:

- Official statistics are statistics collected by governments and their various government departments. Statistics can be useful because they are easily obtainable and comprehensive sources of information that usually cover long periods of time.
- Scientific reports are accounts of work done on research projects. They are written to provide a summary of the results collected to research institutions, governments and other interested researchers. A report may originate from completed research or ongoing research projects.
- Scholarly journals contain original research or experimentation written by experts in specific fields. Articles in scholarly journals usually undergo a peer review, which involves other experts in the same field reviewing the content of the article for accuracy, originality and relevance.
- Literature review articles assemble and review original research dealing with a specific topic. Reviews are usually written by experts in the field and may be the first written overview of a topic area. Review articles discuss and list all the relevant publications from which the information is derived.
- Science magazines contain articles that discuss practical information concerning various fields. These magazines provide people in these fields with information pertaining to that field or trade.
- Reference books provide secondary source material. In many cases, specific facts or a summary of a topic is all that is included. Textbooks, student activity books and dictionaries are all considered reference books.

Web searches through Google Scholar are an excellent way to search for reliable data and research, because this search engine only searches through academic literature and includes information from academic publishers, universities and government websites.

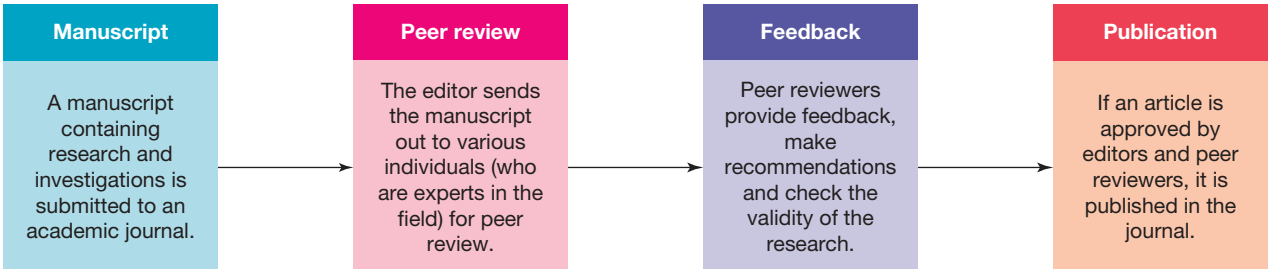
Peer-reviewed work

A large amount of information is available to assist with your research investigation.

For stronger evidence, it is important to use work that has been peer reviewed. In science, this is work that has been reviewed for quality of research, experimental reproducibility, accuracy and validity, and adheres to the required standards of a journal. Peer-reviewed work is checked for bias before publication is submitted, usually by a board of reviewers or independent experts who have no conflict of interest with the work being examined but are experts in the specific field of research. The process for peer reviewing is outlined in figure 9.9.

Peer-reviewed work provides stronger evidence because it has been accepted to be scientifically valid, to be based on facts and to have drawn appropriate conclusions.

FIGURE 9.9 Process for articles to be peer reviewed



Unfortunately, while peer-reviewed articles are useful, they are not perfect. Usually, during reviewing, recommendations are made for required improvements and further explorations. In the case of articles in which in-depth experimental work has been completed, it is often not possible for a reviewer to replicate this work to test its validity and, therefore, it can be hard to determine the accuracy of the outlined conclusions. It is good practice to find multiple, recent peer-reviewed journals that have investigated similar ideas.

Research articles may also cite other peer-reviewed articles, which may be useful as part of your research investigation.

Many journals are now moving to an open-access format, which allows free access to everyone.

Some examples of peer-reviewed chemistry journals include:

- *Chemical Reviews*
- *Nature Reviews Chemistry*
- *Chemical Society Reviews*
- *Nature Materials*
- *Nature Chemistry*
- *Accounts of Chemical Research*
- *Trends in Chemistry*
- *Journal of the American Chemical Society*
- *Nature Communications* (open access)
- *ACS Central Science* (open access)
- *Open Chemistry* (open access)
- *RSC Advances* (open access)

FIGURE 9.10 Some examples of journals with an open-access format

9.3.3 Analysing and evaluating secondary data

The key to analysing and evaluating secondary data is the ability to judge the quality of the data or information that has been gathered. The following tips will help you assess the quality of the data.

Determine the original purpose and author of the data collection

It is important to consider the purpose of the data or publication and the authority of the source of the data.

Consider the following:

- Is it a government document or statistic?
- Was the data collected for corporate and/or marketing purposes?
- Was the data collected by a research institution? Were the scientists undertaking research funded by the government or a commercial business?

It is also important to identify the author's or source's credentials — including educational background, past publications and experience — in this area. For example, reliable sources of data and information include research reports that document findings from chemical research published by peer-reviewed publications.

Understanding the author and purpose of the data collected will help to guide you to evaluate the quality of the data, and discern the potential level of bias and conflicts of interest. Many journal articles and some popular media articles now include a disclosure statement addressing any conflicts of interest.

Methodology

Does the article being reviewed have a 'Methods' section that discusses the methods used to conduct the research? If it does not, you should look for additional supporting information or data. However, if the research methods are discussed, review them to ascertain the quality of the study. Methods sections are important tools to determine whether the results are reproducible — that is, the experiment shows reliability.

Date of publication

When was the article published? Is the source current (last five years) or older? These are important considerations, particularly if your research topic is in an area of continuous and rapid development, such as green chemistry or the circular economy, where older articles can quickly become outdated.

Reference list

All peer-reviewed journal articles will include a comprehensive reference list of all sources quoted in the study, including name, date and journal of publication. If you are using a source other than a peer-reviewed article, it should still include its sources of data. If not, the credibility of the data is not considered reliable. Without appropriate references to sources of information, it is impossible to judge the quality and validity of the information reported.

Primary or secondary source

Primary sources are the raw material of the research process; they represent the data as first described. Secondary sources are based on primary sources. These sources analyse, describe and synthesise the primary or original data. Caution needs to be exercised if the article uses secondary data to ensure that it accurately relates information from primary source(s).

Identifying the source

When using a website for information, a clue to the source — and potentially the reliability — of the website can be found in the generic top-level domain name. These are the abbreviations found after the 'dot' in a website address. Table 9.3 provides some examples of common generic top-level domain names. Remember that a website address ending in '.au' means that it is registered in Australia, but some non-governmental Australian websites are registered in the United States and will not include the '.au'.

It is important to search not just the internet, but also podcasts, journals, videos and books. When using online resources, make use of an advanced search to find a key phrase, recent document or a particular file type such as a PowerPoint (.pptx) or PDF (.pdf).

Recognising error and bias

If you are using a peer-reviewed article as a secondary source, it is likely that any potential errors or biases were corrected during the peer-review and manuscript development phases before publication. However, you should always evaluate your source material for error and bias.

TABLE 9.3 Examples of generic top-level domain names

Domain name contains	Source
.gov.au	Australian federal, state, territory and local government entities; cannot be edited by the general public
.gov	United States government entities; cannot be edited by the general public
.edu.au	Limited to Australian educational institutions, including schools and universities; cannot be edited by the general public
.edu	Largely limited to American higher educational institutions (universities and colleges; not schools); cannot be edited by the general public
.org	Intended for use by not-for-profit organisations; some webpages can be edited by the public (for example, Wikipedia.org)
.com.au	Intended for commercial use, but now used within Australia across a range of sectors including businesses, not-for-profit organisations, schools and individual people
.com	Intended for commercial use, but now used within the United States across a range of sectors including businesses, not-for-profit organisations, schools and individual people
.net	Intended for networking technology organisations such as internet service providers, but now used across a range of sectors including businesses, not-for-profit organisations, schools and individual people

Other considerations

You may also wish to consider the intended audience of your source. Is the publication aimed at a specialised or a general audience? If the source is too mainstream or aimed at the public, it may not provide the necessary level of detail required for your investigation.

Does the report substantiate other materials, differ to the materials previously reviewed, or has the report provided you with new information on the topic area?

9.3.4 The influence of social, economic, ethical and other factors

Many factors can have a great impact on the information obtained and the presentation of your research investigation, other than just scientific factors. It is important to incorporate these factors in to your investigation and understand the ways they affect findings and results.

Social factors

In your investigation, you need to have an understanding of factors that underpin the development, use and modification of chemicals and materials. One of the largest aspects to consider is **social factors**. These factors shape the behaviours and thoughts of individuals. These include the historical, cultural and political aspects of a population. Understanding how new discoveries or changes to current discoveries affect both individuals and society as a whole is vital.

social factors factors that influence both a whole society and the individuals within the society

When creating new chemicals and materials, researchers need to understand society in order to enhance marketability of the product. They need to make sure that the product will have long- and short-term benefits for individuals. This may be done using surveys, focus groups and opinion polls.

Social factors also affect ideas such as education of individuals and their understanding and interpretation of information being presented to them. When researching, it is vital to gain an understanding of how society will be involved with the use of the product, because this in turn will affect the supply and demand of a new development.

FIGURE 9.11 Social factors influence all research and development involving humans.

An example of this may be in the creation of a new material that can be used as a contraceptive. This would affect societies in various ways — some individuals would be opposed to the use of contraception for cultural or religious reasons, or political interference may prevent the new material being imported into their country. It may be that the contraceptive was designed to better protect against HIV, which is prevalent in different levels in various populations, so the product may have very different demands based on societal requirements. Understanding social factors is vital to understanding the viability of the manufacturing and altering of chemicals or materials, because the behaviours and thoughts of society can be a huge barrier.

Economic factors

Economic factors are those related to the economy and the market. These factors relate to the cost of not only conducting investigations but also acquiring talent, identifying, applying and developing chemicals, and analysing scientific information to draw conclusions.

Research is very reliant on economic factors, in the form of grants to conduct investigations. It is important to be aware of the economic situation of where you are getting your information from, because research may be incomplete or rushed due to the loss of funding.

It is also important to consider **bias** in research due to financial incentives and funding from external companies. Sometimes, data is skewed or misrepresented due to financial reasons — it may be that research was funded by a source that wanted data represented in a way that is favourable for their intentions. This greatly affects the validity of data collected and research conducted. When looking at research, it is important to check that no conflict of interest exists that may lead to bias.

Economic factors often can be linked to social factors and classed as **socioeconomic**. An example of a socioeconomic factor in chemistry research is the access to different research based on cost (some academic journals are quite expensive to download, while others are free). This can affect the availability of information to different socioeconomic groups.

Finally, in investigating your topic, it is vital to understand the economic viability of the development of new materials and chemicals — will they provide significant benefit to others to an extent that they are profitable?

economic factors factors that are related to the financial implications of an action or decision; these may be on an individual level or on a larger community or national level

bias the intentional or unintentional influence on a research investigation

socioeconomic the interaction of social and economic factors

Legal factors

Legal factors are directly related to the law, and what is deemed to be a legal or an illegal action rather than being right or wrong. The law provides a scaffold through which the environment can be protected from pollution and waste. A number of both federal and state and territory Acts help provide a framework to support protection of the environment.

FIGURE 9.12 The law should provide equal treatment for all, including in environmental and chemical issues.

Federal and state Acts include the:

- *Environment Protection and Biodiversity Conservation Act 1999*
- *Environment Protection Act 2017*
- *Pollution of Waters by Oils and Noxious Substances Act 1986*
- *National Environment Protection Council (Victoria) Act 1995*.

These laws are essential — when synthesising and developing new chemicals and materials, the impact on the environment must be considered. This is evident in our search for new sources of fuel that are renewable, and in the development of new materials that are biodegradable and do not build up in the environment and cause harm. However, sometimes environmental issues are unintentional, in which chemicals or materials designed for beneficial purpose are found to have drastic effects in the long-term because they are retained in the environment. This is important to consider in law, because often long-term effects are difficult to study and so can be neglected.

Legal factors also include the need for the law to be enforced through health and safety codes, professional bodies, licensing boards, government committees, and sets of regulations, rules and laws. Such regulations, agreements and codes of conduct enhance public trust in the profession through transparent expectations and behaviour.

Political factors

Political factors are those related to the actions of government. Political factors and government decisions influence each other in a variety of ways. The Australian political system allows citizens to exercise freedom of speech and to think freely, and encourages public engagement and opinion. Collectively, these ideas allow for the development of shared or socially acceptable moral standards, which in turn inform the development of laws and government regulations. These standards include not only the rights of citizens, but also the limitations and responsibilities placed on them in any given situation. This protection and regulation of rights and responsibilities of citizens includes matters directly related to the chemical industry. For example, when we look at climate change and the corresponding politics to reduce carbon emissions, we can see the protection of the environment is competing with political factors such as potential job losses and removal of certain industries. While the political system can protect the rights of people, unfortunately politics is not separate from science; both are intertwined. Climate change is managed by individual governments, and the provision of a fair and equitable reduction in emissions is extremely difficult due to the economic factors of each country, which in turn then become a political issue, where the actions of governments are influenced by upcoming, or the outcomes of previous, elections.

legal factors factors that are related to the legality of an action

political factors factors that are related to the actions of government

Ethical factors

When conducting research, ethics are vital to consider; that is, the morals around what actions are correct or incorrect. In the development of new materials and chemicals, and the adaptations of current materials and chemicals, it is important to take into consideration ethical implications of both the results and the method of the research being conducted. This includes the process of testing and checking that informed consent and confidentiality were provided to those involved in the study. It is important that the risks involved are made exceptionally clear to individuals taking part in studies.

FIGURE 9.13 The impact of chemicals on the environment must be considered at all stages, from manufacture to disposal.

Ethics are also involved in the data and research presented. Authors of research may choose to select the best information to include in their data, leading to a misrepresentation of the information and conclusions drawn. It is important to be able to carefully analyse the research you are investigating for bias (both economic and ethical) that may lead to a lack of validity of results. Research that has been conducted unethically should not be used for your investigation.

An example of this was in the MMR (mumps, measles and rubella) vaccine and autism controversy, in which Andrew Wakefield published a paper arguing that a link existed between the MMR vaccination and autism. In this paper (since redacted), data was falsified to show that the MMR vaccination caused autism.

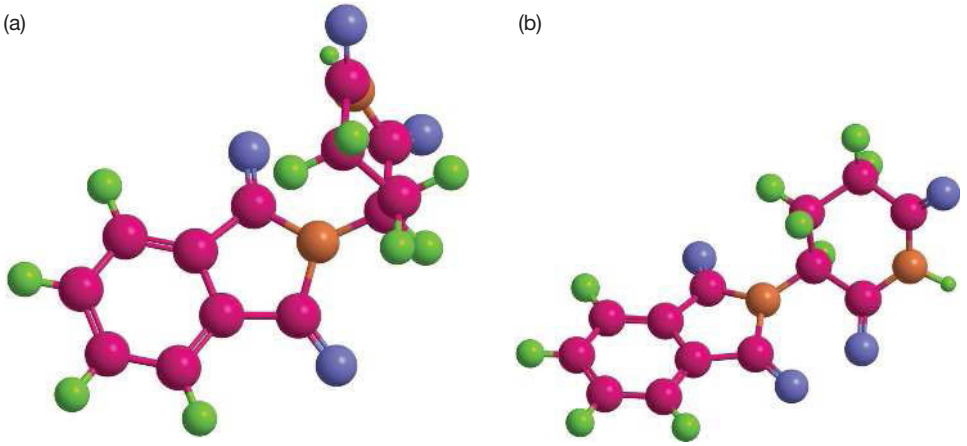
A significant bias existed in the individuals chosen to examine, along with issues relating to the limited sample size (only 12 individuals were used in the sample). However, as a result of this paper, vaccination rates began to drop. Despite the paper being redacted and many elements shown to be false, the impact of this unethical paper is still felt today. This shows how important it is to carefully read not just the conclusions drawn in a paper, but also the way the data was collected and research carried out.

Use of the chemical thalidomide demonstrates very serious ethical issues in research. In the 1950s, thalidomide was used to treat morning sickness in tens of thousands of pregnant women. Soon after it was introduced, it was found to cause significant birth defects in babies whose mothers took thalidomide during certain stages of pregnancy (see figure 9.14).

FIGURE 9.14 An X-ray showing a child whose mother took thalidomide during key stages of the child's development

This tragedy highlights the very serious ethical concerns of releasing an inadequately tested drug to the general public. The impact of this drug completely changed the ways new drugs are tested, particularly the ethical considerations in drug trials. It has inadvertently created another ethical concern; because of this disaster, more animal testing is now conducted.

FIGURE 9.15 Thalidomide exists in two mirror image forms: (a) the R-isomer and (b) S-isomer. One has sedative actions and the other causes birth defects. However, in the body, the two isomers interconvert.

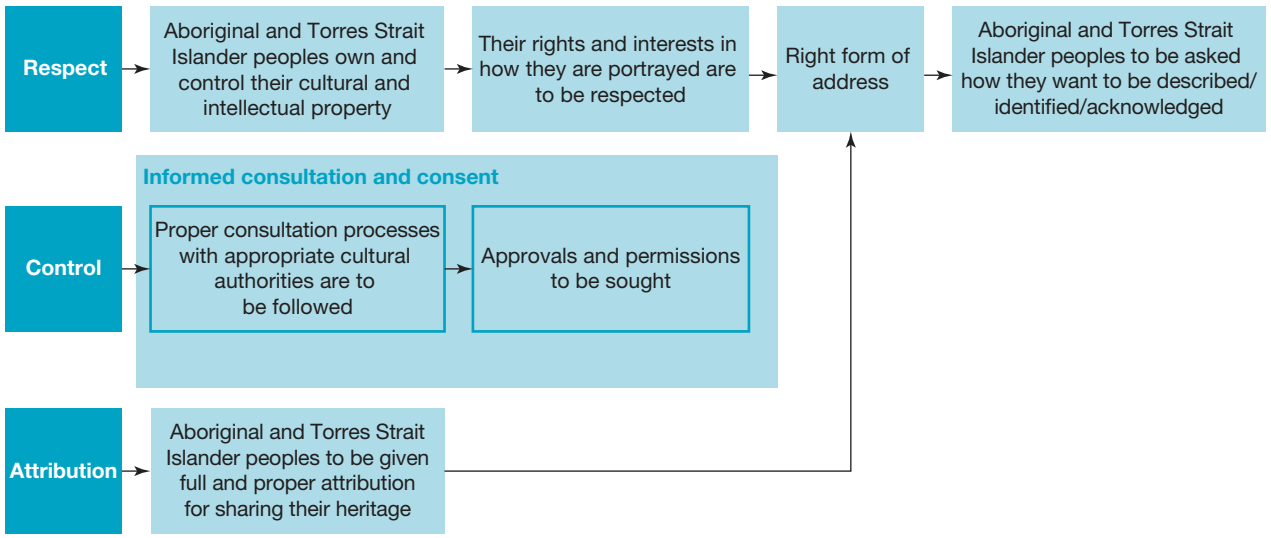


9.3.5 Aboriginal and Torres Strait Islander cultural protocols

If your investigation is focused on Aboriginal and Torres Strait Islander peoples' traditional knowledge, techniques or artefacts, you will need to apply appropriate cultural protocols to access information. These are ethical principles guiding your behaviour in a particular situation. They are designed to protect Aboriginal and Torres Strait Islander peoples' intellectual property rights, and help you respect Aboriginal and Torres Strait Islander cultural beliefs and practices.

Be mindful, however, that no set rules are in place for interacting with Aboriginal and Torres Strait Islander peoples. Some of the values and principles of these protocols are outlined in figure 9.16. You can use these as general guidelines.

FIGURE 9.16 Framework to protect the cultural and intellectual property rights of Aboriginal and Torres Strait Islander Peoples



EXTENSION: The right to control the use of traditional knowledge

The **United Nations Declaration on the Rights of Indigenous Peoples** was adopted by the General Assembly in 2007. (At the time, Australia was among only four countries that voted against it; however, in 2009, the Australian Government endorsed the declaration.)

Article 31 of the declaration states that:

1. Indigenous peoples have the right to maintain, control, protect and develop their cultural heritage, traditional knowledge and traditional cultural expressions, as well as the manifestations of their sciences, technologies and cultures, including human and genetic resources, seeds, medicines, knowledge of the properties of fauna and flora, oral traditions, literatures, designs, sports and traditional games and visual and performing arts. They also have the right to maintain, control, protect and develop their intellectual property over such cultural heritage, traditional knowledge, and traditional cultural expressions.
2. In conjunction with indigenous peoples, States shall take effective measures to recognize and protect the exercise of these rights.

Source: Article 31 of the UN Declaration on the Rights of Indigenous Peoples

on Resources

Weblink Victorian Aboriginal Education Association

9.3.6 Using a logbook to authenticate collated secondary data

As part of your scientific investigation (as well as all practical experiments throughout the year), you are required to keep a logbook. Usually this logbook is a bound exercise book (however, your teacher may request a digital logbook instead). It is vital to show all aspects of your investigation within your logbook using the scientific approach. This logbook will be assessed by your teacher. You must date all work in your logbook to show when it was completed and to assist in validating your work.

Authenticating secondary data

In researching your investigation, it is important to record any secondary data sourced and used within the investigation, because this will then need to be used within your written report.

It is important to organise and categorise the information obtained while researching, because this will enable you to locate the information easily if required when writing your final report. You may wish to organise each piece of your research under specific headings in a table, as shown in table 9.4.

TABLE 9.4 Organising your research

Logbook headings	Notes
Title of research	
Date accessed	
Key findings (in your own words)	
Issues to consider (social, legal, economical, biological)	
Evidence (source type)	
Reference (for bibliography)	

Remember to note any gaps in the information or the presence of competing opinions. You can organise the information using highlighters to collate different groups of ideas or sections. Write a plan showing the flow of the report and where each piece of information will be presented.

United Nations Declaration on the Rights of Indigenous Peoples a universal framework of minimum standards for the survival, dignity and wellbeing of the Indigenous peoples of the world

9.3.7 Interpreting your information and drawing conclusions

As part of your research, it is important to analyse and interpret the information you have — both information that is prior knowledge, and information that you have researched. This process should involve you analysing and assessing data, exploring trends and patterns, making generalisations and drawing conclusions. Be sure to make connections between the various sources of your information and clearly link research back to the question you are investigating. Interpret your findings — what do they suggest about the topic you are exploring? Do they provide answers to your question of investigation? Once you have recorded information from a few sources, read your notes carefully and write out information as if you were explaining it to another person, using your own words and sentence structure. You may also consider using a mind map to better understand the relationships between different pieces of evidence. It is important that you understand the content and make links between different lines of evidence. This will help you come to a conclusion. When drawing conclusions, you should summarise and tie together the findings of your report and reflect back on your introductory paragraph. This might provide the opportunity to suggest areas for further investigation and reflect on your learning, but it should not include any new information.

9.3 Activities

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9.3 Quick quiz on | 9.3 Exercise | 9.3 Exam questions

9.3 Exercise

1. Describe three aspects found in strong evidence as opposed to weak evidence.
2. Describe the difference between scientific and non-scientific ideas.
3. Why is it useful to use journal articles that are peer reviewed?
4. Compare and contrast anecdotes and opinions.
5. Describe the difference between legal and political factors.

9.3 Exam questions

Question 1 (1 mark)

MC Which of the following is the most reliable source of secondary data?

A. A subject-matter expert speaking on a breakfast television program
B. An internet search with many results that report the same data
C. A peer-reviewed article in an academic journal
D. An article in a popular science magazine

Question 2 (1 mark)

MC Websites containing which of the following top-level domain names would provide the most reliable sources of information?

A. .net
B. .org
C. .org.au
D. .gov.au

Question 3 (2 marks)

Explain why some webpages now include a disclaimer statement addressing potential conflicts of interest.

Question 4 (4 marks)

a. Identify one advantage of primary data. **(1 mark)**
b. Identify one advantage of secondary data. **(1 mark)**
c. Suggest two ways to ensure that any secondary data used provides strong evidence for your investigation. **(2 marks)**

More exam questions are available in your learnON title.

9.4 Models and theories to understand observed phenomena

KEY KNOWLEDGE

- The use of data representations, models and theories in organising and explaining observed phenomena and chemical concepts, and their limitations

KEY SCIENCE SKILL

- Analyse, evaluate and communicate scientific ideas

9.4.1 Models

It is important to be able to use models and theories to allow us to understand a variety of observed phenomena.

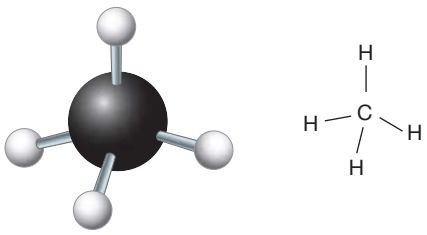
Models are representations of ideas, phenomena or scientific processes. They can be physical models, mathematical models or conceptual models. Models can provide an explanatory framework that explains observed phenomena, and can help with the understanding of abstract concept. In chemistry, many concepts are hard to visualise, so models help contextualise the idea on a smaller and simpler scale.

model representation of ideas, phenomena or scientific processes; can be physical models, mathematical models or conceptual models

While they are very useful, models have their limitations:

- Models cannot include all the details of the processes or the things that they represent because of the complexity of the processes.
- Models do have some limits in their accuracy and are often simplified and stylised. For example, a ball-and-stick model of methane (a gas involved in the greenhouse effect — see figure 9.17) is useful, but is a very highly simplified and stylised representation that reduces covalent bonds to sticks and atoms to solid balls. How do you think this differs from a 'real' methane molecule?

FIGURE 9.17 Different models of methane



- Models are approximations of the real world; for example, we can model various lattices (as seen in figure 9.18) or bonding, but we often assume ideal conditions around temperature and pressure, which is not realistic about what happens in the real world.
- Models are based on current observations and knowledge at the time. This means that they aren't definite and can change as new observations and experiments allow different ideas to evolve. Models that we currently believe to be accurate will be disproven in the future.

Models are constantly being refined as new information becomes available. One model that has continuously changed is that of the atom (see figure 9.19). These changes are described further in table 9.5.

FIGURE 9.18 Model showing the lattice of diamond

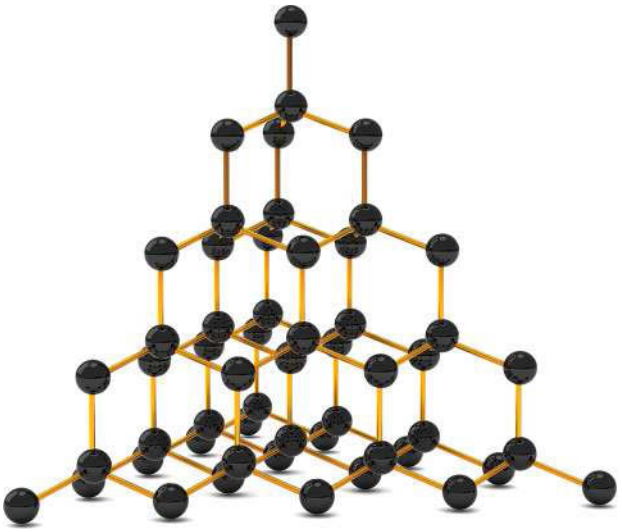
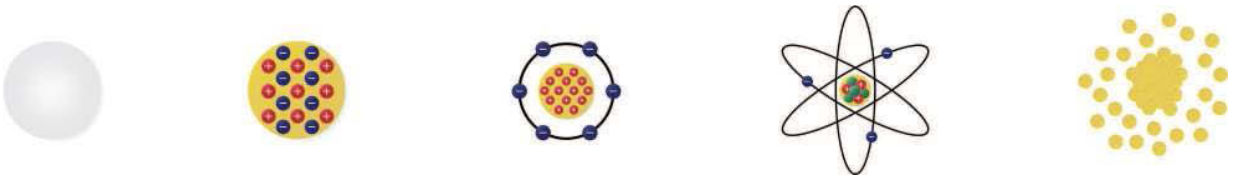


FIGURE 9.19 Various models of the atom, including John Dalton's model and the 'plum pudding' model by J.J. Thomson



1803
John Dalton

1904
J.J. Thomson

1911
Ernest Rutherford

1913
Niels Bohr

1926
Erwin Schrödinger

TABLE 9.5 Models are challenged and adjusted

Model	What was the model?	What observations challenged this?
Solid sphere model (John Dalton)	Atoms from different elements are different to each other. Atoms are indivisible and not made of smaller components.	The finding that atoms are made of subatomic particles known as protons, neutrons and electrons.
Plum pudding model (J.J. Thomson)	Negatively charged electrons float around in a large section of positive charge.	The gold foil experiment: when positively charged particles were fired at a very thin layer of gold, only some deflected back, showing the atom is mostly empty space.
Nuclear model (Ernest Rutherford)	The atom is mostly empty space, with a positively charged nucleus in the centre.	Moving electrons: the model didn't explain their orbit around the nucleus.
Planetary model (Niels Bohr)	Electrons move around the nucleus in orbits of fixed size and energy.	Heavier atoms: these would collapse with moving electrons and energy.
Quantum model (Erwin Schrödinger)	Electrons do not move in set orbits, but instead move in waves, or clouds known as orbitals.	Currently supported model

As seen in table 9.5, new observations and experiments often lead to the development and changing of previously supported models. In the future, more discoveries could come to light that change models that we currently believe to be accurate. This shows how important it is to make sure that any research you conduct is current and relevant.

9.4.2 Theories

A **theory** is a well-supported explanation of a phenomenon. It is based on the interpretation of facts that have been obtained through investigations, research and observations.

Limitations can exist with theories. These may include:

- A reliance on theories rather than observations made during practical investigations. Often, individuals manipulate results to match a theory, failing to realise the power of observations. Often, theories are treated as perfect when, in fact, they are able to be disproven as further observations and evidence come to light.
- Theories often rely on a very specific set of conditions to be met. For example, the Law of Conservation of Mass cannot be used in systems that are open systems (not isolated) or systems that involve very large amounts of energy.
- Often, a phenomenon needs to be described using multiple theories. Sometimes, aspects of theories may have some contradictions to each other.

Theories can change overnight or take a very long time to change. Theories that were once popular and well accepted may be discarded when too much evidence builds up against them. They are replaced by a theory that better fits the observations.

Some examples of theories that have been superseded by other theories are:

- miasma theory of disease (the idea that diseases were caused by bad air)
- the theory that the sun and planets orbit around the Earth (geocentrism)
- the flat Earth theory.

In the geocentric model of the Earth, the idea that the planets and sun rotated around the Earth was the prominent theory for a long time, supported by many philosophers and statements in the Bible. The heliocentric theory was proposed quite early in history, but only gained traction when a mathematical model of this system was showed by Nicolaus Copernicus. Individuals such as Johannes Kepler developed this further with the idea that orbits were elliptical and not spherical. These ideas were further shown and developed by Galileo, through the use of a telescope.

The flat Earth theory argued that the Earth was a plane or a disc. Hundreds of years ago, this was thought to be accurate. The idea that Earth was a sphere was suggested by Pythagoras, and later evidenced by Aristotle. This theory then replaced the flat Earth theory.

theory a well-supported explanation of a phenomena, based on facts that have been obtained through investigations, research and observations

FIGURE 9.20 The flat Earth theory has been replaced with the idea that the Earth is spherical.

How do refinements of existing theory come about?

A scientific theory is based upon substantial evidence but, of equal importance, it must be *tested repeatedly for validity*. Before an idea is tested, it is not a theory in the scientific sense, but a **hypothesis**. A hypothesis is a tentative, testable and falsifiable statement that predicts the relationship between two variables or predicts the outcome of an investigation. It provides a possible explanation that can then be tested experimentally or, if not possible, evaluated through reasoning.

After gathering data, testing, and getting consistent and irrefutable results, a hypothesis can *eventually* become a theory; however, most hypotheses do not. As new results or evidence arise, many hypotheses are discarded when found to be incorrect, and new ones take their place. It is important to remember that a scientific theory cannot be proven to be true, but can be proven to be false.

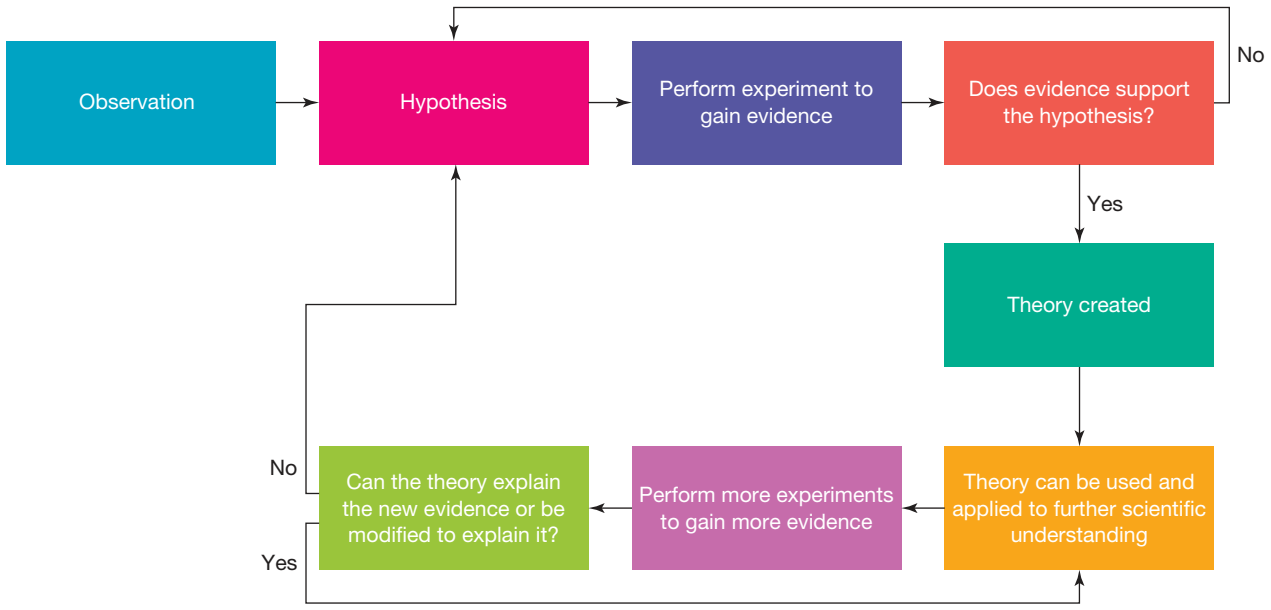
How a hypothesis becomes a theory

The establishment of a theory is an extremely rigorous process in science. Because theories describe the nature of facts, they must undergo extensive testing and reflect the most accurate and up-to-date knowledge in the field.

All theories start with an observation. This is then developed into a hypothesis to explain what was observed. This hypothesis then undergoes many rounds of testing, with no result disproving the hypothesis. This does not mean that *new* facts and information cannot change a theory; in fact, the ability to be modified with further information is an essential characteristic of scientific theories! The highest level of regard and prestige an idea can achieve is to become a theory. A summary of this process can be found in figure 9.21.

hypothesis a tentative, testable and falsifiable statement for an observed phenomenon that acts as a prediction for the investigation

FIGURE 9.21 Establishing a theory through ongoing collection of evidence



9.4 Activities

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9.4 Quick quiz on | **9.4 Exercise** | **9.4 Exam questions**

9.4 Exercise

1. Describe three ways in which models can be useful.
2. Describe three limitations of models.
3. What is the difference between a model and a theory?
4. How can theories help us understand phenomena?
5. Why is it important in science that models and theories are constantly being challenged and revised?

9.4 Exam questions

Question 1 (1 mark)

MC Which of the following does not provide strong evidence that can be used to support models and theories?

A. Minimal bias and high validity
B. Supporting research from journals that are not peer reviewed
C. Based on scientific evidence
D. Reproductible and reliable methods

Question 2

MC A _____ is not considered to be proven true because new evidence may come to light that later leads to its rejection.

A. theory
B. model
C. aim
D. hypothesis

Question 3 (1 mark)

Can a hypothesis be proven? Justify your response.

Question 4 (2 marks)

Describe why it is important to use theories in research investigations.

More exam questions are available in your learnON title.

9.5 Effective science communication

KEY KNOWLEDGE

- Chemical concepts specific to the investigation: definitions of key terms; and use of appropriate chemical terminology, conventions and representations
- Characteristics of effective science communication: accuracy of chemical information; clarity of explanation of chemical concepts, ideas and models; contextual clarity with reference to importance and implications of findings; conciseness and coherence; and appropriateness for purpose and audience
- The use of data representations, models and theories in organising and explaining observed phenomena and chemical concepts, and their limitations
- The influence of social, economic, legal and/or political factors relevant to the selected research question
- Conventions for referencing and acknowledging sources of information

KEY SCIENCE SKILLS

- Construct evidence-based arguments and draw conclusions
- Analyse, evaluate and communicate scientific ideas

Source: VCE Chemistry Study Design (2023–2027) extracts © VCAA; reproduced by permission.

9.5.1 Characteristics of effective science communication

Communicating effectively in science allows trends and patterns to be clearly seen, provides support for research and deepens conceptual understanding. Clear communication also provides context regarding the importance and implication of findings. Science communication is effective when others can read a report and gain clarity and understanding of the information and research undertaken.

Methods that assist in effective communication include the following:

- using correct and relevant terminology and explaining all scientific terms
- dividing the report into paragraphs and, where necessary, using subheadings to separate ideas
- taking care with punctuation, spelling and grammar
- having accurate chemical information, including using correct formulas
- making sure that all information has context and is important to your findings
- being concise and sticking to provided word limits — it is better to have a succinct 100-word summary than 500 words of irrelevant information where your key ideas are hard to distinguish
- offering clear explanations of chemical concepts, ideas and models found in your research and your report
- considering the audience you are writing for when establishing the ways you should communicate your findings.

Using chemical concepts specific to investigations

As part of an investigation, it is vital to include key chemistry concepts that are relevant and to clearly explain their significance. This shows a clear link to your understanding of an investigation, and allows others to see the connection between theory and practical applications.

Using chemical concepts

Concepts should be researched prior to commencing your investigation and recorded in your logbook (and referenced). This background information will also form part of your introduction in your report or poster.

In a specialised field such as chemistry, communication can be improved by providing definitions of key terms, and using appropriate chemical terminology, conventions, units, representations, models and theories. This helps prevent misunderstandings in your communication and allows ideas to be presented clearly. Chemistry frequently uses capital letters, superscripts and subscripts to convey information and these should be applied with care. For example, CO is carbon monoxide, while Co is the element cobalt. Chemical formulas should always include the use of subscripts and superscripts where appropriate; for example, the formula for water should be written as H_2O, not H2O.

Each of these ideas is outlined in table 9.6.

TABLE 9.6 Chemical concepts specific to investigations

Chemical concept	Use	Limitations
Definition of key terms	Allows for a clear understanding of the meaning of terminology in the context of the investigation	Uncertainty exists about which specific words need to be defined.
Chemical terminology	Allows for individuals reading a report to contextualise the information to chemistry concepts	Similar words can be used to describe an identical concept or a phenomenon, which leads to confusion.
Conventions	Provides a set of standards followed in reporting and investigations (e.g. including an introduction, aim and conclusion)	Some conventions are unclear and so are changed, leading to confusion when different conventions are used.
Units	Provides an international set of units (SI) for measurements, allowing for easy comparisons between different research investigations	Units may need to be converted (i.e. from kilograms to grams); care should be taken to avoid a misinterpretation of data.
Representations	Shows certain chemistry aspects using symbols and models that make communication easier and clearer	This can lead to oversimplification of concepts by only representing them rather than showing all detail.
Models	Provides an explanatory framework to assist with understanding observed phenomena and abstract concepts	This oversimplifies concepts because they cannot be true representations of real-world conditions or show all details.
Theories	Outlines a well-supported explanation of phenomena, based on investigation, research and observations	Can rely on specific conditions being met that are not always realistic. Phenomena often need to be described with multiple theories.

Conciseness, coherence and appropriateness for purpose and audience

Conciseness is defined as the quality of communicating clear information in as few words as possible. Coherence is defined as the quality of being logical and consistent.

The goal when conveying information is to ensure the information is as concise and coherent as possible. It is important to choose words and phrases that the audience will understand. While technical and scientific terms are appropriate for academic audiences, common terms that have a similar meaning are more suitable for the public. Using highly technical language will cause the public to form misunderstandings and misconceptions of the information presented. Therefore, important chemical writing needs to get straight to the point so the audience can easily comprehend it and avoid confusion.

Techniques to communicate (in verbal or written form) successfully

1. *Know your audience:* Research and understand your audience; always think from a reader's perspective to ensure your communication is interesting, engaging and understandable for your audience's literacy level.
2. *Don't use jargon:* Clarity involves avoiding jargon, uncommon abbreviations and references, because this can create confusion with your audience.
3. *Get to the point:* Choose your words carefully so that they communicate the exact meaning of your research. The key points or findings need to be delivered very early while the audience is still engaged, and it needs to be concise. Though it is harder, it is more meaningful to describe something with fewer words.
4. *Use analogies:* Analogies are a powerful tool when trying to explain something complex. Many gruelling and complex scientific concepts can be clarified using an analogy, because everyone in the audience can relate to their previous experiences.

9.5.2 Terminology and representations

Symbols

It is important throughout your report to use clear and concise terminology relevant to the related chemistry concept.

Due to the sheer quantity of variables we have to represent, it is important to note that the uppercase and lowercase versions of a letter usually represent different things. As well as that, sometimes the same symbol is used to represent different variables.

For example, μ can be used to represent:

- mean in statistics
- micro in measurement
- coefficient of friction.

Similarly, C can be used to represent:

- carbon (C)
- concentration (c).

It is important to carefully use the correct symbol in your report to minimise confusion. Table 9.7 shows some symbols used in Units 1 and 2. It is important to know these and other symbols, including those used in equations.

TABLE 9.7 Some common symbols used in chemistry

Symbol	Representation
Δ	Change in variable
$\rightleftharpoons$	Equilibrium
N_A	Avogadro's constant
N	Number of particles
n	Amount in mol
M	Molar mass
m	Mass
c	Concentration
V	Volume
V_m	Molar volume

(continued)

TABLE 9.7 Some common symbols used in chemistry *(continued)*

Symbol	Representation
e^-	Electron
A_r	Relative atomic mass
K_W	Ionisation constant
(g)	Gas (used in chemical formulas)
(aq)	Aqueous (used in chemical formulas)
(l)	Liquid (used in chemical formulas)
(s)	Solid (used in chemical formulas)

Note: symbols are often used to represent specific variables, and different elements. Many symbols are also specific to drawing the skeletal structure of molecules.

Standard abbreviations

Often, it is appropriate to abbreviate terms rather than writing them out in full on each appearance, which can make a report bulky and hard to follow.

In chemistry, we often shorten formulas to their chemical formula instead of writing it out in full. This is particularly true with molecules such as water, in which the abbreviation H_2O is universally known.

Some common abbreviations used in Units 1 and 2 Chemistry are:

- IUPAC: International Union of Pure and Applied Chemistry
- pH: potential of hydrogen
- UV: ultraviolet
- AAS: atomic absorption spectroscopy
- HPLC: high-performance liquid chromatography
- ppm: parts per million
- ppb: parts per billion.

When using an abbreviation, on the first appearance write the word out in full followed by its abbreviation in brackets; then, in subsequent uses, the abbreviation alone may be used. An example of this is as follows:

> Atomic absorption spectroscopy (AAS) was developed in Australia. In AAS, the concentrations of metal ions are detected.

9.5.3 Referencing and acknowledgment of sources

In-depth scientific reporting requires a depth of research for concepts relating to an investigation. This research may include:

- using other sources for definitions and background material
- finding examples of similar investigations
- research on the obtained results to link to scientific understanding.

Acknowledgements are usually used to thank individuals or organisations that provided assistance — whether this assistance was the provision of specific materials, experimental assistance or intellectual assistance.

References are used when you are sourcing information and intellectual property that is not your own. You will find that your references will be longer than your acknowledgements.

If you use any material that is the work of another person, you must reference its source. Do not claim it as your own work.

Acknowledgments and references come in two formats:

- a short version when they occur in the body of your poster (known as in-text referencing)
- a longer version when they occur in the *Reference and acknowledgments* section at the end of your poster or scientific investigation.

You should include both forms of referencing.

Such references and acknowledgments can be included in many different ways, and various institutions and publications use different systems. Details of these systems can be found online and can be quite complicated. You should check with your teacher as to how your references are expected to be included. Many online generators can also assist you with referencing in the correct style.

Acknowledging sources within your report: in-text referencing

An in-text reference is a shortened or abbreviated form of a reference and should be used in the body of your report in the location in which the sourced information is referred to. This is used for not only direct quotes, but also tables, images and any information that has been paraphrased.

You can do this numerous ways and it depends on what style you are using so, again, check with your teacher.

Author–date system

The author–date system lists the last name of the author and the year of publication. This style of in-text referencing is more commonly used, particularly in the APA and Harvard styles of referencing. As well as the shortened in-text reference, a full reference is included in the reference list.

The in-text reference appears directly after the end of the information being used. This may mean the reference appears in the middle of a sentence. For example:

> . . . over the past 10 years, the number of eligible children has increased (Kringle, 2008) and a need has, therefore, developed for sleighs to travel faster to meet the required delivery schedule. More efficient fuels are required for this purpose.

Items such as tables, diagrams and graphs that are inserted without being substantially altered can often be acknowledged by stating the details directly below them. The reference at the end of table 9.8 is an example.

TABLE 9.8 Energy content of commonly available sleigh fuels

Fuel	Energy content (kJ L^{-1})
Rudolphene	45
Polar plus	29
Super sleigh	53

(Claus, 2016, p. 45)

Note the following for the author–date system:

- If the online article is undated, put (n. d.) in place of the date.
- If no author is listed, use the title in place of the author's name.
- If the source has up to three authors, list them all.
- If the source has more than three authors, only use the name of the first author and follow it by the phrase 'et al.' (meaning 'and others').
- If you quote directly from an author or want to cite a specific idea or piece of information from the source, also include the page number of the quote in your in-text reference (see the example in table 9.8).

Footnotes system

This style of in-text referencing is usually used in the Chicago and Vancouver style of referencing.

In this style, the full citation is included at the bottom of the page the referenced material appears on, with a superscript number showing the point in the text where the material has been used. For example:

> . . . over the past 10 years, the number of eligible children has increased[5] and a need has, therefore, developed for sleighs to travel faster to meet the required delivery schedule. More efficient fuels are required for this purpose.

In the footnotes section of the relevant page (depending on the format of the report), the work is then referenced alongside the corresponding number, and then again included in the reference list.

[5] Kringle, K (2008). *Journal of Polar Transport*, Vol. 34, p. 356.

Acknowledging sources at the end of your report: Reference list

At the end of a scientific report or poster, a reference list is included.

In your reference list, references should be listed alphabetically. If the footnotes version of in-text referencing was used, references should be listed in order of footnote number.

Table 9.9 provides examples of the use of Harvard style in creating references. You may be required to reference many other types of information.

TABLE 9.9 Different forms of referencing using the Harvard style

Type	Format	Example
Book	Author surname(s), initial(s) (Year published). *Title*. Edition (if applicable). Place of publication, publisher.	Langley, P & Jones, M (2013). *The search for Richard III: The king's grave*. 6E. London, John Murray.
Journal	Author surname(s), initial(s) (Year published). 'Title of article'. *Title of Journal*. Volume number, Issue number, Page numbers, DOI (if available)	Naranjo-Ortiz, MA & Gabaldon, T (2019). 'Fungal evolution: diversity, taxonomy and phylogeny of the Fungi'. *Biological Reviews*, vol. 94, no 6., pp. 2101–2137, doi: 10.1111/brv.12550
TV program	Title of program (date) (TV program) Channel identification.	Gene editing made simple (2nd April 2016), (TV program), ABC Sydney.
Website	Author surname(s), initial(s) (Year published). *Title of page*. Name of website, date and website of retrieval	Australian Government (2020). *Coronavirus (COVID-19) health alert*. Department of Health, retrieved 19th of March, 2020 from https://www.health.gov.au/news/health-alerts/novel-coronavirus-2019-ncov-health-alert

Notes

- Do not use 'et al.' in your reference list. This is only appropriate in your in-text referencing.
- A DOI is a permanent identifier for a journal article and is often used in place of a URL. If a DOI is not available, you can list the database it was retrieved from instead (for example, Wiley Online Library).
- When using websites, you should only be using those written for an academic target audience in your report.
- If no author is available, the company responsible for the webpage is fine.
- If no date is available, use 'nd'.

Acknowledging tables and images

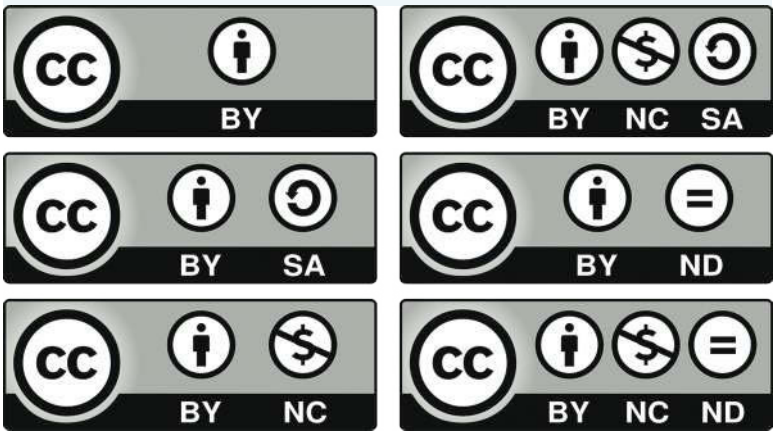
Tables and images can be very effective at summarising large amounts of data and providing added interest; however, they need to be relevant and referenced. Images and data may be covered by copyright.

Keep the following in mind:

- Data should always be accompanied by a source line to demonstrate where it was found.
- Copyright-free images may be identified by a Creative Commons icon or using 'usage rights' in an advanced Google search engine.
- Many websites also have Creative Commons usage rights, such as Wikimedia Commons.

Many different icons in Creative Commons indicate different licences that determine how you can use an image, as shown in figure 9.22. Because you will be using your images in a non-commercial setting it is often easier to get permission, but it is important to note what is required when an image (or other work) is used.

FIGURE 9.22 Different licences available in Creative Commons



Creative Commons requirements may include the following:

- *Attribution (BY):* referencing the image with the copyright holder's identity is required.
- *No Derivative (ND):* altering or remixing the image (or other work) is not permitted.
- *Non-commercial (NC):* the image (or other work) is available only for non-commercial purposes (not for financial gain).
- *Share-alike (SA):* works that are adaptations or derivatives of the work need to be under the same licence as the original.

Some work is in the public domain and may be used freely, without any issues regarding attribution and derivatives. The image for the public domain is shown in figure 9.23. Some examples of websites that have images in the public domain include Pexels, Pixabay and Unsplash. While including attribution or crediting the owner of the images from these sites is not required, it is always good practice to acknowledge your sources.

FIGURE 9.23 Images and media in the public domain can be used without copyright issues. Sometimes, only the icon with the crossed-out C is shown.

9.5.4 Present a report of your research investigation

Decide how you will present your report. Various ways of presenting a report include:

- literature review
- poster
- slideshow
- video
- animation
- oral presentation
- information pamphlet
- community campaign.

9.5.5 Research investigation checklist

Before completing your research investigation, use the following as your final checklist.

Have you:

- ☐ understood your topic sufficiently
- ☐ used your own words
- ☐ included an introduction with an outline of what the topic is about
- ☐ communicated the information clearly, logically and accurately
- ☐ used correct chemical units and terminology
- ☐ explained important scientific terms
- ☐ referred to relevant chemical theories or models
- ☐ checked whether your information is reliable and trustworthy
- ☐ considered related ethical, environmental and social impacts where necessary
- ☐ referenced all diagrams and tables
- ☐ presented the report in an appropriate and succinct form that would be understood by a fellow student at this level
- ☐ written a concluding paragraph that relates back to the introduction
- ☐ provided a detailed bibliography?

9.5 Activities

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9.5 Quick quiz on | **9.5 Exercise** | **9.5 Exam questions**

9.5 Exercise

1. Describe why effective science communication is required for your research investigation.
2. A student submits their research investigation and their teacher comments that their task was extremely concise but lacked coherence. What did the teacher mean by this, and what improvements would you suggest to the student to improve coherence?
3. Why is it important to include a reference list in your report?
4. Write a reference in Harvard style for this textbook. Show both the in-text referencing and the full reference for the reference list.
5. You find an image that you would love to include on your report. Under copyright, it has the letters 'BY' and 'NC'. What does this mean and would you be able to use the image in your report?

9.5 Exam questions

Question 1

MC You find an incredible journal article that relates to your investigation. Six authors are listed.

How should the article be referenced in your report?

A. Only write the first author's name followed by 'et al.' in both your in-text referencing and reference list.
B. Write the names of all authors in both your in-text referencing and your reference list.
C. Write the names of all authors in your reference list, but only the first author followed by 'et al.' in your in-text referencing.
D. Write the names of the first three authors in both your in-text referencing and your reference list.

Question 2 (1 mark)

MC An image is described as being in the public domain.

What does this mean?

A. You must reference the image source.
B. You can only use it for non-commercial use.
C. You can freely use this in your report on your investigation.
D. You must pay for its use.

Question 3 (3 marks)

a. Define the term *jargon*. **(2 marks)**
b. Should you use jargon in your report? Justify your response. **(1 mark)**

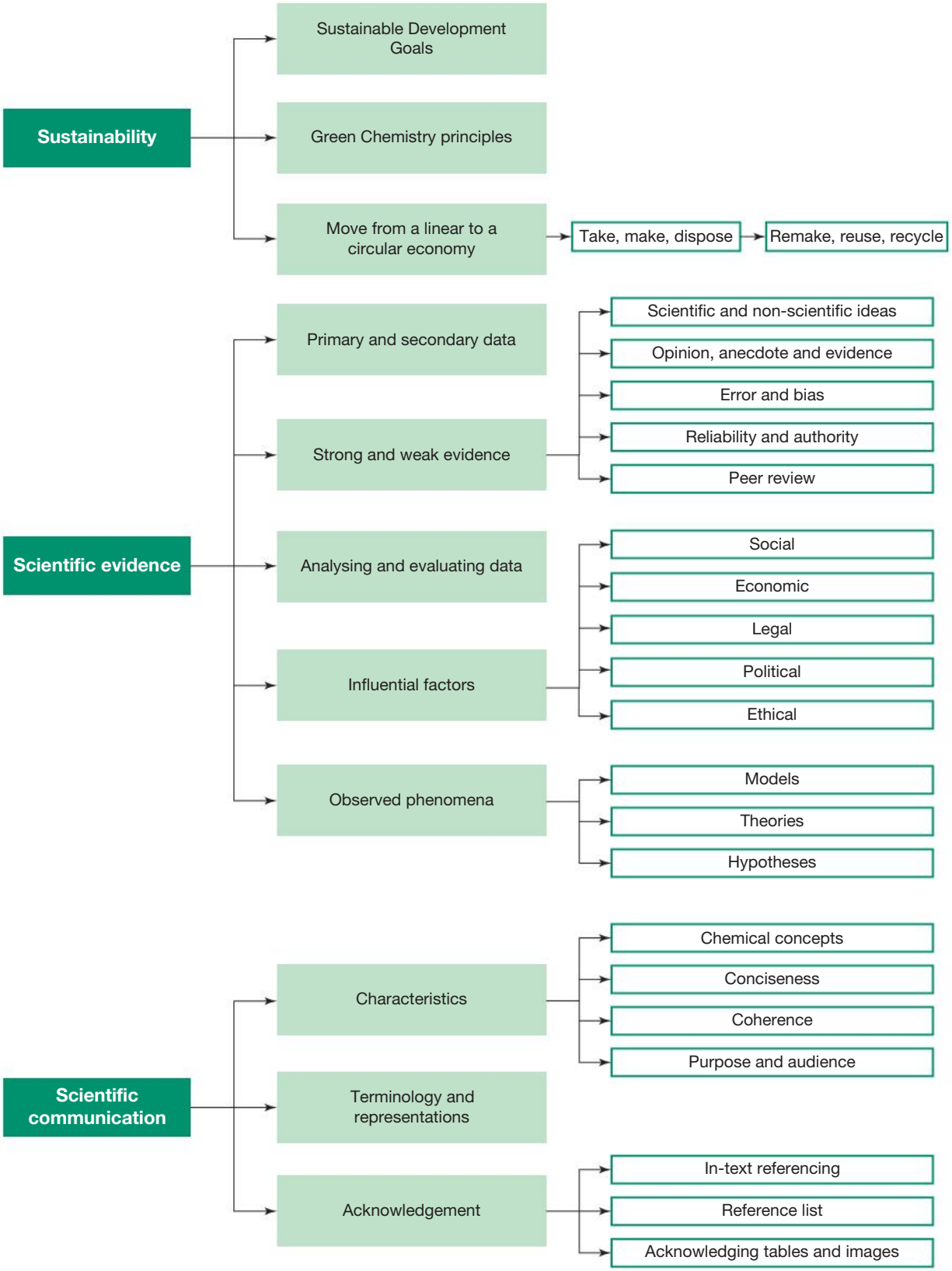
Question 4 (1 mark)

In reports, you may use models and analogies as part of your communication. Explain why these can be useful in scientific communication.

More exam questions are available in your learnON title.

9.6 Review

9.6.1 Topic summary



Hey students! Now that it's time to revise this topic, go online to:

Access the topic summary **Review your results** **Watch teacher-led videos** **Practise exam questions**

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9.6.2 Key ideas summary

online only

9.6.3 Key terms glossary

online only

on Resources

Solutions	Solutions — Topic 9 (sol-0808)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 9 (doc-37083) Key idea summary — Topic 9 (doc-37084)
Exam question booklet	Exam question booklet — Topic 9 (eqb-0090)

9.6 Activities

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9.6 Review questions

1. Explain the difference between the following.
   a. Acknowledgements and references
   b. Conciseness and coherence
   c. Social and economic factors
2. a. Differentiate between opinion, anecdote and evidence.
   b. Give an example of when an opinion could be used as strong evidence.
   c. Describe why many popular media articles include an author disclaimer.
   d. Suggest three ways to determine whether a secondary source contains scientific or non-scientific ideas.
3. An investigation is conducted into the synthesis of a new product. For this research investigation, identify one of each of the following: social, economic and ethical factors.
4. You are using a webpage as a reference in your scientific report. Suggest a reason why you need to include in your bibliography the date that you accessed the webpage.
5. Outline one of the greatest challenges to sustainability, and discuss potential solutions to this challenge.

9.6 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC Why are the 12 Principles of Green Chemistry necessary?

A. The chemical industry has a sustainable manufacturing processes
B. To save our depleting natural resources
C. To ensure the future of the next generation
D. All of the above

Question 2

MC Which of the following is true?

I. An example of primary data is the results published in the paper from a government census.
II. Secondary data include comments on, or summaries of, primary data.
III. Information found on the internet cannot be edited by the general public.

A. I and II
B. II only
C. I and III
D. III only

Question 3

MC The reference section of a scientific report must include a list of all what?

A. Equipment used in the experiment
B. People who contributed to carrying out the experiment
C. Companies who provided financial support for the experiment
D. Sources cited in the report

Question 4

MC Which of the following is the best definition of a theory?

A. Concepts that were once accurate, but have now been rejected
B. Diagrammatic representations of abstract concepts
C. Ideas that are unable to be rejected due to the amount of evidence they have
D. Well-supported ideas where evidence has been gained from investigations, research and observations

Question 5

MC A peer-reviewed journal article

A. has been proofread for spelling mistakes by an editor.
B. has undergone review by experts in the field to ensure that it is credible.
C. only contains secondary data.
D. must be funded by government agencies.

Section B — Short answer questions

Question 6 (1 mark)

How does the equation for the production of ethanol from the hydration of ethene, $C_2H_4(g) + H_2O(g) \rightarrow CH_3CH_2OH(g)$, show that this reaction has a high atom economy?

Question 7 (2 marks)

Some authors and creators allow their work to be used freely, without a need to reference or include acknowledgement. This is called *public domain information* and is copyright-free. It may be shown using a Creative Commons image. Other work becomes part of the public domain once certain criteria are met.

Give two examples of when information will become copyright-free and enter the public domain.

Question 8 (2 marks)

Describe two considerations when choosing to use secondary data. Why are these important?

Question 9 (4 marks)

An investigation is being conducted on the use of nuclear energy instead of coal in Australia. Identify two examples of primary data and two examples of secondary data you may collect that is relevant to this investigation.

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UNIT 2

How do chemical reactions shape the natural world?

AREA OF STUDY 1

How do chemicals interact with water?

OUTCOME 1

Explain the properties of water in terms of structure and bonding, and experimentally investigate and analyse applications of acid–base and redox reactions in society.

10 Water as a unique chemical 369

11 Acid–base (proton transfer) reactions 405

12 Redox (electron transfer) reactions 463

AREA OF STUDY 2

How are chemicals measured and analysed?

OUTCOME 2

Calculate solution concentrations and predict solubilities, use volumetric analysis and instrumental techniques to analyse for acids, bases and salts, and apply stoichiometry to calculate chemical quantities.

13 Measuring solubility and concentration 513

14 Analysis for acids and bases 557

15 Measuring gases 587

16 Analysis for salts 615

AREA OF STUDY 3

How do quantitative scientific investigations develop our understanding of chemical reactions?

OUTCOME 3

Draw an evidence-based conclusion from primary data generated from a student-adapted or student-designed scientific investigation related to the production of gases, acid–base or redox reactions, or the analysis of substances in water.

17 Scientific investigations ..... online only

Source: VCE Chemistry Study Design (2023–2027) extracts © VCAA; reproduced by permission.

10 Water as a unique chemical

KEY KNOWLEDGE

In this topic you will investigate:

Water as a unique chemical

- existence of water in all three states on Earth's surface, including the distribution and proportion of available drinking water
- explanation of the anomalous properties of H_2O (ice and water) with reference to hydrogen bonding:
  - trends in the boiling points of group 16 hydrides
  - the density of solid ice compared with liquid water at low temperatures
  - specific heat capacity of water including units and symbols
- the relatively high latent heat of vaporisation of water and its impact on the regulation of the temperature of the oceans and aquatic life.

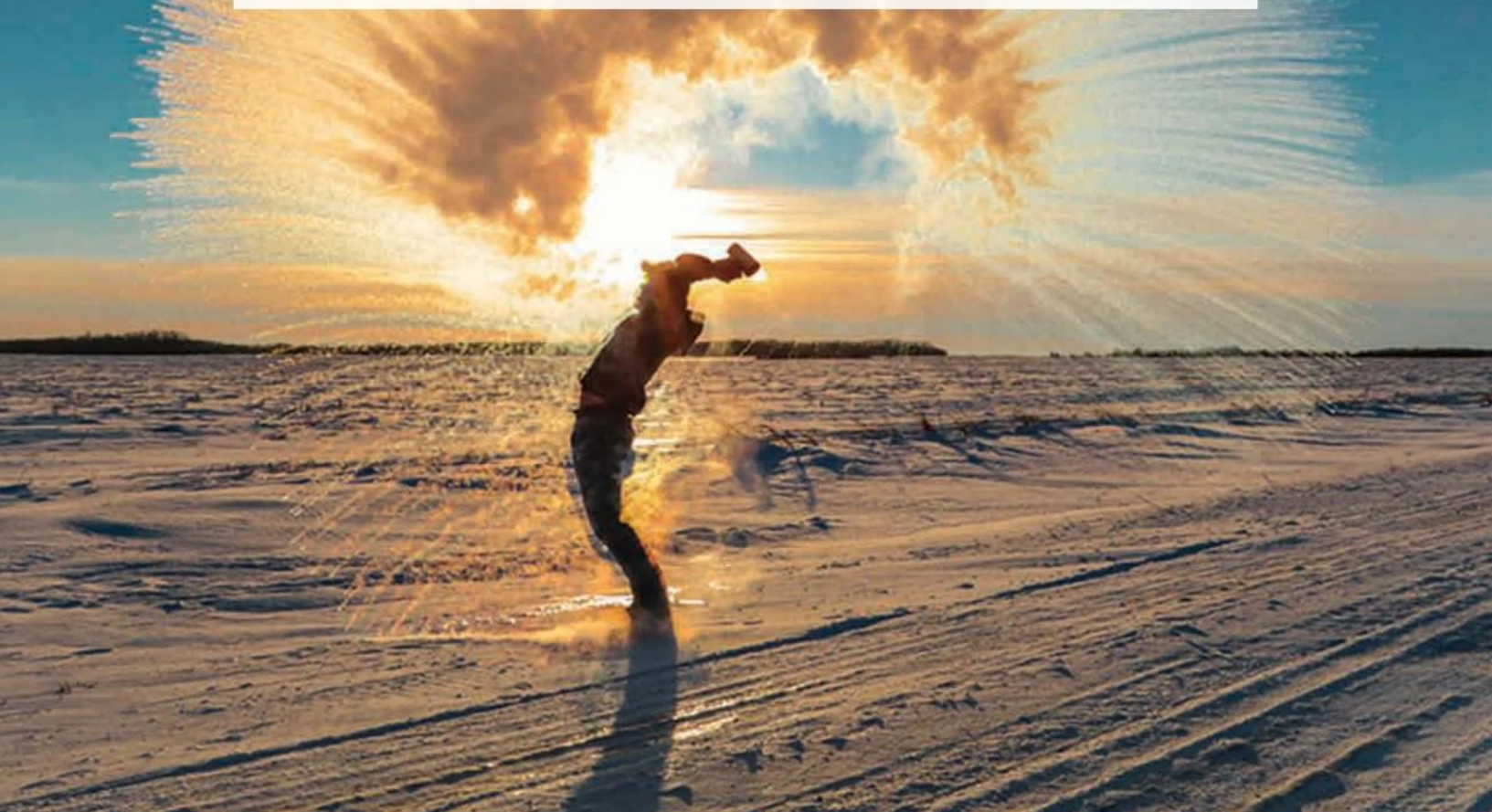
Source: VCE Chemistry Study Design (2023–2027) extracts © VCAA; reproduced by permission.

PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of VCE Chemistry. Experiments and investigations, supported by a practical investigation eLogbook and teacher-led videos, are included in this topic to provide opportunities to undertake investigations and communicate findings.

EXAM PREPARATION

Access exam-style questions and their video solutions in every lesson, to ensure you are ready.



10.1 Overview

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10.1.1 Introduction

More than 70 per cent of Earth's surface is covered by water, but only a small percentage of this water is fit to drink. Water is critical to life and plays an important role in moderating our climate. Within the normal temperature range of our planet, water exists in all three states: solid, liquid and as water vapour in the atmosphere.

FIGURE 10.1 Since 1990, 2.6 billion people have gained access to an improved drinking water source.

Why does ice float on water? Why do ponds freeze over in winter, forming a layer of ice on their surfaces, instead of freezing solid? How is it that aquatic organisms can survive in waters under the ice? To answer these questions, it is necessary to examine how hydrogen and oxygen atoms join together to form water molecules, and how water molecules then interact with each other. Water has many interesting properties owing to its structure — water can absorb a relatively large amount of heat with a correspondingly low temperature rise, which aids the oceans' role as regulators of our climate. This is particularly significant as global temperatures rise due to climate change.

This topic will explore the significance of these concepts for processes in living systems and the environment, and look at current global initiatives and practices to ensure safely managed drinking water will be available for all people in the near future.

LEARNING SEQUENCE

10.1 Overview 370
10.2 Water on Earth 371
10.3 Properties of water 380
10.4 Heat capacity and latent heat 388
10.5 Review 398

on Resources

Solutions	Solutions — Topic 10 (sol-0809)
Practical investigation eLogbook	Practical investigation eLogbook — Topic 10 (elog-1615)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 10 (doc-37085) Key ideas summary — Topic 10 (doc-37086)
Exam question booklet	Exam question booklet — Topic 10 (eqb-0091)

10.2 Water on Earth

KEY KNOWLEDGE

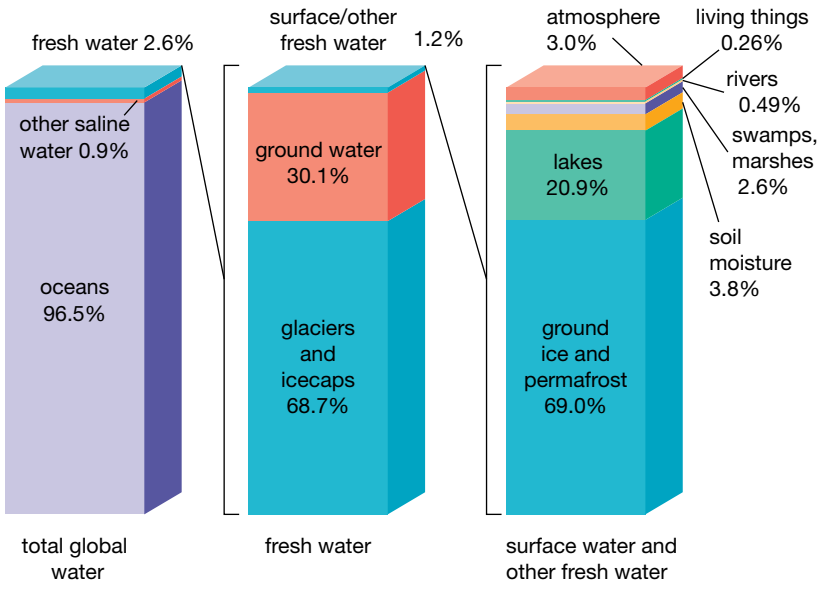
- Existence of water in all three states at Earth's surface including the distribution and proportion of available drinking water

Source: VCE Chemistry Study Design (2023–2027) extracts © VCAA; reproduced by permission.

10.2.1 Using water

Nearly all the water that is used around the world every day, including that used for drinking, is not pure. This is because, as we have seen in previous topics, water is an excellent solvent. In the environment, water comes into contact with many substances that dissolve in it. As water falls through the atmosphere in the form of rain, gases from the atmosphere, both natural and pollutants, dissolve in it to some extent. When water travels over Earth's surface, or soaks into the ground, naturally occurring salts also dissolve. As demonstrated in figure 10.2 and table 10.1, only approximately 2.6 per cent of all water on Earth is fresh water, and of this 2.6 percent only 1.2 per cent is surface or other fresh water. This means only ~0.03 per cent of all water on Earth must sustain all land-based life, including more than 7 billion people.

FIGURE 10.2 Global distribution of water



Because all water eventually flows into the oceans as part of the water cycle, this is precisely why the ocean is so salty and contains a wide range of dissolved substances. The major ions in sea water are shown in table 10.2.

The United Nations has estimated that water usage has grown at more than twice the rate of population increase in the last century, and by 2025 an estimated 1.8 billion people will live in water-scarce regions. Increased water usage is exacerbated by population growth and climate change, and by 2025 two-thirds of the world's population will live in water-stressed regions. Globally, water stress is caused by water sources drying up and/or becoming more polluted, inadequate government investment in water and sanitation combined with reduced international cooperation on transboundary (across country borders) waters. Therefore, this emphasises more than ever our need to cooperate and act together as a global community in partnership and implement strategies targeting the United Nations Sustainability Goals, including Goal 6 (Clean Water and Sanitation) and Goal 14 (Life Below Water). Achieving these goals will require action on behalf of individuals, governments and scientists to develop clean water solutions. These goals are discussed in detail later in this subtopic.

TABLE 10.1 Approximate global distribution of water by percentage and state

	Location of water	State of water	Percentage of total water on Earth
Salt water 97.4%	Oceans	Liquid	96.5%
	Other saline water	Liquid	0.9%
Fresh water 2.6%	Glaciers and icecaps	Solid	1.72%
	Ground water	Liquid	0.753%
	Ground ice and permafrost	Solid	0.0207%
	Lakes	Liquid	0.00627%
	Soil moisture	Liquid	0.00114%
	Atmospheric water vapour	Gas	0.0009%
	Swamps and marshes	Liquid	0.00078%
	Rivers	Liquid	0.000147%
	Living things	Liquid and gas	0.000078%

TABLE 10.2 Major ions in sea water

Ion	Typical concentration ($mg\ L^{-1}$)
Chloride, Cl^-	18 980
Sodium, Na^+	10 556
Sulfate, SO_4^{2-}	2 649
Magnesium, Mg^{2+}	1 262
Calcium, Ca^{2+}	400
Potassium, K^+	380
Bicarbonate, HCO_3^-	140
Strontium, Sr^{2+}	13
Bromide, Br^-	65
Borate, BO_3^{3-}	26

10.2.2 Water use in Australia

Each state in Australia uses one or more sources of water depending on geography, available sources and population. Australia does not have a large number of natural lakes to assist in water storage and, as the driest inhabited continent with highly variable rainfall and run-off, water management is a significant undertaking. In general, the sources of water in Australia are from:

- water catchments — protected and open
- lakes
- rivers and creeks
- groundwater (bore water)
- desalinated water
- recycled water.

About 14 270 billion litres of water was used for consumptive purposes across Australia in 2019–2020. Agriculture was the biggest use of this water (67 per cent). This is a decrease of 11 per cent from 2018 to 2019 due to dry conditions and low water availability in the Murray–Darling Basin. Urban use was 22 per cent and industrial use was 11 per cent.

Water in urban areas of Australia is sourced from a mains water supply, called **reticulated water**. Reticulated water is controlled by local water authorities and is carried via a pipe from a reservoir. These reservoirs may source their water from a few sources. For example, Melbourne's source of water is mainly from surface water assisted by desalinated water; Perth uses both groundwater and desalinated water; Adelaide relies on water from rivers and desalinated water.

reticulated water water provided through a network of pipes

CASE STUDY: Melbourne's water sources

In Melbourne, Melbourne Water is the authority that treats and provides high-quality water across the Port Phillip and Westernport region. Figure 10.3 shows the sources of drinking water for Melbourne, including ten catchment reservoirs that are linked through a network of pipes. These reservoirs depend on rainfall on a catchment area or diversion from a river. To diversify the water supply system for a growing population, a rainfall-independent desalination plant was built. The plant assists with maintaining water storage volumes and recovery after drought periods, and is part of planning for increased water resources. Figure 10.4 demonstrates the long-term planning for Melbourne's water supply, which is based on modelling of climate change impacts and potential demand variations.

FIGURE 10.3 Melbourne water supply system (Melbourne Water)

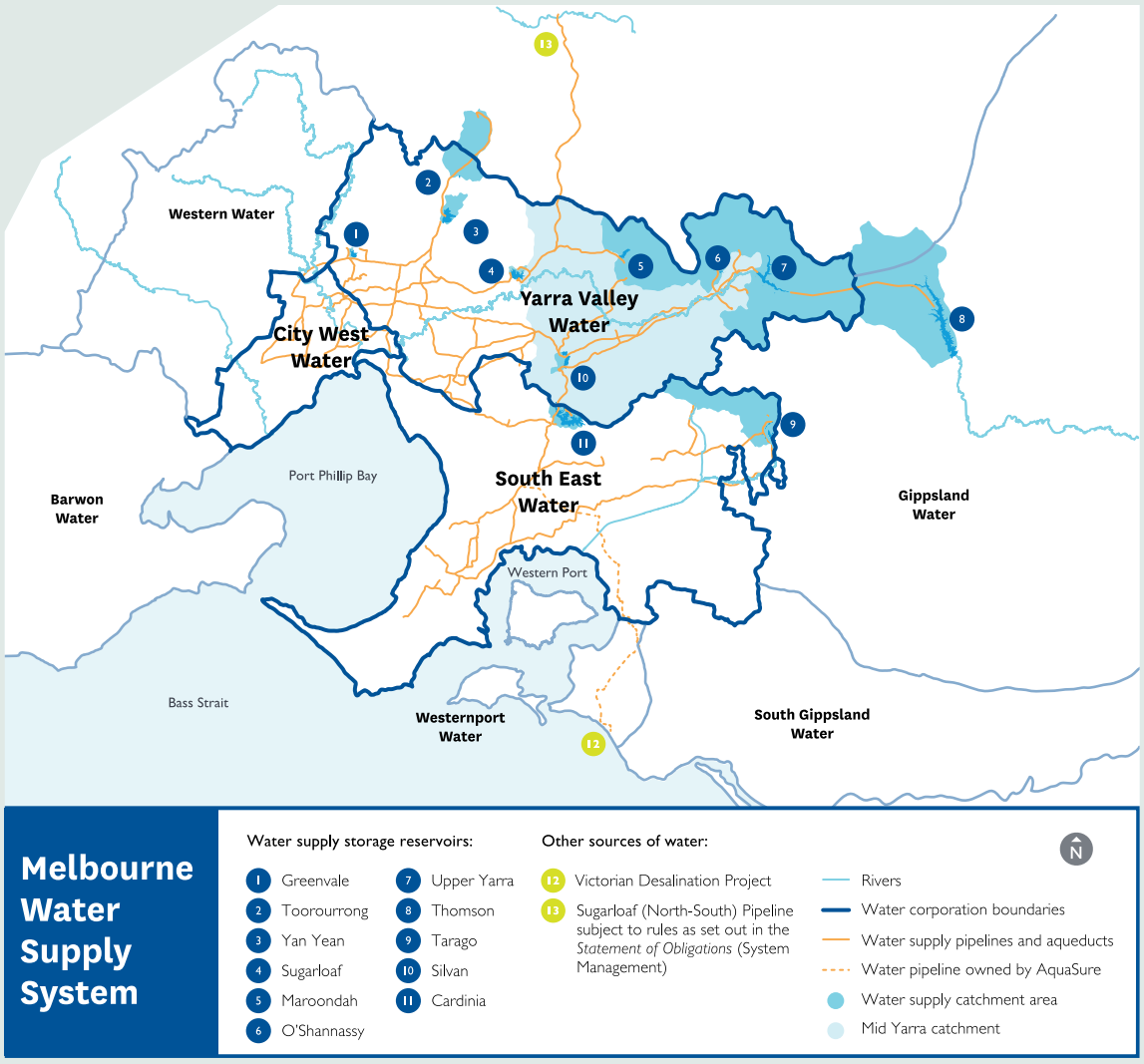
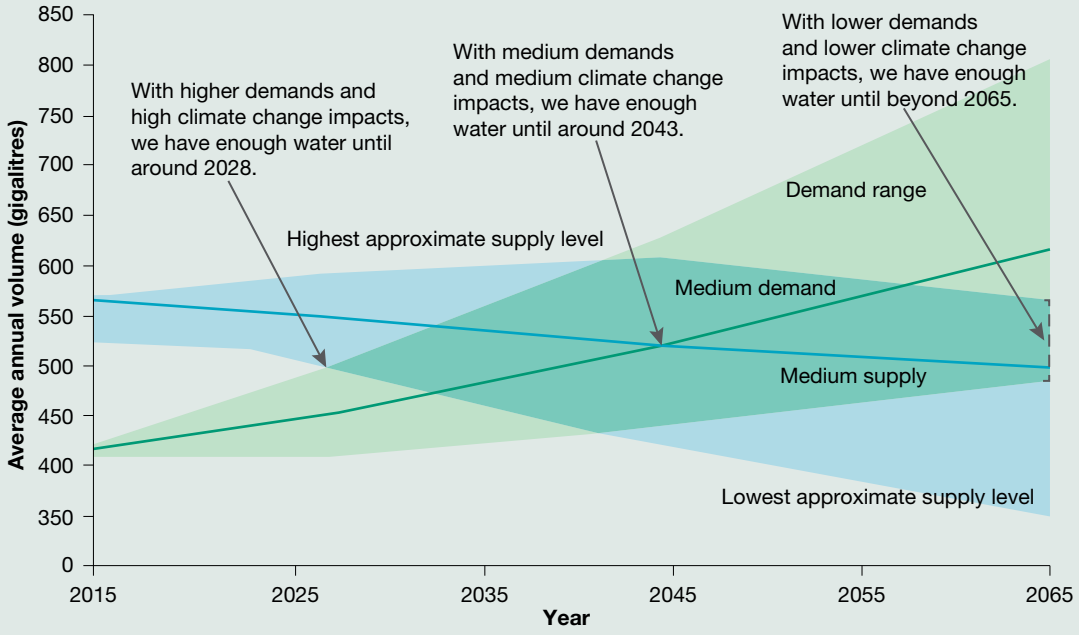


FIGURE 10.4 Long-term water supply and demand for Melbourne



on Resources

Weblinks Melbourne's water supply
Water storage dashboard

CASE STUDY: Wurundjeri creation story of the Yarra river

First Nations Australians have inhabited the Australian landmass for at least 60 000 years. Australia's First Nations peoples have the responsibility to look after the land, waterways and skies. Before the arrival of the Europeans, First Nations Australians ran sophisticated and structured societies using sustainable land management practices. For tens of thousands of years they have been caretakers of Country and have maintained the stories and lore of many special places. Australia's First Nations peoples are made of many distinct groups, each with their own languages, cultures, beliefs and practices. Some people chose to identify with their traditional clan; for example, the Wurundjeri, Woi Wurrung people of the Kulin nation are from Melbourne and the surrounding areas. People may also describe themselves in ways that relate to Country, such as 'saltwater people' for those who live near the coast. Different nation groups have different creation and Dreaming stories of how the land and geographical features came to be.

FIGURE 10.5 The Yarra River in at Moo-rool-bark (Warrandyte): where the wide waters meet

Water moulds the landscape over thousands of years, and this has become integrated into Dreaming Stories. This version of a Wurundjeri, Woi Wurrung creation story is attributed to Elder Billi-Billeri, of the Woi Wurrung Kulin, and describes how the Yarra River used to be locked in by the mountains as a large water mass. This was

called Moorool — Great Water. Mo-yarra, the headman of the Wurundjeri, Woi wurrung people wanted to free the Country of water to give them more land to hunt, so he cut a channel through to the hills towards Bunurong/Boon Wurrung and the area known today as Western Port Bay. However, the path closed over leaving Mo-yarra in Koo Wee Rup. Later, Bar-wool became the headman of the Wurundjeri, Woi wurrung people and also wanted to release the water. He cut a channel up the valley with his stone axe; however, Baw-Baw the mountain stood in his way. He then headed north, but Donna Buang and his brothers stopped him, so he went west and cut through the hills to Warr-an-dyte. At this time he met Yan-yan, a Wurundjeri, Woi Wurrung man who was cutting a path to free Morang, the great water where he lived. Together, they combined their efforts at Moo-rool-bark: where the wide waters meet. When they got to land between the Darebin and Merri creeks, the land was extremely hard so the water ways become narrow and twisted as they were looking for softer ground. They finally reached the bay, releasing the water, filling Port Phillip Bay and freeing their land.

10.2.3 Sustainable Development Goal 6 — clean water and sanitation

The sixth goal of the United Nations **Sustainable Development Goals** (SDGs) is to 'ensure availability and sustainable management of water and sanitation for all' by 2030. Between 2015 and 2020, improvements have been achieved to people gaining access to safely managed drinking water services (70.2 to 74.3 per cent), safely managed **sanitation** services (47.1 to 54.0 per cent) and basic **hygiene** needs (67.3 to 70.7 per cent); however, the world is currently not on track to achieve this goal by 2030, indicating the current rate of progress must double.

FIGURE 10.6 West African girls carrying water

In 2020, 771 million people were still without even basic drinking water, with half of these people (387 million) living in sub-Saharan Africa. As well as this, 46 per cent of the world's population still do not have access to safely managed sanitation. Water access is often through a community bore, because many homes do not have access to reticulated water. This task of collecting water may take many hours and, because it is typically the responsibility of women and girls, this task often prevents girls from attending school. In 2019, only two in three schools globally met basic drinking water and sanitation needs, while three in five schools had basic hygiene services.

In Australia, we are fortunate to have access to clean, reticulated water in our cities and towns. However, Australia's 2020 progress report of SDGs only reported against two of the eight targets within Goal 6 (cost and consumption per capita — see table 10.3). While the cost of water in Australia needs improvement, we are on track to achieve water consumption per capita. Australia's water consumption per capita has been declining since 2001 (with slight increases since 2011 — see figure 10.7). This demonstrates both individuals' commitment to reduced water usage, and government policies and schemes to reduce water use.

Sustainable Development Goals adopted by all United Nations Member States in 2015, consists of 17 goals that address 169 urgent calls for action to global challenges with the overall goal of a sustainable future for all

sanitation conditions of public health, specifically the delivery of clean drinking water and adequate sewage disposal

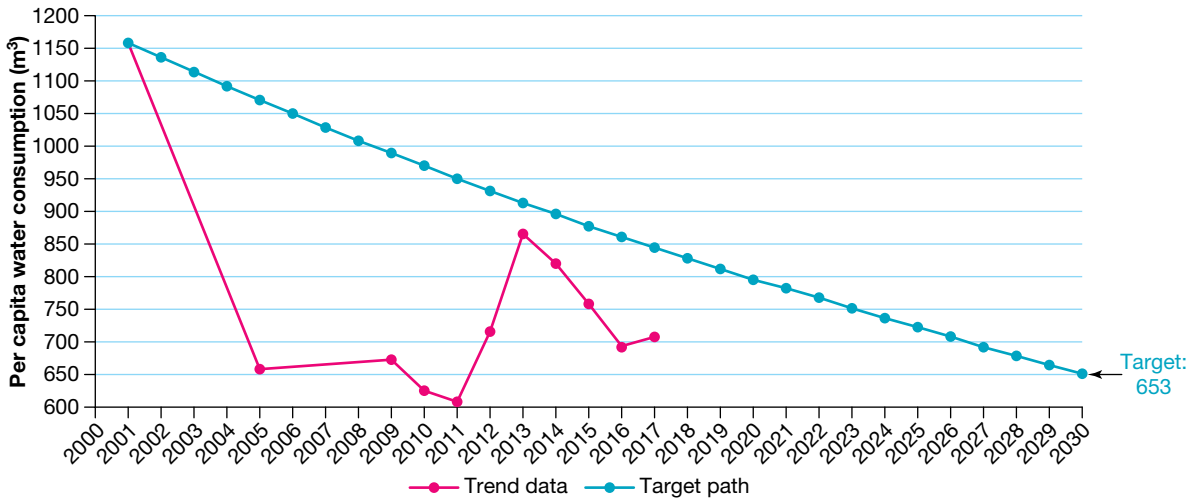
hygiene conditions or practices beneficial to maintaining health and preventing disease, especially through cleanliness

TABLE 10.3 Australia's reported water expenditure needs improvement, while consumption per capita is 'on track' from the 2020 progress report.

Target	Indicator	Status	Outcome
6.1	Average weekly household expenditure on water, sewage and wastewater	Needs improvement*	A gradual increase in cost has occurred between 2010 and 2018; however, on average, the cost is still below the global benchmark of 3% of the average household income. For lower income families, this can increase up to 5.8%.
6.4	Water consumption per capita	On track*	Water consumption has declined significantly since 2001. While slight increases have occurred since 2011, consumption is still below the global benchmark trend.

* On track = >90% of the desired rate of change to meet the target; Needs improvement = current value is better than target value, but trend >0% in wrong direction OR 50–90% of the desired rate of change to meet target (adapted from Transforming Australia SDG Progress Report 2020 update).

FIGURE 10.7 Australia's performance against UN Target 6.4: water consumption per capita



Source: www.sdgtransformingaustralia.com/explore-by-goal/#/1247/1271//, based on QECD data on total water abstractions per capita.

10.2.4 Sustainable Development Goal 14 — life below water

Oceans take up 96.5 per cent of the world's water (refer to figure 10.2), cover three quarters of the world's surface and, by volume, represent 99 per cent of the living space of the planet. Three billion people rely on ocean resources for their livelihoods. But more significant is the role oceans play in absorbing both heat and carbon dioxide from the environment. The oceans absorb more than 90 per cent of excess heat in the climate, and 30 per cent of the carbon dioxide produced by humans. An increase in carbon dioxide emissions has led to:

- an increase in ocean water temperatures (and, as a result, global temperatures). As the oceans absorb more carbon dioxide, they reach a point where they can no longer absorb this excess carbon dioxide, and it leaks out of the ocean and re-enters Earth's atmosphere, contributing to the enhanced greenhouse effect. The oceans can no longer act as a buffer to the high levels of carbon dioxide compared to pre-industrial levels. Increases in atmospheric temperatures result in increasing surface water temperatures, which causes ocean currents to slow. The waters start to form layers, with the surface layers becoming saturated in carbon dioxide. This stagnant water has significant impacts on marine life, and also slows the carbon dioxide uptake from photosynthesis, further reducing the amount of carbon dioxide the ocean can absorb.

- **ocean acidification** (see subtopic 11.7). As fossil fuels are burnt and atmospheric carbon dioxide levels increase, the ocean absorbs more of this carbon dioxide to maintain balance, through a series of reactions between water, carbon dioxide and carbonate (from marine life) producing bicarbonate ions. As you will see in topic 11, the carbonic acid and bicarbonate dissociates, increasing the levels of hydrogen ions (H^+). This lowers the ocean's pH, resulting in increased acidity. This series of reactions can be summarised as

$$H_2O + CO_2 + CO_3^{2-} \rightarrow 2HCO_3^- \rightarrow H^+ + CO_3^{2-}$$

Other threats to water quality and marine life includes electronegativity (deoxygenation). This is the excessive growth of plants and algae, which seriously affects water quality. **Eutrophication** is usually due to still water, sunlight and an excess of nutrients (such as nitrates and phosphates) required for growth. The source of these excess nutrients is human activity (domestic wastewater, agricultural practices and sewage). When the excessive growth of algae and plants die, their subsequent decomposition seriously depletes the level of dissolved oxygen in the water. As a result of this, animals, especially fish, and even plants may die due to a lack of oxygen required for respiration. Additionally, all this decay may produce biotoxins, which can be a serious health hazard to any organism that consumes this water.

In Australia, these impacts can be seen in the rapid changes in the Great Barrier Reef's ecosystem.

The percentage of hard coral declining since 2000 due to ocean acidification (see figure 10.8) is classified as 'off-track' from its goal. The threats to the Great Barrier Reef can be summarised as the following:

- coral bleaching due to heat stress caused by climate change
- decreased water quality due to increased nutrients, contaminants, rising sea temperatures, acidification and the impact of coastal human development
- introduced species, such as the crown-of-thorns starfish.

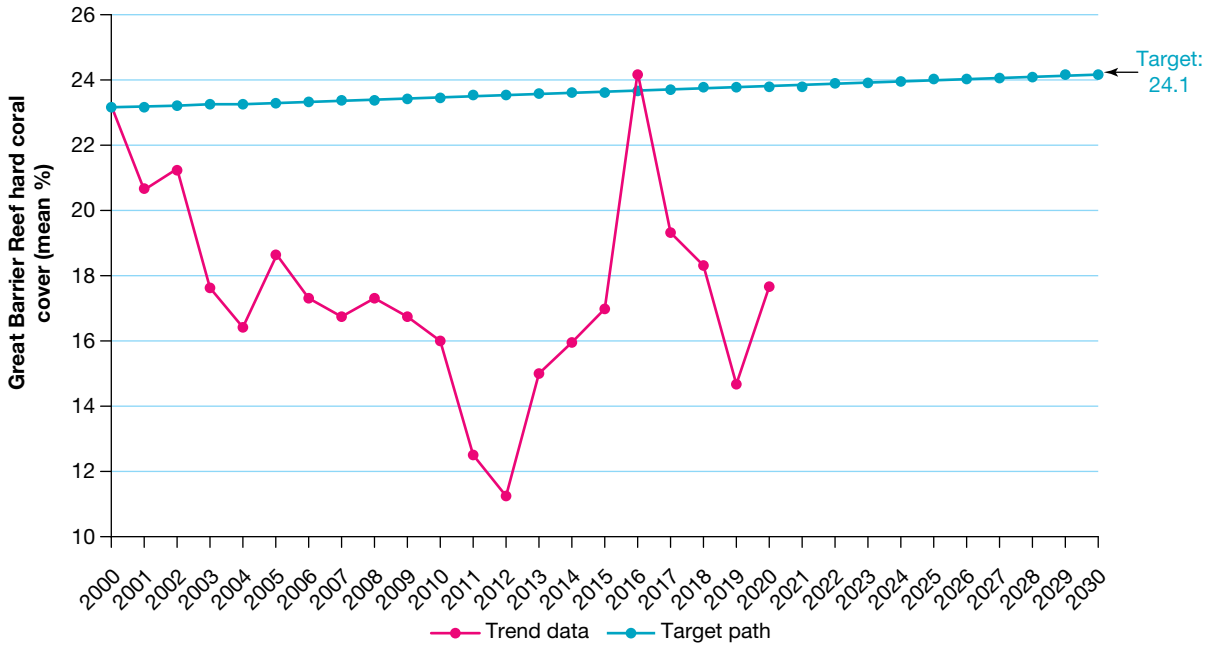
ocean acidification the decrease in the pH of Earth's oceans, due to increased absorption of carbon dioxide (CO_2) from the atmosphere

eutrophication form of water pollution involving excess nutrients, such as nitrogen and phosphorus, leaching from soils, typically resulting in excessive growth of algae

FIGURE 10.8 Coral bleaching on the Great Barrier Reef

Reporting against the SDGs has resulted in increased public awareness of the significant threats of climate change. The Global Day of Action against climate change saw school children become active citizens protesting the current global state of climate change. This global rally was instigated by the Swedish teenage climate activist Greta Thunberg, who has been nominated for the Nobel Peace Prize in 2019, 2020 and 2021. Goal 14 is just one of the UN SDGs being fought for by the youth of today.

FIGURE 10.9 Australia's reporting of Target 14.2: hard coral cover shows a decrease of cover and is listed as off track for this target.



Source: www.sdgtransformingaustralia.com/explore-by-goal/#/1255/1376//, based on data from the Australian Institute of Marine Science (AIMS) Long-term Reef Monitoring Program.

FIGURE 10.10 Greta Thunberg started a global movement for students to protest the inaction of governments on climate change.

Resources

Weblinks Sustainable Development Goal 6
Sustainable Development Goal 14
Sustainable Development Goals Australia
Climate activism
Sustainable Development Goals Australia progress report

10.2 Activities

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10.2 Quick quiz on | 10.2 Exercise | 10.2 Exam questions

10.2 Exercise

1. Australia has the highest per capita water storage capacity in the world. Why?
2. MC 'Potable water' is a term used to describe water fit for drinking. Brackish water is slightly salty water. Water is classified according to its salinity into the following groups.

	Fresh water	Brackish water	Saline water	Hypersaline water
%(w/w)	0.05	0.05–3.0	3.0–5.0	> 5.0

If a water sample is found to have 4.0×10^4 mg L^{-1} of dissolved salt, what would it be classed as?

A. Fresh **B.** Brackish
C. Saline **D.** Salt water

3. Water on Earth exists in all three states of matter: solid, liquid and gas.
   a. What percentage of all water on Earth is found in its solid form?
   b. Presuming that gaseous water in living things is negligible, what is the total percentage of water in its liquid form on Earth?
4. What are the United Nations Sustainable Development Goals?
5. The United Nations adopted 17 Sustainable Development Goals encompassing 169 urgent calls for action targets in a global partnership with the overall goal of a sustainable future for all.
   a. What is Goal 6 and how many targets are there within it?
   b. What aspect of sustainability is this goal targeting?
   c. Since the implementation of these SDGs in 2015, have any improvements been seen in Goal 6 targets in the following areas?
      i. Globally
      ii. In Australia

10.2 Exam questions

Question 1 (1 mark)

MC How is reticulated water delivered?

A. Via an underground bore **B.** From a pump in a river
C. Via a network of pipes **D.** From a desalination plant

Question 2 (6 marks)

a. Only 2.6 per cent of the world's water is freshwater. Why is this such a low number? **(1 mark)**
b. Where is the majority of the world's fresh water found? **(1 mark)**
c. i. List four sources of drinking water in Australia. **(2 marks)**
   ii. Suggest two methods Australia can use to protect its water security. **(2 marks)**

Question 3 (3 marks)

You went for a walk around Lake Louise in Canada in the spring on a cold day and made the realisation that you were able to identify water in all of its states of matter. What states of matter could you identify? Provide an example of each.

Question 4 (1 mark)

MC Which of the following does not directly relate to eutrophication?

A. Results in the deoxygenation of the water supply
B. Results from a combination of excessive sunlight and excess nutrient supply
C. Results from a farmer not having a water management system incorporated onto his properties, and any excess water just draining into the nearby creeks
D. Pteropod shells are disintegrating

Question 5 (3 marks)

How does an increase in atmospheric CO_2 levels lead to an increase in ocean acidification?

More exam questions are available in your learnON title.

10.3 Properties of water

KEY KNOWLEDGE

- Explanation of the anomalous properties of H_2O (ice and water) with reference to hydrogen bonding:
  - trends in the boiling points of Group 16 hydrides
  - the density of solid ice compared with liquid water at low temperatures

Source: Adapted from VCE Chemistry Study Design (2023–2027) extracts © VCAA; reproduced by permission.

10.3.1 Bonding in water

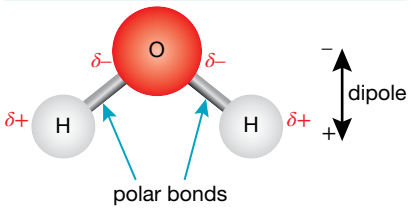
Intramolecular forces in water

Water is a small molecule that contains only three atoms — one oxygen and two hydrogen — giving water the chemical formula H_2O. **Covalent bonds** hold these atoms together. As these bonds occur within a molecule, they are **intramolecular bonds**.

Water is a **polar bond** molecule due to the following features of these bonds:

- There are two covalent bonds in a water molecule — one between each of the hydrogen atoms and the oxygen atom.
- Hydrogen shares its only electron with oxygen, while oxygen shares an electron with each of the hydrogen atoms. Therefore, the oxygen atom gains a greater share of the bonding electrons and becomes slightly more negatively charged (represented by the symbol $\delta -$, 'delta negative') than the hydrogen atom. The hydrogen atom then becomes slightly positively charged (represented by the symbol $\delta +$, 'delta positive'). Consequently, a difference in charge occurs at either end of the water molecule, resulting in water being a polar molecule (see figure 10.11).
- The two non-bonding pairs of electrons force the V-shape or bent-shape of the water molecule.

FIGURE 10.11 A polar water molecule showing polar bonds



Intermolecular forces in water

Two types of **intermolecular bonds** occur between water molecules:

- **Dispersion forces**: the bond between adjacent molecules formed by instantaneous dipoles through the constant movement of negatively charged electrons. Dispersion forces are due to the constant movement of electrons.

covalent bond sharing of electrons between nuclei that bonds them together in a molecule

intramolecular bond internal bond within a molecule

polar bond bond formed when two atoms that have different electronegativities share electrons unequally

intermolecular bond bond between molecules

dispersion force the bond between adjacent molecules formed by instantaneous dipoles

- **Hydrogen bonds**: the bond created between a molecule with a hydrogen atom covalently bonded to an atom of F, O or N and another molecule that also contains an atom of H, F, O or N. Hydrogen bonds are a type of **dipole–dipole bond** and exist between the δ^+ hydrogen atom of one water molecule and the δ^- oxygen atom of another water molecule (see figure 10.12).

When a molecule consists of atoms with different electronegativities such as water, the bonding electrons are attracted toward the atom with the stronger **electronegativity** (oxygen electronegativity is 3.4, compared to hydrogen electronegativity of 2.2), and this uneven distribution of charge is called a **dipole**.

FIGURE 10.12 Hydrogen bond between two water molecules

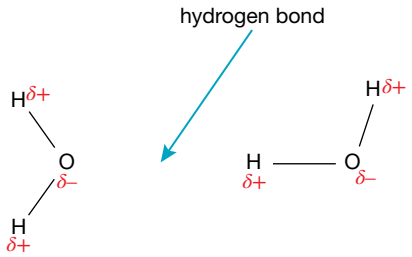


TABLE 10.4 Summary of bonding within and between water molecules

Type of bond	Description of bond	Where the bond is found relative to the molecular structure of water	Relative bond strength (in order from strongest to weakest)
Covalent bond	Intramolecular	Between hydrogen and oxygen within the water molecule	Strongest ↓ Weakest
Hydrogen bond	Intermolecular	Between the delta-positive hydrogen of one water molecule and the delta-negative oxygen atom of another water molecule	
Dipole–dipole bond	Intermolecular	Hydrogen bond is a type of dipole–dipole bond; only hydrogen bonding dipole–dipole bonds exist due to water only having hydrogen and oxygen atoms	
Dispersion force	Intermolecular	Present between all molecules; however, due to the small size of a water molecule, this force is extremely weak	

Hydrogen bonding is fundamental for life processes; these interactions hold the two strands together in the double helix that forms our DNA, for example. Hydrogen bonding significantly affects the melting and boiling points and other physical properties of many compounds, including water.

Bonding in and between water molecules

Water molecules are not only polar, but also contain hydrogen–oxygen bonds, creating hydrogen bonding between the water molecules. This bonding is the reason for the important properties of water.

hydrogen bond the bond between a hydrogen atom covalently bonded to an atom of F, O or N and another molecule that also contains an atom of H, F, O or N

dipole–dipole bond weak bonding caused by the positive end of one dipole attracting the negative end of another dipole

electronegativity the electron-attracting power of an atom

dipole unequal sharing of electrons between atoms in a molecule

on Resources

Video eLesson Hydrogen bond between two water molecules (eles-2483)

10.3.2 Melting and boiling temperatures of water

The melting and boiling temperatures of a compound depend on the strength of the intermolecular bonding between the molecules. Water has relatively high melting and boiling points compared with similarly sized molecular substances due to the presence of strong hydrogen bonds. More heat is required to enable the molecules to gain sufficient kinetic energy to break free of the hydrogen bonds, which are 10 times stronger than dipole–dipole bonds and significantly stronger than dispersion forces.

The **boiling temperature** of a substance is the temperature at which it changes from a liquid to a gas. An input of energy is required to cause this change of state. Energy must be provided in this **endothermic** process to enable the molecules to gain sufficient energy to break the intermolecular bonds. The stronger the intermolecular bonds, the higher the boiling temperature. It is well known that the boiling temperature of pure water when measured at sea level is 100 °C. Water boils at lower temperatures at higher altitudes due to the decrease in atmospheric pressure. It boils at higher temperatures if, for example, it contains dissolved salt.

The **melting temperature** of ice is 0 °C, but the presence of hydrogen bonding between water molecules significantly affects what happens at low temperatures. Most substances become denser as their temperature decreases because the particles move closer together. Water, however, is quite different.

boiling temperature temperature at which a liquid boils and turns to vapour

endothermic chemical process or reaction that absorbs energy

melting temperature temperature at which state changes from solid to liquid

exothermic chemical process or reaction that releases energy through light or heat

crystalline lattice structure of particles that are held together in an ordered, three-dimensional arrangement

Changing states

Water molecules in liquid form have more kinetic energy and move more quickly than water molecules in ice. To change state from liquid to solid, energy must be removed — making this an **exothermic** process (figure 10.13). The density of water is greatest at 4 °C (see the next section). As the temperature continues to decrease, the denser water at 4 °C sinks and eventually the temperature at the surface becomes 0 °C and freezes. Ice forms in an open, hexagonal **crystalline lattice** that places the water molecules further apart than occurs in the liquid state (figure 10.14). Each water molecule is linked to four other water molecules by hydrogen bonding in a tetrahedral arrangement (figure 10.15). This results in water expanding upon freezing, because it is less dense as a solid compared to a liquid. Ice, therefore, floats on water.

FIGURE 10.13 Changing state from a liquid to solid water is an exothermic reaction process because energy is given out.

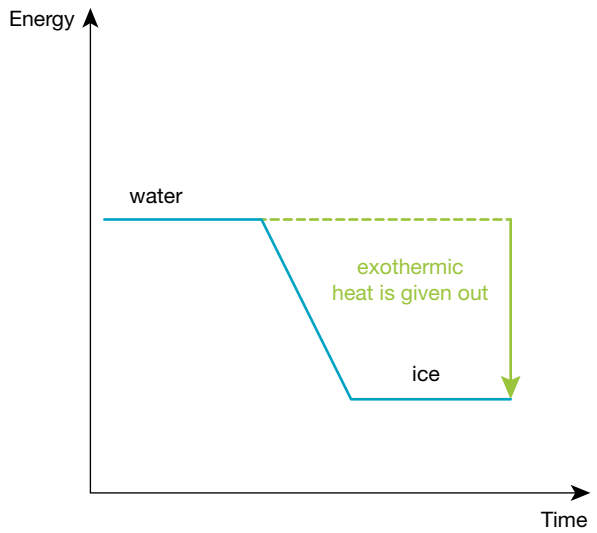
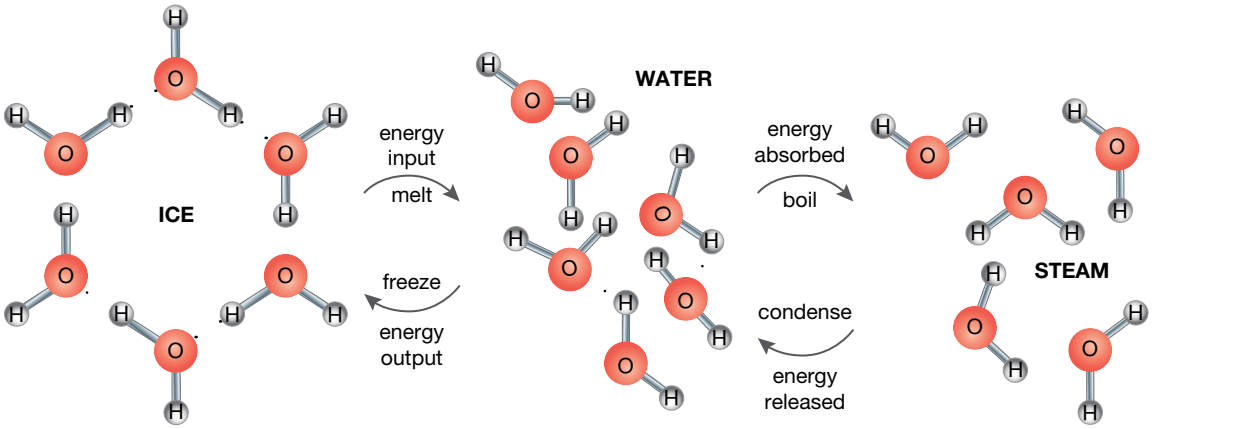


FIGURE 10.14 Arrangement of water molecules in its three different states

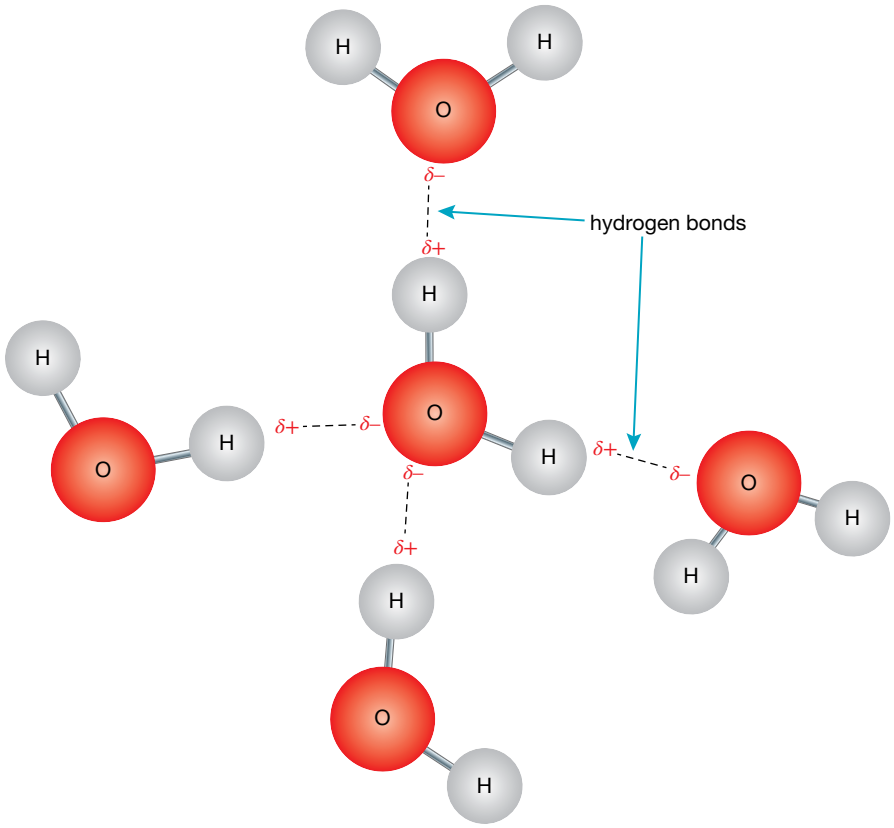


When in solid form (ice), water molecules join together to form hexagonal crystalline structures.

Water molecules move too quickly when in liquid form to form regular structures; however, they still remain attached.

In gaseous form, water molecules move very quickly and are far apart from each other, preventing the formation of significant intermolecular forces.

FIGURE 10.15 Each water molecule forms tetrahedral hydrogen bonds to four other water molecules.



Density of water and ice at different temperatures

The **density** of a substance is a measure of the amount of mass that is contained in a certain volume of that substance. It is calculated by using the formula $d = \frac{m}{V}$ (where d is the density of the substance, m is the mass of the substance measured in grams (g) or kilograms (kg) and V is its volume measured in litres (L), millilitres (mL) or cubic metres (m^{-3})). Typical units are g mL^{-1} and kg m^{-3}.

density physical property found by dividing the mass of an object by its volume

tlvd-3378

SAMPLE PROBLEM 1 Calculating the mass of a given volume of water

If the density of water at 25.0 °C is 0.997 g mL^{-1}, calculate the mass, in grams, of 200.0 mL of water.

THINK	WRITE
1. State the given information and identify the quantity required.	$d(H_2O) = 0.997$ g mL^{-1} $V(H_2O) = 200.0$ mL $m(H_2O) = ?$
2. Recall the formula to calculate density is $d = \frac{m}{V}$ and rearrange the formula to find the mass.	$d = \frac{m}{V}$ $m = dV$
3. Substitute known values to calculate mass. **TIP** Always ensure the answer is given to the least number of significant figures and the units are correct.	$m(H_2O) = 0.997 \times 200.0$ $= 199$ g

PRACTICE PROBLEM 1

Calculate the mass, in grams, of 150 mL water at 25.0 °C given that the density of water is 0.997 g mL^{-1}.

The density of water varies across its three states due to expansion and contraction, and this depends on the water's temperature. Table 10.5 shows the density of ice and water at selected temperatures, and this is demonstrated graphically in figure 10.16.

TABLE 10.5 Density of ice and water at selected temperatures

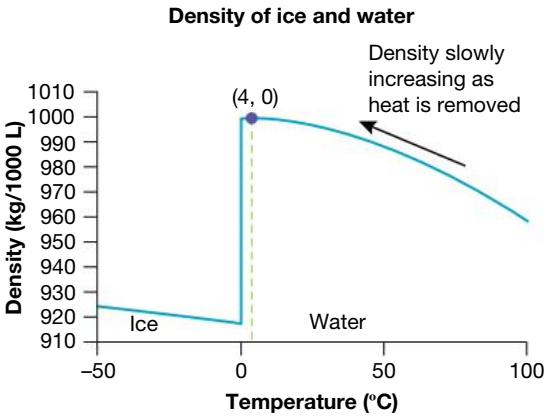
Ice		Liquid water	
Temperature (°C)	Density (ρ) (g mL)	Temperature (°C)	Density (ρ) (g mL)
−50	0.922	0	0.999 84
−40	0.921	2	0.999 94
−30	0.920	4*	0.999 98
−20	0.919	6	0.999 94
−10	0.919	8	0.999 85
0	0.916	10	0.999 7
		20	0.998 2
		25	0.997 1
		30	0.995 7

* Water has its highest density at 3.98 °C.

An inspection of table 10.5 and figure 10.16 reveals two interesting facts about water:

1. Water's maximum density is not at its freezing point but at 4 °C (see table 10.5). This means that water that is approaching 0 °C, and is therefore close to freezing, rises to the top because it is less dense than the surrounding water.
2. Ice has a density that is significantly less than water at the temperatures shown. This is the reason that ice floats on water.

FIGURE 10.16 The density of water changes as it approaches freezing point.



Why ice floats

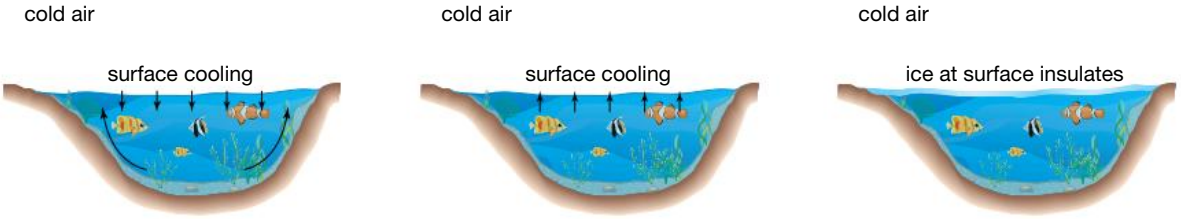
Water has an unusual property: its solid form, ice, is less dense than its liquid form. The molecules in the solid form are arranged in a more structured hexagonal way. This means that ice floats on water and that water freezes from the surface down.

Biological importance of less dense ice

The fact that water freezes with a layer of ice across the top is attributable to both these properties about its density. The water that is about to freeze is at the surface and, when it does turn into ice, its lower density keeps it there. The layer of ice that forms then acts as an insulator, preventing the water below from freezing, and allows aquatic life to survive in sub-zero conditions. This decrease in density from liquid to solid water means that the volume of ice is greater than the corresponding volume of water. This explains why bottles of water crack when frozen. It also results in an important weathering phenomenon when water freezes after seeping into cracks in rocks. As it freezes and expands, enough pressure can build up to crack the rocks and contribute to their eventual erosion.

FIGURE 10.17 When ice forms on the surface of water, it acts as an insulator, preventing the water below from freezing.

FIGURE 10.18 The density of water is greatest at 4 °C.



At 4 °C, water sinks and is replaced by water over 4 °C.

Water below 4 °C rises to the surface.

Water at the surface freezes, but deeper water is 0–4 °C so aquatic life survives.

elog-1802

EXPERIMENT 10.1 online only

A water modelling exercise

Aim

To prepare a bonding model for the water molecule, and to use the model to explain the intermolecular bonding between water molecules in the liquid and solid states

10.3.3 Melting point and boiling point of group 16 hydrides

The impact of hydrogen bonding in water can be seen more clearly when we compare melting and boiling temperatures for other **hydrides** in group 16. These temperatures are demonstrated in table 10.6 and figure 10.19.

hydride binary compound containing hydrogen

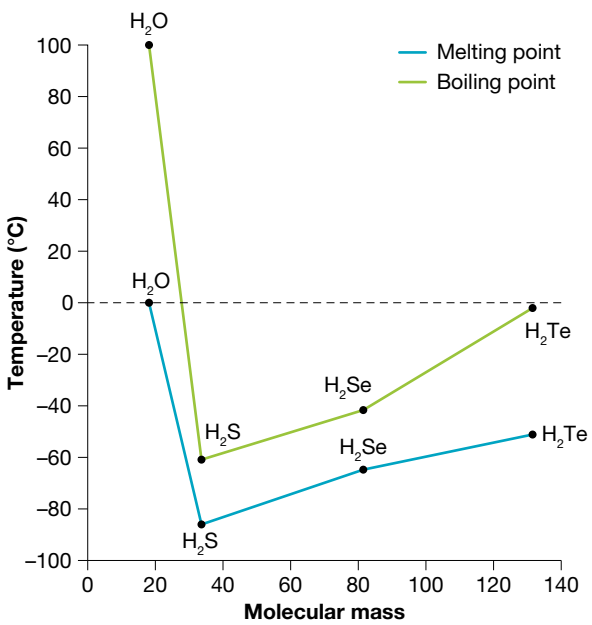
TABLE 10.6 Melting and boiling temperatures of group 16 hydrides

Hydride	Formula	Molar mass (g mol–1)	Melting temperature (°C)	Boiling temperature (°C)
Water (hydrogen oxide)	H_2O	18	0	100
Hydrogen sulfide	H_2S	34	–86	–61
Hydrogen selenide	H_2Se	81	–65	–42
Hydrogen telluride	H_2Te	129.6	–51	–2

If either melting or boiling temperatures are compared, water is an anomaly in the group 16 hydrides. If water is ignored, an obvious trend can be seen of increasing temperatures within the group. The larger the molecule, the higher its melting and boiling temperatures. This is exactly what would be expected if dispersion forces were the only significant intermolecular force present. The values for water, however, indicate that the intermolecular forces are much stronger than expected for just dispersion forces alone.

This is because hydrogen bonding is also present as an additional, much stronger, force between the molecules. Water contains hydrogen–oxygen bonds, which, as mentioned previously, allow for the formation of comparatively strong permanent dipoles called hydrogen bonds.

FIGURE 10.19 Melting and boiling temperatures of group 16 hydrides



10.3 Activities

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10.3 Quick quiz on | 10.3 Exercise | 10.3 Exam questions

10.3 Exercise

1. Why does water have a higher boiling point than other group 16 hydrides?
2. Arrange the following interactions from weakest to strongest: covalent bonding, dispersion forces, hydrogen bonding and dipole–dipole forces.
3. **a.** Describe the structure of a water molecule.
   b. A water molecule is described as polar. Explain this statement.
   c. Draw two water molecules showing correct orientation, and label the intramolecular and intermolecular bonds.
4. Changing state from solid to liquid is an endothermic process. Explain whether changing state from gas to liquid is an exothermic or endothermic process.
5. Explain whether covalent bonds or hydrogen bonds are broken when liquid water evaporates into steam.
6. **a.** What variables are required to find density?
   b. State a unit for density.
   c. Is the density of water at 5 °C higher or lower than the density of water at 25 °C?
7. Use the density of water at 25 °C provided in table 10.5 to find the mass (in g) of 65.5 mL of water.
8. Explain why ice floats on water.
9. Use table 10.6 and the graph in figure 10.19 to predict possible melting and boiling points for hydrogen polonide, H_2Po, molar mass (211 g mol^{-1}).
10. The boiling points of water, H_2O, and hydrogen sulfide, H_2S, are 100 °C and –60 °C respectively. Account for the large difference in the two temperatures.

10.3 Exam questions

Question 1 (1 mark)

MC Which of the following equations represents condensation of water on a cold surface?

A. $H_2O(l) \rightarrow H_2O(g)$
B. $H_2O(g) \rightarrow H_2O(aq)$
C. $H_2O(s) \rightarrow H_2O(g)$
D. $H_2O(g) \rightarrow H_2O(l)$

Question 2 (1 mark)

MC Hydrogen bonding plays a significant role in the properties of water. Which of the following is not a consequence of hydrogen bonding?

A. High boiling temperature compared with similar sized molecules
B. Very high decomposition temperature
C. Expansion of water as it freezes
D. High specific heat capacity

Question 3 (2 marks)

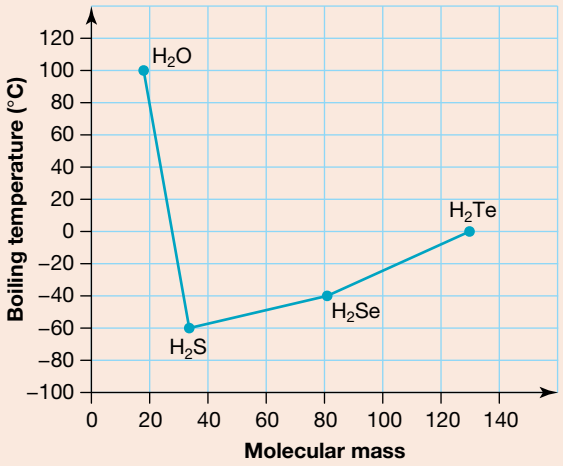
Define the term ‘hydrogen bonding’.

Question 4 (1 mark)

MC Which of the following molecules does not exhibit hydrogen bonding within its solid or liquid state?

A. NH_3 **B.** HNO_3 **C.** SO_2 **D.** H_2SO_4

Question 5 (2 marks)



Use the graph to account for the following.

a. The sharp fall in boiling temperature for H_2S compared with H_2O **(1 mark)**

b. The rise in boiling temperature from H_2S to H_2Te **(1 mark)**

More exam questions are available in your learnON title.

10.4 Heat capacity and latent heat

KEY KNOWLEDGE

- Explanation of the anomalous properties of H_2O (ice and water) with reference to hydrogen bonding: specific heat capacity of water including units and symbols
- The relatively high latent heat of vaporisation of water and its impact on the regulation of the temperature of the oceans and aquatic life

Source: Adapted from VCE Chemistry Study Design (2023–2027) extracts © VCAA; reproduced by permission.

10.4.1 Specific heat capacity

The **specific heat capacity** (c) of a substance is the amount of energy needed to raise the temperature of one gram of the substance by one degree Celsius. It is, therefore, a measure of how easy or how difficult it is to change the temperature of a substance. Water has a high specific heat capacity ($4.18\ J\ g^{-1}\ °C^{-1}$ or $4.18\ J\ g^{-1}\ K^{-1}$), which means that its temperature is difficult to change — that is, more heat energy is required to change ocean water temperature by one degree Celsius. This is why the beach is popular on hot days. The water can absorb heat from its warmer surroundings, which in turn lowers the air temperature by the seashore.

FIGURE 10.20 Water has cooling properties.

specific heat capacity energy (measured in joules) needed to raise the temperature of 1 g of a pure substance by 1 °C

Calculating the energy required to raise the temperature of a substance

The energy required to raise the temperature of a particular mass of a substance can be calculated using:

$q = mc\Delta T$

where:

q is the energy required in joules (J)

m is the mass in grams (g)

c is the specific heat capacity of the substance (J g^{-1} $°C^{-1}$ or J g^{-1} K^{-1})

ΔT is the temperature change (°C or K).

TIP: This formula can be found in Table 3 of the VCE Chemistry Data Book. Note that heat energy is calculated in joule (not kilojoule) and mass in grams. The symbol Δ (the upper case Greek letter delta) is used to represent a change in a quantity — in this case, a change in temperature.

Water has a high heat capacity

Water has a high heat capacity due to the large amount of hydrogen bonding between water molecules.

Because temperature is relative to the kinetic energy of the atomic particles (particle speed), large amounts of energy (for example, heat energy) are required to overcome the strength of the hydrogen bonding, to increase the particles' speed and, hence, heat water.

- Water is an excellent insulator of heat.
- With a heat capacity almost five times higher than soil or rock (see table 10.7), water retains heat five times more effectively than land, but may take five times longer than land to heat up. This means water can absorb large amounts of heat energy with only a minimal increase in temperature.

You will also notice that the beach sand is much hotter than the water, despite receiving approximately the same amount of heat energy. We can say that the sand has a lower specific heat capacity than water because less energy is required to heat one gram of sand by one degree Celsius compared to water (table 10.7).

TABLE 10.7 Specific heat capacity of some common substances

Substance	Specific heat capacity (c) (J g^{-1} $°C^{-1}$)
Water	4.18
Ethanol	2.5
Vegetable oil	2.0
Soil (wet)	1.5
Aluminium	0.90
Concrete	0.88
Glass	0.84
Sand (dry)	0.80
Soil (dry)	0.80
Copper	0.39
Gold	0.13
Lead	0.13

Calculations using specific heat capacity

If you were to heat 1 gram of all the substances from table 10.7 with the same amount of heat energy, in which substance would you record the greatest temperature change?

It is possible to measure the energy transferred to water when it is being heated using the formula provided.

tlvd-0558

SAMPLE PROBLEM 2 Calculating the heat energy to raise the temperature of water

Calculate how much heat energy (in kJ) is needed to raise 250.0 grams of water from 20.0 °C to 100.0 °C. The specific heat capacity of water is 4.18 J g^{-1} $°C^{-1}$.

THINK	WRITE
1. State the given information and identify the quantity required.	$m(H_2O) = 250.0$ g $c(H_2O) = 4.18$ J g^{-1} C^{-1} $\Delta T(H_2O) = 100.0\,°C - 20.0°C$ $= 80.0\,°C$ $q(H_2O) = ?$
2. Recall the formula to calculate energy is $q = mc\Delta T$.	$q = 250.0 \times 4.18 \times 80.0\,°C$ $= 8.36 \times 10^4$ J
3. Convert J to kJ by dividing by 1000. Check that significant figures and the unit are correct.	$q = \frac{8.36 \times 10^4}{1000}$ $= 83.6$ kJ

PRACTICE PROBLEM 2

Calculate the amount of heat energy (in kJ) required to warm 1000 grams of water from fridge temperature at 4.0 °C to room temperature, 25.0 °C.

elog-1803

tlvd-0627

EXPERIMENT 10.2 online only

Measuring specific heat capacity

Aim

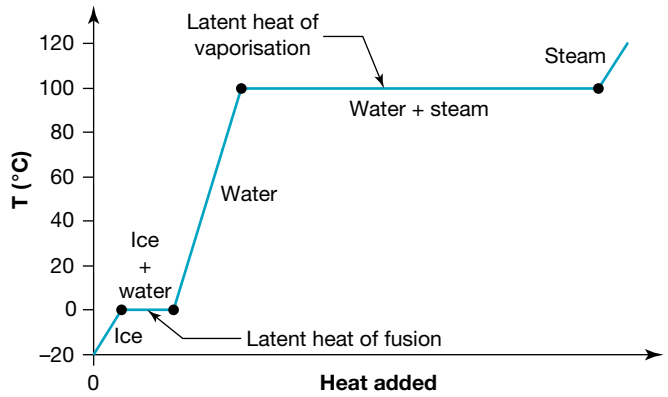
To measure and compare the specific heat capacity of a number of different liquids

10.4.2 Latent heat

When a substance changes phase (for example, when ice melts and becomes water, or water becomes water vapour), energy is needed. This energy comes from the surrounding atmosphere. The energy needed to change the phase of a substance at its melting or boiling temperature is called the **latent heat** of that substance. When water is changing state, its temperature remains constant, so no change occurs in the kinetic energy of the water; the energy comes from or goes to the potential energy stored in the hydrogen bonds between the molecules. That is why it is called latent, or hidden, heat. When ice is melting, the energy supplied is used to break bonds between the molecules and the temperature only begins to rise again once the change of state is complete. The symbol for latent heat is *L*. Common units are kJ mol^{-1} and J kg^{-1}.

latent heat energy needed to change state

FIGURE 10.21 The latent heat of water. Energy is needed to break the bonds between the water molecules. This causes the temperature to remain constant even though the water is being heated.



- The **latent heat of fusion (L_f)** of water is the amount of energy needed to change a fixed amount of water from a solid to liquid phase at 0 °C. To change a fixed amount of water from a solid to liquid phase at 0 °C:
  - The energy required (the latent heat of fusion) breaks the hydrogen bonds between the ice molecules in the crystal lattice.
  - This allows the water molecules to move around more freely in a liquid state.
  - The latent heat of fusion of water is 6.02 kJ mol^{-1}, meaning 6.02 kJ of energy must be supplied to change each mole of water from solid to liquid.
- The **latent heat of vaporisation (L_v)** of water is the amount of energy needed to change a fixed amount of water from a liquid to a gas at 100 °C. To change a fixed amount of water from a liquid to a gas at 100 °C:
  - The energy required (the latent heat of vaporisation) breaks the remaining forces holding the water molecules together.
  - This allows the molecules to move around freely as a gas.
  - The latent heat of vaporisation is considerably greater than the latent heat of fusion, because the molecules must be completely separated from each other when changing from liquid to gas and not just slightly moved apart as when moving from the crystal lattice in ice to liquid water.
  - The latent heat of vaporisation of water is 40.7 kJ mol^{-1}, meaning 40.7 kJ energy must be supplied to vaporise one mole of water.

latent heat of fusion, (L_f) energy needed to change a fixed amount from the solid to the liquid state

latent heat of vaporisation, (L_v) energy needed to change a fixed mass from the liquid to the gaseous state

Biological importance of latent heat

The relatively high latent heat of vaporisation of water is very useful in keeping organisms cool through perspiration. In humans, when a person perspires, droplets of liquid water on the skin evaporate. This change from liquid water to water vapour requires a large amount of heat energy. This is energy absorbed (or taken) from the body heat, making perspiration an efficient mechanism for removing unwanted heat and, therefore, cooling the body down. Interestingly, if a lot of water vapour is already in the air (as is the case when humidity is high — imagine being in Darwin during the wet season!), this process is slow, so people often feel hot and uncomfortable on such days. Condensation, the reverse process of evaporation, requires the release of 40.7 kJ of energy per mole of water.

FIGURE 10.22 Sweating uses heat from the body to evaporate the moisture.

The higher the latent heat of a substance, the greater the amount of energy it needs when it changes state.

on Resources

Video eLesson Phase changes (eles-2487)

Calculations using latent heat

The heat energy required to change the state of a substance is given by the formula:

$$q = nL$$

where:

q is the heat energy required or released (kJ)
n is the amount of substance (mol)
L is the latent heat.

TABLE 10.8 Latent heat values

Substance	Melting point °C	Latent heat of fusion (L_f) kJ mol–1	Boiling point °C	Latent heat of vaporisation (L_v) kJ mol^{-1}
Ethanol, C_2H_5OH	–114	5.02	78.4	38.6
Methanol, CH_3OH	–97.6	3.18	64.7	35.2
Methane, CH_4	–182	0.94	–165	8.18
Ammonia, NH_3	–77.7	5.65	– 33.3	23.4
Water, H_2O	0	6.01	100	40.7

tlvd-0557

SAMPLE PROBLEM 3 Calculating the heat energy to melt ice

Calculate the heat energy (in kJ) required to fully convert 100 grams of ice to liquid water.

THINK

1. State the given information and identify the quantity required.
2. To use the latent heat formula, $q = nL_f$, first calculate the number of mole of water, n, where $n = \frac{m}{M}$
3. Find $q = nL_f$. The latent heat of fusion for water can be found in table 10.8, $L_f = 6.01$ kJ mol^{-1}. Check significant figures and units are correct.

WRITE

$m(H_2O) = 100$ g

$q(H_2O) = ?$

$$n(H_2O) = \frac{m}{M} = \frac{100}{(2 \times 1.0 + 16.0)} = 5.56 \text{ mol}$$

$$q(H_2O) = 5.56 \times 6.01 = 33.4 \text{ kJ}$$

PRACTICE PROBLEM 3

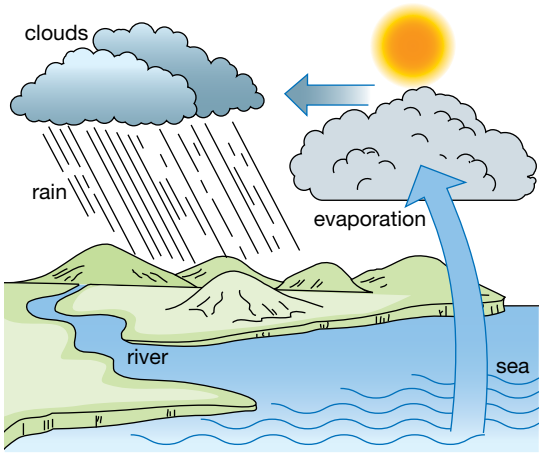
Calculate the heat energy, in kJ, required to fully convert 100 grams of liquid water to steam.

10.4.3 Impacts on ocean temperature and ocean life

As discussed in subtopic 10.2, 96.5 per cent of Earth's water is stored in the ocean. Together with the heat energy from the sun, this creates the driving force of our global weather and climate. The oceans act as a buffer to control global temperatures for the following reasons:

- Water has much higher latent heat values than other substances of a similar size owing to the stronger intermolecular hydrogen bonds within its solid and liquid states. This means that water evaporates more slowly than many other liquids.
- The ocean and the atmosphere continuously exchange heat, water vapour and carbon, forming the water cycle and weather patterns (figure 10.23). Due to the high specific heat capacity of water, the oceans can release the heat over a long period of time. On a global scale, short-term extremes in temperature can be evened out as the oceans absorb heat from the atmosphere in the tropics and then release it when ocean currents take this water to the polar regions.

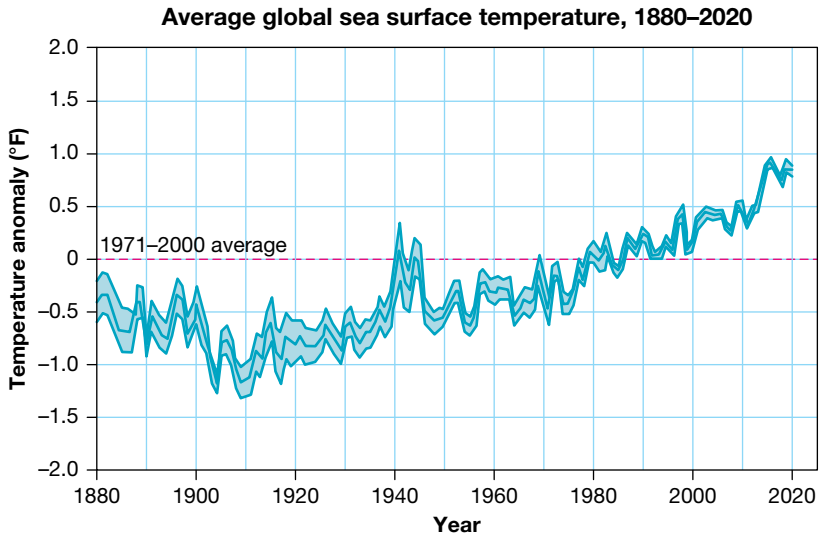
FIGURE 10.23 The relatively high latent heat of vaporisation of water, 40.7 kJ mol^{-1}, means that a large amount of energy from the sun and surroundings is needed for evaporation to occur in the water cycle.



However, human activities — including the burning of fossil fuels, agriculture and land clearing — are increasing the amount of greenhouse gases in the atmosphere (see topic 15). This results in increased heat being trapped within Earth's atmosphere through the enhanced greenhouse effect. As average global temperatures rise, the oceans are struggling to maintain stable atmospheric and sea temperatures. At the time of writing, 2016 was the warmest year on record for the world's average sea surface temperatures (see figure 10.24a, which also provides additional evidence for climate change).

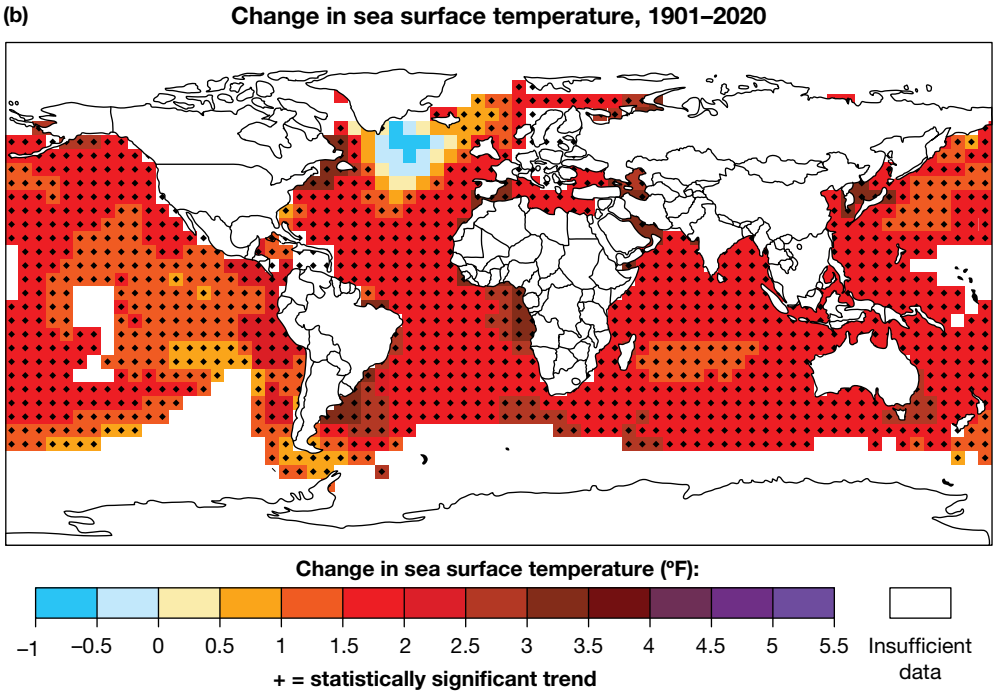
FIGURE 10.24 (a) United States National Oceanic and Atmospheric Administration (NOAA) data highlighting the average global sea surface temperatures for the period 1880 to 2020, and (b) a heat map produced by a collaboration between the Intergovernmental Panel on Climate Change and NOAA, focusing on the global distribution of changes in sea surface temperature.

(a)



Data source: NOAA (National Oceanic and Atmospheric Administration). 2021. Extended reconstructed sea surface temperature (ERSST.v5). National Centers for Environmental Information. Accessed February 2021. www.ncdc.noaa.gov/data-access/marineocean-data/extended-reconstructed-sea-surface-temperature-ersst.

(b)



Data sources:

- IPCC (Intergovernmental Panel on Climate Change). 2013. Climate change 2013: The physical science basis. Working Group I contribution to the IPCC Fifth Assessment Report. Cambridge, United Kingdom: Cambridge University Press. www.ipcc.ch/report/ar5/wg1.
- NOAA (National Oceanic and Atmospheric Administration). 2021. NOAA Merged Land Ocean Global Surface Temperature Analysis (NOAAGlobalTemp). Accessed March 2021. www.ncdc.noaa.gov/data-access/marineocean-data/noaa-global-surface-temperature-noaaglobaltemp.

Source: NOAA, February 2021, https://www.epa.gov/climate-indicators/climate-change-indicators-sea-surface-temperature.

Warming sea temperatures are affecting the following areas:

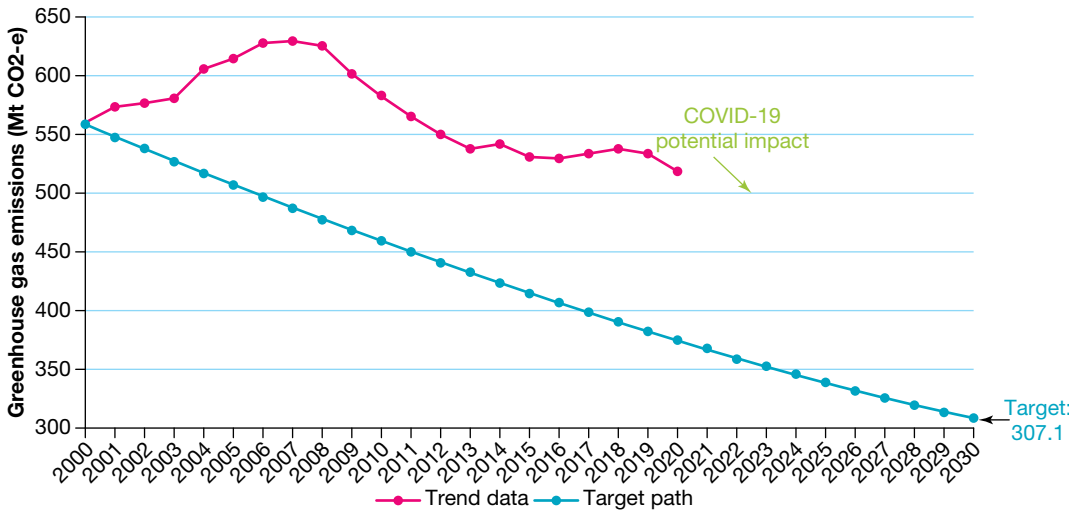
- *Rainfall patterns:* a redistribution of heat spots on the ocean surface water is directly influencing a change in rainfall patterns. This has serious implications, not only for ecosystems but also for human activity such as agriculture.
- *Increased severity of storms:* increased sea surface temperatures increase the amount of evaporation (atmospheric water vapour) over the oceans, leading to increased risk of heavier rainfall. With additional energy in the environment, storms are also more severe.
- *Impacts on ecosystems and biological diversity:* different aquatic organisms have different optimal temperatures for survival and, as such, spatial changes in ocean temperatures are starting to have an impact on biological diversity through shifting oceanic ecosystems. For example, plankton are at the bottom of the marine food chain yet make up 90 per cent of the living mass in oceans. Phytoplankton play a significant role in regulating the climate as they absorb carbon dioxide and release oxygen via photosynthesis. With ocean temperatures rising, the plankton are migrating to colder waters and, consequently, the whole food chain must shift too (see figure 10.25). This also affects food and job security for humans.

FIGURE 10.25 A phytoplankton bloom seen from space off the coast of France

Sustainable Development Goal 13 — Climate action

The 13th goal of the United Nations Sustainable Development Goals (SDGs) is to 'take urgent action to combat climate change and its impacts' by 2030. This goal encompasses five targets with nine performance indicators, with a large focus on ensuring climate strategies are integrated into national policies while also increasing public awareness through national education policies and curriculums. A major indicator in monitoring climate change is the total greenhouse gas emissions per year. At the time of writing, globally these are continuing to increase. Australia has seen only slight improvements in emissions since 2000, and has been classified as 'breakthrough needed' to be able to reach our emissions target by 2030 (see figure 10.26).

FIGURE 10.26 Australia total green gas emissions have only slightly declined since 2000.



Source: www.sdgtransformingaustralia.com/explore-by-goal/#/1254/1327//, based on data from the Australian Department of Energy and Environment, Quarterly Emissions Update; National GHG Inventory.

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Weblink Sustainable Development Goal 13

10.4 Activities

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10.4 Quick quiz on | 10.4 Exercise | 10.4 Exam questions

10.4 Exercise

1. What is the name of the energy required to change liquid to gas and vice versa without any change in temperature?
2. Heat energy is continually supplied to keep water boiling at 100 °C. Explain why the temperature remains constant.
3. Water has a specific heat capacity of 4.18 J g^{-1} $°C^{-1}$. Explain what this means.
4. On a hot day at the beach, the sand feels very warm, but the sea is cool. Explain why.
5. Refer to table 10.7 and state whether greater energy is required to raise the temperature of 10 g of copper or 10 g of gold by 1.00 °C.
6. Calculate the amount of energy needed to change the temperature of 300 g of water by 30.2 °C (specific heat capacity of water is 4.18 J g^{-1} $°C^{-1}$).
7. At 25.0 °C, 50.0 g of water absorbs 6274 J of heat. What is its final temperature?
8. A 5.10 kg lump of copper is given 30.0 kJ of energy and it rises in temperature by 15.0 °C. Calculate the specific heat capacity of copper.
9. Calculate the amount of energy needed to change the temperature of 250 g of aluminium from 20 °C to 45 °C. (The specific heat capacity of aluminium is 0.90 J g^{-1} $°C^{-1}$.)
10. Calculate the amount of heat energy (in kJ) that must be released from 152 g of liquid water to convert it to ice.
11. 120 kJ of energy is used to heat 500 mL of water initially at 23.0 °C. Assuming that all the energy was absorbed by the water, what is the final temperature of the water?
12. 50.0 g samples of oil and water, both at an initial temperature of 20.0 °C, were poured into identical beakers and heated together on a hotplate. At the end of 5 minutes, the temperature of the water was 51.4 °C and the oil was 85.6 °C. What is the heat capacity of the oil?

10.4 Exam questions

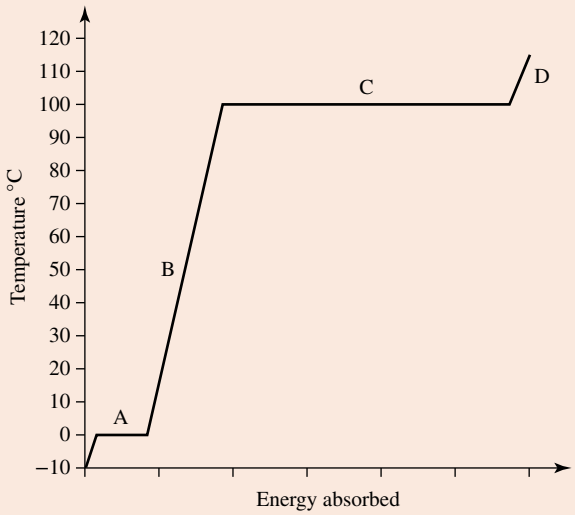
Question 1 (1 mark)

MC Water has a high latent heat of vaporisation. What happens during this process?

A. Covalent bonds within water molecules are broken.
B. Hydrogen bonds within water molecules are broken.
C. Covalent bonds between water molecules are broken.
D. Hydrogen bonds between water molecules are broken.

Question 2 (1 mark)

MC The provided graph shows the temperature of a sample of ice at –10 °C heated constantly until the temperature reaches 120 °C. Which of the letters (A to D) represents water absorbing the latent heat of vaporisation?



A. A
B. B
C. C
D. D

Question 3 (1 mark)

MC Liquid nitrogen consists of non-polar N_2 molecules. Compared with that of water, what should the latent heat of vaporisation of 1 mole of liquid nitrogen be?

A. Higher due to the strong triple bond between nitrogen atoms in N_2 molecules
B. Lower because the intermolecular bonding between N_2 molecules is weaker than that of water
C. Equal to that of water since they are the same physical process
D. Unable to be classed as higher or lower without further information

Question 4 (3 marks)

a. Explain, by referring to the specific heat capacity of water, why coastal regions do not experience the extremes of temperature experienced by inland regions. **(1 mark)**

b. Explain, with reference to the latent heat of vaporisation, why a rise in surface sea temperatures results in increasingly severe storms. **(2 marks)**

Question 5 (6 marks)

A pair of students carried out an experiment to determine the heat capacity of a sample of cooking oil. They designed their experiment so that the heat capacity of the oil could be found by comparing the temperature changes of the oil with those of a sample of water heated in a similar manner.

a. Outline a procedure that could be used for this experiment, including the specific types of measurements that need to be taken, as well as the appropriate procedure for a valid result. **(5 marks)**

b. Give one safety precaution that would be required for this experiment. **(1 mark)**

More exam questions are available in your learnON title.

10.5 Review

Hey students! Now that it's time to revise this topic, go online to:

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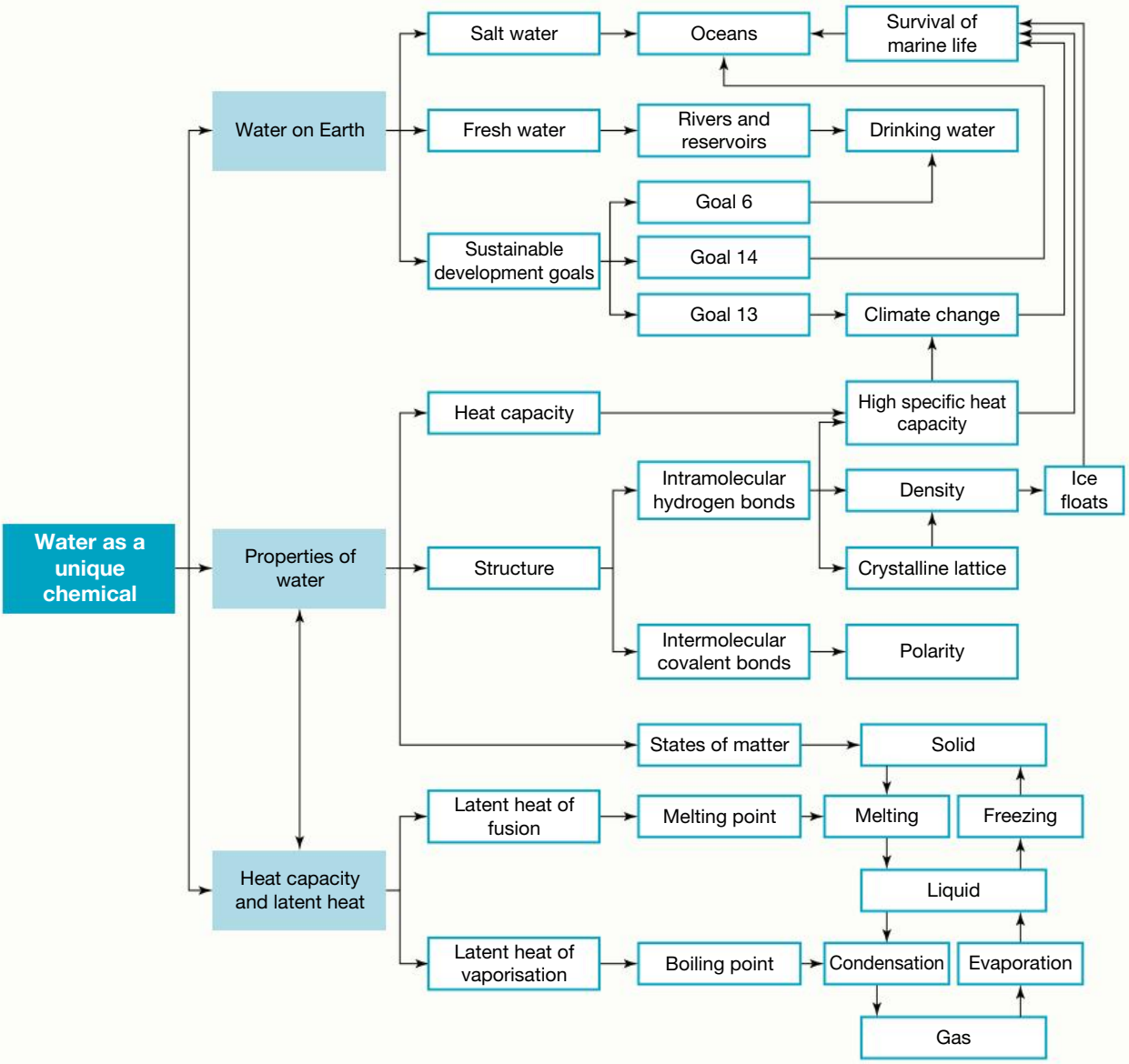
10.5.1 Topic summary



10.5.2 Key ideas summary

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10.5.3 Key terms glossary

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Solutions	Solutions — Topic 10 (sol-0809)
Practical investigation eLogbook	Practical investigation eLogbook — Topic 10 (elog-1615)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 10 (doc-37085) Key ideas summary — Topic 10 (doc-37086)
Exam question booklet	Exam question booklet — Topic 10 (eqb-0091)

10.5 Activities

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10.5 Review questions

1. What property of water causes tumble dryers to use so much energy to dry clothes?
2. a. What is specific heat capacity?
   b. Explain how water is unusual with respect to this physical property.
   c. Give an example where this property is used.
3. Refer to the graph in figure 10.19 and explain why the boiling points of group 16 hydrides decrease and then increase.
4. a. The temperature of 435 g of water changed from 33.5 °C to 86.4 °C. How much heat energy (in kJ) did this sample absorb?
   b. When 925 J of heat energy is added to a mass of water, the temperature changed by 5.00 °C. Calculate the mass of water used.
5. 6.10 mole of water at 12.0 °C is heated until its temperature is 37.0 °C. Calculate the amount of heat energy, in kJ, needed to cause this rise in temperature.
6. Cells contain mainly water. If lettuce leaves are placed in the freezer and then defrosted, they become soft and soggy. Explain this observation.
7. What is the final temperature of a cup (250 mL) of water, initially at 22.0 °C, after passing 50.5 kJ of heat energy into it?

8. Three metal objects of equal mass were heated together on the same hotplate. The specific heat capacities of the three metals are given in the provided table.

Metal	Specific heat capacity (J g^{-1} °C^{-1})
Silver	0.233
Aluminium	0.900
Copper	0.386

Place the metals in order from hottest to coolest after 10 minutes had passed.

9. a. Use table 10.8 to calculate the quantity of heat energy (in kJ) given out when 220 g of steam at 100 °C condenses.
   b. Why is it more dangerous to receive a burn by steam at 100 °C than one by water at 100 °C?
10. a. What is latent heat?
    b. The latent heat of fusion for water is 6.01 kJ mol^{-1} and the latent heat of vaporisation is 40.7 kJ mol^{-1}. Explain the large difference between these two values.
    c. i. Explain how the high latent heat of vaporisation of water affects the regulation of ocean temperature.
    ii. Give two examples of negative impacts of rises in sea temperatures.

10.5 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC The sixth goal of the United Nations SDGs is the availability and sustainable management of water and sanitation for all people by 2030.

In 2020, how many people still did not have access to basic drinking water?

A. 387 million
B. 1 billion
C. 771 million
D. 52 million

Question 2

MC In 2020, what was Australia's progress to achieve SDG 6.4 Water consumption per capita rated as?

A. Breakthrough needed
B. Off track
C. Needs improvement
D. On track

Question 3

MC Which one of the following types of bonding would be least likely to be present in a sample of water?

A. Ionic
B. Covalent
C. Dispersion
D. Hydrogen

Question 4

MC Which of the following processes require the absorption of heat?

A. Evaporation
B. Condensation
C. The transition from a gas to a liquid
D. The transition from a liquid to a solid

Question 5

MC The density of water is 0.997 gmL^{-1}, so what will the mass of 100 mL of water be?

A. 0.997 grams
B. 9.97 grams
C. 99.7 grams
D. 997 grams

Question 6

MC What is the latent heat of fusion?

A. The amount of heat required to raise the temperature of a substance by 1 °C
B. The amount of heat required to increase the temperature of a 1 g of a substance by 1 °C
C. The amount of heat required to change the state of a substance from solid to liquid without any change in temperature
D. The amount of heat required to change the phase of a substance from liquid to gas without any change in temperature

Question 7

MC Why does water have a high heat capacity?

A. It can boil at higher temperatures than other liquids.
B. It can form hydrogen bonds with any compound.
C. It can absorb large amounts of heat before the temperature changes.
D. It has a low density as a solid.

Question 8

MC What property of water helps organisms to survive in the ocean even when the temperature is below the freezing point of water?

A. Water has a low melting point.
B. Ice is denser than water.
C. Water has a high heat capacity.
D. Ice is less dense than water.

Question 9

MC What is the amount of heat (in J) required to heat 1.5 L of water from 20.0 °C to 95.0 °C?

A. 75.0 **B.** 470 **C.** 4.7×10^5 **D.** 6.0×10^5

Question 10

MC The latent heat of vaporisation of water is 44.0 kJ per mole (18.0 g) of water. This amount of energy is equivalent to heating how many grams of water by 50 °C?

A. 0.21 **B.** 4.75 **C.** 47.5 **D.** 211

Section B — Short answer questions

Question 11 (7 marks)

The density of ethanol, C_2H_5OH, is 0.789 g mL^{-1} and specific heat capacity is 2.5 J g^{-1} $°C^{-1}$.

a. Calculate the mass of the ethanol that fills a 250.0-mL container. **(2 marks)**

b. Define the term 'specific heat capacity'. **(1 mark)**

c. If an equal amount of energy is used to heat 100 g of water and to heat 100 g of ethanol, explain which substance would experience the higher increase in temperature. **(2 marks)**

d. Calculate the energy (in kJ) required to raise the temperature of 180 g of ethanol by 17 °C. **(2 marks)**

Question 12 (4 marks)

a. 50.0 g samples of oil and water, both at an initial temperature of 20.0 °C, were poured into identical beakers and heated together on a hotplate. At the end of 5 minutes, the temperature of the water was 51.4 °C and the oil was 85.6 °C. What is the heat capacity of the oil? **(2 marks)**

b. Identify the independent and dependent variable in the experiment. **(2 marks)**

Question 13 (10 marks)

Water has a number of unusual properties.

a. Explain what happens to the density of water as the temperature decreases by referring to the changes in arrangement of its molecules. **(5 marks)**

b. Why is this significant for organisms that live in the sea? **(2 marks)**

c. Describe an experiment to compare the effect on density of freezing water and the effect of freezing oil. **(3 marks)**

Question 14 (6 marks)

Explain what is occurring at the molecular level when the following observations are made.

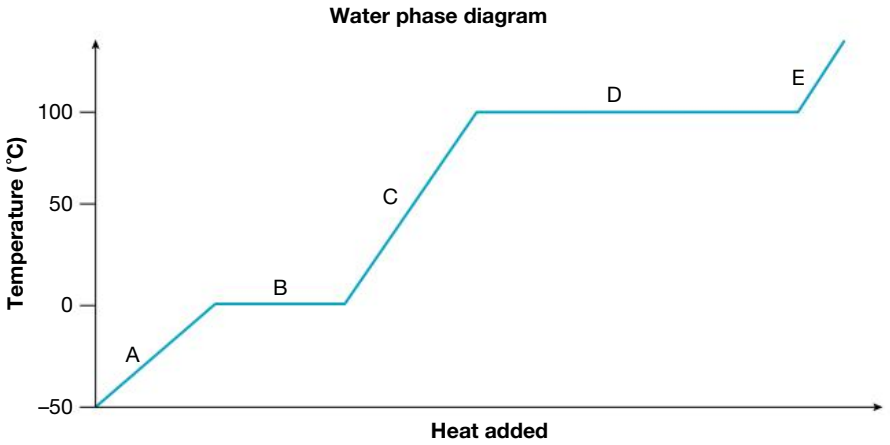
a. A damp cloth placed around a water bottle helps to keep it cool. **(2 marks)**

b. Water droplets form on the sides of a can when it is removed from the fridge. **(2 marks)**

c. Fresh fish is placed on a bed of crushed ice in the supermarket. **(2 marks)**

Question 15 (6 marks)

The following graph indicates what occurs as water is heated over time. Match the letters to the statements, noting that each letter may be used more than once.



a. Molecules have the greatest kinetic energy _____ **(1 mark)**

b. Represents the highest latent heat value _____ **(1 mark)**

c. Both ice and water present _____ **(1 mark)**

d. Latent heat of fusion _____ **(1 mark)**

e. Corresponds to boiling occurring _____ **(1 mark)**

f. Has the greatest amount of hydrogen bonding _____ **(1 mark)**

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11 Acid–base (proton transfer) reactions

KEY KNOWLEDGE

In this topic you will investigate:

Acid–base (proton transfer) reactions

- the Brønsted–Lowry theory of acids and bases, including polyprotic acids and amphiprotic species, and the writing of balanced ionic and full equations with states for their reactions in water
- the distinction between strong and weak acids and strong and weak bases, and between concentrated and dilute acids and bases, including common examples
- neutralisation reactions to produce salts:
  - reactions of acids with metal carbonates and hydroxides, including balanced full and ionic equations, with states
  - types of antacids and their use in the neutralisation of stomach acid
- use of the logarithmic pH scale to rank solutions from most acidic to most basic; calculation of pH for strong acid and strong base solutions of known concentration using the ionic product of water (K_w at a given temperature)
- accuracy and precision in measurement as illustrated by the comparison of natural indicators, commercial indicators, and pH meters to determine the relative strengths of acidic and basic solutions
- applications of acid–base reactions in society; for example, natural acidity of rain due to dissolved CO_2 and the distinction between the natural acidity of rain and acid rain, or the action of CO_2 in forming a weak acid in oceans and the consequences for shell growth in marine invertebrates.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of VCE Chemistry. Experiments and investigations, supported by a practical investigation eLogbook and teacher-led videos, are included in this topic to provide opportunities to undertake investigations and communicate findings.

EXAM PREPARATION

Access exam-style questions and their video solutions in every lesson, to ensure you are ready.

11.1 Overview

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11.1.1 Introduction

The notion of acids and their presence in citrus fruits with their sour taste is familiar to most people, as is acid in the stomach to help break down food. Less commonly known is information about bases. Bases neutralise acids and are also regularly used in everyday life. Baking soda, $NaHCO_3$, otherwise known as sodium hydrogen carbonate, has a multitude of uses from cooking to cleaning. Magnesium hydroxide, $Mg(OH)_2$, is taken as an antacid and ammonia, NH_3, is widely used as a cleaning agent. You can find many acids and bases just in your kitchen alone. But did you know that you can even extract a pH indicator in your very own kitchen to test the pH of these products? Red cabbage and blueberries, for example, have a molecule that changes colour depending on whether it is in an acidic or basic environment. The image that opens this topic is a test tube showing a range of colours indicating a range of pH values. This colour banding is obtained using universal indicator, with a sodium carbonate crystal in the bottom of a test tube of water and acid at the top. This banding develops as the alkali (dark purple) and acid (red) meet. The neutral region is indicated by the narrow green band, while yellow is slightly acidic.

FIGURE 11.1 Red cabbage pH indicator solution: acidic lemon juice (red), alkali sodium bicarbonate (blue) and neutral water (purple)

This topic provides definitions of acids and bases, and explains how acidity is measured, including the use of natural pH indicators from red cabbage, and how acids and bases react. The distinction between strong and concentrated acids and bases in terms of numbers and types of particles present is also explained. The reactions of acids and bases with water and with each other are closely linked to their interactions with water, which were discussed in the previous topic.

LEARNING SEQUENCE

11.1 Overview 406
11.2 Acids and bases 407
11.3 Concentration and strength of acids and bases 415
11.4 The pH scale 420
11.5 Measuring pH 431
11.6 Neutralisation reactions to produce salts 443
11.7 Applications of acid–base reactions in society 452
11.8 Review 456

on Resources

Solutions	Solutions — Topic 11 (sol-0810)
Practical investigation eLogbook	Practical investigation eLogbook — Topic 11 (elog-1616)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 11 (doc-37087) Key ideas summary — Topic 11 (doc-37088)
Exam question booklet	Exam question booklet — Topic 11 (eqb-0092)

11.2 Acids and bases

KEY KNOWLEDGE

- The Brønsted-Lowry theory of acids and bases including polyprotic acids and amphiprotic species, and the writing of balanced ionic and full equations with states for their reactions in water

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BACKGROUNG KNOWLEDGE: Acids and bases

- All **acids** share common properties, including a sour taste, will turn litmus red and neutralise bases.
- All **bases** share common properties, including a bitter taste, will turn litmus blue and neutralise acids. **Alkalis** are bases that are soluble in water.
- Acids and bases react in **neutralisation** reactions to produce a **salt** and water.

TABLE 11.1 Common acids and their uses

Acid	Formula	Occurrence or use
Nitric	HNO_3	Used to etch copper and manufacture fertilisers
Hydrochloric	HCl	Produced in the stomach; used to clean bricks and concrete
Sulfuric	H_2SO_4	Used in car batteries, plastics, insecticides, detergents and pharmaceuticals
Ethanoic	CH_3COOH	Found in vinegar; used in pickling and fermentation of foods
Citric	$C_6H_8O_7$	Found in lemon juice
Carbonic	H_2CO_3	Found in soft drinks
Lactic	$C_3H_6O_3$	Makes milk sour; used in wrinkle-smoothing face cream
Ascorbic	$C_3H_8O_6$	Found in citrus fruits

TABLE 11.2 Common bases and their uses

Base	Formula	Use
Ammonia	NH_3	Fertilisers and detergents
Sodium hydroxide (caustic soda)	NaOH	Soaps and detergents
Sodium carbonate	Na_2CO_3	Manufacture of glass; washing powder and detergents
Calcium oxide (quicklime)	CaO	Bricklayers' mortar
Lead(II) oxide	PbO	House paint (now phased out)
Calcium hydroxide (slaked lime)	$Ca(OH)_2$	Garden lime, plaster and cement
Ammonium hydroxide	NH_4OH	Cleaning agents
Magnesium hydroxide	$Mg(OH)_2$	Milk of magnesia (for treatment of indigestion)

FIGURE 11.2 Caustic soda is another name for sodium hydroxide, a strong base.

acid substance that can donate a proton to a base

base substance that can accept a proton from an acid

alkali soluble base

neutralisation chemical reaction in which an acid and a base react to form a salt and water

salt ionic compound consisting of a metal ion and a non-metal ion, except oxides and hydroxides

Safety with acids and bases

Even so-called weak acids can be dangerous. A concentrated sample of ethanoic acid (acetic acid or vinegar), which is regarded as weak, can be extremely harmful to eyes and mucous membranes. Strong acids such as nitric acid, sulfuric acid and hydrochloric acid are extremely dangerous in concentrated form and should be handled with more care than their more dilute forms. When diluting a concentrated acid, **always** add acid slowly to water. Adding water directly to concentrated acids may cause them to splatter, and severe burns may result. Strong bases such as sodium hydroxide and potassium hydroxide are also very dangerous, particularly in concentrated form and when hot. Acid and base spills on skin should be rinsed with large quantities of running water.

For more information, download the **Properties of acids and bases** document from the Resources panel.

on Resources

Digital document Properties of acids and bases (doc-37664)

11.2.1 Brønsted–Lowry theory of acids and bases

An acid–base theory suggested independently by Brønsted and Lowry in 1923 is still used today. This theory is summarised as follows.

The Brønsted–Lowry theory definition of acids and bases

- An acid is a substance that donates a proton (H^+ ion) to another substance.
- A base is a substance that accepts a proton (H^+ ion) from another substance.

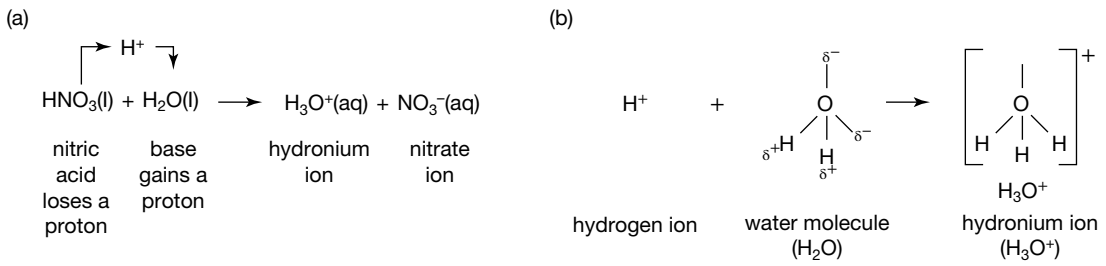
A hydrogen ion, H^+, is formed when an electron is removed from a hydrogen atom, H.

A proton is simply a hydrogen ion, H^+.

A reaction between an acid and a base involves a proton transfer from an acid to a base. In other words, the substance that loses a hydrogen ion, H^+, is the acid and the substance that accepts the hydrogen ion, H^+, is the base. This is illustrated in figure 11.3(a) where nitric acid, HNO_3, loses a proton to water, which creates a **hydronium ion**, H_3O^+, seen in figure 11.3(b). According to the Brønsted–Lowry theory, the water in the equation shown in figure 11.3 is acting as a base because it accepts a proton.

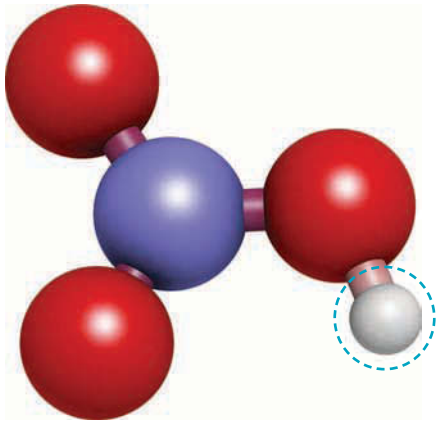
hydronium ion H_3O^+, combination of a proton with a water molecule

FIGURE 11.3 (a) Nitric acid, HNO_3, loses a proton to water, creating a hydronium ion and nitrate ion, and (b) a water molecule accepts a proton to create a hydronium ion, H_3O^+.



The hydronium ion, $H_3O^+(aq)$, is often shortened to $H^+(aq)$ but it is important to realise that a water molecule is usually attached to the proton.

FIGURE 11.4 A nitric acid molecule, with the proton to donate circled.



on Resources

Video eLesson The Brønsted–Lowry theory (eles-2490)

11.2.2 Conjugate acid–base pairs

When chemists want to refer to particles that could be atoms, ions or molecules, the term that is used is **species**. The **conjugate base** of a Brønsted–Lowry acid is the species formed after the proton is donated. In the following example, the conjugate base of the acid HNO_3 is NO_3^-.

The **conjugate acid** of a base is the species formed after a base accepts a proton. The conjugate acid of the base H_2O is H_3O^+.

In this case, the conjugate pairs are HNO_3/ NO_3^- and H_3O^+/H_2O.

$$\underset{\text{acid}_1}{HNO_3(aq)} + \underset{\text{base}_2}{H_2O(l)} \longrightarrow \underset{\text{acid}_2}{H_3O^+(aq)} + \underset{\text{base}_1}{NO_3^-(aq)}$$

The conjugate pairs in the following reaction are NH_4^+/NH_3 and H_2O/OH^-.

$$\underset{\text{base}_1}{NH_3(g)} + \underset{\text{acid}_2}{H_2O(l)} \longrightarrow \underset{\text{acid}_1}{NH_4^+(aq)} + \underset{\text{base}_2}{OH^-(aq)}$$

Conjugate acid–base pairs differ by a proton, H^+.

- To find the conjugate base of an acid, subtract one H^+.
- To find the conjugate acid of a base, add one H^+.

The convention is that the acid is always written first in the conjugate pairs.

$$\text{conjugate acid} \underset{-H^+}{\overset{+H^+}{\rightleftharpoons}} \text{conjugate base}$$

For example:

conjugate acid		conjugate base
H_3PO_4	$\underset{-H^+}{\overset{+H^+}{\rightleftharpoons}}$	$H_2PO_4^-$

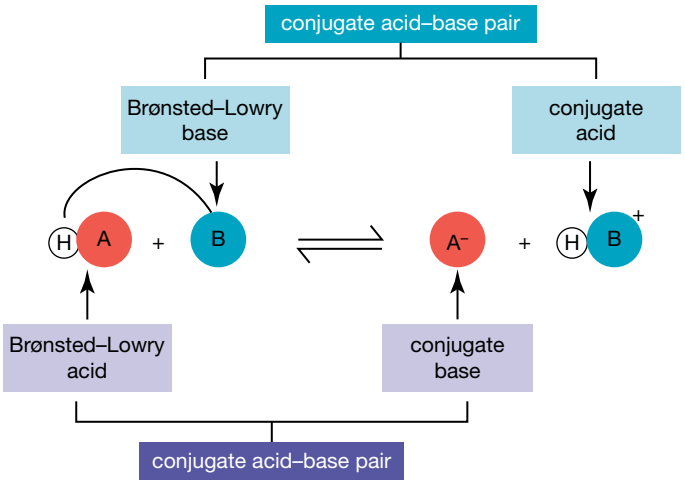
Note that a conjugate acid and a conjugate base can re-form the original acid and base by once again transferring a proton in a 'backwards' reaction according to our definitions.

species general term that could refer to atoms, molecules, molecular fragments or ions

conjugate base the product that an acid forms when it has donated a proton to a base

conjugate acid the product that a base forms when it has accepted a proton from an acid

FIGURE 11.5 Conjugate acids and bases



Identifying conjugate acid–base pairs

$$\text{Conjugate base} = \text{acid} - H^+$$

$$\text{Conjugate acid} = \text{base} + H^+$$

tlvd-3406

SAMPLE PROBLEM 1 Identifying conjugate pairs

Show that the following reaction is a proton transfer reaction and state the acid–base conjugate pairs: $HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$.

THINK	WRITE
1. Decide whether each of the reactants has donated or gained a proton.	HCl has become Cl^-; it has lost a proton (H^+). H_2O has become H_3O^+; it has gained a proton, (H^+).
2. The substance that has donated a proton is the acid.	HCl is an acid.
3. The substance that has gained a proton is the base.	H_2O is a base in this reaction.
4. Conjugate base = acid – H^+. Conjugate acid = base + H^+. **TIP:** Remember that the acid should be put first in each pair.	Conjugate base is Cl^-. Conjugate acid is H_3O^+. Conjugate pairs are HCl/ Cl^- and H_3O^+/ H_2O.

PRACTICE PROBLEM 1

Show that the following reaction is a proton transfer reaction and state the acid–base conjugate pairs:

$$NH_3(aq) + HNO_3(aq) \rightarrow NH_4^+(aq) + NO_3^-(aq)$$

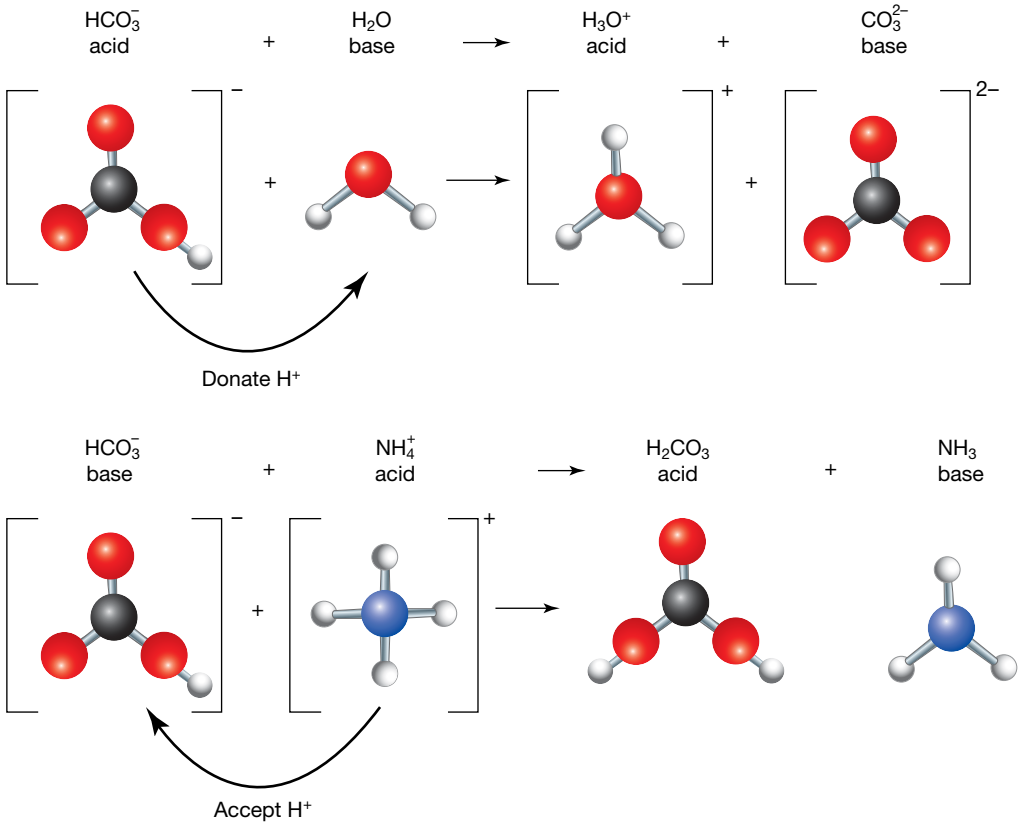
11.2.3 Amphiprotic species

You may have noticed in the previous section that water is acting as a base in the first reaction and then acting as an acid in the second reaction. Some substances can act as acids or bases, according to their chemical environment. This means that they can donate or accept protons. Such substances are described as **amphiprotic**.

Water can act as an acid or a base depending on whether it is reacting with a stronger acid or base. This is illustrated in the examples in the previous section (also see the video eLesson 'The Brønsted–Lowry theory'). Other examples of amphiprotic substances are the hydrogen sulfate ion, HSO_4^-, hydrogen carbonate ion, HCO_3^- and hydrogen sulfide ion, HS^-. These ions have both a proton, H^+, that they can donate and a negative ion to attract a proton, H^+.

An example of the hydrogen carbonate ion, HCO_3^-, acting as an amphiprotic species is shown in figure 11.6.

FIGURE 11.6 Hydrogen carbonate is an amphiprotic molecule.



Amphiprotic species

Amphiprotic species can both donate and accept a proton, H^+.

amphiprotic substance that can both accept and donate protons

11.2.4 Polyprotic acids

Acids can be classified as monoprotic, diprotic or triprotic depending on the actual number of protons that can be donated.

- A **monoprotic** acid is one that can only donate one proton, H^+, because 'mono' means one.
- A **polyprotic** acid is one that can donate more than one proton, H^+.
- A **diprotic** acid can donate two protons.
- A **triprotic** acid can donate three protons.

TABLE 11.3 Common monoprotic and polyprotic acids (acidic protons in bold)

Monoprotic acids	Polyprotic acids	
	Diprotic acids	Triprotic acids
*Ethanoic acid, $CH_3COO\mathbf{H}$	Ascorbic acid, $\mathbf{H_2}C_6H_6O_6$	Citric acid, $\mathbf{H_3}C_6H_5O_7$
Hydrochloric acid, **H**Cl	Carbonic acid, $\mathbf{H_2}CO_3$	Boric acid, $\mathbf{H_3}BO_3$
Hydrocyanic acid, **H**CN	Sulfuric acid, $\mathbf{H_2}SO_4$	Phosphoric acid, $\mathbf{H_3}PO_4$
Hydrofluoric acid, **H**F	Tartaric acid, $\mathbf{H_2}C_4H_4O_6$	
Nitric acid, $\mathbf{H}NO_3$		

*Contains four hydrogen atoms, but can only donate one proton.

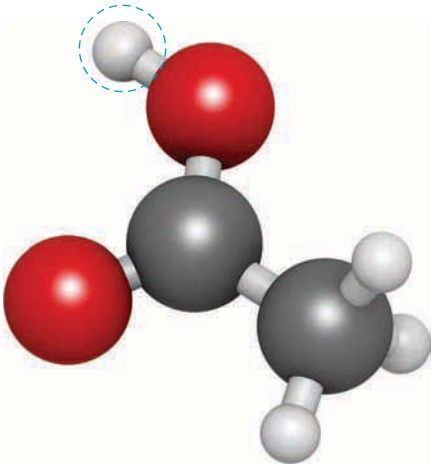
Monoprotic acids

Hydrochloric acid, HCl, is an example of a monoprotic acid because it can donate only one proton. Ethanoic acid, CH_3COOH, despite the fact that it has four hydrogen atoms, is also monoprotic because it can only donate the acidic proton of the carboxylic acid functional group, –COOH (see figure 11.7).

Ethanoic acid + water → Hydronium ion + ethanoate ion.

$$CH_3COOH\ (aq) + H_2O\ (l) \rightarrow H_3O^+\ (aq) + CH_3COO^-\ (aq)$$

FIGURE 11.7 Ethanoic acid, CH_3COOH, is a monoprotic acid (with its acidic proton circled).



Diprotic acids

An acid such as sulfuric acid, H_2SO_4, is called a diprotic acid because it can donate two protons (refer to table 11.3). This occurs in two stages, as shown in the following equations:

$$\text{Stage 1: } H_2SO_4(aq) + H_2O(1) \rightarrow H_3O^+(aq) + HSO_4^-(aq)$$

$$\text{Stage 2: } HSO_4^-(aq) + H_2O(1) \rightleftharpoons H_3O^+(aq) + SO_4^{2-}(aq)$$

The first stage is more extensive than the second, as evident by the different type of reaction arrows used. The single arrow shows that virtually all of the H_2SO_4 molecules ionise because it is a strong acid. In stage two, the **double arrow** indicates only partial ionisation, because hydrogen sulfate is only a weak acid. This means that while some HSO_4^- ions are reacting with water to form SO_4^{2-}, some SO_4^{2-} ions are also recombining with water to form HSO_4^- ions. HSO_4^- is a weak acid because it only partially ionised.

monoprotic (of an acid) can donate one proton

polyprotic (of an acid) able to donate more than one proton

diprotic (of an acid) can donate two protons

triprotic (of an acid) can donate three protons

double arrow used to show a chemical equation is reversible

FIGURE 11.8 Sulfuric acid, H_2SO_4, can donate two protons (circled) so is a diprotic acid.

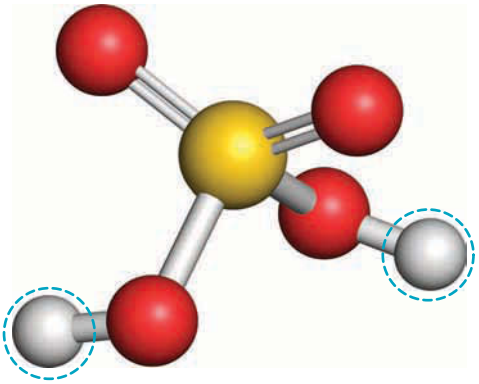
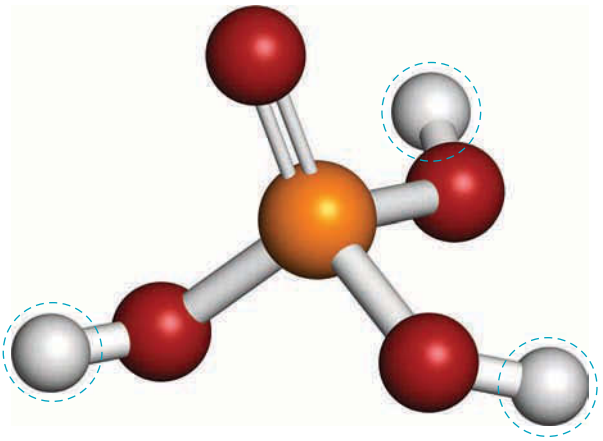


FIGURE 11.9 Phosphoric acid, H_3PO_4, can donate three protons (circled) so is a triprotic acid.



Triprotic acids

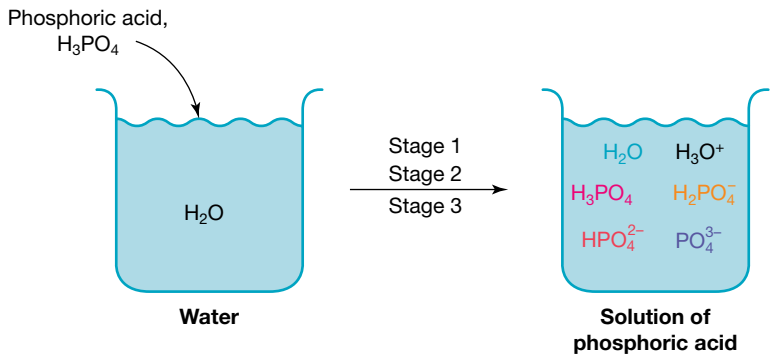
Phosphoric acid, H_3PO_4, is an example of a triprotic acid because it has three protons that it can donate (figure 11.9). This means that three successive ionisation stages are taking place simultaneously, albeit to different ionisation levels. Phosphoric acid is a weak acid so is not completely ionised in stage 1; therefore, stage 2 and 3 occur to a much lesser extent. This also means that in a solution of phosphoric acid, you will find H_2O and ions of H_3O^+, H_3PO_4, $H_2PO_4^-$, HPO_4^{2-} and PO_4^{3-} (see figure 11.10).

Stage 1: $H_3PO_4(aq) + H_2O(l) \rightleftharpoons H_2PO_4^-(aq) + H_3O^+(aq)$

Stage 2: $H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$

Stage 3: $HPO_4^{2-}(aq) + H_2O(I) \rightleftharpoons PO_4^{3-}(aq) + H_3O^+(aq)$

FIGURE 11.10 Ions present in a solution of phosphoric acid showing incomplete ionisation of the acid



Polyprotic acids

Polyprotic acids can donate more than one proton.

11.2 Activities

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11.2 Quick quiz on | **11.2 Exercise** | **11.2 Exam questions**

11.2 Exercise

1. In the following reaction, which substance is acting as an acid and which is acting as a base?

$$H_2SO_4\,(l) + H_2O\,(l) \rightarrow H_3O^+\,(aq) + HSO_4^-\,(aq)$$

2. Identify the reactant that is acting as an acid in each of the following.
   - **a.** $OH^-(aq) + HBr(aq) \rightarrow H_2O(l) + Br^-(aq)$
   - **b.** $HSO_4^-\,(aq) + OH^-(aq) \rightleftharpoons SO_4^{2-}\,(aq) + H_2O(l)$
   - **c.** $HCO_3^-\,(aq) + HSO_4^-\,(aq) \rightleftharpoons H_2O(l) + CO_2(g) + SO_4^{2-}\,(aq)$
   - **d.** $H_3O^+(aq) + OH^-(aq) \rightleftharpoons 2H_2O(l)$
3. Which of the following reactions can be classified as proton-transfer reactions?
   - **a.** $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
   - **b.** $H_3PO_4(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + H_2PO_4^-\,(aq)$
   - **c.** $O^{2-}(aq) + H_2O(l) \rightarrow 2OH^-(aq)$
   - **d.** $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
   - **e.** $CO_3{}^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-\,(aq) + OH^-(aq)$
4. Classify each of the following as a Brønsted–Lowry acid, base or amphiprotic substance.
   - **a.** Cl^- **b.** H_2O **c.** HSO_4^- **d.** HNO_3 **e.** OH^-
5. **a.** Write equations to illustrate H_3PO_4 acting as a triprotic acid in water.
   b. Identify the conjugate acid–base pairs for the equations you wrote in part **a**.
6. Write equations to illustrate the amphiprotic nature of the following species.
   - **a.** HS^- **b.** HSO_4^-
7. Classify each of the following as monoprotic, diprotic, triprotic or amphiprotic.
   - **a.** H_3PO_4 **b.** CH_3COOH **c.** HCO_3^{2-} **d.** H_2CO_3 **e.** HNO_3 **f.** H_2O

11.2 Exam questions

Question 1 (2 marks)

Write equations to show that H_2S is a weak diprotic acid.

Question 2 (1 mark)

Explain why sulfuric acid, H_2SO_4, is a diprotic acid but ethanoic acid, CH_3COOH, has more hydrogen atoms but is only monoprotic.

Question 3 (1 mark)

Which species can be both the conjugate acid of a sulfate ion, SO_4^{2-}, and the conjugate base of sulfuric acid, H_2SO_4?

Question 4 (1 mark)

MC In the following reactions, which of the underlined species is acting as a Brønsted–Lowry base?

- **A.** $NH_3(aq) + \underline{HCO_3^-}(aq) \rightleftharpoons NH_4^+(aq) + CO_3^{2-}(aq)$
- **B.** $\underline{HCN}(aq) + \underline{NaOH}(aq) \rightarrow NaCN(aq) + H_2O(l)$
- **C.** $CH_3COOH(aq) + \underline{HCO_3^-}(aq) \rightleftharpoons CH_3COO^-(aq) + H_2CO_3(aq)$
- **D.** $NH_3(aq) + \underline{HCl}(aq) \rightleftharpoons NH_4^+(aq) + Cl^-(aq)$

Question 5 (1 mark)

MC In the following reactions, which of the underlined species is acting as a Brønsted–Lowry acid?

- **A.** $\underline{HS^-}(aq) + HCl(aq) \rightarrow H_2S(g) + Cl^-(aq)$
- **B.** $\underline{2NaOH}(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
- **C.** $HCO_3^-(aq) + \underline{NH_3}(aq) \rightleftharpoons NH_4^+(aq) + CO_3^{2-}(aq)$
- **D.** $NH_3(aq) + \underline{H_2O}(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

More exam questions are available in your learnON title.

11.3 Concentration and strength of acids and bases

KEY KNOWLEDGE

- The distinction between strong and weak acids and strong and weak bases, and between concentrated and dilute acids and bases, including common examples

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11.3.1 Concentrated versus dilute solutions

A **concentrated** acid has a large amount of acid dissolved in a given volume, and a **dilute** acid has a small amount of acid dissolved in a given volume. The concentration of a solution can be changed to become more dilute by adding water.

Concentrated and dilute solutions

Concentrated solutions have a large number of solute particles in a given volume. Dilute solutions have a small number of solute particles in a given volume.

Calculating concentration using molarity

When working with solutions, the concentration can have implications on how they react.

- Concentration is the quantity of solute in a particular volume of solution.
- In chemistry, it is usually measured in units of moles of solute per litre (mol L^{-1} or M) of solution. This is known as the **molarity** of a solution.

You may have already noticed this on bottles of chemicals that you have used. The molarity of solutions is one of the most commonly used measures in chemistry. For example, a 3.0 M HCl solution contains 3.0 moles of HCl dissolved in 1 litre of solution, and a 0.5 M NaOH solution has 0.5 mole of sodium hydroxide dissolved in 1 litre of solution.

concentrated a large quantity of solute in a known volume of solution

dilute not concentrated; has a small quantity of solute in a known volume of solution

molarity concentration measured in units of moles of solute per litre of solution, with units M or mol L^{-1}

Calculating concentration and the number of moles

Concentration is number of moles of solute per litre (mol L^{-1}); also known as molarity (M).

$$c = \frac{n}{V}$$

Alternatively, the moles of solute in a given volume of solution of a known concentration:

$$n = c \times V$$

where:

n is the number of moles of solute

c is the concentration or molarity (mol L^{-1} or M)

V is the volume (L).

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SAMPLE PROBLEM 2 Calculating the number of moles in a solution

Calculate the number of moles present in 250 mL of 3.00 M HCl.

THINK	WRITE
1. Identify the given information and compare units given to units required. The volume of the solution is given in mL, while the concentration is calculated in L. Convert from mL to L. Identify the unknown quantity.	$c(\text{HCl}) = 3.00\text{M}$ $V(\text{HCl}) = \frac{250\text{ mL}}{1000}$ $= 0.250\text{ L}$ $n(\text{HCl}) = ?$
2. Use the formula to calculate the number of moles: $n = c \times V$ **TIP:** Remember to check significant figures and units.	$n(\text{HCl}) = 3.00 \times 0.250$ $= 0.750\text{ mol}$

PRACTICE PROBLEM 2

Calculate the number of moles present in 100 mL of 1.35 M NaOH.

If you are required to make up a solution of a particular concentration, you will need to start with a mass of the substance because it is not possible to measure out moles directly. In this case, you need to use the formula introduced in topic 6 to calculate n:

$$n = \frac{m}{M}$$

where m is the mass in grams and M is the molar mass in g mol^{-1}. These calculations are explored in topic 13.

TIP: Remember that uppercase M represents the molar concentration of a solution in mol L^{-1}. Uppercase M (in italics) is molar mass in g mol^{-1}.

11.3.2 Strength of acids and bases

The strength of an acid or base is related to the ease with which it donates or accepts a proton, H^+. It is possible to have a concentrated strong acid or strong base as well as a concentrated weak acid or weak base. Similarly, you can have a dilute strong acid or strong base and a dilute weak acid or weak base.

Strong acids and bases

A **strong acid** donates protons readily.

A **strong base** accepts protons readily.

Common strong laboratory acids include hydrochloric acid, HNO_3, and sulfuric acid, H_2SO_4. Only a few strong acids exist and they fully **ionise** in water, producing many ions. Their solutions, therefore, are strong **electrolytes**. All protons are donated to water, forming hydronium ions and the anion of the acid.

strong acid acid that fully ionises in water

strong base base that fully dissociates in water

ionise reaction in which a substance reacts with water to form ions

electrolyte solution or liquid that can conduct electricity

Hydrochloric acid fully reacts with water to produce a hydronium ion and a chloride ion.

$$\underset{\text{strong acid}}{HCl\,(g)} + H_2O\,(l) \xrightarrow{\text{virtually complete}} H_3O^+(aq) + Cl^-(aq)$$

FIGURE 11.11 Hydrochloric acid must be handled with care.

FIGURE 11.12 Strong electrolytes show high conductivity because they produce many ions in solution; water is classified as a non-conductor because it has very few ions.

Some strong bases include lithium hydroxide, LiOH, sodium hydroxide, NaOH, and potassium hydroxide, KOH. Strong bases are generally group 1 metal hydroxides of dissociation. Strong bases fully **dissociate** in water producing many ions. Their solutions, therefore, are also strong electrolytes. Sodium hydroxide dissociates in water to form sodium ions and hydroxide ions.

$$NaOH(s) \xrightarrow{H_2O(l)} Na^+(aq) + OH^-(aq)$$

Strong acids and bases

Strong acids completely ionise in water. Strong bases completely dissociate in water.

Weak acids and bases

Examples of **weak acids** are ethanoic acid, lactic acid and citric acid. Weak acids only **partially ionise** in water, resulting in a solution that is a poor conductor of electricity because not many ions are formed. Ethanoic acid (discussed in section 11.2.4 as a monoprotic acid), will only partially ionise in water to form some hydronium ions and ethanoate ions. Double arrows are used to show that these are partial reactions. That is, they indicate that reactants will also be present in the solution.

$$CH_3COOH(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$$

However, a **weak base**, such as ammonia, NH_3, does not readily ionise and accept protons. It forms very few ions in solution. Ammonia is a poor conductor of electricity.

Weak acids and bases

Weak acids partially ionise in water. Weak bases partially dissociate in water.

dissociate the process by which ions separate when an ionic compound dissolves

weak acid acid that partially ionises in water

partially ionise only a small fraction of the species will break to produce ions

weak base base that partially dissociates in water

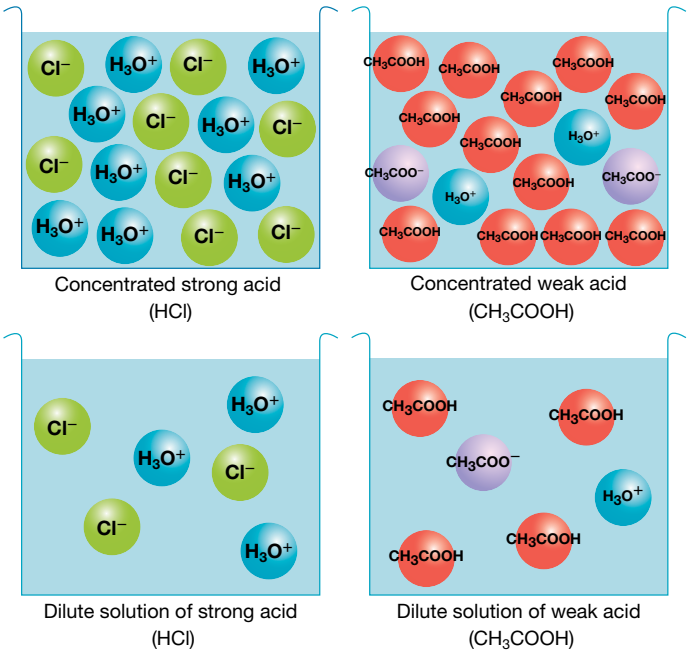
11.3.3 Concentration versus strength

In summary, the concentration of an acid or base refers to the quantity of solute in a given volume of solution, whereas the strength of a solution refers to the extent of ionisation or dissociation of the acid or base. A strong acid fully ionises in solution and a strong base fully dissociates in solution.

TABLE 11.4 Relative strengths of common acids and their conjugate bases

		Name of acid	Formula	Conjugate base	Name of base		
very strong ↓ decreasing strength of acids ↓ very weak	COMMON STRONG ACIDS	Hydrochloric	HCl	Cl^-	Chloride ion		very weak ↑ increasing strength of bases ↑ very strong
		Nitric	HNO_3	NO_3^-	Nitrate ion		
		Sulfuric	H_2SO_4	HSO_4^-	Hydrogen sulfite ion		
		Hydronium ion	H_3O^+	H_2O	Water		
strong ↓ weak	COMMON WEAK ACIDS	Phosphoric	H_3PO_4	$H_2PO_4^-$	Dihydrogen phosphate ion		weak ↑ strong
		Hydrofluoric	HF	F^-	Fluoride ion		
		Ethanoic (acetic)	CH_3COOH	CH_3COO^-	Ethanoate (acetate) ion		
		Carbonic	H_2CO_3	HCO_3^-	Hydrogen carbonate ion		
		Hydrogen sulfide	H_2S	HS^-	Hydrogen sulfide ion		
		Ammonium ion	NH_4^+	NH_3	Ammonia	COMMON WEAK BASES	
		Hydrogen carbonate ion	HCO_3^-	CO_3^{2-}	Carbonate ion		
		Hydrogen sulfide ion	HS^-	S^{2-}	Sulfide ion		
		Water	H_2O	OH^-	Hydroxide ion	COMMON STRONG BASES	
		Hydroxide ion	OH^-	O^{2-}	Oxide ion		
		Hydrogen	H_2	H^-	Hydride ion		

FIGURE 11.13 Concentrated and dilute weak and strong acids



Concentrated strong acid (HCl)

Concentrated weak acid (CH_3COOH)

Dilute solution of strong acid (HCl)

Dilute solution of weak acid (CH_3COOH)

REMEMBER

Strength

- In a strong acid or base solution, you will only find the ionised products because it is completely ionised.
- In a weak acid or base solution, you will find both the reactants and products because it is only partially ionised.

Concentration

- In a concentrated solution, you will have a large number of solute particles.
- In a dilute solution, you will have fewer solute particles.

11.3 Activities

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Find all this and MORE in jacPLUS

11.3 Quick quiz on | **11.3 Exercise** | **11.3 Exam questions**

11.3 Exercise

1. Calculate the number of mole present in the following.
   a. 0.100 L of 1.30 M NaOH. **b.** 120 mL of 2.00 M HCl **c.** 50.0 mL of 0.950 M NHO_3
2. If a 0.0850 litre solution contains 0.0750 moles of H_2SO_4, what is the molarity?
3. Calculate the volume, in L, of each of the following solutions that contain 2.50 moles of the substance.
   a. 3.50M NaOH **b.** 0.500 M KOH **c.** 0.240 M CH_3COOH
4. How many moles of H_3O^+ ions are present in 10.5 mL of 0.180 M HCl?
5. The acids found in fruit juice are weak acids. Explain what this means.
6. Dimethylamine, $HN(CH_3)_2$, is a weak base that reacts in water like ammonia, NH_3. Write an equation to illustrate the reaction of dimethylamine with water.
7. Which of the following would be expected to produce the highest concentration of hydronium ions, H_3O^+? Explain your reasoning.
   0.10 M HCl, 0.10 M HCOOH, 0.5 M HNO_3, 0.5 M CH_3COOH

11.3 Exam questions

Question 1 (1 mark)

MC Which of the following best explains the difference between strong and weak acids?

A. Weak acids have a lower solute : solvent ratio than strong acids.
B. Weak acids are amphiprotic whereas strong acids are not.
C. Strong acids have more acidic protons than weak acids do.
D. Strong acids fully react with water to produce hydronium ions but only a small percentage of weak acids do.

Question 2 (1 mark)

MC Which of the following solutions would be expected to contain the lowest concentration of hydronium ions, H_3O^+?

A. 1 M CH_3COOH **B.** 1 M HNO_3 **C.** 2 M HCOOH **D.** 2 M HCl

Question 3 (1 mark)

Propanoic acid, like other carboxylic acids, is classed as a weak acid. Write an equation showing the ionisation of propanoic acid in water.

Question 4 (2 marks)

Using HBr and HS^- as examples, write equations, with appropriate arrows, to illustrate the difference between a strong acid, HBr, and a weak acid, HS^-. Explain the difference in strength in terms of the conductivity of the resulting solution.

Question 5 (2 marks)

Acids such as nitric and hydrochloric are used by artists to make engravings on metal plates made from copper, zinc or iron.

a. Use table 11.4 to find the conjugate pairs of the acids used by the artists for their etchings. **(1 mark)**

b. What safety precautions would you recommend an artist working with strong acids should take? **(1 mark)**

More exam questions are available in your learnON title.

11.4 The pH scale

KEY KNOWLEDGE

- Use of the logarithmic pH scale to rank solutions from most acidic to most basic; calculation of pH for strong acid and strong base solutions of known concentration using the ionic product of water (K_w at a given temperature)

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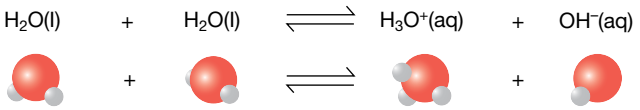
11.4.1 Ionic product of water

Because water is amphiprotic and can act as both an acid and a base, it can also react with itself (self-ionisation) according to the equation shown in figure 11.14.

FIGURE 11.14 Self-ionisation of water



Careful measurements have demonstrated the concentration of both the hydronium ion, H_3O^+, and hydroxide ion, OH^-, at 25 °C is 10^{-7} M. Such values indicate that this reaction occurs only to an extremely small extent.

Furthermore, measurements have shown that, even when dissolved acids and bases are present, the mathematical product of these two ions is always 10^{-14}. If the concentration of H_3O^+ ions increases due to the presence of a dissolved acid, the concentration of OH^- ions decreases such that the product is still 10^{-14}. The reverse applies if a source of OH^- ions (a dissolved base) is present.

For any acidic, basic or neutral solution, use square brackets to show concentration in mol L^{-1}. For example, $[H_3O^+]$ is read as 'the concentration of hydronium ions'.

K_w is termed the **self-ionisation constant** for water, and $[H_3O^+][OH^-]$ is called the ionic product, where the concentration of hydronium ions is multiplied by the concentration of hydroxide ions.

It follows that:

- if $[H_3O^+] > [OH^-]$, the solution is acidic
- if $[OH^-] > [H_3O^+]$, the solution is basic
- if $[H_3O^+] = [OH^-]$, the solution is neutral.

However, in all cases, the product $[H_3O^+][OH^-]$ equals 10^{-14} at 25 °C.

This information is useful when calculating the pH of basic solutions.

self-ionisation constant K_w: in all aqueous solutions this value is constant and results from the self-ionisation of water

Self-ionisation constant, K_w

For all aqueous solutions $K_w = [H_3O^+][OH^-] = 10^{-14}\ M^2$ at 25 °C.

11.4.2 The pH scale

The **pH scale** is used to measure the acidity of a solution. It is logarithmic and inversely indicates the concentration of hydrogen ions in solution (see figure 11.15) and is usually applied over a range from 1 to 14. Using this scale:

- if pH = 7, the solution is neutral
- if pH <7, the solution is acidic
- if pH >7, the solution is alkaline (basic).

The stronger the acid, the lower the pH value. For example, a solution of pH 3 is stronger than a solution of pH 4. Any pH value greater than 7 indicates an alkaline solution. In this case, the higher the value, the stronger the base. It is interesting to note that pH does not have a unit because it is just a number on a scale, and the temperature is specified because the pH can change slightly if the temperature changes. The pH scale of some common acids and bases is shown in figure 11.16.

FIGURE 11.15 The pH scale is logarithmic.

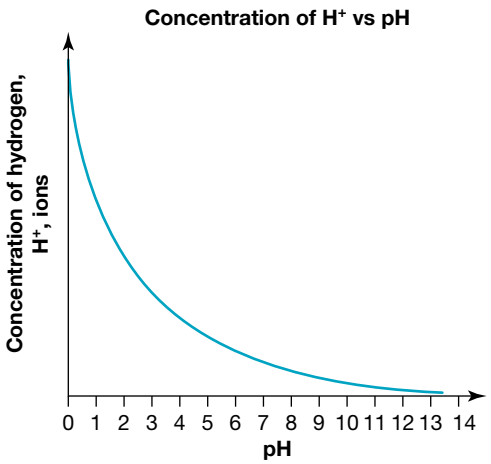
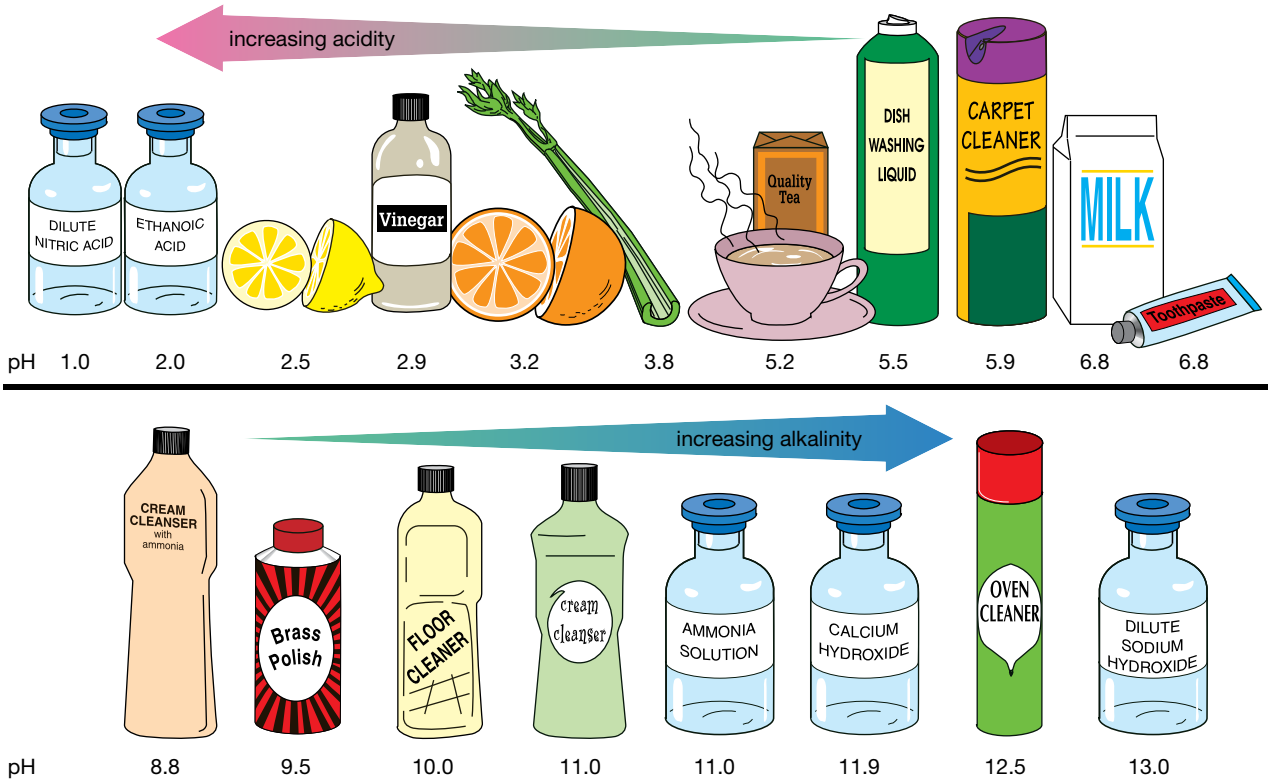


FIGURE 11.16 The pH of common acids and bases



pH scale used to measure acidity; pH can be calculated using $pH = -\log_{10}[H_3O^+]$

CASE STUDY: The chemistry of hair

Hair is made from amino acids that polymerise into long chains held together by hydrogen bonding, salt bridges between the acid and base groups, and disulfide bonds. Hair is normally acidic (pH 4 to 5), and for this reason alkaline shampoos, which clean the hair, are followed by an 'acidic rinse' or conditioner. However, both strongly acidic solutions and slightly alkaline solutions can damage the hair. Strongly acidic solutions (pH 1 to 2) break the H–bonding and salt bridges of the hair. Slightly alkaline (pH 8.5) solutions break some of the disulfide bonds, causing the outer surface of hair strands to become ruffled and appear dull, as light is unevenly reflected from their surfaces.

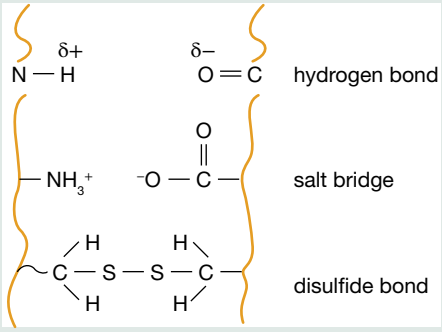
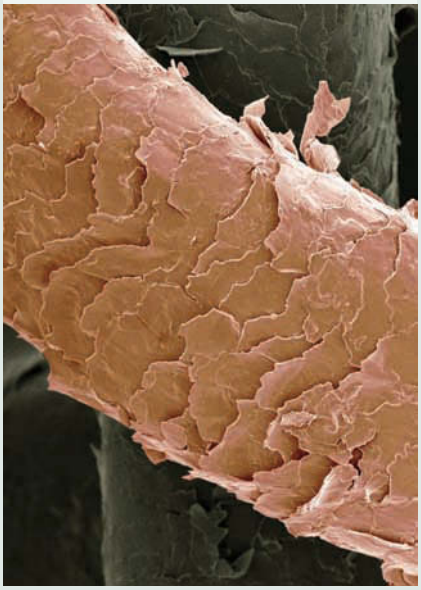


FIGURE 11.17 The bonds holding hair proteins together can be damaged by both alkaline and strongly acidic solutions.



11.4.3 Calculating pH of strong acids

If we know the concentration of the hydronium ion $[H_3O^+]$ in moles per litre of solution, we can calculate the pH using the following formula.

$$pH = -\log_{10}\left[H_3O^+\right]$$

This means that pH is the negative logarithm of the molar hydronium concentration. This formula enables pH to be calculated for any known H_3O^+ or OH^- concentration by using a calculator with a **logarithm** (or log) function. In this case the base is 10, so the logarithm is the power of 10. So, this formula can be rearranged to find the concentration of acid when given the pH:

$$\left[H_3O^+\right] = 10^{-pH}$$

If the $[H_3O^+]$ is easily converted into a power of 10, a calculator is not needed. This formula can only be used for strong acids because strong acids are completely ionised. Consider a strong acid, HA, where H represents hydrogen and A is its anion. So for a strong acid, HA:

$$\left[H_3O^+\right] = [HA]$$

This is not true for weak acids because if the acid, HA, is weak, it is only partially ionised and the $[H_3O^+] \ll [HA]$ and cannot be calculated using this formula.

Calculating the pH of strong acids, HA

$$[HA] = \left[H_3O^+\right] = 10^{-pH}$$

$$pH = -\log_{10}\left[H_3O^+\right]$$

logarithm quantity representing the power to which a fixed number (the base) must be raised to produce a given number

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SAMPLE PROBLEM 3 Calculating the pH of a solution when $[H_3O^+]$ is known

Calculate the pH of a 0.0001 M solution of HCl at 25 °C.

THINK	WRITE
1. Identify the given information and the unknown information.	$[HCl] = 0.0001\ M$ $pH = ?$
2. HCl is a strong acid and so is fully ionised. This means the concentration of H_3O^+ = [HCl].	$[H_3O^+] = [HCl]$ $= 0.0001\ M$
3. Convert this concentration to a power of 10.	$[H_3O^+] = 10^{-4}\ M$
4. Find the pH using $pH = -\log_{10}[H_3O+]$. Log_{10} cancels the base 10, and the two negatives cancel each other, leaving 4.	$pH = -\log_{10}(10^{-4})$ $pH = -\cancel{\log_{10}}\left(\cancel{10}^{-4}\right)$ $= -(-4)$ $= 4$

PRACTICE PROBLEM 3

A solution is found to have a H_3O^+ concentration of 0.01 M at 25 °C. Calculate its pH.

If the $[H_3O^+]$ is not a number that can easily be converted into a power of 10, the procedure is simple using a scientific calculator.

tlvd-0566

SAMPLE PROBLEM 4 Calculating the pH of an acid when the concentration cannot be converted into a power of 10

Calculate the pH of 0.054 M solution of HCl at 25 °C.

THINK	WRITE
1. Identify the given information and the unknown information.	$[HCl] = 0.054\ M$ $pH = ?$
2. HCl is a strong acid and so is fully ionised. This means the concentration of H_3O^+ = [HCl].	$[H_3O^+] = [HCl]$ $[H_3O^+] = 0.054\ M$
3. Find the pH using $pH = -\log_{10}[H_3O^+]$. This time a scientific calculator is needed: • Press the negative key • Press the log key • Key in the number • Press enter **TIP:** Remember to give your answer to the correct number of significant figures.	$pH = -\log_{10}(0.054)$ $pH = 1.3$

PRACTICE PROBLEM 4

Calculate the pH of 0.037 M solution of HNO_3 at 25 °C.

If you need to find the concentration of $[H_3O^+]$, the pH formula must be rearranged.

tlvd-3409

SAMPLE PROBLEM 5 Calculating the concentration of a solution when $[H_3O^+]$ is not known

Calculate the concentration of a solution of HCl with a pH of 2.1 at 25 °C.

THINK	WRITE
1. Identify the given information and the unknown information.	$pH = 2.1$ $[HCl] = ?$
2. HCl is a strong acid and so is fully ionised. This means $[HCl] = [H_3O^+]$. Find $[H_3O^+]$ using $[H_3O^+] = 10^{-pH}$.	$[H_3O^+] = 10^{-2.1}$ M
3. Calculate using a scientific calculator: • Press 10 • Press ^ • Press –2.1 • Press enter	$[H_3O^+] = 7.9 \times 10^{-3}$ M
4. Since HCl is a strong acid $[HCl] = [H_3O^+]$	$[HCl] = [H_3O^+] = 7.9 \times 10^{-3}$ M

PRACTICE PROBLEM 5

Calculate the concentration of a solution of HNO_3 with a pH of 1.8 at 25 °C.

Relative strengths of acids and bases and pH

The strength of an acid or base is measured by its tendency to donate (in the case of acids) or accept (in the case of bases) protons (see section 11.3.2). The strength of an acid has to be taken into consideration when calculating the pH, because it determines the acid's ability to ionise.

Experiments have found that not all acids of 0.1 M concentration have the same pH, and do not conduct electricity to the same extent. Remember that the conductivity of a solution is directly related to the number of ions in solution. Consider two acids, hydrochloric acid, HCl, and ethanoic acid, CH_3COOH, with the same concentration of 0.1 M at 25 °C. Experimental results confirm that 0.1 M hydrochloric acid has a pH of 1; however, 0.1 M ethanoic acid has a pH of 3. This is explained by the differing abilities of the two acids to ionise.

Almost every HCl molecule donates protons to water. This is consistent with the fact that HCl is a strong acid and ionises virtually completely to produce a good electrolyte. CH_3COOH, on the other hand, is a weak acid. Only one in every hundred CH_3COOH molecules loses a proton and, therefore, a solution of ethanoic acid is a poor electrolyte and hence a poor conductor of electricity. This variability is shown in table 11.5.

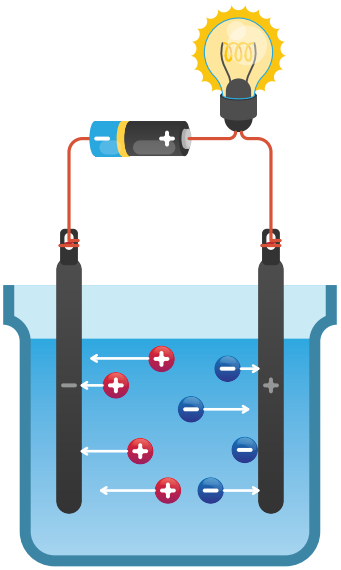

FIGURE 11.18 Strong acids and bases produce many ions, which makes the solution a good conductor of electricity.

TABLE 11.5 pH values for solutions of a range of acids

Acid	Formula	pH (0.01 M solution)	pH (0.001 M solution)
Benzoic acid	C_6H_5COOH	3.1	3.6
Ethanoic acid	CH_3COOH	3.4	3.9
Hypochlorous acid	HOCl	4.8	5.3
Lactic acid	$HC_3H_5O_3$	2.9	3.4
Methanoic acid	HCOOH	2.9	3.4
Nitrous acid	HNO_2	2.6	3.1

The pH of a basic solution is also affected by the ability of the base to dissolve or dissociate in water. For example, a 0.1 M solution of NaOH, a strong base, has a higher pH than a 0.1 M solution of NH_3, a weak base. The sodium hydroxide solution is, therefore, a stronger base than the ammonia solution.

This dependence of pH on acid strength can be used to compare the relative strengths of a range of acids. Provided that the concentrations of the acids are all the same, the lower the pH the stronger the acid is (refer to table 11.5). Note that nitrous acid is the strongest acid of those listed (lowest pH) while hypochlorous acid (highest pH) is the weakest. It does not matter which concentration is used to make this comparison, as long as it is the same for all the acids that you are comparing.

11.4.4 Calculating the pH of strong bases

The definition of pH is based on the concentration of the hydronium ion H_3O^+, so to find the pH of an alkaline solution it is necessary to use the expression given for the self-ionisation of water, remembering that in aqueous solutions $K_w = [H_3O^+][OH^-] = 10^{-14}\ M^2$.

Therefore, by rearranging this formula, the concentration of hydronium ion can be found and the pH can then be found using a calculator as shown in sample problem 5.

Calculating the pH of strong bases

$$[H_3O^+] = \frac{10^{-14}}{[OH^-]}$$

$$pH = -\log_{10}[H_3O^+]$$

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SAMPLE PROBLEM 6 Calculating the pH of a solution when $[OH^-]$ is known

Calculate the pH of a solution of 1.5 M NaOH solution at 25 °C.

THINK	WRITE
1. Identify the given information and the unknown information.	$[NaOH] = 1.5\ M$ $pH = ?$
2. $[OH^-] = [NaOH]$ because NaOH is a strong base.	$[OH^-] = [NaOH]$ $[OH^-] = 1.5\ M$
3. To find $[H_3O^+]$ use the formula from the self-ionisation of water: $[H_3O^+] = \frac{10^{-14}}{[OH^-]}$	$[H_3O^+] = \frac{10^{-14}\ M}{1.5}$

4. To find $[H_3O^+]$ use a scientific calculator:
   - Press 10
   - Press ^
   - Press −14
   - Press enter
   - Press ÷ 1.5

$$[H_3O^+] = \frac{10^{-14}\ M}{1.5}$$
$$[H_3O^+] = 6.7 \times 10^{-15}\ M$$

5. Now to find the pH, apply the formula:
   $pH = -\log_{10}[H_3O^+]$
   And use a scientific calculator:
   - Press −log (2nd ans)
   - Press enter

$$pH = -\log_{10}[6.7 \times 10^{-15}]$$

6. Check that the pH is greater than 7 (because NaOH is a base).

$$pH = 14.17$$

PRACTICE PROBLEM 6

Calculate the pH of a solution of 2.4×10^{-2} M NaOH solution at 25 °C.

If you are required to find the concentration of a base given the pH, the method is shown in sample problem 7.

tlvd-0569

SAMPLE PROBLEM 7 Calculating the concentration solution when pH is known

Calculate the concentration of a solution of NaOH with a pH of 12.5 at 25 °C.

THINK	WRITE
1. Identify the given information and the unknown information.	$pH = 12.5$ $[NaOH] = ?$
2. Recall the formula $[H_3O^+] = 10^{-pH}$.	$[H_3O^+] = 10^{-12.5}\ M$
3. To find $[OH^-]$ use the formula from the self-ionisation of water: $[H_3O^+] = \frac{10^{-14}}{[OH^-]}$	$[OH^-] = \frac{10^{-14}}{10^{-12.5}}$ $= 10^{-1.5}\ M$
4. Use a scientific calculator: • Press 10 • Press ^ • Press −1.5 • Press enter	$[OH^-] = 0.032\ M$
5. $[NaOH] = [OH^-]$ because NaOH is a strong base.	$[NaOH] = 0.032\ M$

PRACTICE PROBLEM 7

Calculate the concentration of a solution of KOH with a pH of 11.8 at 25 °C.

CASE STUDY: Controlling the pH of swimming pools

The growth of algae and bacteria is controlled in swimming pools through the use of chlorine compounds. Sodium hypochlorite, NaOCl, and calcium hypochlorite, $Ca(OCl)_2$, are used, undergoing hydrolysis to produce the weak acid hypochlorous acid, HOCl, according to the equation:

$$OCl^-(aq) + H_2O(l) \rightleftharpoons HOCl(aq) + OH^-(aq)$$

However, if the pH is too high, the concentration and effectiveness of HOCl are reduced. To lower the pH, an acid solution (such as hydrochloric acid or solid sodium hydrogen sulfate) can be used to neutralise the excess hydroxide ions. If the pH is too low, the acid content of the water can cause eye and skin irritation and can damage pool fittings. The pH can be raised by neutralising with sodium carbonate or sodium hydrogen carbonate.

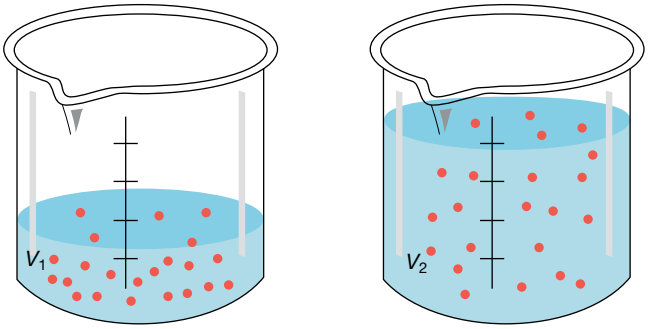
FIGURE 11.19 Pools are kept free of algae and bacteria by managing pH levels.

11.4.5 Dilution of solutions

Chemicals are often bought and stored in concentrated form known as **stock solutions**. They are then diluted to the concentration required. Dilution is the process of adding more water to a solution. In this process, the volume increases and the concentration decreases but the number of mole has not changed because no solute has been added or removed.

When using this process, it is essential to be able to accurately calculate the concentration of the diluted solution or the volume of water required to be added to produce a desired concentration.

FIGURE 11.20 Before and after a dilution; the volume changes but the number of moles remains the same.



Consider a beaker with volume V_1 and concentration c_1. The number of moles present is $n = c_1V_1$. If water is added to this beaker, there is a new volume V_2 and a different concentration c_2; however, the moles present are the same, $n = c_2V_2$ (see figure 11.20).

Calculating concentrations after dilution of solutions

When a solution is diluted, the new volume (V) or concentration (c) can be determined by:

$$c_1V_1 = c_2V_2$$

The units of volume can be either mL or L but both V_1 and V_2 must be the same unit.

stock solution concentrated solution that will be diluted as required for later use

tlvd-0570

SAMPLE PROBLEM 8 Calculating concentration after dilution

If 50.0 mL of a 2.50 M solution is diluted to 85.0 mL, what is the final concentration?

THINK

1. Identify the given information and the unknown information.
   TIP: As long as your units of volume are consistent, the final concentration will be correct. In this case the volumes are given in mL so mL is used, but using 0.05 L and 0.085 L would also give the same result.
2. Rearrange the dilution formula, $c_1V_1 = c_2V_2$, to find c_2.
   TIP: Remember to check significant figures and units and that the concentration is lower after dilution.

WRITE

$c_1 = 2.50$ M

$V_1 = 50.0$ mL

$V_2 = 85.0$ mL

$c_2 = ?$

$$c_2 = \frac{c_1V_1}{V_2}$$

$$c_2 = \frac{2.50 \times 50.0}{85.0}$$

$$= 1.47 \text{ M}$$

PRACTICE PROBLEM 8

If 125.0 mL of a 1.80 M solution is diluted to 220 mL, what is the final concentration?

To calculate the water added to achieve a new concentration, remember to subtract the initial volume.

Water added = $V_2 - V_1$

tlvd-0571

SAMPLE PROBLEM 9 Calculating the volume of water required to obtain a new concentration

How much water would be added to 500 mL of a 2.20 M HCl solution to make a 1.50 M solution?

THINK

1. Identify the given information and the unknown information.
2. Rearrange the dilution formula, $c_1V_1 = c_2V_2$, to find V_2.
3. Determine amount of water added. Remember to check for significant figures and units.

WRITE

$c_1 = 2.20$ M

$V_1 = 500.0$ mL

$c_2 = 1.50$ M

$V_2 = ?$

Water added = ?

$$V_2 = \frac{c_1V_1}{c_2}$$

$$V_2 = \frac{2.20 \times 500}{1.50}$$

$$= 733 \text{ mL}$$

$V_2 - V_1 = 733 - 500$

$= 233$ mL

PRACTICE PROBLEM 9

How much water would need to be added to 750 mL of a 3.20 M HCl solution to make a 2.00 M solution?

Calculating the pH after diluting solutions

If pH values are provided instead of concentrations, it is necessary to convert the pH into a concentration of acid or base (as shown in section 11.4.3) before proceeding with the dilution formula.

For solutions dilutions that involve acidic or basic solutions, this affects the concentration of H_3O^+ ions that are present and, hence, the pH. Because pH is on a logarithmic scale, the factor by which H_3O^+ changes is not the same as that by which pH changes. A change in pH of 1 means a change in concentration of a factor of 10.

This means that a solution of pH 3 is 10 times more dilute than a solution of pH of 2.

tlvd-0572

SAMPLE PROBLEM 10 Calculating the pH of a solution after dilution

60.0 mL of a solution of hydrochloric acid of concentration 0.0550 M is diluted by adding 40.0 mL of water. Calculate the pH of the resulting solution.

THINK	WRITE
1. Identify the given information and the unknown information.	$c_1 = 0.0550$ M $V_1 = 60.0$ mL $V_2 = 100.0$ mL pH = ?
2. To determine the pH, the concentration must first be determined by applying the dilution formula: $c_1V_1 = c_2V_2$	$c_1V_1 = c_2V_2$ $c_2 = \frac{c_1V_1}{V_2}$ $= \frac{0.0550 \times 60.0}{100.0}$ $= 0.0330$ M $c_2 = [HCl] = 0.0330$ M
3. $[H_3O^+] = [HCl]$ because HCl is a strong acid. Use the formula $pH = -\log_{10}[H_3O^+]$ to find pH of the diluted solution.	$pH = -\log_{10}[H_3O^+]$ $= -\log_{10}(0.033)$ $= 1.48$

PRACTICE PROBLEM 10

20.0 mL of a solution of hydrochloric acid of concentration 0.250 M is diluted by adding 40.0 mL of water. Calculate the pH of the resulting solution.

11.4 Activities

learnon

Students, these questions are even better in jacPLUS

Receive immediate feedback and access sample responses

Access additional questions

Track your results and progress

Find all this and MORE in jacPLUS

11.4 Quick quiz on | 11.4 Exercise | 11.4 Exam questions

11.4 Exercise

1. a. Refer to the ionic product of water and determine $[OH^-]$ if the concentration of $[H_3O^+] = 10^{-5}$ M.
   b. Is this solution acidic or basic?
2. What is the pH scale and how can pH be calculated?
3. In a pH test on a swimming pool, the water is found to have a pH of 4.5.
   a. What would you add to the water to raise the pH?
   b. The pH should be about 7.5 for best conditions. Is this acidic or alkaline?
4. Determine the pH of the following solution given that $[H_3O^+]$ is:
   a. 10^{-5} M. b. 0.1 M.
5. Determine the pH of the strong acid $HClO_4$, which has a concentration of 0.0015 M.
6. Determine the pH of the following solution given that $[OH^-]$ is:
   a. 10^{-4} M. b. 0.1 M.
7. Determine the pH of a solution where:
   a. $[H_3O^+] = 9.2 \times 10^{-4}$ M. b. $[OH^-] = 1.8 \times 10^{-5}$ M.
8. Using the formula for pH, explain why a neutral solution has a pH of 7 (at 25 °C).
9. Hydrogen fluoride is a weak acid, while HCl is a strong acid. Explain how a solution of HF could have the same pH as a solution of HCl.
10. a. Calculate the pH of solutions that have the following H_3O^+ concentrations.
    i. 0.0010 M ii. 0.00050 M
    b. Calculate the pH of solutions that have the following OH^- ion concentrations.
    i. 0.0010 M ii. 0.000500 M
11. Nitric acid, HNO_3, is a corrosive acid. When purchased from a chemical company, its molarity is 16 M. How much of the stock solution of nitric acid needs to be used to make 2.5 L of a 0.20 M solution?

11.4 Exam questions

Question 1 (4 marks)

Answer the following.

a. If $[OH^-] = 10^{-6}$ M, calculate (i) $[H_3O^+]$ and (ii) pH. **(2 marks)**
b. If $[OH^-] = 0.025$ M, calculate (i) $[H_3O^+]$ and (ii) pH. **(2 marks)**

Question 2 (3 marks)

Calculate the pH for each of the following.

a. $[H_3O^+] = 0.000\ 10$ M **(1 mark)**
b. $[H_3O^+] = 1.5 \times 10^{-5}$ M **(1 mark)**
c. $[H_3O^+] = 2.0$ M **(1 mark)**

Question 3 (1 mark)

A solution of HCl has a pH of 2.6. Calculate $[H_3O^+]$.

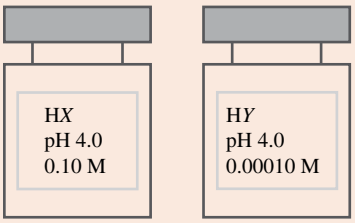
Question 4 (1 mark)

MC Which of the following solutions would *not* have a pH of 3?

A. 10.0 mL of an HCl solution with pH 2 diluted by a factor of 10
B. 50.0 mL of 0.0010 M HNO_3
C. The solution formed when 10 mL of 0.1 M HCl is made up to 1 litre with water
D. 10.0 mL of a 1.0×10^{-4} M solution added to 90.0 mL of water

Question 5 (1 mark)

MC Two acids, H*X* and H*Y*, are shown with their concentrations and pH values.



Which of the following statements is/are correct?

- Statement I: H*X* is a stronger acid than H*Y*.
- Statement II: Both acids ionise to the same extent in water.
- Statement III: The reaction showing the ionisation of H*Y* in water should be written with a double arrow.

A. Statement I only
B. Statements I and III only
C. All statements are correct.
D. No statements are correct.

More exam questions are available in your learnON title.

11.5 Measuring pH

KEY KNOWLEDGE

- Accuracy and precision in measurement: comparison of natural indicators, commercial indicators, and pH meters to determine the relative strengths of acidic and basic solutions

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11.5.1 Ways to measure pH

We now know that pH is a measure of acidity and, specifically, the amount of hydronium ions (protons) in a solution measured as moles per litre (mol L^{-1} or M).

pH can be measured:

- **quantitatively** using a pH meter to give an **accurate** measurement
- **qualitatively** using pH indicators.

Indicators are generally weak acids or weak bases that only partially ionise in solution. Depending on the colour change after adding the indictor to the solution, you can determine an approximate pH of that solution by comparing it to a pre-determined pH-colour chart.

11.5.2 Quantitative measurements of pH

The pH meter

Most laboratories in industry have a pH meter that is used to make rapid, accurate measurements of the pH of a solution.

- It can be connected to a computer to monitor pH changes continuously (figure 11.21a).
- The voltage of the electrode changes with the $[H_3O^+]$ in the solution into which it is dipped.
- Values of pH obtained by a pH meter are accurate to within 0.01 units of the true pH and are not affected by the colour and cloudiness of the unknown solution — an advantage over qualitative measurements using pH indicators.

Hospitals use pH meters to find small but significant changes of pH in blood and other body fluids. Soils, ponds, rivers, sewage and industrial effluents are also easily monitored with a pH meter — either in the lab using the collected samples, or at the collection site using a mobile pH meter (figure 11.21b).

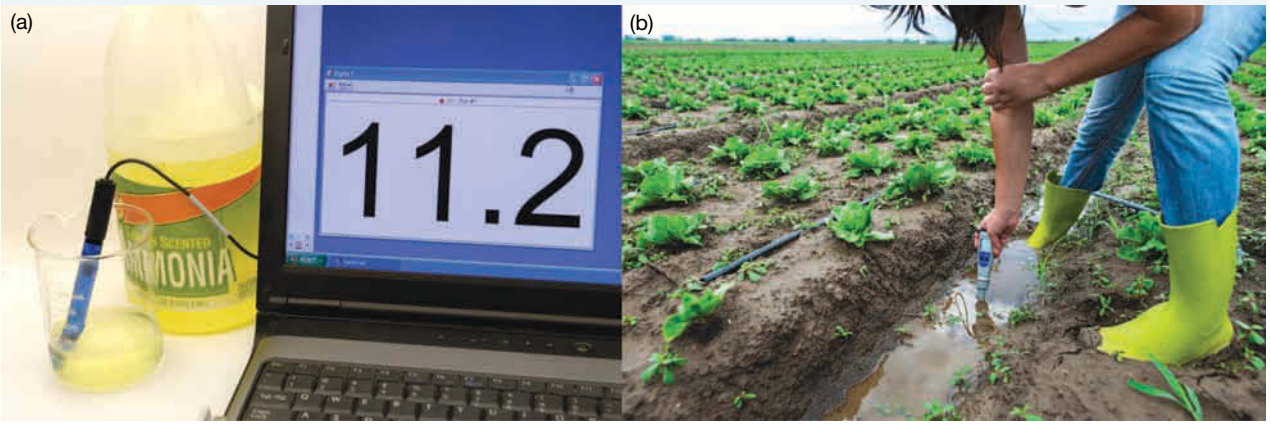
quantitatively measurement of the quantity of something (e.g. length or time); always numerical

accurate when an experimental measurement is close to a known value

qualitatively measurement or observation that describes the qualities of something; labels or categories rather than numbers

indicator substances that undergo an observable change in an acidic or basic solution

FIGURE 11.21 Digital pH meter (a) in the laboratory measuring the pH of the base ammonia (note this accuracy is to 0.1 units), and (b) measuring the pH of irrigation water in the watering canal between rows of lettuce plants.



pH meters are useful pH measurement tools, although they do have certain limitations of accuracy:

- pH meters need to be **calibrated**, such that they are tested against buffered solutions of an exact, known pH, and adjusted to ensure they read that same pH.
- pH is dependent on the temperature. Some pH meters have in-built thermometers that automatically correct for the change in temperature, while others require the user to calibrate the meter each time to account for temperature variations.

Commercial pH indicators

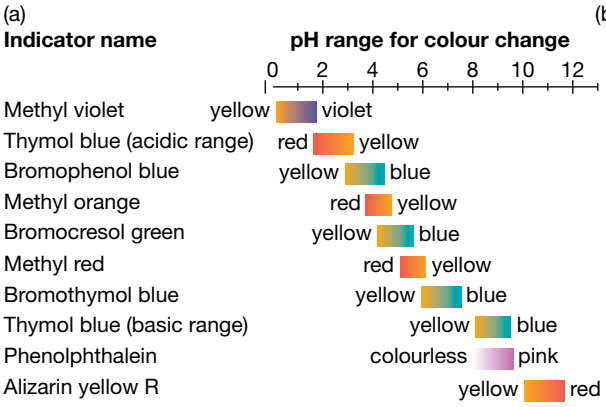
Indicators are chemical dyes that are themselves weak acids or bases.

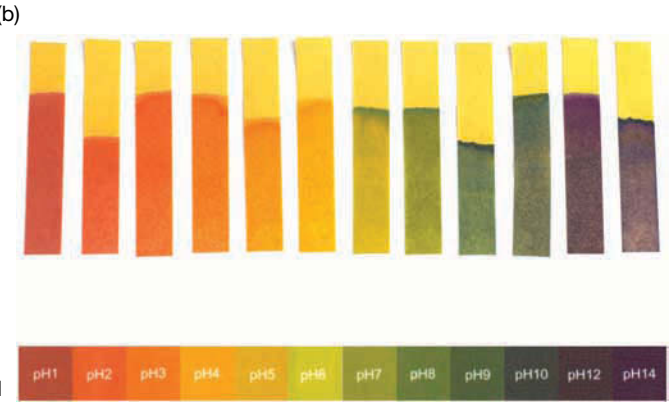
- They show a different colour in an acidic solution than in an alkaline solution. This is because their conjugate base has a different colour to their conjugate acid.
- They can be inexpensively used to determine whether a substance is acidic or alkaline, and how acidic or alkaline it is, without the requirement of determining how many hydrogen ions are present.
- Each indicator changes colour at a certain pH, so it is important to choose an indicator that is appropriate for a particular reaction (figure 11.22a).

Litmus, phenolphthalein, methyl orange and universal indicator are the most commonly used commercial indicators. Universal indicator is a mixture of several indicators and so changes colour gradually from red to green to violet as the solution changes from acidic to neutral to alkaline. The pH of a solution can be measured by adding universal indicator and matching the colour shade obtained against a special chart (see figure 11.22b).

calibrated to adjust an instrument using standards of known measurements to ensure an instrument's accuracy

FIGURE 11.22 (a) Colour changes for different indicators across a range of pH. (b) Strips of universal indicator paper that have been dipped into solutions of various pH values.





Indicators are useful pH measurement tools, although they do have certain limitations of accuracy:

- The indicator colours found in charts are quoted for 25 °C; at other temperatures, an indicator may change colour at a different pH.
- If a solution being tested has a colour of its own, the colour of the indicator may be distorted. For example, if your solution is pink to start with and the pH should be reading yellow, it may appear a little more orange.
- The naked eye is used to make the comparison between the indicator colour and the pH chart, so subjectivity will always be involved.
- Indicators do not give accurate measurements; more accurate readings can be obtained using specially designed pH meters.

FIGURE 11.23 Universal indicator paper

elog-1806

EXPERIMENT 11.1 online only

Indicators and pH

Aim

To investigate the colours of various indicators in solutions of various acids and bases, and use a pH chart to estimate the pH of the substances tested

elog-1807

EXPERIMENT 11.2 online only

Finding the pH of common household substances

Aim

To investigate the pH of common household substances

11.5.3 Qualitative measurements — natural pH indicators

CASE STUDY: Gardening with acids and bases

Have you ever wondered why hydrangeas can bloom a different shade from year to year? The colour of the hydrangea bloom depends on the acidity of the soil. If the soil becomes too acidic, pink hydrangea blooms turn blue. To correct this, calcium oxide, CaO, generally known as 'lime' or 'quicklime' is added. Calcium oxide is a base that dissolves in water to form calcium hydroxide, $Ca(OH)_2$, or 'slaked lime'. This base changes the colour of the blossoms back to pink.

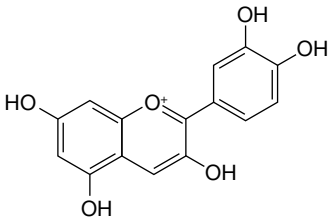
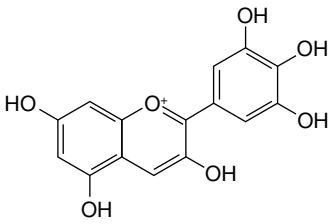
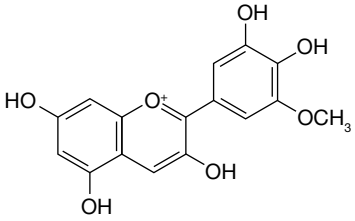
FIGURE 11.24 Blue hydrangeas grown in acidic oil and pink hydrangeas grown in more basic soil

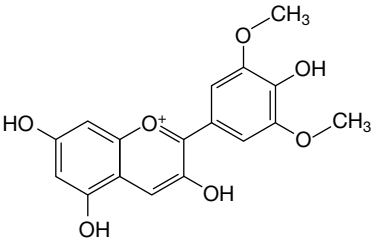
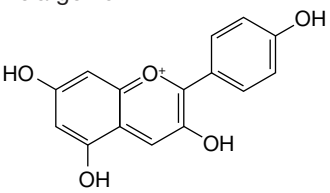
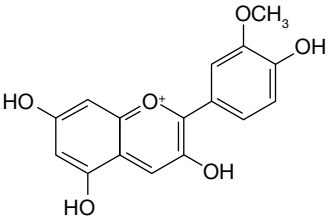
The most common and well-known plant pigments are chlorophyll, responsible for their role in photosynthesis, producing energy for the plant and causing the plant to be a shade of green. However, various edible fruits and vegetables and some non-edible leaves and flowers have other pigments in the vacuoles of their cells that can change colour with different pH levels (see table 11.6). When isolated from the plants, we can use these pigments as natural pH indicators. Most of these pH-sensitive molecules are a variant of the water-soluble **anthocyanin** molecule belonging to a class of molecules called flavonoids. The presence of the hydroxyl groups (–OH) explains their high solubility in water as they form hydrogen bonds with the water molecules and provide a source of **acidic protons** (see section 11.2.4). Like all pigments, these molecules form part of a complex conjugated system that absorb different wavelengths of visible light that have a similar wave frequency at or near that of the electron vibration frequency of the pigment molecule. Consequently, light waves that are not absorbed are reflected, and this determines the colour that we perceive.

anthocyanin pH-sensitive pigment molecule found in plants that will change colour in response to different acidic and alkaline environments

acidic protons the hydrogen atoms that will readily dissociate from the acidic molecule

TABLE 11.6 Sources of natural pH indicators — edible and non-edible plant sources containing anthocyanin molecules

Edible plant examples	Non-edible plant examples	Anthocyanin variant	pH of colour change*
Blackberries, raspberries, elderberries	Poinsettia	Cyanidin 	pH < 3 = red pH 3–4 = clear** (although crude extractions can be a combination of purple/pink shades) pH 4–7 = violet pH 7–8 = blue pH > 8 = greenish yellow (degradation of pigment)
Red (purple) cabbage	Butterfly pea flower (can be made into a tea)		
Cherries			
Blueberries	Delphinium petals, hydrangeas	Delphinidin  Petunidin 	pH < 2.8 = red pH 2.8–3.2 = blue pH > 5 = greenish yellow (degradation of pigment)

Edible plant examples	Non-edible plant examples	Anthocyanin variant	pH of colour change*
Red and purple grapes	Morning glories, blue primrose	Malvidin 	pH < 6 = red pH > 8 = violet
Red wine (derived from the fermentation of red grapes)			
Strawberries, radishes	Geranium petals	Pelargonidin 	pH < 6 = orange/red pH > 8 = violet
Cranberries	Peonies	Peonidin 	pH < 3 = cherry red pH > 8 = deep blue

* Note that some of these edible and non-edible plant sources can contain more than one anthocyanin variant causing slight variations to the pH they indicate through the change of colour.

** A pure extract will produce a cyanidin molecular structure that does not transmit light at pH 3–4 and is, therefore, a clear solution; however, this may not occur for all crude pH indicator extractions from plants due to the presence of more than one variant of anthocyanin in that plant.

Red cabbage contains cyanidin — a natural pH indicator

You may have made red cabbage indicator in junior science and watched all the colours change as you tested it in different acids and bases. Now you will be able to explain at the molecular and atomic level what is happening to cause that change in colour. Your extracted red cabbage 'juice' contains the cyanidin molecule (figure 11.25), an anthocyanin variant. The natural indicator properties of this extract and the process of colour change in different acidic and alkaline environments can be best explained by pH levels.

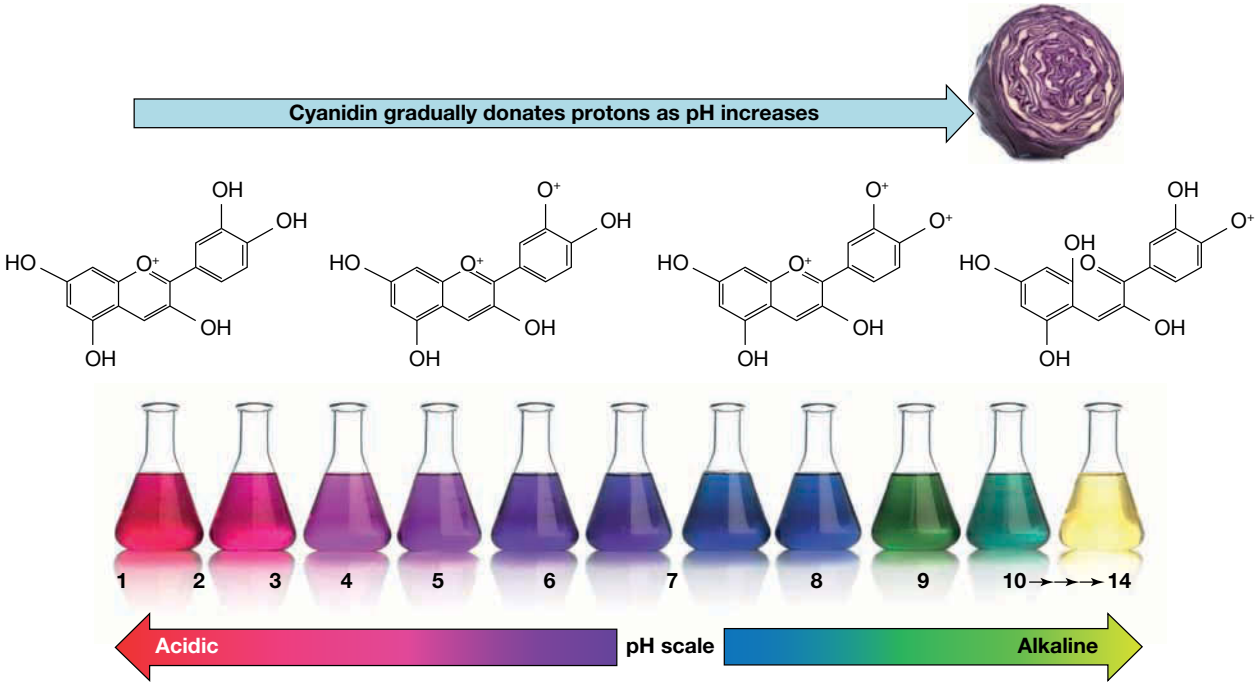
- At a low pH (< pH 3.0), the cyanidin molecule exists as a flavylium cation that is protonated (contains the hydrogen atoms) and has an overall positive charge (refer to table 11.6). This molecule absorbs green light and so appears red in colour.
- Increasing the pH results in **deprotonation**, producing an initial clear solution (complete decolouration; see **note attached to table 11.6).

deprotonation the donation or dissociation of a proton, H^+, from the acidic molecule

- Further deprotonation results in a purple solution at pH 6–7, due to it absorbing yellow colours.
- At pH 7–8, further deprotonation results in a negatively charged anion that absorbs orange light and transmits blue light.
- For the less stable molecules in an alkaline environment (pH > 8), the pigments begin to degrade and absorb red light and transmit green to yellow light.

In summary, as the pH increases, the cyanidin molecule gradually donates its protons, changing the colour of light absorbed from red (750 nm) to green (490 nm), resulting in the natural red cabbage pH indicator appearing various shades of red, violet, blue, green and yellow (see figure 11.25).

FIGURE 11.25 Cyanidin, a natural pH indicator extracted from red cabbage, indicates red solutions at pH < 3, gradually donating protons as the pH increases to change to pinks, purples, blues, green and yellows.



tlvd-3410

SAMPLE PROBLEM 11 Investigating natural indicators

You and a friend go out for afternoon tea one sunny day. She orders a regular green tea and you decide to try the butterfly pea tea. Both are served with a slice of lemon on the side.
Your tea arrives, and you love the vibrant blue colour. But when you add a squeeze of lemon and stir the tea, it begins to change colour and become more purple-pink.

a. Explain why your tea started off blue and changed colour after a squeeze of the lemon.
b. Determine the pH of each colour liquid observed.

THINK

a. 1. Recall the main ingredient in the butterfly pea tea that results in it being a blue solution.

WRITE

1. The butterfly pea tea was made using butterfly pea flowers that contain anthocyanins including the cyanidin variant. These are pigments capable of changing colours with pH. These molecules are known as natural pH indicators.

	2. Lemon juice contains citric acid. Adding an acid changes the pH of a solution.	2. Lemon juice is very high in citric acid and, therefore, an acidic solution. Adding an acidic solution (lemon juice) to a slightly alkaline solution (tea), will cause the tea to become more acidic. This causes the protonation (adding hydrogen ions) to the anthocyanin pigment, changing the light that is absorbed and reflected as pink/purple.
b.	1. Identify the pH of the blue tea by considering figure 11.25.	1. The deep blue colour indicates the tea was between pH 7 and 8 and, therefore, considered relatively neutral or slightly alkaline.
	2. Identify the pH of the purple-pink tea.	2. The final pink-purple tea would have a pH closer to 4.

PRACTICE PROBLEM 11

Following your studies of natural indicators, you decide to test various items around your house because you know they are not all neutral. You test tap water, carbonated water, bicarb soda, vinegar, drain cleaner and some Pool Clear. You have a red cabbage and extract its juice. You place a small amount of each test liquid in a cup, and then add some of your red cabbage extract.

(*Note:* If you are going to do this at home, please make sure you have permission and supervision from a parent or carer, and ensure you are wearing appropriate personal protective gear such as gloves, goggles and protective clothing.)

a. What are your predicted results (pH and expected colour) for each of the items you are testing?

b. What is the main general molecule in your red cabbage indicator causing the colour change? Explain in which testing agent you would predict the molecule to be in its most protonated state.

elog-1808

EXPERIMENT 11.3 online only

Using natural indicators

Aim

To change the colour of common foods using natural indicators

on Resources

Weblinks Changing colour with cabbage
Blue dyes from cabbage
Colour-changing noodles

11.5.4 Measuring pH with accuracy and precision

When conducting experiments, we need to consider how accurate and how **precise** our measurements are going to be following the experimental design and with the equipment we are using.

If we consider testing for pH, accuracy refers to how close your measured pH values are to the actual **true value** of pH for that solution, while precision refers to how close your repeated pH measures are to each other. The final **measurement result** will generally be an average of precise measurements.

Tips to remember accuracy and precision

A good way to remember the difference between accuracy and precision is to use the following piece of word association:

Accurate data is close to the **A**ctual, true value of that measurement.

Precise data is when all the different data **P**oints are close together, reducing the numerical range of your data measurements.

Sources of error in pH measurements

A calibrated pH meter is going to provide you with the most accurate and the most precise pH values. This is due to the equipment being designed to specifically measure the concentration of hydrogen ions that correlates to a pH value using the pH equation. Therefore, **repeatability** should achieve precise data with a very low range between your data points.

Systematic errors can occur with a non-calibrated pH meter because, while it is likely to still produce precise pH values, the accuracy of these repeated measures to the true pH value are skewed — hence the need to calibrate the equipment.

Using either commercial or natural pH indicators are just that – an *indication* of pH. Therefore, they are not going to be exact measurements. Instead, they provide an indication of your solution being acidic or alkaline and roughly what pH it is. You would not use a pH indicator if you want your solution to be exactly pH 7.62, for example. You would use a pH meter for this.

Random errors can result in pH values that are not accurate and not precise. Mistakes or **personal errors** should not be included in reporting and analysis, as part of the ethical consideration of data handling. Rather, the experiment should be repeated correctly. Other errors in technique could also result in a spread of data points — for example, by:

- not stirring your solution while measuring the pH — this would mean that your results may fluctuate (with not precise values) because the solution is not homogenous
- using a faulty pH meter — the equipment may not have been taken care of properly, such as not being cleaned correctly, not topping up potassium chloride in the glass bulb, or being too rough with the glass bulb causing it to break or crack
- using a pH indicator not suitable for the reaction you are testing.

Random errors and mistakes can produce **uncertainty** in the results collected.

precise how close multiple measurements of the same investigation are to each other, a measure of repeatability or reproducibility

true value the value that would be found if the quantity could be measured perfectly

measurement result a final result; usually the average of several measurement values

repeatability how close the results of successive measurements are to each other in the exact same conditions

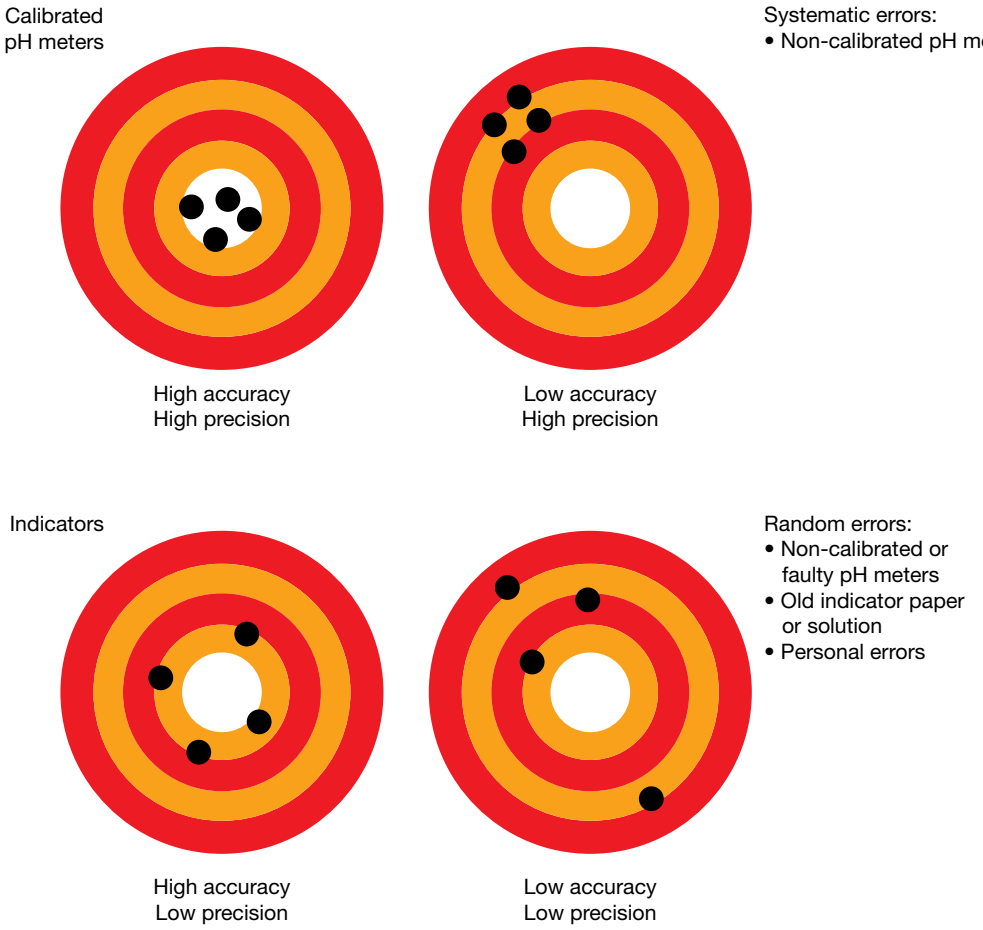
systematic errors errors that affect the accuracy of a measurement and cannot be improved by repeating an experiment; usually due to equipment or system errors

random errors chance variations in measurements; result in a spread of readings

personal errors human errors or mistakes that can affect results but should not be included in analysis

uncertainty a limit to the precision of data obtained; a range within which a measurement lies

FIGURE 11.26 Accuracy and precision of pH measurements



tlvd-3411

SAMPLE PROBLEM 12 Comparing accuracy and precision

Students conducted an experiment to compare the pH of tap water (given as pH 7.35) to carbonated water after using a SodaStream on its highest bubble level. Students used a pH meter. The expected reading of the carbonated water was pH 4.70
They repeated the experiment four times and achieved the following results.

TABLE pH of carbonated water

Student 1	Student 2	Student 3
pH 4.68	pH 4.21	pH 4.35
pH 5.08	pH 4.14	pH 4.70
pH 4.73	pH 4.23	pH 5.11
pH 4.84	pH 4.19	pH 4.79

a. **Which student had the least accurate data?**
b. **Which student had the least precise data?**
c. **Was the student with the most precise data also the student with the most accurate data? Explain your answer.**
d. **Suggest what Student 2 could do to improve their pH values.**

THINK	WRITE
a. 1. Recall what accuracy means.	Accuracy is how close a measurement is to a known value.
2. Explore the data of the three students.	• Student 1 had one pH value that was 0.02 below the expected value and one value 0.38 above the expected value. Their average value was 4.83, which is 0.13 above the expected value. • Student 2 had pH value results that were consistently below the expected value. Their average value was 4.19, which is 0.51 below the expected value. • Student 3 had one pH value that was 0.35 below the expected value and one value 0.41 above the expected value. Their average value was 4.74, which is 0.04 above the expected value.
3. Determine the student with the least accurate pH values.	Student 2 had the least accurate data, because their values were furthest away from the actual value.
b. 1. Review what precision means.	Precision is how close multiple measurements of the same investigation are to one another.
2. Explore the data of the three students.	• Student 1 had a pH data range of 0.40 (5.08 – 4.68) • Student 2 had a pH data range of 0.09 (4.23 – 4.14) • Student 3 had a pH data range of 0.76 (5.11 – 4.35)
3. Determine which student had the least precise data.	Student 3 had the least precise data, because they had a larger range between values.
c. 1. Identify the students with the most accurate data and the most precise data.	Using the results from **a** and **b**, it can be seen that Student 1 had the most accurate data and Student 2 had the most precise data.
2. Respond to the question and explain your answer.	The student who had the most precise data was not the student who had the most accurate data. Students may have pH values very close together (precise), but the data may not be accurate. This may be due to systematic errors of their pH meter. Data may also be accurate without being precise; you can be close to the target but have inconsistent readings.
d. 1. Identify Student 2's results.	Student 2 has precise data (data range = 0.09), but not accurate data (0.51 below the expected value). This suggests a possible systematic error with the pH meter.
2. Respond to the question and explain your answer.	Given Student 2 has a possible systematic error with their pH meter, the student should re-calibrate the pH meter using solutions of known pH, and then re-measure the carbonated water. This should bring the precise data close to the expected value.

PRACTICE PROBLEM 12

Students conducted an experiment to determine the pH of their school's local creek to determine the health of the waterway. Students used a pH meter, with the reading expected to be pH 8.25.

They repeated the measurements four times each and achieved the following results.

TABLE pH of creek water

Student 1	Student 2	Student 3
pH 8.23	pH 8.78	pH 8.90
pH 8.18	pH 8.89	pH 8.91
pH 8.21	pH 8.30	pH 8.87
pH 8.32	pH 8.45	pH 8.89

a. Which student had the least accurate data?
b. Which student had the least precise data?
c. Was the student with the most precise data also the student with the most accurate data? Explain your answer.

11.5 Activities

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11.5 Quick quiz on | 11.5 Exercise | 11.5 Exam questions on

11.5 Exercise

1. MC Identify the main chemical pigments in some fruits, vegetables and flowers that provide their vibrant colours of reds, purple and blues and are capable of changing colours with a change in hydrogen ion concentrations.
   A. Anthocyanin B. Betanin C. Beta carotene D. Melanin
2. MC Of the commercial pH indicators used within the school laboratory, which one more closely represents the natural pH indicator qualities of anthocyanin molecules?
   A. Phenolphthalein B. Bromothymol blue C. Universal indicator D. Methyl orange
3. Explain two limitations of accuracy when using pH meters.
4. Explain two limitations of accuracy when using pH indicators.
5. Predict your results to the following sequential scenarios of testing your pool water that has an expected pH of 7.40.
   a. You boiled raspberries and strained the juice to use as an indicator. You collected pool water samples in plastic cups in triplicate. You added some of your raspberry indicator to each of your pool water samples.
   b. Using a commercial phenolphthalein indicator, you repeatedly tested your pool water that had a pH of around 7. You counted at least 8 sample cups.
   c. Using methyl red, you tested your pool water again.
   d. You tested your pool using a calibrated pH meter.

11.5 Exam questions

Question 1 (1 mark)

MC Red cabbage contains anthocyanins, a class of pigment capable of changing colour with different pH environments. Red cabbage was boiled for 15 minutes in distilled water before being strained to produce a natural pH indicator that appears purple at a neutral pH. When testing the pH indicator in a white vinegar (ethanoic acid), identify the colour of the indicator and if the anthocyanin molecules are being protonated or deprotonated.

A. Blue, protonated **B.** Blue, deprotonated **C.** Red, protonated **D.** Red, deprotonated

Question 2 (1 mark)

Thymol blue (abbreviated as HTb) is an indicator that, despite its name, changes colour from red to yellow according to the equation:

$$\underset{\text{red}}{HTb(aq)} + H_2O(1) \rightleftharpoons \underset{\text{yellow}}{Tb^-(aq)} + H_3O^+(aq)$$

What colour is the indicator when the concentration of HTb is approximately equal to the concentration of Tb^-?

Question 3 (1 mark)

MC If you use a phenolphthalein pH indicator and add it to a colourless solution, and that solution turns a deep crimson, what can be determined?

A. The solution is acidic.
B. The solution is neutral.
C. The solution is alkaline.
D. The result is inconclusive.

Question 4 (6 marks)

During a dinner party, one of your guests accidently spilt some red wine on your white carpet. You had heard that blotting the area with water and white vinegar (ethanoic acid) worked well. After trying this, your carpet still had a bright red stain and was now quite wet. You had also heard that sprinkling baking soda onto the carpet helps to dry the carpet and pull the stain out. However, after trying this, before your eyes, the stain started to turn blue!

a. What is the molecule in the red wine responsible for this colour change? **(1 mark)**
b. Given that red wine is red in colour, what would its pH be? **(1 mark)**
c. What is the chemical name for baking soda and what is its pH? **(2 marks)**
d. What happens to the pH-sensitive molecule when baking soda is sprinkled on top? **(2 marks)**

Question 5 (11 marks)

You have prepared a natural pH indicator from red cabbage by boiling the cabbage for 15 minutes in distilled water, and then straining out the solids so you are left with the indicator solution.

a. If an aqueous solution turns your red cabbage natural pH indicator blue, describe the relationship between the hydronium ion [H_3O^+] and hydroxide ion [OH^-] concentrations using the < (less than), > (greater than) or = (equals) signs. **(1 mark)**
b. **i.** What type of solution would cause the indicator to turn blue? **(1 mark)**
ii. Give an example. **(1 mark)**
c. You decided to blow bubbles through a straw within your solution that appeared blue with the red cabbage pH indicator.
i. Describe the colour change that you would expect to observe. **(1 mark)**
ii. Is the solution becoming more acidic or more alkaline? **(1 mark)**
d. At the chemical level, describe what is happening to cause this colour change (include CO_2, H_2O and the anthocyanin molecule). **(2 marks)**
e. If you added some soda water to your solution from c., what would you expect to observe and why? **(2 marks)**
f. Your favourite colour was the starting blue colour. In general, what could you do to get your solution back to that colour? **(2 marks)**

More exam questions are available in your learnON title.

11.6 Neutralisation reactions to produce salts

KEY KNOWLEDGE

- Neutralisation reactions to produce salts:
  - reactions of acids with metal carbonates and hydroxides including balanced full and ionic equations, with states
  - types of antacids and their use in the neutralisation of stomach acid

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11.6.1 Salts

Acid reactions occur all around us and in us, and are used in agriculture and industry. Many industrial products rely on the use of acids as raw materials. The physical properties of acids and bases have been discussed, so it is appropriate to look at some chemical properties of acids, in particular. In chemical reactions, a new product is formed and, as with many other aspects of chemistry, patterns in the products make it easier to learn what happens. A salt solution is often a product of a reaction of an acid with another substance. If the solid salt is required, the solution can be heated to evaporate the water. The specific reactions that will be studied are reactions of acids with metals, carbonates and hydroxides.

Salt is commonly thought of as a substance that is added to food. This salt has been used for flavouring and as a food preservative for thousands of years. In chemistry, however, other salts exist besides table salt, sodium chloride.

- A salt is an ionic compound formed when an acid is neutralised by a base.
- Ionic compounds consist of a positive ion (cation), usually a metal, that comes from the base and a negative ion (anion), usually a non-metal, that comes from the acid. These compounds and their reactions were studied in topic 4. Examples of salts include KCl, Na_2CO_3 and $Zn(NO_3)_2$ (see table 11.7).

TABLE 11.7 Some common salts and their applications

Name	Formula	Application
Ammonium sulfate	$(NH_4)_2SO_4$	Fertiliser
Barium sulfate	$BaSO_4$	X-ray of digestive tract
Calcium sulfate dihydrate	$CaSO_4 \cdot 2H_2O$	Plasterboard
Copper sulfate pentahydrate	$CuSO_4 \cdot 5H_2O$	Dyeing; fungicides
Sodium chloride	NaCl	Table salt
Potassium chloride	KCl	Sodium-free salt substitute

11.6.2 Reactions of acids with metals

The reaction of an acid on a metal depends on the type of metal, and the strength and concentration of the acid. Many metals, but not all, react with acids and the reactions can be very vigorous (Na, K, Li) to very slow (Ni) or not at all (Cu, Hg, Au, Ag). Other metals that react with acid include Ca, Mg, Mn, Zn, and Fe. The general reaction for an acid and a metal is as follows.

Reactions of acids and metals

$$\text{Acid} + \text{metal} \rightarrow \text{salt} + \text{hydrogen}$$

The salt produced depends on the acid used:

- If the acid is hydrochloric acid, the salt is a chloride salt.
- If the acid is nitric acid, the salt is a nitrate salt.
- If the acid used is sulfuric acid, the salt is a sulfate.

For example:

Hydrochloric acid	+	magnesium	→	magnesium chloride	+	hydrogen
$2HCl(aq)$	+	$Mg(s)$	→	$MgCl_2(aq)$	+	$H_2(g)$

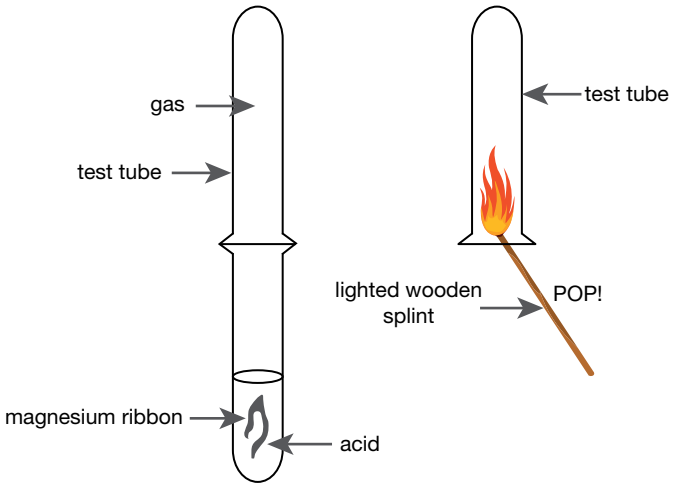
Even though this reaction involves an acid, it is not officially classified as an acid–base reaction because no proton transfer occurs. It is a redox reaction, which involves electron transfer. Redox reactions are discussed in topic 12.

The production of hydrogen can be tested in this experiment using the **'pop' test**. The gas produced can be collected above the reaction text tube (because hydrogen is lighter than air) and then a lit match or ignited splint can be placed at the opening. A 'pop' sound is heard if enough gas is collected (figure 11.28), showing the presence of hydrogen gas.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

FIGURE 11.27 Measuring the volume of gas produced during a reaction between an acid and zinc

FIGURE 11.28 Pop test for hydrogen



Ionic equations

As with **precipitation reactions**, it is possible to write ionic equations for reactions of acids. Ionic equations include only the species that have taken part in the reaction and omit spectator ions.

For the reaction of magnesium and hydrochloric acid, the acid is ionised and so exists as separate ions in solution. The product, magnesium chloride, is an ionic compound that will be dissociated in solution.

Inspection of this expanded equation shows that the chloride ions are the only species that have not taken part in the reaction. This means that they can be omitted from the ionic equation.

$$2H^+(aq) + \cancel{2Cl^-(aq)} + Mg(s) \rightarrow Mg^{2+}(aq) + \cancel{2Cl^-(aq)} + H_2(g)$$

The ionic equation for this reaction is $2H^+(aq) + Mg(s) \rightarrow Mg^{2+}(aq) + H_2(g)$.

'pop' test a test where a lighted wooden splint is placed in the vicinity of captured hydrogen gas, causing a 'pop' sound in its presence

precipitation reaction reaction where two solutions are mixed and an insoluble solid is formed

tlvd-3412

SAMPLE PROBLEM 13 Writing full and ionic equations for a reaction between an acid and a metal

For the reaction between dilute nitric acid and zinc metal, write the following.

a. The full equation representing the reaction

b. The ionic equation representing the reaction

THINK	WRITE
a. 1. Write out the correct formulas of the reactants. The general equation for this reaction is acid + metal → salt + hydrogen.	Reactants are HNO_3 and Zn. Acid: HNO_3 Metal: Zn
2. The salt formed from nitric acid is a nitrate salt. Write out the correct formulas for the products.	$Zn(NO_3)_2 + H_2$
3. Write out a skeleton equation.	$HNO_3 + Zn \rightarrow Zn(NO_3)_2 + H_2$
4. Balance the equation.	$2HNO_3 + Zn \rightarrow Zn(NO_3)_2 + H_2$
5. Add states to complete the full equation.	$2HNO_3(aq) + Zn(s) \rightarrow Zn(NO_3)_2(aq) + H_2(g)$
b. 1. Write out the equation showing ions present.	$2H^+(aq) + 2NO_3^-(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + 2NO_3^-(aq) + H_2(g)$
2. Write the ionic equation by omitting the species that have not reacted — in this case, the NO_3^- ions.	$2H^+(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + H_2(g)$

PRACTICE PROBLEM 13

For the reaction between magnesium metal and dilute sulfuric acid, write the following.

a. The full equation representing the reaction

b. The ionic equation representing the reaction

11.6.3 Reactions of acids with carbonates

When having a break from writing equations, you may decide to have a piece of cake. Consider while you are eating your cake that it is likely that the cake is light and fluffy because a reaction occurred while it was cooking between sodium hydrogen carbonate (baking soda), $NaHCO_3$, and an acid to produce carbon dioxide, CO_2. The carbon dioxide made the cake rise.

Like reactions of acids with metals, the reaction between acids and metal carbonates produce salt and a gas but this time the gas is carbon dioxide and not hydrogen. The salt is the same and water is also produced.

If the metal carbonate is in solution, the equations for the reaction are as follows.

Reactions of acids and carbonates

$$\text{Acid} + \text{metal carbonate} \rightarrow \text{salt} + \text{carbon dioxide} + \text{water}$$

For example: $2HCl\,(aq) + Na_2CO_3\,(aq) \rightarrow 2NaCl\,(aq) + CO_2\,(g) + H_2O\,(l)$

We can now convert the aqueous ionic compounds into their respective cation and anions and so this equation can be written as follows.

$$2H^+\,(aq) + \cancel{2Cl^-\,(aq)} + \cancel{2Na^+\,(aq)} + CO_3^{2-}\,(aq) \rightarrow \cancel{2Na^+\,(aq)} + \cancel{2Cl^-\,(aq)} + CO_2\,(g) + H_2O\,(l)$$

The sodium and chloride ions have not reacted and so are spectator ions, which we can cross out as shown. The resulting ionic equation is

$$2H^+\,(aq) + CO_3^{2-}\,(aq) \rightarrow CO_2\,(g) + H_2O\,(l)$$

If the metal carbonate is in *solid* form, the equation would be as follows.

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

This equation can be written as the ions are not dissociated.

$$2H^+\,(aq) + \cancel{2Cl^-\,(aq)} + Na_2CO_3\,(s) \rightarrow 2Na^+\,(aq) + \cancel{2Cl^-\,(aq)} + CO_2\,(g) + H_2O\,(l)$$

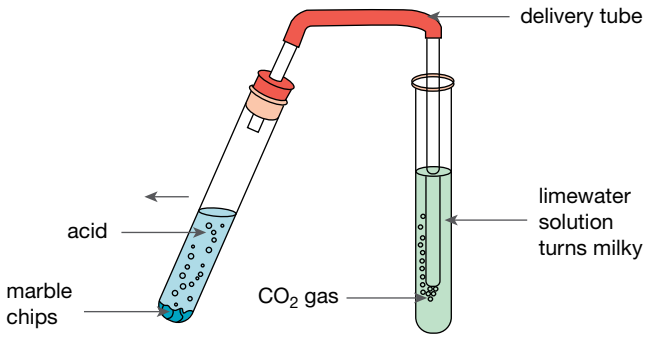
Only the chloride ions have not reacted and so are spectator ions. The ionic equation is

$$2H^+\,(aq) + Na_2CO_3\,(s) \rightarrow Na^+\,(aq) + CO_2\,(g) + H_2O\,(l)$$

To test for carbon dioxide, direct the gas produced (carbon dioxide is heavier than air) into a tube of limewater, $Ca(OH)_2$. If enough gas is produced, a precipitate of calcium carbonate forms.

$$Ca(OH)_2\,(aq) + CO_2\,(g) \rightarrow CaCO_3\,(s) + H_2O\,(l)$$

FIGURE 11.29 Test for carbon dioxide



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SAMPLE PROBLEM 14 Writing full and ionic equations for a reaction between an acid and a carbonate

For the reaction between dilute sulfuric acid and solid potassium carbonate, write the following.

a. The full equation representing the reaction

b. The ionic equation representing the reaction

THINK	WRITE
a. 1. Write out the correct formulas of the reactants. The general equation is acid + metal carbonate → salt + carbon dioxide + water.	Reactants: H_2SO_4 and K_2CO_3 Acid: H_2SO_4 Metal carbonate:K_2CO_3

2.	The salt formed from sulfuric acid is a sulfate salt. Write out the correct formulas for the products.	$K_2SO_4 + CO_2 + H_2O$
3.	Write out a skeleton equation.	$H_2SO_4 + K_2CO_3 \rightarrow K_2SO_4 + CO_2 + H_2O$
4.	Balance the equation.	$H_2SO_4 + K_2CO_3 \rightarrow K_2SO_4 + CO_2 + H_2O$
5.	Add states to complete full equation.	$H_2SO_4\ (aq) + K_2CO_3\ (s) \rightarrow K_2SO_4\ (aq) + CO_2\ (g) + H_2O\ (l)$
b. 1.	Write out the expanded equation to show ions present.	$2H^+\ (aq) + SO_4^{2-}\ (aq) + K_2CO_3\ (s) \rightarrow 2K^+\ (aq) + SO_4^{2-} + CO_2\ (g) + H_2O\ (l)$
2.	Write the ionic equation by omitting the spectator ions — in this case, the SO_4^{2-} ions.	$2H^+\ (aq) + K_2CO_3\ (s) \rightarrow 2K^+\ (aq) + CO_2\ (g) + H_2O\ (l)$

PRACTICE PROBLEM 14

For the reaction between dilute ethanoic acid and solid sodium carbonate, write the following.

a. The full equation representing the reaction

b. The ionic equation representing the reaction

Metal hydrogen carbonates also react with acids to produce a salt, carbon dioxide and water; for example, the reaction of hydrochloric acid and sodium hydrogen carbonate is as follows.

$$HCl(aq) + NaHCO_3(s) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

11.6.4 Reactions of acids with hydroxides

When an acid and metal hydroxide react, they usually produce a salt and water. This is called neutralisation. The general equation is as follows.

Reactions of acids and hydroxides

$$\text{Acid} + \text{metal hydroxide} \rightarrow \text{salt} + \text{water}$$

FIGURE 11.30 Indicator changing colour showing neutralisation has occurred

For example: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

This equation can be written as follows.

$$H^+(aq) + \cancel{Cl^-(aq)} + \cancel{Na^+(aq)} + OH^-(aq) \rightarrow \cancel{Na^+(aq)} + \cancel{Cl^-(aq)} + H_2O(l)$$

If the spectator ions are omitted, the ionic equation is as follows.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Metal oxides also react with acids to produce a salt and water.

tlvd-3414

SAMPLE PROBLEM 15 Writing full and ionic equations for a reaction between an acid and a hydroxide

For the reaction between dilute nitric acid and potassium hydroxide solution, write the following.
a. The full equation representing the reaction
b. The ionic equation representing the reaction

THINK	WRITE
a. 1. Write the correct formulas of the reactants. The general equation is acid + metal hydroxide → salt + water.	Reactants: HNO_3 and KOH Acid: HNO_3 Metal hydroxide: KOH
2. The salt formed from nitric acid is a nitrate salt. Write out the correct formulas for the products.	$KNO_3 + H_2O$
3. Write out a skeleton equation.	$HNO_3 + KOH \rightarrow KNO_3 + H_2O$
4. Balance the equation.	$HNO_3 + KOH \rightarrow KNO_3 + H_2O$
5. Add states to complete full equation.	$HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$
b. 1. Write out the equation showing all of the ions present.	$H^+(aq) + NO_3^-(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + NO_3^-(aq) + OH^-(aq) + H_2O(l)$
2. Write the ionic equation by omitting the spectator ions — in this case, the K^+ and NO_3^- ions.	$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

PRACTICE PROBLEM 15

For the reaction between dilute hydrochloric acid and magnesium hydroxide solution, write the following.
a. The full equation representing the reaction
b. The ionic equation representing the reaction

An unusual acid–base reaction does not involve water. Pure hydrogen chloride, HCl, is a gas; it can be dissolved in water to produce a concentrated solution. Both concentrated HCl and ammonium hydroxide, NH_4OH, produce fumes. These fumes react to produce a fine white solid, ammonium chloride, NH_4Cl.

Summary of general reactions for acids

- Acid + metal → salt and hydrogen
- Acid + metal carbonate → salt + water + carbon dioxide
- Acid + metal hydroxide → salt + water

FIGURE 11.31 Hydrogen chloride gas and ammonia gas reacting to form a salt

elog-1809

tlvd-3952

EXPERIMENT 11.4 online only

The reactions of an acid

Aim

To investigate the reactions of a typical acid (dilute hydrochloric acid) with metals, metal hydroxides and metal carbonates

elog-1810

tlvd-0633

EXPERIMENT 11.5 online only

Making copper sulfate crystals

Aim

To observe the reaction between copper oxide and sulfuric acid

elog-1811

tlvd-0634

EXPERIMENT 11.6 online only

The reactivity of metals and their salt

Aim

To react various metals and metal salts with one another to determine a simple order of reactivity

Neutralising acid reflux with antacids

Hydrochloric acid, HCl, is produced in our stomachs to assist enzymes to catalyse the digestion of the proteins in the food that we eat. If too much acid is produced, we can feel pain, often referred to as acid reflux or heartburn. Relief of this discomfort can be achieved by using antacids, which neutralise some of the excess acid. Different types of antacids are available, which ultimately influences how long it takes for them to start working, how long they are biologically active for, what other medications they may interact with and, therefore, which types of patients they may be suitable for.

Antacids can contain bases such as the following, either singularly or in combination:

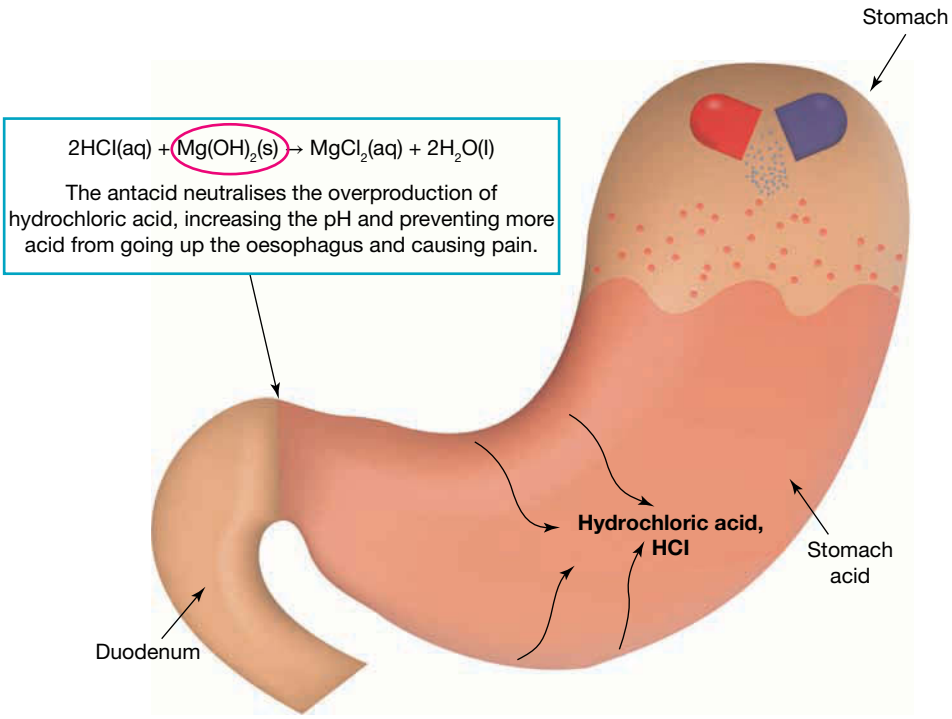
- aluminium hydroxide
- magnesium hydroxide
- magnesium carbonate
- calcium carbonate
- sodium bicarbonate.

Magnesium hydroxide is an insoluble powder; however, it will react with the hydrochloric acid in the stomach to form aqueous magnesium chloride and water (see figure 11.32), according to the reaction:

$$2HCl\,(aq) + Mg(OH)_2(s) \rightarrow MgCl_2(aq) + 2H_2O\,(l)$$

Some antacids, such as Gaviscon, are sometimes combined with alginates involving polymer chemistry. These alginates form cross-linked polymers within the stomach that create a floating barrier on the surface of the stomach fluid. This prevents any further stomach acid from travelling back up the oesophagus and causing the painful ‘heartburn’ symptoms. This works together with the neutralisation reactions to relieve the symptoms.

FIGURE 11.32 Neutralisation of hydrochloric acid in the stomach by the antacid magnesium hydroxide



11.6 Activities

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11.6 Quick quiz on | **11.6 Exercise** | **11.6 Exam questions**

11.6 Exercise

1. Complete the following table.

Acid	Formula of acid	Type of salt formed from the acid	Anion present
Hydrochloric			
		Nitrate	
			SO_4^{2-}
Ethanoic			

2. Write the general equations for the reactions of acids with the following.
   - **a.** Metals
   - **b.** Metal carbonates
   - **c.** Metal hydroxides
3. Complete and balance the following equations.
   - **a.** Zn (s)+ HBr (aq)
   - **b.** Ca (s)+ H_2SO_4 (aq)
   - **c.** K (s)+ HNO_3 (aq)
4. Complete and balance the following equations.
   - **a.** $MgCO_3(s) + CH_3COOH(aq)$
   - **b.** $Na_2CO_3(s) + HCl(aq)$
   - **c.** $MnCO_3(s) + HNO_3(aq)$

5. Complete and balance the following equations.
   a. $NaOH(aq) + H_2SO_4(aq)$ **b.** $Ca(OH)_2(aq) + HCl(aq)$ **c.** $Mg(OH)_2(aq) + HNO_3(aq)$
6. Write ionic equations for part a of questions 3, 4 and 5. State the spectator ions.
7. For each of the following reactions, write full equations and ionic equations.
   a. The acid and solid metal that could be used to produce calcium sulfate
   b. The acid and metal carbonate that could be combined to produce zinc nitrate
   c. The acid and metal carbonate solution that could be combined to produce sodium chloride
8. **a.** Describe how to collect and test for hydrogen gas.
   b. Write a balanced equation for this reaction.
   c. If no reaction was observed when this test was applied, does this imply that no hydrogen was produced?
9. **a.** Describe how to collect and test for carbon dioxide gas.
   b. Write a balanced equation for this reaction.
10. Write both full and ionic equations for the reactions that result when the following substances are mixed. State the spectator ions. Assume all reactions go to completion.
    a. Nickel metal and dilute sulfuric acid
    b. Aluminium metal and dilute sulfuric acid
    c. Solid calcium carbonate and dilute nitric acid

11.6 Exam questions

Question 1 (1 mark)

MC Which of the following reactants would produce the same salt?

Reactants I: sodium hydroxide and nitric acid

Reactants II: nitric acid and magnesium carbonate

Reactants III: magnesium and nitric acid

A. Reactants I and II
B. Reactants II and III
C. Reactants I and III
D. All will result in a different salt since different reactants are used.

Question 2 (1 mark)

MC Which of the following shows the balanced equation for the reaction between aqueous methanoic acid and magnesium powder?

A. $HCOOH(aq) + Mg(s) \rightarrow Mg(HCO_2)_2(aq) + H_2O(l)$
B. $2HCOOH(aq) + Mg^{2+}(s) \rightarrow Mg(HCO_2)_2(aq) + H_2(g)$
C. $2HCOOH(aq) + Mg(s) \rightarrow Mg(HCO_2)_2(aq) + H_2(g)$
D. $HCOOH(aq) + Mg(s) \rightarrow Mg(HCO_2)(aq) + H_2O(l)$

Question 3 (1 mark)

Write the equation for the reaction between sulfuric acid and potassium hydroxide solutions. Include symbols of state.

Question 4 (1 mark)

Write the equation for the reaction between aqueous nitric acid and chips of calcium carbonate. Include symbols of state.

Question 5 (6 marks)

Write the full and ionic equations to demonstrate the reactions of how the following antacids neutralise stomach acid.

a. Magnesium hydroxide **(2 marks)**
b. Aluminium hydroxide **(2 marks)**
c. Magnesium carbonate **(2 marks)**

More exam questions are available in your learnON title.

11.7 Applications of acid–base reactions in society

KEY KNOWLEDGE

- Applications of acid–base reactions in society; for example, natural acidity of rain due to dissolved CO_2 and the distinction between the natural acidity of rain and acid rain, or the action of CO_2 forming a weak acid in oceans and the consequences for shell growth in marine invertebrates

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11.7.1 Acidification of the atmosphere

Acidic oxides react with water to form acids, and are usually oxides of non-metals. Some acidic oxides, such as CO_2, SO_2 and NO_x, are released into the atmosphere as a result of human activity and have an impact on the environment.

Natural acidity of rain

Normal rain is slightly acidic due to dissolution of carbon dioxide in water according to the following equation.

$$CO_2(g) + H_2O(g) \rightleftharpoons H_2CO_3(aq)$$

The product is the weak acid carbonic acid, which reacts slightly with water to produce a few hydronium ions, as per the following equation.

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

Acid rain

Many industrial processes involve the burning of coal, oil or some other fossil fuel. Many of these fuels contain sulfur in varying amounts. When sulfur is burned in air, it forms sulfur dioxide, SO_2. This gas is often released into the air in vast quantities.

It must be noted that sulfur dioxide is released in many natural processes as well. In particular, active volcanoes release a large amount of sulfur dioxide into the air. The atmosphere can cope with large quantities of sulfur dioxide if it is given time to disperse the gas and break it down. When a large number of industries are all producing the gas over a small area, however, it cannot disperse in the air fast enough and becomes too concentrated to be safe.

Sulfur dioxide, SO_2, can react with water to produce sulfurous acid (a weak acid). It can also react with oxygen in the air to produce sulfur trioxide, SO_3, which then reacts with water to produce the strong acid sulfuric acid. The equations for these three processes are:

$$SO_2(g) + H_2O(l) \rightleftharpoons H_2SO_3(aq)$$
$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$
$$SO_3(g) + H_2O(l) \rightleftharpoons H_2SO_4(aq)$$

Other gases contribute to the **acid rain** problem and many of these are produced by industry and by every car that uses the internal combustion engine.

In a car engine, the temperature produced by the spark plug is sufficient to cause the nitrogen and oxygen in the air to combine and form nitrogen monoxide, NO. Nitrogen monoxide combines spontaneously with oxygen in the air to form nitrogen dioxide, NO_2. This nitrogen dioxide reacts with rainwater to form nitric acid, another contributor to the problem known as acid rain.

acidic oxide oxide of a non-metal, such as CO_2 or SO_2

acid rain rain made acidic by mixing with atmospheric pollution, resulting in a pH less than 5

The equation for this reaction is:

$$3NO_2(g) + H_2O(g) \rightleftharpoons 2HNO_3(aq) + NO(g)$$

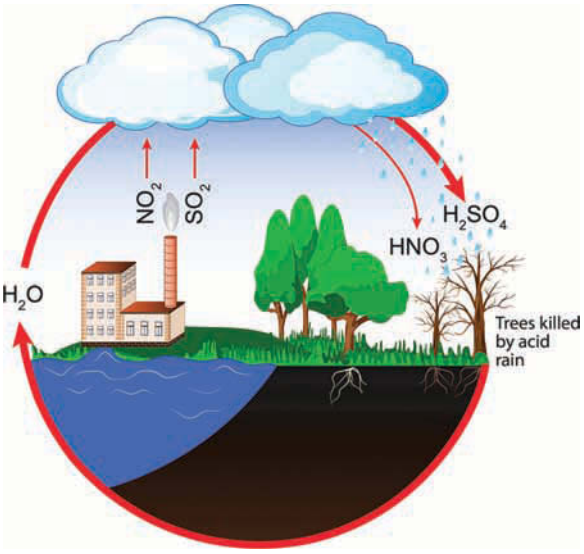
It must be noted, however, that carbonic acid and the oxides of sulfur are largely responsible for acid rain.

Impacts and control

Acid rain has been responsible for the defoliation of significant amounts of forest in both Europe and North America. In Australia, it is one of the principal causes of the 'lunar' landscape that occurred around Queenstown in Tasmania where large amounts of sulfur dioxide were produced as a result of copper ore smelting during the last century.

Where acid rain has run off into lakes, the pH of such lakes has in some instances dropped to the point where they have been unable to sustain life.

FIGURE 11.33 The production of acid rain



A range of responses has now evolved to reduce the input of gases into the atmosphere that cause acid rain. These include:

- catalytic converters in car exhausts to change nitrogen oxides into nitrogen and oxygen
- the use of natural gas as a fuel (natural gas contains lower levels of sulfur impurity) in place of other fuels such as oil and coal
- using renewable energy for electricity to replace coal-fired power stations
- taking measures to assist the dispersal of SO_2, such as by using taller chimney stacks
- the use of scrubbing, which is where exhaust gases are treated to remove undesirable emissions.

An increasing number of processes are now removing SO_2 from their emissions and concentrating it so that it can be used as a feedstock to produce sulfuric acid. This has an extra advantage because sulfuric acid is a valuable industrial acid and can be sold to offset the costs of the original process.

11.7.2 Acidification of the ocean

As we saw in subtopic 10.2, carbon dioxide dissolves in water to produce the weak acid carbonic acid, H_2CO_3. The oceans are slightly alkaline, but research has shown that, since the start of the Industrial Revolution, their pH has dropped from 8.2 to 8.1. While many localised variations occur in these figures, such a change is significant. This is directly attributable to the increased level of carbon dioxide in the atmosphere.

Increase in atmospheric CO_2 levels leads to the following flow-on effects:

- atmospheric CO_2 dissolves in ocean water to produce the weak acid carbonic acid, H_2CO_3
- carbonic acid, H_2CO_3, dissociates into bicarbonate ions, HCO_3^-, and hydrogen ions, H^+
- this increases the H^+ ion levels in oceans
- this decreases the pH in the ocean, leading to ocean acidity.

This sequence can be summarised according to the following equation:

$$\underset{\text{carbon dioxide}}{CO_2(g)} + \underset{\text{water}}{H_2O(l)} \rightleftharpoons \underset{\text{carbonic acid}}{H_2CO_3(aq)} \rightleftharpoons \underset{\text{bicarbonate}}{HCO_3^-(aq)} + \underset{\text{hydrogen ion}}{H^+(aq)}$$

Although the effect of this is the subject of ongoing research, problems that are the result of this acidification include:

- coral bleaching
- interference with marine organisms, particularly those at the lower end of the food chain such as oysters and pteropods

- reproductive disorders in certain marine organisms
- interference with shell building in shellfish and other similar organisms
- changes to food chains in ecosystems, potentially leading to dead zones.

The shells of many marine species have calcium carbonate as an important constituent. Others, including corals, form skeletons using calcium carbonate. Although calcium carbonate is virtually insoluble, Ca^{2+} and CO_3^{2-} ions are present in sea water at concentrations that are low but nevertheless important. However, carbonate ions can be removed by a reaction between themselves, water and carbon dioxide according to:

$$CO_2(g) + H_2O(l) + CO_3^{2-}(aq) \rightleftharpoons 2HCO_3^-(aq)$$

If the carbonate ion concentration is lowered too much, the calcium carbonate in the shells of marine organisms will dissolve in an attempt to replace the removed carbonate ions.

$$CaCO_3(s) \rightarrow Ca^{2+}(aq) + {}^{2-}_{3}(aq)$$

FIGURE 11.34 Laboratory experiment demonstrating the effect of acidic water on marine animal shells

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Weblink Ocean acidification

11.7 Activities

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11.7 Quick quiz on | **11.7 Exercise** | **11.7 Exam questions**

11.7 Exercise

1. What is acid rain?
2. Write a word and symbol equation showing the formation of carbonic acid.
3. Explain whether carbonic acid is a strong or weak acid. Use an equation in your answer.

4. Show using equations how the presence of SO_2 in the atmosphere results in the production of sulfuric acid.
5. List some sources of SO_2 in the atmosphere.
6. Describe three possible methods of controlling the input of SO_2 in the atmosphere.
7. Explain how acid–base reactions are involved in the acidification of the ocean.
8. Calculate the concentration of hydroxide ions and hydrogen ions in the ocean where the pH is 8.1.
9. Identify two problems that might occur as oceans become less basic.
10. Outline the effect the decrease in pH has on shell fish, including an equation to illustrate your answer.

11.7 Exam questions

Question 1 (1 mark)

MC Which of the following is unlikely to be a contributing factor to acid rain?

A. Combustion of hydrocarbons in cars and factories
B. Release of sulfur emissions from refining of metal ores
C. Addition of metal contaminants to the waterways
D. Emissions related to volcanic activity

Question 2 (1 mark)

MC Nitrogen dioxide, NO_2, is a brown gas produced in industrial areas. It is responsible for the brown colour of smog and produces acid rain through the formation of both nitric acid and the weak acid nitrous acid, HNO_2. Which of the following equations best shows this reaction?

A. $NO_2(g) + H_2O(l) \rightleftharpoons HNO_3(aq) + HNO_2(aq)$
B. $2NO_2(g) + H_2O(l) \rightleftharpoons HNO_3(aq) + HNO_2(aq)$
C. $NO_2(g) + H_2O(l) \rightleftharpoons H_2NO_3(aq)$
D. $2NO_2(g) + H_2O(l) \rightleftharpoons 2HNO_3(aq)$

Question 3 (1 mark)

MC Which of the following processes produces the least acid rain?

A. Combustion of hydrocarbons in cars and factories
B. Increased numbers of grazing animals producing methane
C. Release of sulfur emissions from refining of metal ores
D. Emission of gases from coal-fired electrical power plants

Question 4 (4 marks)

Freshwater lakes typically have a pH between 6 and 8. However, some lakes have a much lower pH. The crater lake on the Kawah Ijen volcano in Indonesia has a pH of 0.3.

a. What is the hydronium concentration in the Kawah Ijen crater lake? **(1 mark)**
b. Suggest a reason the pH of this lake is so low. **(1 mark)**
c. Would it be possible to cross the lake in an aluminium boat? Justify your response. **(2 marks)**

Question 5 (6 marks)

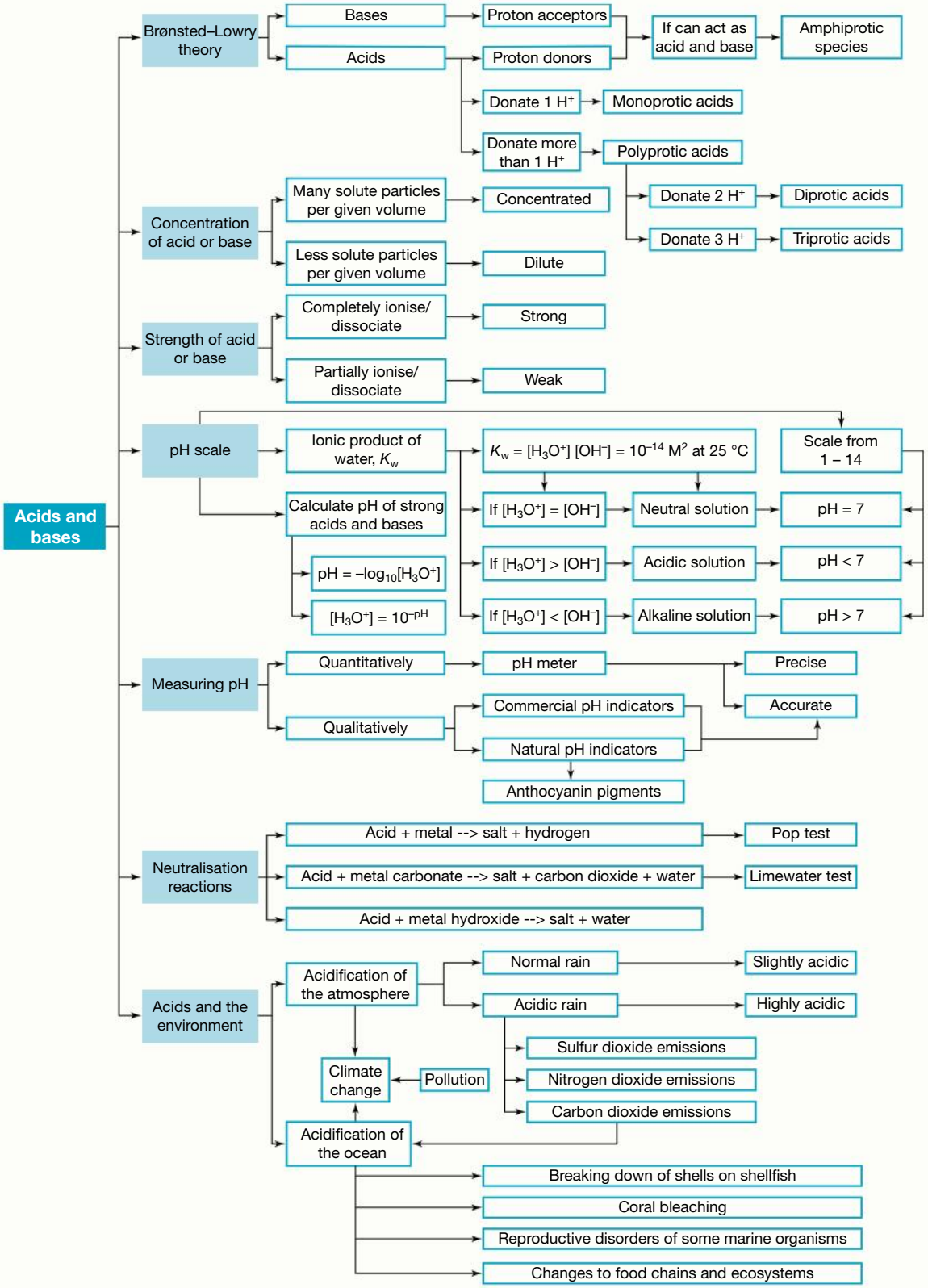
Increasing levels of carbon dioxide in the atmosphere are leading to ocean acidification.

a. Explain, using an equation, why this occurs. **(2 marks)**
b. Describe two effects of ocean acidification on marine life. **(2 marks)**
c. Describe two effects of ocean acidification on human life. **(2 marks)**

More exam questions are available in your learnON title.

11.8 Review

11.8.1 Topic summary



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11.8.2 Key ideas summary

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11.8.3 Key terms glossary

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Solutions	Solutions — Topic 11 (sol-0810)
Practical investigation eLogbook	Practical investigation eLogbook — Topic 11 (elog-1616)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 11 (doc-37087) Key ideas summary — Topic 11 (doc-37088)
Exam question booklet	Exam question booklet — Topic 11 (eqb-0092)

11.8 Activities

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11.8 Review questions

1. a. In terms of the Brønsted–Lowry theory, define an acid, a base and a proton-transfer reaction. You may use an equation in your explanation.
   b. You have been provided with a 0.5 M acid that is not labelled. It could be ethanoic acid or it could be hydrochloric acid. Describe how you could determine which it is purely by using your own home-made red cabbage indicator.
   c. Regarding your red cabbage indicator, what general molecule is present that enables it to act as an indicator and how does this work?
   d. What is a more accurate and precise method to measure the pH of the acid?
   e. Your acid turned purplish-pink in your red cabbage indicator. What is your acid? Write the reaction between your acid and water.
2. Antacid tablets are used to neutralise the hydrochloric acid in the stomach. Choose two bases mentioned in the discussion of antacids in section 11.6.4 to write neutralisation equations for the reactions that may happen in our bodies.
3. Write an equation to illustrate the amphiprotic nature of H_2O.
4. Sulfuric acid is a strong diprotic acid. Explain, using equations, what this term means.
5. In a pH test on a swimming pool, the water is found to have a pH of 4.5.
   a. What would you add to the water to raise the pH?
   b. The pH should be about 7.5 for best conditions. Is this solution acidic or alkaline?

6. **a.** What are conjugate acid–base pairs?
   b. State the formulas of the conjugate acid of the following bases:
   i. Cl^- **ii.** CO_3^{2-} **iii.** HSO_4^- **iv.** H_2O **v.** OH^-.
   c. State the formulas of the conjugate base of the following acids:
   i. HF **ii.** HCO_3^- **iii.** HSO_4^- **iv.** H_2O **v.** H_2S
7. Write equations to show the following.
   a. The hydrogen carbonate ion, HCO_3^-, is amphiprotic.
   b. Hydrogen sulfide is a weak diprotic acid.
   c. The reaction between sulfuric acid and sodium hydroxide.
8. Write both full and ionic equations for the reactions that result when the following substances are mixed. Assume all reactions go to completion.
   a. Solid sodium sulfide and phosphoric acid, H_3PO_4
   b. Solid copper(II) oxide and hydrochloric acid
   c. Potassium hydroxide solution and sulfuric acid
9. A solution of sodium hydroxide is prepared by dissolving 0.120 g in 250 mL of water.
   a. **i.** What is the concentration of sodium hydroxide in M in this solution?
   ii. What is the concentration of hydronium ions in this solution (M)?
   iii. Calculate the pH of the solution.
   b. The sodium hydroxide solution is diluted by a factor of 10.
   i. Will the solution become more acidic, more basic, less acidic or less basic?
   ii. What is the pH of the final solution?
   c. Sodium hydroxide is considered to be a strong base.
   i. Give an example of a strong monoprotic acid and a weak monoprotic acid.
   ii. Indicate which of these two acids would have a higher acidity constant, K_a.
10. **a.** What is the definition of the following?
   i. Accuracy
   ii. Precision
   b. How does this relate to the measurement of pH using either a pH meter or an indicator?

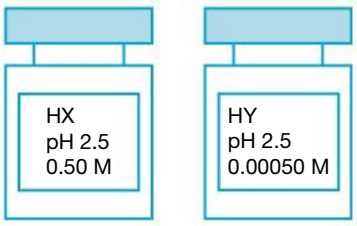
11.8 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC Two acids, HX and HY, are shown in the figure with their concentrations and pH values.



Which of the following statements is/are correct?

i. HX is a stronger acid than HY.
ii. Both acids ionise to the same extent in water.
iii. The reaction showing the ionisation of HY in water should be written with a double arrow.

A. (i) only
B. (ii) and (iii) only
C. (ii) and (iii) are correct
D. No statements are correct.

Question 2

MC Which of the following is the best definition of an acid?

A. An acid is an electron acceptor.
B. An acid is a substance that tastes sour.
C. An acid is a substance that donates H^+.
D. An acid is a substance that accepts H^+.

Question 3

MC Hydrogen sulfide, H_2S, is a gas at room temperature and dissolves in water. When this solution is tested, it conducts electricity and turns litmus solution red. Why is the solution acidic?

A. The H_2S has gained a proton.
B. The H_2O has gained a proton.
C. The H_2S and H_2O have gained protons.
D. The H_2S and H_2O have lost protons.

Question 4

MC Consider the following equations.

i. $Zn(s) + 2H_3O^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g) + 2H_2O(l)$
ii. $Cl^-(aq) + H_3O^+(aq) \rightleftharpoons HCl(aq) + H_2O(l)$
iii. $H_3O^+(aq) + OH^-(aq) \rightleftharpoons 2H_2O(l)$
Which can represent acid–base reactions?

A. (i) and (ii)
B. (i) and (iii)
C. (ii) and (iii)
D. (i), (ii) and (iii)

Question 5

MC When hydrogen chloride gas is dissolved in water, the resulting solution can conduct an electric current. What is the best explanation for this?

A. Ions in the hydrogen chloride gas are released when dissolved in water.
B. The water reacts with the hydrogen chloride to form ions.
C. The water molecules carry the electric current in one direction while the hydrogen chloride molecules carry it in the other direction.
D. The water contains impurities that react with the hydrogen chloride to form ions.

Question 6

MC Consider the reaction between methanoic acid and water:

$$HCOOH(aq) + H_2O(l) \rightleftharpoons HCOO^-(aq) + H_3O^+(aq)$$

Which of the following pairs are Brønsted–Lowry acids?

A. H_2O, HCOOH
B. HCOOH, H_3O^+
C. H_2O, H_3O^+
D. HCOOH, $HCOO^-$

Question 7

MC At 25 °C, which equation represents the self-ionisation of water?

A. $H_2O(l) + H_2O(l) \rightleftharpoons H_2O_2(aq) + O_2(g)$
B. $H_2O(l) + H_2O(l) \rightleftharpoons 4H^+(aq) + 2O^{2-}(aq)$
C. $H_2O(l) + H_2O(l) \rightleftharpoons 2H_2(g) + O_2(g)$
D. $H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

Question 8

MC What is the strength of an acid determined by?

A. The number of hydrogen ions present
B. Its concentration
C. Its degree of ionisation in aqueous solution
D. Its ability to change the colour of litmus

Question 9

MC What is the pH of 0.0001 M HCl solution?

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

Question 10

MC Which of the following reactions takes place when a metal carbonate reacts with an acid?

A. Acid + metal carbonate → salt + water + carbon dioxide
B. Acid + metal carbonate → salt + water + oxygen
C. Acid + metal carbonate → salt + water + hydrogen
D. Acid + metal carbonate → metal oxide + carbon dioxide

Section B — Short answer questions

Question 11 (8 marks)

Butanoic acid is a weak monoprotic acid.

a. What is meant by the term 'monoprotic'? **(1 mark)**
b. Write the equation for the ionisation of butanoic acid in water. **(1 mark)**
c. Write an equation for the neutralisation of butanoic acid by sodium hydroxide. **(1 mark)**
d. What is meant by the term 'weak'? Use your equation from part **b** to illustrate your answer. **(2 marks)**
e. Describe a test that could be used to show that butanoic acid is a weak acid. **(1 mark)**
f. Draw the structure of an butanoic acid molecule. Given that it is monoprotic, circle the hydrogen atom that is donated when it reacts with a base. **(1 mark)**

Question 12 (4 marks)

The dihydrogen phosphate ion, $H_2PO_4^-$, is amphiprotic.

a. What is meant by the term 'amphiprotic'? **(1 mark)**
b. Write two different equations for the reaction of the dihydrogen phosphate ion with water that demonstrate its amphiprotic nature. **(2 marks)**
c. Solutions of the dihydrogen phosphate ion in water are slightly acidic. Comment on this with respect to the two equations that you have written for part (**b**). **(1 mark)**

Question 13 (5 marks)

Ammonium sulfate is used as a fertiliser.

a. State two compounds that could be mixed to produce ammonium sulfate. **(1 mark)**

b. Write the full equation for this reaction. **(1 mark)**

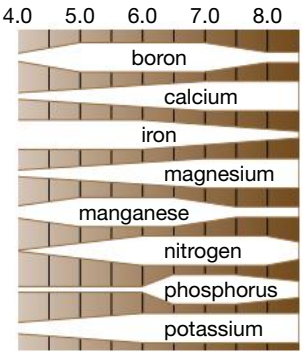
c. Write the ionic equation for this reaction. **(1 mark)**

d. How would solid ammonium sulfate be obtained? **(1 mark)**

e. When the solid fertiliser dissolves in water it dissociates. Write an equation showing this process. **(1 mark)**

Question 14 (11 marks)

Soil contains a number of different elements. Plants need to take up these elements in order to grow and flower. If the pH of soil is too high or too low, plants cannot take up these elements. The figure shows how the pH of the soil influences the amount of an element that can be taken up by a plant. The narrower the bar, the harder it is for the plant to take up the element.



a. Which elements can easily be absorbed at a pH below 4.5? **(1 mark)**

b. Which elements can be absorbed at a pH of 8? **(1 mark)**

c. Which elements cannot be taken up easily if the pH is 6? **(1 mark)**

d. Which elements cannot be taken up easily if the soil pH is 8? **(1 mark)**

e. Fuchsias require the soil to be acidic. Which elements would be most important to these plants? **(1 mark)**

f. Azaleas require a pH of around 5.5. Which elements are not important to these plants? **(1 mark)**

g. A gardener comes to you with a problem. He says that he cannot grow lettuce successfully any longer. It is not an insect problem or a water problem. Design an experiment you could do to find out the problem with his soil. (*Note:* lettuces need a lot of iron.) **(5 marks)**

Question 15 (6 marks)

The increase in carbon dioxide in the atmosphere is causing the pH of rain water to decrease.

a. Write the equation for this reaction. **(1 mark)**

b. Explain using a balanced equation how this acid can produce the ion that causes the decrease in pH. **(2 marks)**

c. Acid rain causes significant damage to limestone, $CaCO_3$, used in buildings. Write an equation showing the reaction of sulfuric acid with limestone. **(1 mark)**

d. Name three acidic oxides that contribute to acid rain and provide their formulas. **(2 marks)**

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12 Redox (electron transfer) reactions

KEY KNOWLEDGE

In this topic you will investigate:

Redox (electron transfer) reactions

- oxidising and reducing agents, and redox reactions including writing of balanced half and overall redox equations (including in acidic conditions), with states indicated
- the reactivity series of metals and metal displacement reactions, including balanced redox equations with states
- applications of redox reactions in society; for example, corrosion or the use of simple primary cells in the production of electrical energy from chemical energy.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of VCE Chemistry. Experiments and investigations, supported by a **practical investigation eLogbook** and **teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

EXAM PREPARATION

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12.1 Overview

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12.1.1 Introduction

Oxidation–reduction, or redox, reactions continually occur all around us and are one of the largest groups of chemical reactions. In redox reactions, one or more electrons are transferred between reactants. Oxidation and reduction reactions always happen together.

FIGURE 12.1 Fireworks are very rapid redox reactions.

Redox chemistry is a fundamental part of life — photosynthesis and respiration are redox reactions. Our bodies work by redox reactions — the food we eat is oxidised to enable us to obtain the energy we need to live. If we break a bone, electrical currents are generated around the damaged area as part of the healing process. The conduction of impulses in nerves and the supply of battery power in our mobile phones, calculators, computers and cars rely on spontaneous redox reactions. Many metals are extracted by reduction reactions, and municipal water supplies and swimming pools are treated with oxidising agents that act as bactericides and algaecides. Corrosion of metals, combustion of fuels and launching rockets into space relies on these reactions. Bushfires are uncontrolled redox reactions on a large scale, while fireworks and explosions are very fast redox reactions. Redox reactions are investigated in this topic as well as their applications in society.

LEARNING SEQUENCE

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12.2 Redox reactions 465
12.3 EXTENSION: Oxidation numbers 475
12.4 Reactivity series of metals 481
12.5 Applications of redox reactions 486
12.6 Review 498

on Resources

Solutions	Solutions — Topic 12 (sol-0811)
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Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 12 (doc-37089) Key idea summary — Topic 12 (doc-37090)
Exam question booklet	Exam question booklet — Topic 12 (eqb-0093)

12.2 Redox reactions

KEY KNOWLEDGE

- Oxidising and reducing agents, and redox reactions including writing of balanced half and overall redox equations (including in acidic conditions), with states

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12.2.1 Introduction to redox reactions

The word **redox** is a combination of shortened forms of the words '**reduction**' and '**oxidation**'. Redox reactions (reduction–oxidation reactions) were originally defined in terms of loss or gain of oxygen but are now characterised as *electron transfer reactions*.

- Oxidation is the loss of electrons.
- Reduction is the gain of electrons.

Equations for redox reactions can be split into separate **half-equations**, one showing oxidation and the other showing reduction. These can be balanced using a series of steps that are demonstrated in this topic. The reactivity of metals is explored as an example of redox reactions and important real-life applications, primary cells and corrosion, are investigated.

Early chemists knew that a gas called oxygen was essential for the survival of living things. This prompted them to study the reactions of oxygen with other substances in an effort to find out more about its behaviour. They found that many substances combined with oxygen, so the term *oxidation* was used to describe these reactions. The combustion reactions of hydrocarbons such as the burning of propane:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

and the burning of iron in air:

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

were described as oxidation reactions because a reactant, or some part of it, combined with oxygen.

Conversely, reactions that involved the decomposition of a compound, with the loss of oxygen, were called reduction reactions because the compound seemed to be 'reduced' to something simpler. For example, copper(II) oxide may be reduced to copper by hydrogen:

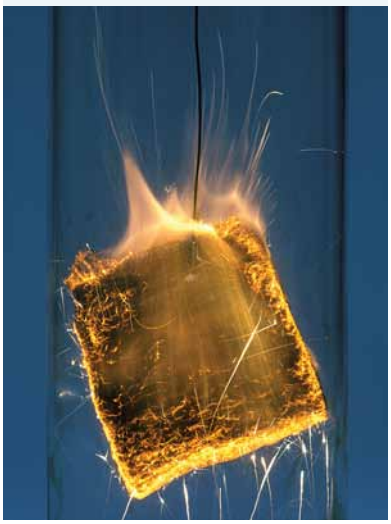
$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$$

Similarly, iron(III) oxide is reduced to iron by carbon monoxide in a blast furnace:

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

As one reactant is reduced, the other reactant is oxidised. Oxidation and reduction must occur at the same time.

FIGURE 12.2 Steel wool (iron) burning in air to form iron(III) oxide



redox a reaction that involves the transfer of one or more electrons between chemical species

reduction a gain of electrons (decrease in the oxidation number)

oxidation an increase in the oxidation number; a loss of one or more electrons

half-equation chemical equation showing either oxidation or reduction in a redox reaction

Redox reactions

- The word *redox* is a combination of shortened forms of the words 'reduction' and 'oxidation'.
- Redox reactions (oxidation–reduction reactions) were originally defined in terms of loss or gain of oxygen, but are now characterised as electron transfer reactions.
- Oxidation is the loss of electrons and reduction is the gain of electrons.

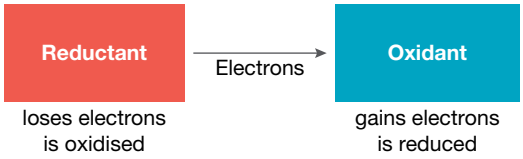
Oxidising and reducing agents

The substance that is being reduced is called the **oxidising agent** and the substance being oxidised is called the **reducing agent**. Although this sounds confusing, with practice you will become familiar with these terms.

- The oxidising agent *causes* the oxidation of *another substance* and is *itself* reduced.
- The reducing agent *causes* the reduction of *another substance* and is *itself* oxidised.

Oxidising agents are also known as **oxidants** and reducing agents are known as **reductants**.

FIGURE 12.3 Oxidant and reductant in redox reaction

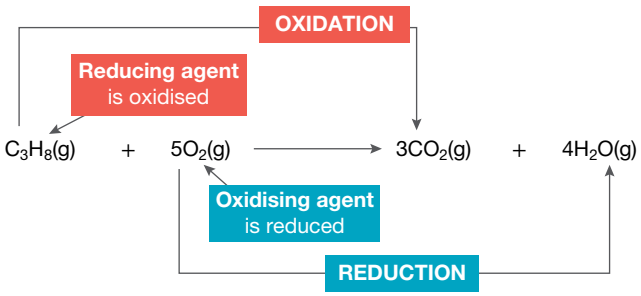


The oxidising agent is the species that is reduced.

The reducing agent is the species that is oxidised.

These terms are demonstrated in the reaction $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$ in figure 12.4.

FIGURE 12.4 Oxidation–reduction terminology in the equation $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$



oxidising agent the substance that causes oxidation of another substance; it accepts electrons and is reduced

reducing agent the substance that causes reduction of another substance; it donates electrons and is oxidised

oxidant the substance that causes oxidation of another substance; it accepts electrons and is reduced

reductant the substance that causes reduction of another substance; it donates electrons and is oxidised

12.2.2 Modern definition of oxidation and reduction

The modern definition of oxidation and reduction processes describes redox reactions in terms of electron transfer. For example, the combustion of magnesium in oxygen to form magnesium oxide involves the oxidation of magnesium.

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

A similar reaction occurs when magnesium chloride is formed by the combustion of magnesium in chlorine, but in this case no oxygen is involved.

$$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$$

Analysis of these two reactions *in terms of a transfer of electrons* reveals that they are essentially the same process. Both **products** are ionic substances and so contain Mg^{2+} ions. In the process of oxidation, each atom of magnesium has lost two electrons.

$$Mg \rightarrow Mg^{2+} + 2e^-$$

In general, a substance that is oxidised is one that loses electrons and is, therefore, an electron donor.

When magnesium reacts in this way with either oxygen or chlorine, these atoms accept these electrons and become O^{2-} and Cl^- anions in the ionic lattice of the products.

$$O_2 + 4e^- \rightarrow 2O^{2-}$$
$$Cl_2 + 2e^- \rightarrow 2Cl^-$$

In general, a substance that is reduced is one that gains electrons and is, therefore, an electron acceptor.

Redox reactions, therefore, are those in which electrons are transferred from one reactant to another.

Redox reactions are electron transfer reactions

- Oxidation is defined as a loss of electrons to another substance.
- Reduction is defined as a gain of electrons from another substance.
- This can be remembered using the acronym 'OIL RIG':

Oxidation

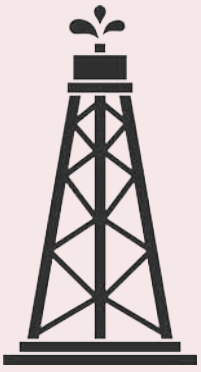
Is

Loss of electrons

Reduction

Is

Gain of electrons

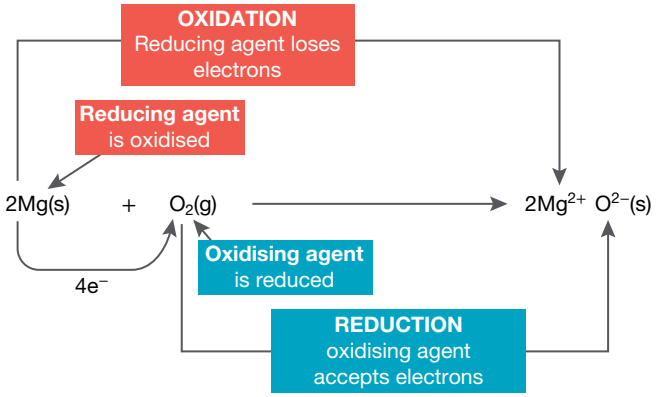


Oxidation must be accompanied by reduction; that is, oxidation cannot occur unless reduction occurs simultaneously.

product chemical species obtained as the result of a chemical reaction

In summary, the reactant that is losing electrons (undergoing oxidation) is called the reducing agent or reductant, because it transfers electrons to another substance and causes it to be reduced. The reactant that is gaining electrons (undergoing reduction) is called the oxidising agent or oxidant, because it accepts the electrons from the reducing agent and causes it to be oxidised. This is shown in figure 12.5.

FIGURE 12.5 Oxidation and reduction in terms of electron transfer



on Resources

Video eLesson Redox-electron transfer (eles-2495)

tlvd-3480

SAMPLE PROBLEM 1 Identifying species in a redox reaction

Explain why the following reaction is described as a redox reaction and identify the species oxidised and reduced.

$$Fe(s) + 2Ag^+(aq) \rightarrow 2Ag(s) + Fe^{2+}(aq)$$

THINK	WRITE
1. Redox reactions are electron transfer reactions. To decide if electrons have been transferred, compare reactants and products and identify the changes.	Fe(s) has become $Fe^{2+}(aq)$. $Ag^+(aq)$ has become Ag(s).
2. Identify which reactant has lost electrons and which reactant has gained electrons. The species that has lost electrons is oxidised. The species that has gained electrons is reduced. **TIP** When naming ions, it is important to refer to them as ions and not just state the metal. In this case, the silver **ion** is reduced.	Fe has lost two electrons to become Fe^{2+}. Ag^+ has gained an electron to become Ag. Fe is oxidised Ag^+ is reduced; that is, the silver ion is reduced.
3. Electron transfer has occurred.	This is a redox reaction because electron transfer has occurred from Fe to Ag^+.

PRACTICE PROBLEM 1

Explain why the following reaction is described as a redox reaction and identify the species oxidised and reduced.

$$Cl_2(aq) + 2Br^-(aq) \rightarrow Br_2(aq) + 2Cl^-(aq)$$

12.2.3 Redox equations

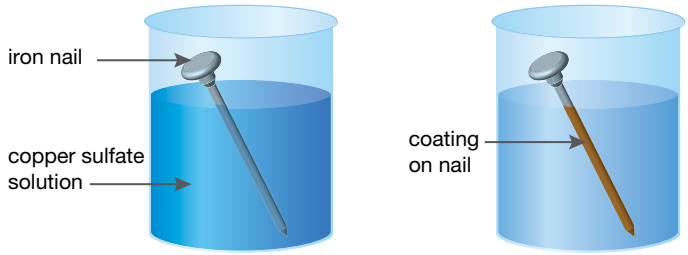
When writing overall equations (as opposed to half-equations) for redox reactions, we do not show electrons because all electrons given off during oxidation are taken in during reduction. However, if we are considering oxidation or reduction reactions separately, it is appropriate (and necessary) to write reactions that do show the electrons in half-equations as has been done in the earlier examples. It should be noted if you multiply the two half-equations by a factor, so the same number of electrons are on each side, and then add the half-equations together, you do get the overall ionic redox equation. There should be no electrons in the overall ionic equation.

Half-equations

Half-equations are a useful way of understanding the processes involved in a redox reaction. The following example describes how half-equations are written.

When an iron nail is placed in a blue copper sulfate solution, the nail becomes coated with metallic copper and the blue colour of the solution fades (see figure 12.6).

FIGURE 12.6 An iron nail placed in copper sulfate solution. The solution on the left is bright blue due to the copper sulfate. The one on the right is paler as the copper sulfate is reduced, and a coating of copper appears on the nail.



A redox reaction has taken place because electrons have been transferred from the iron nail to the copper ions in the solution, allowing solid copper to form. This type of reaction is called a **displacement reaction** and will be discussed in subtopic 12.3. The full equation for the reaction is as follows:

$$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$$

Since copper sulfate and iron sulfate are in solution, we can write an equation for this reaction showing all of the ions present.

$$Fe(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Fe^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$$

Then eliminating the sulfate spectator ions, the ionic equation becomes:

$$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$$

displacement reaction chemical reaction in which a more reactive element displaces a less reactive element from its compound

Although the oxidation and reduction reactions occur simultaneously, it is possible to consider the two reactions separately. To do this we separate the conjugate pair of oxidising agent and reducing agent. A **conjugate redox pair** is made up of two species that differ by a certain number of electrons. Each has its own half-equation.

So for the oxidation conjugate pair, we can write:

$$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$$

and for the reduction conjugate pair, we can write:

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

These half-equations are balanced with respect to both atoms and charge. Combining these two half-equations will yield the ionic equation already provided for the reaction as a whole. The conjugate redox pairs are:

$$Fe^{2+}(aq)/Fe(s)$$
$$Cu^{2+}(aq)/Cu(s)$$

conjugate redox pair two species that differ only by a certain number of electrons

Remembering which sides of half-equations electrons are on

$$\overset{+e^-}{RED}/\overset{+e^-}{OX}$$

The electrons are on the left for reduction and the right for oxidation.

tlvd-0577

SAMPLE PROBLEM 2 Writing ionic equations and balanced half-equations

For the following reaction, write the:
a. ionic equation **b. half-equation** **c. conjugate pairs.**

$$2NaBr(aq) + Cl_2(g) \rightarrow 2NaCl(aq) + Br_2(l)$$

THINK	WRITE
a. 1. Show all ions and other species present in the full equation.	$2Na^+(aq) + 2Br^-(aq) + Cl_2(g) \rightarrow 2Na^+(aq) + 2Cl^-(aq) + Br_2(l)$
2. Write the ionic equation, omitting spectator ions (in this case, Na^+).	$2Br^-(aq) + Cl_2(g) \rightarrow 2Cl^-(aq) + Br_2(l)$
b. 1. Write the half-equation for the first reactant, *balancing the charges* by adding electrons to the appropriate side.	$2Br^-(aq) \rightarrow Br_2(l) + 2e^-$
2. Write the half-equation for the second reactant, *balancing the charges* by adding electrons to the appropriate side.	$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$
c. Identify the conjugate pairs from each equation.	$Br_2(l)/Br^-(aq)$ $Cl_2(g)/Cl^-(aq)$

PRACTICE PROBLEM 2

For the following equation, write the:

a. ionic equation
b. half-equation
c. conjugate pairs.

$$2AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$

elog-1842

tlvd-4002

EXPERIMENT 12.1 online only

Simple redox equations

Aim

To investigate a simple redox reaction between iron and copper ions

Writing more complex redox reactions

At first glance, the statement that MnO_4^- is reduced to Mn^{2+} seems incorrect because, according to the charges on the ions, there appears to be a loss, rather than gain, of electrons:

$$MnO_4^-(aq) \rightarrow Mn^{2+}(aq)$$

It is only when the entire half-equation for the change is written that its true nature as reduction becomes obvious, with five electrons being accepted by each permanganate ion:

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$$

The steps involved in obtaining such equations in an acidified solution are provided here. (Reactions in an alkaline solution require a slightly different method, which is not required for this course.)

These equations *must* be split into two half-equations and then these half-equations need to be balanced before being added together.

To balance half-equations, first balance all elements except hydrogen and oxygen. Next, balance the oxygen atoms by adding water molecules, balance the hydrogen atoms by adding hydrogen ions, and balance the difference in charge by adding electrons. Multiply each half-equation by factors that will lead to the same number of electrons in each half-equation.

To balance each separate half-equation, use the KOHES procedure

1. Balance the **K**ey element; that is, the one undergoing oxidation or reduction.
2. Balance the **O**xygen atoms by adding water molecules.
3. Balance the **H**ydrogen atoms by adding hydrogen ions.
4. **E**qualise the charges on both sides of the equation by adding electrons to the more positive side.
5. Add symbols of **S**tate.

Remember to separate complex redox equations into two parts before attempting to balance them.

on Resources

Video eLesson Balancing redox equations (eles-2489)

tlvd-3481

SAMPLE PROBLEM 3 Writing a half-equation for reduction

Write the half-equation for the reduction of permanganate ions, $MnO_4^-(aq)$, to manganese ions, $Mn^{2+}(aq)$.

THINK	WRITE
1. To balance the key element, state the conjugate pair and write it as a skeleton equation, then balance the equation, except for oxygen and hydrogen, which will be balanced next.	$MnO_4^-(aq)/Mn^{2+}(aq)$ $MnO_4^- \rightarrow Mn^{2+}$ Manganese atoms are already balanced.
2. Balance oxygen atoms, where needed, by adding water.	$MnO_4^- \rightarrow Mn^{2+} + 4H_2O$
3. Balance hydrogen atoms, where needed, by adding H+.	$MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O$
4. Equalise the charge of each half-equation by adding electrons. The total charge on the left is $-1 + 8 = +7$. On the right side, the total charge is +2. So, five electrons must be added to the left side. **TIP** Take care when balancing electrons because this is commonly where mistakes are made.	$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$
5. Add states; these reactions generally occur in aqueous solutions so the states are usually aqueous but be aware of common gases (g) and water, which is (l).	$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+} + 4H_2O(l)$

PRACTICE PROBLEM 3

Write the half-equation for the oxidation of sulfur dioxide gas, SO_2, to sulfate ions, SO_4^{2-}, in solution.

The method described here is used to determine one of the two half-equations required to write an overall ionic equation for a redox reaction. The same method can be used to find the other half-equation; in this case, the reduction reaction. If the number of electrons is different in each half-equation, it is necessary to multiply one or both by a factor before adding them together. Electrons should never be in the final overall ionic equation, however, and nor should any spectator ions.

If water molecules or hydrogen ions are on both sides of the equation, cancel the same number of each from both sides to get the lowest number. For example, for the reaction of dichromate ions, $Cr_2O_7^{2-}$, and nitrite ions, NO_2^-, to form chromium, Cr^{3+}, and nitrate ions, NO_3^-, the following half-equations are obtained.

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) \quad [1]$$

$$NO_2^-(aq) + H_2O(l) \rightarrow NO_3^-(aq) + 2H^+(aq) + 2e^- \quad [2]$$

There are $6e^-$ in the first equation and $2e^-$ in the second. To ensure that the electrons cancel out when these equations are added, the second equation must be multiplied by three:

$$3NO_2^-(aq) + 3H_2O(l) \rightarrow 3NO_3^-(aq) + 6H^+(aq) + 6e^- \quad [3]$$

Now equations [1] and [3] can be added together:

$$Cr_2O_7^{2-}(aq) + {}^{8}\cancel{14}H^+(aq) + \cancel{6e^-} + 3NO_2^-(aq) + \cancel{3H_2O(l)} \rightarrow 2Cr^{3+}(aq) + {}^{4}\cancel{7}H_2O(l) + 3NO_3^-(aq) + \cancel{6H^+(aq)} + \cancel{6e^-}$$

By cancelling out and subtracting common terms — that is, $6e^-$, $3H_2O(l)$ and $6H^+(aq)$ — we are left with the final ionic equation:

$$Cr_2O_7^{2-}(aq) + 8H^+(aq) + 3NO_2^-(aq) \rightarrow 2Cr^{3+}(aq) + 4H_2O(l) + 3NO_3^-(aq)$$

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tlvd-0636

EXPERIMENT 12.2 online only

Complex redox reactions

Aim

To react some common oxidants and reductants with one another and write their redox equations

12.2 Activities learnon

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12.2 Quick quiz on | **12.2 Exercise** | **12.2 Exam questions**

12.2 Exercise

1. State two situations where redox reactions might occur.
2. Identify each of the following half-equations as either oxidation or reduction.
   - **a.** $K(s) \rightarrow K^+(aq) + e^-$
   - **b.** $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
   - **c.** $2Br^-(aq) \rightarrow Br_2(l) + 2e^-$
   - **d.** $H_2(g) \rightarrow 2H^+(aq) + 2e^-$
   - **e.** $Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$
3. Complete the following for the redox reaction $Pb(NO_3)_2(aq) + Zn(s) \rightarrow Zn(NO_3)_2(aq) + Pb(s)$.
   - **a.** Write the ionic equation.
   - **b.** Write the two half-equations.
   - **c.** Name the substance that has been oxidised.
   - **d.** Name the substance that has been reduced.
   - **e.** State the oxidising agent.
   - **f.** State the reducing agent.

4. The equation for the reaction between magnesium and hydrochloric acid is as follows.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

This is actually a redox reaction and not an acid–base reaction. Explain why it does not fit the definition of acid–base and how it can be correctly described as a redox reaction. Identify the substances oxidised and reduced, and include the half-equations to support your explanation.

5. The balanced ionic equation for the displacement of silver from an aqueous silver nitrate solution by metallic lead is as follows.

$$2Ag^+(aq) + Pb(s) \rightarrow 2Ag(s) + Pb^{2+}(aq)$$

   a. Write balanced oxidation and reduction half-equations for the reaction.
   b. Which reactant gains electrons?
   c. Which reactant is oxidised? What is it oxidised to?
   d. Which reactant is the reducing agent?
   e. Which reactant loses electrons?
   f. State the conjugate pairs.

6. The reaction of aluminium with hydrogen ions in a dilute solution of hydrochloric acid can be represented by the following half-equations.

$$Al(s) \rightarrow Al^{3+}(aq) + 3e^-$$
$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

   a. Which reactant gains electrons?
   b. Which reactant is reduced? What is it reduced to?
   c. Which reactant is the reducing agent?
   d. Which reactant loses electrons?
   e. Which reactant is oxidised? What is it oxidised to?
   f. Which reactant is the oxidising agent?
   g. State the conjugate pairs.
   h. Write the balanced ionic equation.

7. Write balanced half-equations for the following reactions and state the conjugate pairs.
   a. $Ca(s) + Cl_2(g) \rightarrow CaCl_2(s)$
   b. $Mg(s) + S(s) \rightarrow MgS(s)$
   c. $Al(s) + 3Ag^+(aq) \rightarrow Al^{3+}(aq) + 3Ag(s)$

8. Balance the following more complex half-equations using the KOHES procedure from section 12.2.3 for each.
   a. $NO_3^-(aq) \rightarrow NO_2(g)$ b. $ClO^-(aq) \rightarrow Cl^-(aq)$ c. $BrO_3^-(aq) \rightarrow BrO_2(aq)$

9. For the reaction of dichromate ions, $Cr_2O_7^{2-}$, and nitrite ions, NO_2^-, to form chromium, Cr^{3+}, and nitrate ions, NO_3^-, explain how each of the half-equations were obtained.

10. Balance the following equations by first separating them into half-equations.
   a. $I_2(s) + H_2S(g) \rightarrow I^-(aq) + S(s)$
   b. $SO_3^{2-}(aq) + MnO_4^-(aq) \rightarrow SO_4^{2-}(aq) + Mn^{2+}(aq)$
   c. $Cu(s) + NO_3^-(aq) \rightarrow NO(g) + Cu^{2+}(aq)$

12.2 Exam questions

Question 1 (1 mark)

MC Which of the following statements is correct about the following reaction?

$$Mg(s) + Zn^{2+}(aq) \rightarrow Mg^{2+}(aq) + Zn(s)$$

A. Magnesium is the reducing agent and its conjugate pair is Mg^{2+}.
B. Zn^{2+} is the reducing agent and its conjugate pair is Zn.
C. Magnesium is the reducing agent and its conjugate pair is Zn.
D. Zn^{2+} is the reducing agent and its conjugate pair is Mg.

Question 2 (1 mark)

MC Which statement best describes oxidation–reduction reactions?

A. They involve a proton transfer from an oxidising agent to a reducing agent.
B. They involve an electron transfer from an oxidising agent to a reducing agent.
C. One reactant is oxidised as it loses electrons to the other reactant.
D. Only one electron may be either lost or gained during the process.

Question 3 (7 marks)

a. Complete the following half-equations by balancing as necessary.
  i. $Br_2(l) \rightarrow Br^-(aq)$ **(1 mark)**
  ii. $Co(s) \rightarrow Co^{2+}(aq)$ **(1 mark)**
b. Which of the preceding reactions is the reduction reaction? **(1 mark)**
c. Write the overall equation that would occur if the reactions in part **a** occurred together. **(1 mark)**
d. Which species is the reducing agent in your equation in part **c**? **(1 mark)**
e. Given that bromine is a dark orange liquid, bromide forms a colourless solution, cobalt(II) ions form a pink-red solution, and cobalt is a gray metal, describe the observations that could be seen as the reaction in part **c** progresses. **(2 marks)**

Question 4 (5 marks)

Chlorine is used as a disinfectant and in household bleach.

Chlorine is formed when $KMnO_4$ reacts with hydrochloric acid. The ionic equation for this reaction is as follows.

$$16H^+(aq) + 2MnO_4^-(aq) + 10Cl^-(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5Cl_2(g)$$

a. Write the half-equation for the formation of chlorine from chloride ions. **(1 mark)**
b. Write the equation for the formation of Mn^{2+} from MnO_4^- in an acidic solution. **(2 marks)**
c. Label the reactions as oxidation or reduction. **(2 marks)**

Question 5 (4 marks)

Balance the following equations by first separating them into half-equations.

a. $Br^-(aq) + SO_4^{2-}(aq) \rightarrow SO_2(g) + Br_2(l)$ **(2 marks)**
b. $I_2(s) + H_2S(g) \rightarrow I^-(aq) + S(s) + H^+(aq)$ **(2 marks)**

More exam questions are available in your learnON title.

12.3 EXTENSION: Oxidation numbers

KEY CONCEPT

- The use of oxidation numbers to identify if a redox reaction has occurred

Oxidation numbers are useful for the VCE Chemistry Units 3 & 4 course.

12.3.1 Using oxidation numbers

The concept of oxidation numbers is useful to help understand oxidation and reduction reactions. An **oxidation number**, sometimes called oxidation state, is a theoretical number that corresponds to the charge that an atom would have if the compound was composed of ions.

oxidation number a theoretical number that corresponds to the charge that atom would have if the compound was composed of ions

When determining oxidation numbers, they are written *above* the atom and are not to be confused with the charge on an ion. A set of rules can be used to determine the oxidation number of an atom or ion. These are summarised with examples in table 12.1.

TABLE 12.1 Oxidation number rules and examples

Oxidation number rule	Oxidation number example
1. The oxidation number of an element is 0.	$\overset{0}{Zn}$, $\overset{0}{S}$, $\overset{0}{O_2}$, $\overset{0}{N_2}$
2. In monatomic ions, the oxidation number is the same as the charge on the ion.	$\overset{+2}{Mg^{2+}}$, $\overset{-2}{O^{2-}}$, $\overset{+1}{Na^{+}}$
3. In all compounds of hydrogen, the oxidation number of hydrogen is +1. Except in metal hydrides, where the oxidation number of hydrogen is –1.	$\overset{+1}{H_2}\overset{-2}{O}$, $\overset{+1}{CH_4}$ $\overset{+1}{Na}\ \overset{-1}{H}$
4. The oxidation number of oxygen in a compound is usually –2. Except in: • peroxide compounds, where the oxidation number is –1 • compounds with oxygen bonded to fluorine, where the oxidation number is +2.	$\overset{+1}{H_2}\overset{-2}{O}$ Peroxide compounds: $\overset{+1}{H_2}\overset{-1}{O_2}$ Oxygen difluoride: $\overset{+2}{O}\ \overset{-1}{F_2}$
5. Fluorine always has oxidation number of –1 because it is the most electronegative element.	$\overset{-1}{F}$
6. In a neutral molecule, the sum of all the oxidation numbers must equal 0.	In H_2S, the oxidation number of hydrogen is +1. When adding oxidation numbers, you need to multiply the oxidation number by the number of atoms present, so the total is +2 for the hydrogen atoms. The overall total must add to zero, so the oxidation number of sulfur is –2.
7. In a polyatomic ion, the sum of the oxidation numbers must equal the charge on the ion.	The oxidation number of the chromate ion, $Cr_2O_7^{2-}$, is –2. If the oxidation number of oxygen is –2, the total is –14 so the oxidation numbers of the two chromium atoms must add up to 12 so that the overall total is –2. Each chromium atom must have an oxidation number of +6; that is, $\overset{+6}{Cr_2}\overset{-2}{O_7^{2-}}$.
8. In ionic compounds, it is often easier to separate the compound into its ions first and calculate the oxidation numbers of each ion separately.	$NaCl \rightarrow Na^+ + Cl^-$ $\overset{+1}{Na}$, $\overset{-1}{Cl}$
9. In covalent compounds that do not involve oxygen or hydrogen, the more electronegative element has the negative oxidation number. This is equal to the charge that it would have if it were a negative ion.	In ICl_3, the chlorine is the more electronegative atom. It is, therefore, assigned an oxidation number of –1, because this is the charge on a chloride ion. (*Note:* this is just the way the oxidation number is worked out. This molecule is a covalent, neutral molecule; it does not contain chloride ions.) Using rule 6, we can now calculate that the oxidation number of the iodine in ICl_3 is +3.

tlvd-3482

SAMPLE PROBLEM 4 Determining oxidation numbers

Determine the oxidation numbers of each element in the following examples.

a. Br_2 **b.** Cu^{2+} **c.** Na_2O **d.** HSO_4^-

THINK

For each of the examples, determine if the species is an element, monatomic ion, molecule or polyatomic ion.

a. Br_2 is an element, so its oxidation number is 0. (Rule 1)

b. Cu^{2+} is a monatomic ion, so its oxidation number will be the charge on the ion. (Rule 2)

c. Na_2O is a compound, so the oxidation numbers must add up to zero. (Rule 6)

d. HSO_4^- is a polyatomic ion, so the oxidation numbers must add up to the charge, which in this case is –1.
The oxidation numbers of hydrogen and oxygen are known, so it is necessary to find the oxidation number of sulfur, S. (Rule 7)
Then the oxidation numbers can be stated.

WRITE

$\overset{0}{Br_2}$

$\overset{+2}{Cu^{2+}}$

Na_2O

$\overset{+1\ \ -2}{Na_2O}$ because $(2\times(+1))+(-2)=0$

$\overset{+1\ \ -2}{HSO_4^-}$

Let oxidation number of S be x

$+1+x+(4\times-2)=-1$

$+1+x-8=-1$

$x-7=-1$

$x=+6$

The oxidation number of S is +6.

$\overset{+1+6\ -2}{H\ S\ O_4^-}$

PRACTICE PROBLEM 4

Find the oxidation numbers of the following elements in pink.

a. P_2O_5
b. NO_3^-
c. Cl^-
d. K

Using oxidation numbers to identify redox reactions

Metallic ions always have positive oxidation numbers. Non-metallic ions can have positive or negative oxidation numbers, and the highest oxidation an element can have relates to its group in the periodic table.

The use of oxidation numbers greatly simplifies the identification of redox reactions and the subsequent identification of oxidising agents and reducing agents. If oxidation numbers change during a reaction, it indicates that the reaction is a redox reaction.

Identifying redox reactions

If an oxidation number increases, oxidation has occurred.

If an oxidation number decreases, reduction has occurred.

Consider the reaction $Ca(s) + H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g)$.

We can assign oxidation numbers to all of the atoms present.

$$\overset{0}{Ca}(s) + \overset{+1\ -2}{H_2O}(l) \rightarrow \overset{+2\ -2+1}{Ca(OH)_2}(s) + \overset{0}{H_2}(g)$$

Firstly, the oxidation numbers have changed, so this is a redox reaction.

- The calcium atom has *increased* its oxidation number from 0 to +2; therefore, it has been oxidised and it is the reducing agent.
- The hydrogen atoms have *decreased* their oxidation numbers from +1 to 0; therefore, the water has been reduced and it is the oxidising agent. Note that because the hydrogen is covalently bonded as part of the water molecule, we say that the water is reduced. (For ionic compounds, the ions are separated in solution so only the involved ion is the oxidant or reductant. For example, with NaBr in sample problem 2, Br^- is the reductant.)

tlvd-0580

SAMPLE PROBLEM 5 Using oxidation numbers to explain redox reactions

Use oxidation numbers to explain why the following reaction is a redox reaction and state the reducing agent and the oxidising agent.

$$MnO_2(s) + 2H^+(aq) + NO_2^-(aq) \rightarrow NO_3^-(aq) + Mn^{2+}(aq) + H_2O(l)$$

THINK	WRITE
1. Assign oxidation numbers to all atoms. This is a redox reaction because the oxidation numbers of reactants have changed.	$\overset{+4\ -2}{MnO_2}(s) + 2\overset{+1}{H^{+1}}(aq) + \overset{+3\ -2}{NO_2^-}(aq) \rightarrow \overset{+5\ -2}{NO_3^-}(aq) + \overset{+2}{Mn^{2+}}(aq) + \overset{+1\ -2}{H_2O}(l)$
2. Identify which atom has increased its oxidation number; this belongs to the species being oxidised.	$\overset{+3\ -2}{NO_2^-}$ has been oxidised to $\overset{+5\ -2}{NO_3^-}$.
3. Identify which atom has decreased its oxidation number; this belongs to the species being reduced.	$\overset{+4\ -2}{MnO_2}$ has been reduced to $\overset{+2}{Mn^{2+}}$.
4. The species being oxidised is the reducing agent.	NO_2^- is the reducing agent.
The species being reduced is the oxidising agent.	MnO_2 is the oxidising agent.

PRACTICE PROBLEM 5

Use oxidation numbers to explain why the following reaction is a redox reaction and state the reducing agent and the oxidising agent.

$$Pb(s) + Br_2(l) \rightarrow Pb^{2+}(aq) + 2Br^-(aq).$$

CASE STUDY: Redox reactions in fireworks

Fireworks are spectacular examples of exothermic redox reactions.

To create the redox reactions that occur in fireworks, clearly an oxidising agent and a reducing agent are required. Traditionally, the combination used was gunpowder, which is composed of an oxidising agent (potassium nitrate, also known as saltpetre), sulfur and carbon (in the form of charcoal). The mass ratio of these three substances in gunpowder, which has remained unchanged for over 500 years, is 75 : 15 : 10.

Today the oxidising agents used include nitrates, NO_3^-, chlorates, ClO_3^-, and perchlorates, ClO_4^-. The chlorine in the chlorates help strengthen some colours and, when burnt, potassium perchlorate, $KClO_4$, or ammonium perchlorate, NH_4ClO_4, can reach temperatures of up to 2000°C. Nitrate oxidising agents include potassium nitrate, KNO_3 (saltpetre); strontium nitrate, $Sr(NO_3)_2$, which produces a red colour; and barium nitrate $Ba(NO_3)_2$, which produces a green colour. You may remember seeing these colours when you performed flame tests on metal compounds.

FIGURE 12.7 Different colours in fireworks are due to different redox reactions.

The reducing agents are metals (which have a high heat output when oxidised), carbon (in the form of charcoal or organic compounds) and/or sulfur. The high temperatures of reaction of the oxidising agents allow a simultaneous release of a large amount of oxygen that is rapidly accepted by the reducing agents, oxidising (burning) the fuel, which releases more heat and gas (SO_2 and CO_2) and further propagating the reaction to produce explosive energy and gas — what we see as spectacular explosions of colour.

While elements in the oxidising and reducing agents contribute to the colours seen (for example, burning aluminium creates a white light), the variety and intensity of colours in fireworks is also due to the addition of colour ingredients and intensifiers. Red is created by strontium compounds, orange by calcium, yellow by sodium, green by barium, blue by copper, purple by copper-strontium and silver by magnesium-aluminium compounds.

12.3 Activities

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12.3 Quick quiz on | **12.3 Exercise** | **12.3 Exam questions**

12.3 Exercise

1. Determine the oxidation numbers for each atom in the following examples.
   - **a.** P
   - **b.** F_2
   - **c.** Zn^{2+}
   - **d.** HCO_3^-
   - **e.** $NaNO_3$
   - **f.** BeO
   - **g.** Ag_2S
2. **a.** Determine the oxidation number of nitrogen in the following compounds.
   - **i.** N_2
   - **ii.** HNO_3
   - **iii.** $NH4^+$
   - **iv.** N_2O_4

   b. Which substance listed in **a** has nitrogen in its lowest oxidation state?
3. Arrange these substances in order of the increasing oxidation numbers of nitrogen: NO_3^-, N_2, NH_4^+, NO_2.
4. Sulfur is in group 16. What is the maximum oxidation number that sulfur can have?

5. Use oxidation numbers to identify which of the following equations are redox reactions.
   a. $Br_2(l) + SO_2(g) \rightarrow 2Br^-(aq) + SO_4^{2-}(aq)$
   b. $KOH(aq) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(aq)$
   c. $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$
   d. $5Zn(s) + 2NO_3^-(aq) + 12H^+(aq) \rightarrow 5Zn^{2+}(aq) + N_2(g) + 6H_2O(l)$

12.3 Exam questions

Question 1 (1 mark)

MC Which of the following substances contains vanadium in its highest oxidation state (i.e. highest oxidation number)?

A. V^{3+} **B.** VO_2^+ **C.** V **D.** VO^{2+}

Question 2 (1 mark)

MC Which of the following reactions represents a redox reaction?

A. $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
B. $CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$
C. $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$
D. $CuSO_4(aq) + BaCl_2(aq) \rightarrow CuCl_2(aq) + BaSO_4(s)$

Question 3 (24 marks)

In the unbalanced chemical reactions given, complete the following.

i. Determine the oxidation number of each element.
ii. State which species is oxidised and which is reduced.
iii. State the oxidising agent and the reducing agent.

a. $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$ **(6 marks)**
b. $PbO_2(s) + I^-(aq) \rightarrow Pb^{2+}(aq) + I_2(g)$ **(6 marks)**
c. $I^-(aq) + NO_2^-(aq) \rightarrow I_2(s) + NO(g)$ **(6 marks)**
d. $H_2O_2(aq) + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + O_2(g)$ **(6 marks)**

Question 4 (5 marks)

For the reaction $H_2O_2(l) + Sn^{2+}(aq) + 2H^+ \rightarrow 2H_2O(l) + Sn^{4+}(aq)$, use oxidation numbers to answer the following questions.

a. What is the oxidation number of oxygen in H_2O_2? **(1 mark)**
b. What substance is oxidised? **(1 mark)**
c. What substance is reduced? **(1 mark)**
d. Name the oxidising agent. **(1 mark)**
e. Name the reducing agent. **(1 mark)**

Question 5 (6 marks)

Determine the oxidation numbers of the elements in the following changes. Place a tick in the relevant box to indicate if the changes are oxidation, reduction or neither.

Change	Oxidation	Reduction	Neither
$O_2 \rightarrow O_3$			
$MnO_2 \rightarrow MnO_4^-$			
$P_2O_5 \rightarrow P_4H_{10}$			
$NH_3 \rightarrow NO_2$			
$Fe^{2+} \rightarrow Fe^{3+}$			
$S_2O_3^{2-} \rightarrow S_2O_4^{2-}$			

More exam questions are available in your learnON title.

12.4 Reactivity series of metals

KEY KNOWLEDGE

- The reactivity series of metals and metal displacement reactions, including balanced redox equations with states indicated

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12.4.1 Reactivity of metals

You may have noticed that many of the redox reaction examples involved a metal reacting with a different metal salt solution. Metals have different levels of reactivity, not only with oxygen, water or acids as discussed in topic 3, but also with metal solutions. These reactions of metals in metal salt solutions confirm the order of reactivity discussed in topic 3.

When a copper strip is placed in silver nitrate solution, a reaction occurs, causing silver metal to form on the copper (figure 12.8). This is a displacement reaction: copper removes silver from the solution and, as a result, a deep blue colour of the solution forms as the copper metal forms copper ions. If you place a silver strip, however, in a solution of copper nitrate, no reaction occurs. Therefore, copper is more reactive than silver.

FIGURE 12.8 The flask on the right shows the initial set up. The flask on the left shows the silver forming on the copper strip. The solution has become blue due to the copper ions forming.

Different metals have varying abilities to react with other substances. Potassium is so reactive that it is only found naturally as a compound rather than a pure metal. When extracted, it must be stored in oil to prevent exposure to moisture and oxygen. It reacts very vigorously with water to produce hydrogen gas and potassium hydroxide. The heat given off in this exothermic reaction is sufficient to ignite the hydrogen gas and allows it to burn in air with a bright lilac flame (see figure 12.9).

$$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$$

Gold, on the other hand, is very unreactive. It occurs in nature as a pure element and can be found in its pure form both by mining or panning in rivers. Ornamental gold is valuable because it remains untarnished for centuries. Potassium and gold are at opposite ends of the **reactivity series of metals**, shown in figure 12.10. This series, sometimes called the activity series, lists the metals in order of how readily they react with oxygen, water, steam, dilute acids and salt solutions, and is useful for predicting reactions.

Each of these reactions is a redox reaction, in which the metal is always the reducing agent. A more reactive metal will displace a less reactive metal from solution. In other words, a more reactive metal gives up its electrons more readily and is a stronger reducing agent. Reactivity of metals tends to increase down a group of the periodic table.

reactivity series of metals
an ordered list of how readily the metals react with oxygen, water, steam, dilute acids and salt solutions

A mnemonic using the first letters of the names of the metals can be used to remember the most reactive metals.

Please	Potassium
Stop	Sodium
Calling	Calcium
My	Magnesium
Aunt's	Aluminium
Zebra	Zinc
In	Iron
This	Tin
Lesson	Lead

FIGURE 12.9 Potassium reacting with water

TABLE 12.2 Reactivity of metals

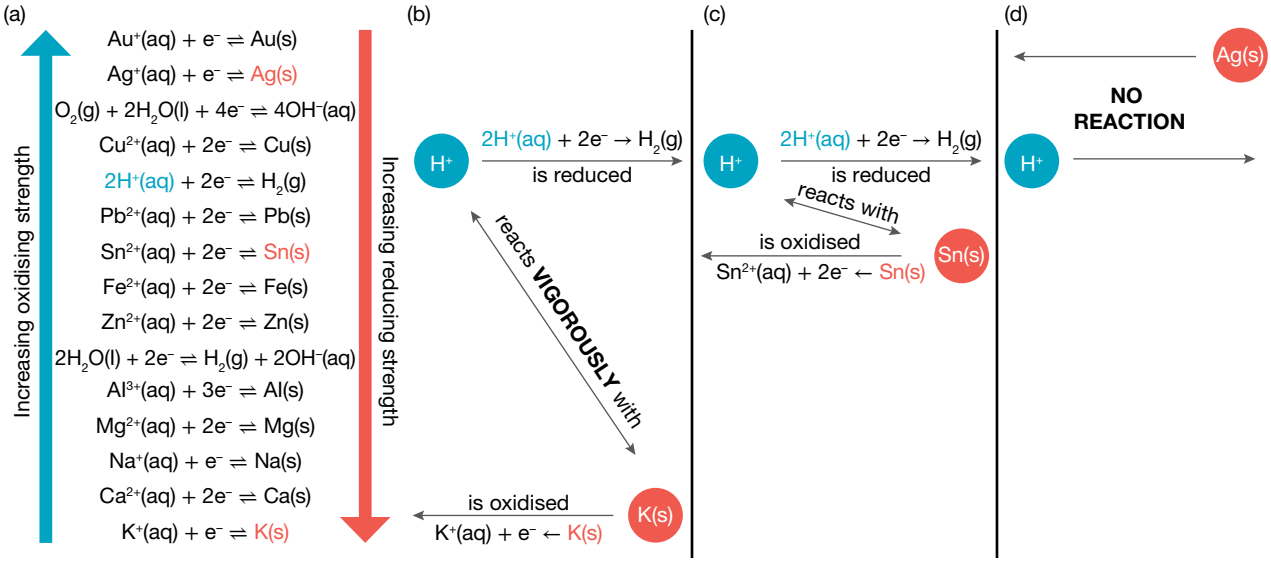
Element	Appearance of metal	Reaction with oxygen	Reaction with water or steam	Displacement of metals	Reaction with dilute acids
Au	Generally shiny	No reaction	No reaction	Any metal in this series displaces any other that occurs lower in the series (i.e. changes the metal ion back to the metal). The further apart the metals are, the easier the displacement.	No reaction
Ag					
Hg					
Cu		Oxidises when heated in air oroxygen to give oxides			
Pb	Generally dull		No reaction at Bunsen burner temperatures		No visible reaction
Sn					Hydrogen formed with cold acid
Fe		Oxidises in air at room temperature to give oxides	Hydrogen formed with steam		
Zn					
Al					
Mg					
Ca			Hydrogen formed from cold water		
Na	Dull; stored under oil				Hydrogen formed violently
K					

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

The reactivity with different species can be predicted from the electrochemical series, found in Table 2 of the VCE Chemistry Data book.

Figure 12.10 shows the reactivity of potassium, tin and silver with an acid. A spontaneous reaction occurs when the reductant (in this case, the metal) is below the oxidant (the acid). The further apart the species are, the more vigorous the reaction. If the reductant is above the oxidant, no reaction occurs.

FIGURE 12.10 Using the electrochemical series (a) to predict the reactivity of metals, such as potassium (b), tin (c) and silver (d), with an acid



Due to its strong tendency to donate electrons, a reactive metal is generally:

- a stronger reducing agent
- more easily oxidised or corroded
- able to donate electrons readily
- able to form a positive ion easily
- more likely to be found naturally as a compound rather than a pure metal
- more difficult to extract from its ore and so requires more energy.

12.4.2 Metal displacement reactions

The reactivity series can be established experimentally, as was shown in topic 3 when the reactions of metals with oxygen, water and acids were compared.

To understand how displacement reactions may be used, consider the example earlier in this subtopic, where copper displaced silver ions from solution to form silver metal and copper ions according to the following equation.

$$Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$$

Silver metal does not displace Cu ions from solution. This demonstrated difference in reactivity forms the basis of the displacement method of determining relative reactivity. Samples of each metal are immersed in separate solutions of metal cations. These cations accept electrons from any other more reactive solid metal atoms that are introduced, forming solid metal atoms themselves. The more reactive metal, therefore, loses electrons and forms ions in solution.

FIGURE 12.11 Reactivity of metals with hydrochloric acid. From left to right: calcium, magnesium, zinc, copper

on Resources

Weblinks Reaction of metals with oxygen and water
Displacement reactions

tlvd-0581

SAMPLE PROBLEM 6 Using the reactivity series to predict metal displacement reactions

Use the reactivity series to consider what happens when a piece of lead is dipped into a solution of silver nitrate.

a. Write an ionic equation.
b. Identify two half-equations for the reaction.
c. Determine the oxidising agent and reducing agent.

THINK	WRITE
a. Recognise that lead is above silver in the activity series. Because lead is more reactive than silver, a displacement reaction will occur between lead metal and silver ions. Lead displaces silver from solution, producing silver metal. The ionic equation omits the NO_3^- ions.	$Pb(s) + 2Ag^+(aq) \rightarrow Pb^{2+}(aq) + 2Ag(s)$
b. The half-equation for oxidation has lead and its conjugate. The half-equation for reduction has silver ion and its conjugate.	$Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$ $Ag^+(aq) + e^- \rightarrow Ag(s)$
c. The substance that is oxidised, Pb, is the reducing agent. The substance that is reduced, Ag+, is the oxidising agent. **TIP** Remember to call this the silver ion and not just silver.	Pb is oxidised to Pb^{2+} so it is the reducing agent. Ag^+ is reduced to Ag so it is the oxidising agent.

PRACTICE PROBLEM 6

Use the reactivity series to consider what happens when a piece of zinc is dipped into a solution of lead nitrate.

a. Write an ionic equation.
b. Identify two half-equations for the reaction.
c. Identify the oxidising agent and reducing agent.

12.4 Activities

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12.4 Quick quiz on | 12.4 Exercise | 12.4 Exam questions

12.4 Exercise

1. **a.** In which part of the periodic table would you find the most reactive metals? Suggest a reason this is the case.
   b. Explain the trend in terms of reactivity down a group of the periodic table.

2. Refer to the activity series of metals in table 12.2 and name the metals that best meet the following.
   a. Do not react at all with oxygen
   b. Will displace tin ions but will be displaced by zinc
   c. React with cold water to form hydrogen gas
   d. Are generally dull and oxidise when heated with oxygen to give oxides
   e. Displace hydrogen from cold water but not reactive enough to be stored under oil
   f. Oxidise when heated in oxygen but are otherwise generally shiny
3. a. State three observations that you would expect to see if iron was placed in a copper sulfate solution and left overnight.
   b. State the products.
4. Use the reactivity series to determine if a reaction occurs when the following are mixed. Write an equation for any reaction that occurs and identify the oxidising agent and the reducing agent.
   a. $Mg(s) + Cu^{2+}(aq)$
   b. $Al(s) + Pb^{2+}(aq)$
   c. $Cu(s) + Fe^{2+}(aq)$
   d. $Ag(s) + Zn^{2+}(aq)$
5. Use the reactivity series to predict what reaction, if any, occurs when the following reagents are mixed. Write half-equations and overall equations for any reaction that you predict.
   a. $Pb^{2+}(aq)$ and Fe(s)
   b. $Sn^{2+}(aq)$ and Zn(s)
   c. $Al^{3+}(aq)$ and Fe(s)
   d. $Zn^{2+}(aq)$ and Al(s)
6. Sharni used the reactivity series to predict that a strip of aluminium metal would react with copper(II) sulfate. When she tried the experiment, she observed no reaction.
   a. Explain her observations.
   b. What steps would she need to take to prove that her prediction was correct?
7. Write a full equation showing the reaction of sodium with water. State the oxidising agent and the reducing agent.
8. Name the product of the reaction of aluminium with oxygen. Write a balanced equation showing this reaction.
9. Identify the products and write the equations for the following reactions.
   a. Magnesium and hydrochloric acid
   b. Magnesium and steam
   c. Magnesium burned in air
10. Matthew recommended that the next set of coins for Australia be made from gold, while Oliver recommended the use of magnesium and Chris suggested that copper would be the best choice. Do you agree with any of these recommendations or do you have a completely different proposal? Include a discussion of the factors that need to be considered when producing coins.

12.4 Exam questions

Question 1 (1 mark)

MC A metal, *M*, undergoes the following reactions.
- It is dissolved when placed in a solution of P^{2+} ions.
- It does not react when placed in a solution of Q^{2+} ions.

In order of increasing reactivity, what is the correct order of the metals?

A. $Q < M < P$ **B.** $P < Q < M$ **C.** $P < M < Q$ **D.** $M < P < Q$

Question 2 (1 mark)

MC A piece of iron metal is placed in a deep green solution of nickel chloride. Which of the following observations would be seen?

A. Iron metal would dissolve and the solution would become a deeper green.
B. Nickel metal would deposit and the solution would become a paler green.
C. Iron metal would dissolve and the solution would remain the same intensity of green.
D. No reaction would be observed.

Question 3 (1 mark)

MC A metal, *M*, is oxidised when placed in a solution containing Ni^{2+}, but no reaction occurs when it is placed in a solution containing Zn^{2+} ions. What does this mean?

A. *M* must be a stronger reducing agent than Zn.
B. *M* must be a stronger oxidising agent than Ni^{2+}.
C. M^{2+} is a stronger oxidising agent than Ni^{2+} but weaker than Zn^{2+}.
D. *M* is a stronger reducing agent than Ni but weaker than Zn.

Question 4 (6 marks)

Predict whether a reaction would occur for the following. If a reaction does occur, write the two half-equations and the overall reaction. If there is no reaction, write 'no reaction'.

a. Aluminium foil is placed in a solution containing tin(II) ions. **(3 marks)**
b. Lead metal is placed into a silver nitrate solution. **(3 marks)**

Question 5 (2 marks)

A metal displacement experiment was carried out where four different metals (*P*, *Q*, *R* and *S*) were tested to see if they would react with solutions of their metal cations (i.e. solutions of P^{2+}, Q^{2+}, R^{2+} and S^{2+}). Metals were not tested in with their own metal cation.

The following results were obtained. (✓ indicates reaction observed, ✘ indicates no reaction, – indicates not tested)

	P^{2+}	Q^{2+}	R^{2+}	S^{2+}
P	–	✓	✓	✓
Q	✘	–	✘	✘
R	✘	✓	–	✓
S	✘	✓	✘	–

Use the table of results to determine the following.

a. The strongest reducing agent present **(1 mark)**
b. The order of oxidising agent from strongest to weakest **(1 mark)**

More exam questions are available in your learnON title.

12.5 Applications of redox reactions

KEY KNOWLEDGE

- Applications of redox reactions in society; for example, corrosion or the use of simple primary cells in the production of electrical energy from chemical energy

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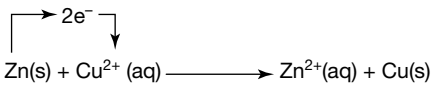
12.5.1 Using redox reactions — primary cells

The displacement reaction between zinc and copper ions, already discussed in this topic, demonstrates that zinc is a better reducing agent than copper and would, therefore, be oxidised more readily than copper. As zinc is oxidised, electrons flow from the zinc metal to the copper ions.



This is a spontaneous reaction and requires no energy; in fact, it releases energy. Energy may be released in several forms, such as heat, light, electricity or sound. The set-up of the displacement reaction determines the form of energy that is released. For example, when zinc metal is placed in direct contact with copper ions, all the chemical energy of the reaction is released as thermal energy (heat) and the transfer of electrons from zinc to copper ions occurs on the surface of the zinc metal. The reaction occurs in the one beaker.

Energy and displacement reactions

Energy is released in displacement reactions.

Primary cells

The energy generated by such reactions can be used, and this is the basis for batteries. Single-use batteries are described as **primary cells** and rechargeable batteries are described as **secondary cells**. To make use of the moving electrons in redox reactions, we need to adjust the set-up of the reaction. A strip of zinc metal is placed in a beaker containing zinc sulfate solution. The strip of metal is called an **electrode** and connects the solution in the beaker to the wire. This electrode is connected by a wire to a strip of copper placed in a beaker containing a copper sulfate solution. The wire provides a pathway for the electrons to pass from the zinc atoms to the copper cations. No reaction is observed, however, because the circuit is not complete.

Converting chemical energy into electrical energy

Chemical energy can be converted into electrical energy in a primary cell.

As the zinc atoms donate electrons and become zinc cations in the first beaker, the electrical neutrality must be maintained, so anions are required. These anions are supplied by the **salt bridge**, which can be a simple filter paper or a U-tube with cotton wool in it. It is soaked in a salt solution, such as potassium nitrate solution, $KNO_3(aq)$, and used to connect the two beakers. Potassium nitrate solution will provide NO_3^- anions for the first beaker. In the second beaker, copper cations are accepting electrons and becoming copper atoms, leaving behind negative sulfate ions that must be balanced. So, the salt bridge supplies cations (in this case, K^+). Electrons carry the current in the wire and ions carry the current in solution. The flow of ions completes the circuit. It is important that the ions in the salt bridge do not react with chemicals in the beakers. The movement of ions in the solutions is called the **internal circuit**.

The salt bridge

The salt bridge provides ions to balance ions consumed or produced in each half-cell.

Each beaker is called a **half-cell**. The metal conducting strips are called electrodes and combined with the wire they are referred to as the **external circuit**. Electrons in the external circuit can be made to do useful work such as lighting a light bulb. Chemical energy has been converted into electrical energy. Solutions that can conduct a current are known as **electrolytes**. The electrode at which oxidation occurs is called the **anode**, and it has a negative charge. The electrode at which reduction occurs is called the **cathode**, and it has a positive charge. All these components together are known as a **galvanic cell**.

primary cell a cell that cannot easily be recharged after use

secondary cell a cell that can be recharged

electrode a solid conductor of electricity in an electrochemical cell

salt bridge provides ions to balance ions consumed or produced in each half-cell

internal circuit the movement of ions in the electrolyte solutions

half-cell one of two containers containing the electrode and electrolyte in an galvanic cell

external circuit the wire joining the two half-cells

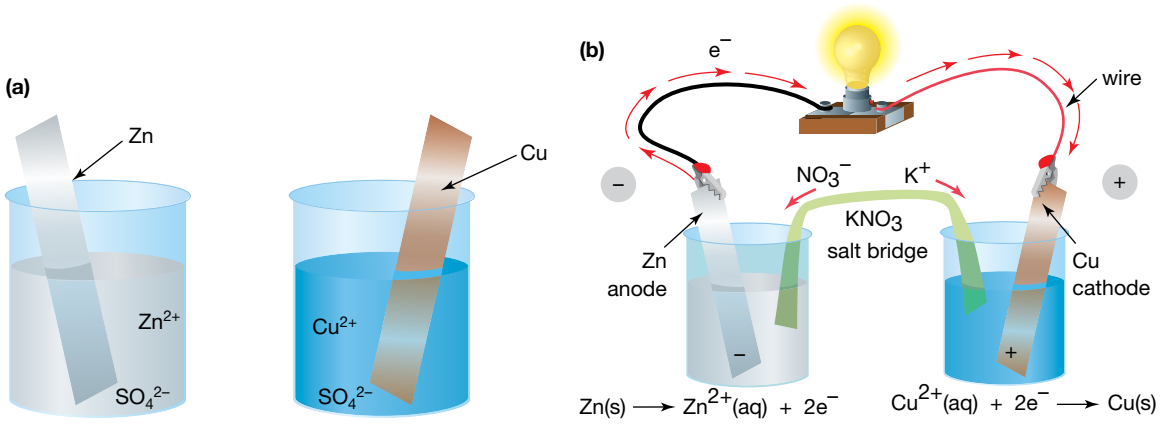
electrolyte solution or liquid that can conduct electricity

anode site of oxidation in an electrochemical cell

cathode site of reduction in an electrochemical cell

galvanic cell electrochemical cell that converts chemical energy into electrical energy

FIGURE 12.12 (a) Two strips of different metals and solutions of each of their ions. (b) With the addition of a wire and a salt bridge, a simple galvanic cell — a device that converts chemical energy into electrical energy — is constructed.



Galvanic cells

A simple galvanic cell consists of:

- two half-cells, containing two electrodes (anode and cathode) and two electrolytes
- a conducting wire (external circuit)
- a salt bridge, containing another electrolyte that will not react with the electrolytes in the half-cells (internal circuit).

A galvanic cell is a type of electrochemical cell, and an electrochemical cell is a cell that can convert chemical energy to electrical energy or use electrical energy to cause a chemical reaction. The most primitive electrochemical cell is the voltaic cell, discovered by Alessandro Volta (1745–1827). The term galvanic cell is named after Luigi Galvani (1737–1798), who is generally believed to have discovered electricity. The first crude version of the zinc–copper cell was constructed by John Frederick Daniell, and for this reason we still call this type of cell the Daniell cell.

Redox reactions within the half-cells

In half-cells, the more reactive metal will always give its electrons to the ions of the less reactive metal. So, the reaction taking place in the zinc half-cell must be oxidation.

$$Zn\,(s) \rightarrow Zn^{2+}\,(aq) + 2e^-$$

The ions of the less reactive metal will gain electrons from the other so the reaction in the copper half-cell must be reduction.

$$Cu^{2+}\,(aq) + 2e^- \rightarrow Cu\,(s)$$

Electrons move through the external circuit from the more reactive electrode to the less reactive electrode.

Features of primary cells

- The more reactive metal gives electrons to the ions of the less reactive metal.
- Electrons always flow to the cathode.
- The **electrode polarity** refers to the sign of the electrodes. In a galvanic cell, the anode is negative, and the cathode is positive.
- Electrodes can be inert like graphite (carbon) or platinum.

electrode polarity the positive or negative sign on the electrodes in a galvanic cell

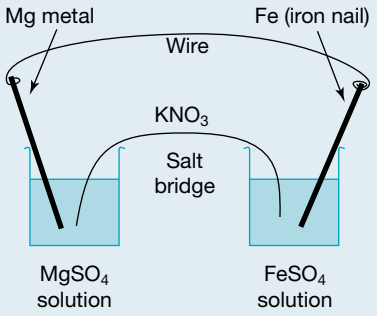
TIP: For all cells

- RedCat: Reduction always occurs at the Cathode
- AnOx: Oxidation always occurs at the Anode

tlvd-3483

SAMPLE PROBLEM 7 Using the reactivity series with galvanic cells

A galvanic cell was set up in the following way. A strip of clean magnesium was placed into a beaker containing a solution of $MgSO_4$ and, in a separate beaker, an iron nail was placed into a solution of $FeSO_4$. The iron nail and magnesium strip were connected with a wire and the circuit was completed with a salt bridge consisting of filter paper saturated with a solution of KNO_3, as shown in the diagram. Use the reactivity series to predict the following.



a. The substance that is oxidised and the one that is reduced
b. The anode and cathode
c. The direction of electron flow
d. The half-cell reactions
e. The overall redox reaction
f. The charge on the electrodes

THINK	WRITE
a. Since magnesium is above iron in the reactivity series, magnesium metal is more reactive. It will donate electrons to the iron ions.	Magnesium metal, Mg, will be the reducing agent and will, therefore, be oxidised. Iron ions, Fe^{2+}, will be reduced.
b. Oxidation occurs at the anode. Reduction occurs at the cathode.	Magnesium will be the anode and iron will be the cathode.
c. Electrons always go to the cathode.	Electrons will flow from the magnesium to the iron electrode.
d. Magnesium is oxidised and iron ions are reduced.	$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$ (oxidation) $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ (reduction)
e. Check that the electrons lost equal the electrons gained and add the two half-equations from step d.	$Mg(s) + Fe^{2+}(aq) \rightarrow Mg^{2+}(aq) + Fe(s)$
f. Oxidation always occurs at the anode, which is negative in galvanic cells.	The magnesium anode is negative, and the iron electrode is positive.

PRACTICE PROBLEM 7

a. Draw and label a galvanic cell for the conjugate redox pairs Cu^{2+}/Cu and Ag^+/Ag. Name and label the following.
  i. Electrolytes
  ii. Electrodes and their polarity
  iii. Salt bridge and direction of ion movement
  iv. Direction of electron flow
b. Write the half-equations for the reaction occurring in each half-cell.
c. Write the overall ionic equation.

Galvanic cells such as these, but modified to be self-contained and compact, are used in calculators, watches, hearing aids, cardiac pacemakers, lawnmowers and power tools. They are convenient, portable sources of power where, using redox reactions, chemical energy is converted into electrical energy. Galvanic cells, including secondary cells, are studied in more depth in Unit 3 Chemistry.

12.5.2 Types of corrosion

Although the ability of metals to act as reducing agent and be oxidised can be useful in cases like batteries, in other situations redox reactions can be quite a problem. **Corrosion** is the oxidation of metals by materials in their environment. When a metal corrodes, it suffers damage due to chemical change. Two types of corrosion are possible:

FIGURE 12.13 Car with corrosion above the wheel

1. **Dry corrosion** occurs when a metal reacts directly with oxygen. Some metals are very reactive and must be stored under oil. Potassium and sodium combine readily with oxygen, so they undergo dry corrosion very quickly. Iron undergoes dry corrosion, but at ambient temperatures its rate of dry corrosion is slow.
2. **Wet corrosion** occurs when a metal reacts directly with oxygen in the presence of water. Water can speed up some corrosion reactions, such as that of iron, which corrodes much more quickly in a moist environment. Corrosion returns metals to their natural states — the original ores. Except for gold and platinum, which are virtually unreactive, most metals are found in the ground as oxidised metals or ores. Since corroded metal often loses its structural purpose and attractiveness, this oxidation reaction has a great economic impact. In Australia, billions of dollars are spent each year preventing and treating corrosion.

Corrosion

Corrosion is the breakdown of a substance, often a metal, due to its reaction with chemicals in its environment.

Wet corrosion process

The wet corrosion of iron is commonly known as **rusting**. Iron rusts when a water drop containing dissolved impurities lands on it (see figure 12.14). Oxidation usually takes place at a 'stress site' on the iron, such as a dent in a car or a scratch on the surface. Iron is oxidised, releasing electrons.

$$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$$

These electrons travel through the metal and are accepted by oxygen, usually at the point where the edge of a water drop is in contact with the air. Oxygen can also come from dissolved oxygen in the water drop. Oxygen accepts electrons and is reduced.

$$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$$

Fe^{2+} and OH^- ions produced migrate toward each other and react to produce an insoluble green iron(II) hydroxide precipitate. The overall equation is as follows.

$$2Fe(s) + O_2(g) + 2H_2O(l) \rightarrow 2Fe(OH)_2(s)$$

corrosion oxidation, generally of metals, by materials in their environment

dry corrosion corrosion of a metal in the air when water is not a reactant

wet corrosion corrosion of a metal by reaction with oxygen in the presence of water

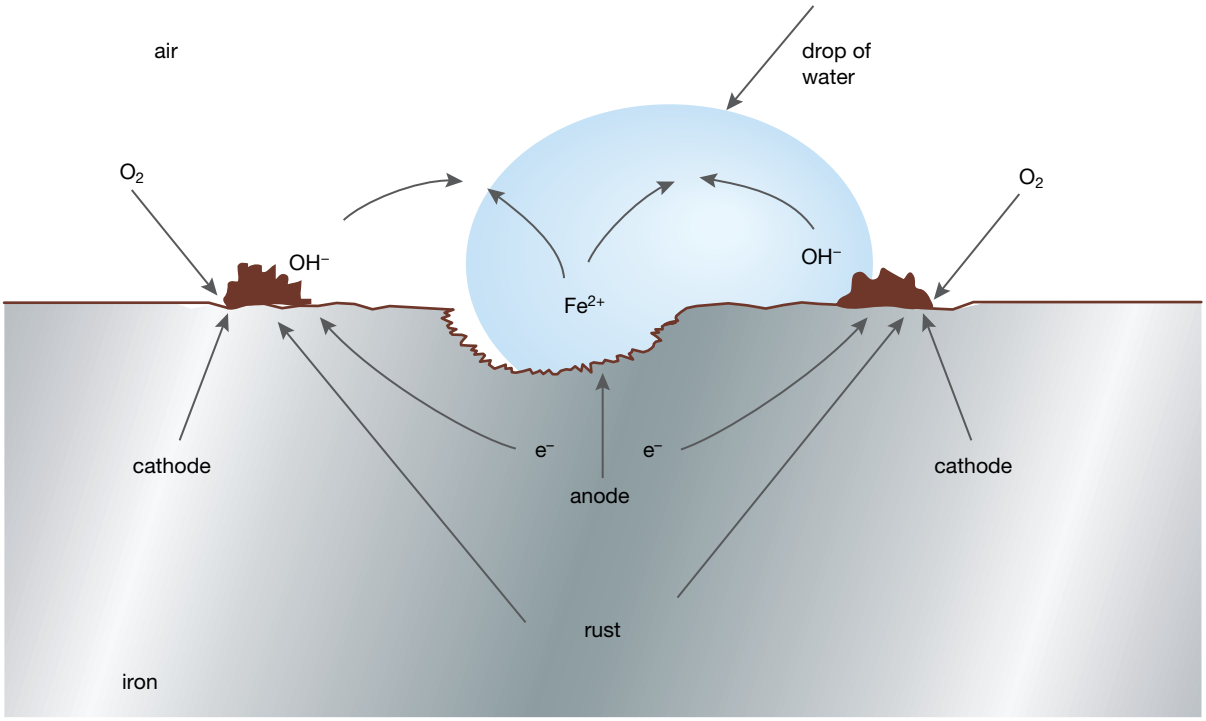
rusting wet corrosion of iron

Iron(II) hydroxide is very unstable in air and quickly reacts with oxygen to produce brown iron(III) hydroxide.

$$4Fe(OH)_2(s) + O_2(g) + 2H_2O(l) \rightarrow 4Fe(OH)_3(s)$$

This brown precipitate then partially dehydrates (loses water) to produce iron(III) oxide, $Fe_2O_3 \,.\, xH_2O$, where x is less than 3. The degree of hydration (the value of x) affects the colour of the rust.

FIGURE 12.14 The corrosion of iron

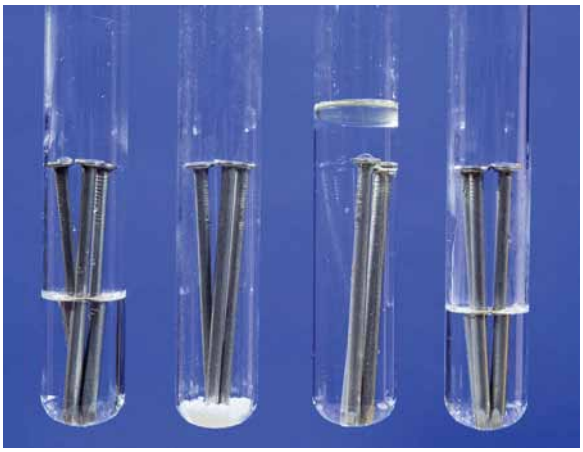


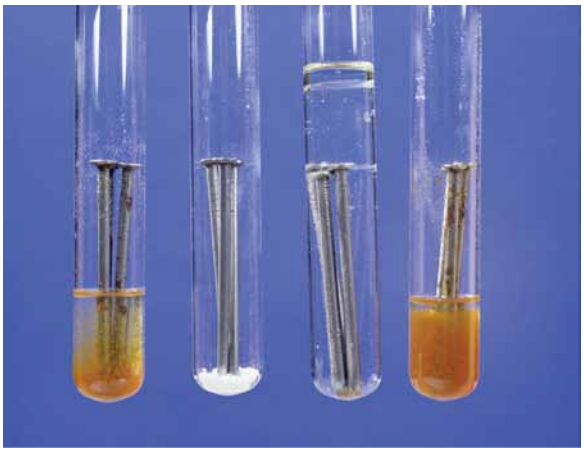
This process is responsible for the corrosion of cars, tools, bridges, buildings, machinery and virtually anything made of iron and steel that is not protected in some way. This is significant because iron is our most commonly used metal. Because it is soft and flaky, rust is easily dislodged, which accelerates further rusting. Water is vital in the process; without a 'salt bridge', iron does not rust. Cars last much longer in dry areas; however, the dissolved salts in the moist air in seaside areas accelerate rusting. The dissolved ions increase the conductivity of the moisture and, therefore, encourage corrosion. Other factors increase corrosion, including the presence of acids and pollutants such as sulfur dioxide and nitrogen dioxide, and imperfections in the metal surface.

Testing the conditions for rusting

The conditions for rusting can be tested experimentally, as demonstrated in figure 12.15. In the first image (a), from left to right, iron nails are placed in water and air, a drying agent, distilled water topped with oil, and salty water. The results after several weeks are shown in the second image (b). In the test tube with water and air (far left), rust (hydrated iron oxide) is seen; this process has been accelerated in the test tube with salt water and air (far right). The nails exposed to air but no water and exposed to water but no air (oxygen) have not rusted (middle test tubes).

FIGURE 12.15 Conditions for rusting; left to right: water and air, drying agent, distilled water and oil, and salty water.

(a) 

(b) 

Adverse effects of corrosion

Corrosion has many adverse effects, including the following.

- Metal oxides, or rust, have less tensile strength and less elasticity than uncorroded metal. Corroded buildings, bridges and machinery may develop weaknesses that result in malfunction or breakage.
- A severely corroded metal cannot conduct electricity. Products of corrosion formed in a car battery, between the terminals and the leads, may cause the electrical system to malfunction.
- Corroded copper pipes and hot water tanks leak as the copper compounds dissolve in the water flowing through them.
- Products of corrosion are brittle and flake off, resulting in holes in car mufflers or iron roofing. In car radiators, iron oxide flakes may cause blockages.
- A corroded metal occupies a larger volume than the original metal. Corroded nuts and bolts may jam machinery.

12.5.3 Corrosion protection

It is important to protect iron and steel from corrosion. Three main methods are used: surface protection, **alloying** and **electrochemical protection**.

Surface protection

If a metal surface can be prevented from coming into contact with oxygen or water, it will not corrode. An obvious way of achieving this is to coat the surface in some way. Commonly used surface coatings are plastic, paint, grease or oil, noble metals and sacrificial metals.

A metal that produces its own surface protection is aluminium and it does this by corroding. It is an unusual metal because even though it is a reactive metal and reacts readily with oxygen in the atmosphere, this produces an oxide layer that provides a thin, hard surface coating and prevents further reaction with the environment. This contrasts with iron, where the rust flakes off exposing more iron to the air.

Alloying

Iron can be alloyed with small quantities of other metals, such as chromium, nickel, manganese and molybdenum, to produce stainless steel. A wide variety of grades of stainless steel is produced for specific purposes. The most widely used stainless steel contains approximately 18 per cent chromium and 8 per cent nickel. This is called 18-8 stainless steel and is used for kitchen sinks and cutlery.

alloying forming a substance with metallic properties that consists of two or more elements, at least one of which is a metal

electrochemical protection protecting a metal from corrosion by connecting it to a power source and another metal that is less active; the electrons feed onto the metal being protected, making it a cathode so that it cannot be oxidised

The formation of a stable film of chromium (III) oxide, Cr_2O_3, provides a very strong surface protection for this alloy, although the metal may still be attacked by chloride ions. It is important, therefore, to choose an appropriate alloy for any specific use. For example, stainless steel containing molybdenum is more resistant to chloride ion attack and is, therefore, recommended for use in sea water; it is used in valves on scuba diving equipment.

Electrochemical protection

Electrochemical protection involves placing a more reactive metal in electrical contact with the metal that needs protection. It involves 'sacrificing' one metal to save another.

If a block of a metal higher in the reactivity series is connected to iron, that metal is corroded while iron remains intact. Pipes may be protected by attaching bags of magnesium scraps at intervals and replacing them when they have been corroded away. The hulls of ships can be protected by attaching blocks of zinc, which are sacrificed to protect the iron. This is sometimes referred to as **sacrificial protection** or **cathodic protection**. It involves the selection of an appropriate sacrificial anode. A **sacrificial anode** is a reactive metal that is used to prevent a less reactive metal from corroding.

sacrificial protection protection of a metal from corrosion by attaching another more active metal to it; the more active metal reacts preferentially with corrosive elements and protects the less reactive metal

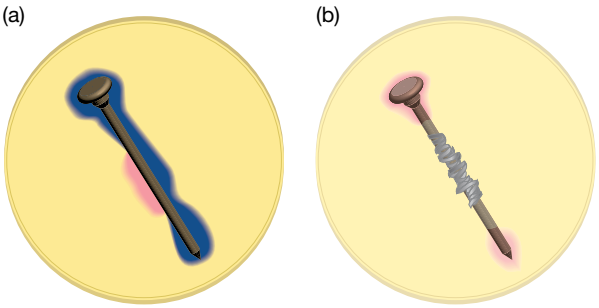
cathodic protection method of protecting a metal from corrosion by connecting it to a power source and another metal that is less active

sacrificial anode a reactive metal used to prevent a less active metal from corroding

A demonstration of cathodic protection can be seen in figure 12.17. Identical iron nails are introduced to two beakers of agar, each containing phenolphthalein indicator and potassium hexacyanoferrate, $K_4Fe(CN)_6$. The first nail (a) is unprotected, while the second nail (b) has magnesium ribbon wrapped around its centre. The pink colouring indicates rusting. For the second nail (b), the magnesium wrapped around the nail protects the nail from corrosion because magnesium is a more reactive metal than iron, meaning it will corrode preferentially, leaving the iron intact.

FIGURE 12.16 Sacrificial anode for cathodic protection on a pipe

FIGURE 12.17 A demonstration of cathodic protection with (a) an unprotected nail and (b) nail with magnesium ribbon wrapped around its centre; the pink colour indicates rusting



Iron pipes exposed to the atmosphere rust very quickly. Even pipes buried beneath the ground rust. To prevent corrosion, blocks of a more reactive metal, such as magnesium, are bolted to aboveground pipes.

Another method of cathodic protection is impressed-current cathodic protection. This is achieved by making iron the cathode that supplies electrons to the protected metal by connecting it to the negative terminal of a battery. A conductor such as graphite is connected to the positive terminal (see figure 12.18). The current is monitored by a reference electrode that keeps the potential difference between the cathode and the anode high enough to ensure protection. The formation of Fe^{2+} ions is inhibited by the negative impressed current on the iron.

Car manufacturers treat newly formed car body shells so that they become resistant to corrosion and attractively coloured. This involves immersing each car body in a series of treatment and paint tanks containing zinc. The body then receives an electrocoat primer on the inside and outside surfaces to provide corrosion protection and prepare the surface for the coloured top-coat. Paint spraying is usually performed by automatic machines using electrostatic attraction.

The car body is then heated to produce a hard, glossy finish.

FIGURE 12.18 Electrochemical protection

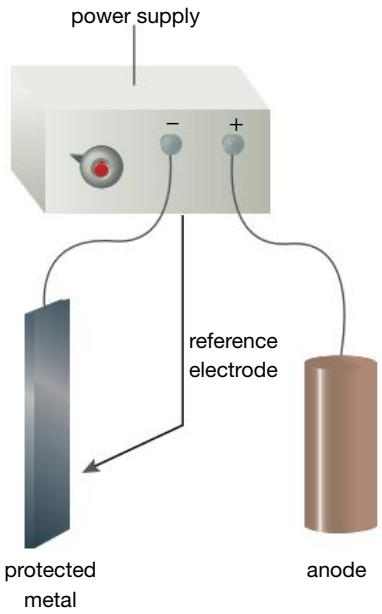


FIGURE 12.19 Car bodies are treated to provide corrosion protection.

CASE STUDY: Napoleon, cans and corrosion

A common example of corrosion prevention is found in your kitchen cupboard. Tin cans have been used as a method for storing food items for about 200 years. The history of this method of food preservation is a fascinating one. In 1795 Napoleon needed a way of feeding his army food that did not spoil while they were away fighting, so the French government offered a large reward to anyone who could provide a solution to Napoleon's problem. After all, he was quoted as saying that 'an army marches on its stomach'.

In 1810 Nicolas Appert, a sweet maker, ultimately won the prize when he proposed putting food in a container, sealing it and then heating it. It was said that he used his method to preserve a whole sheep. His process, however, was not suitable for all purposes because the material he used was glass and prone to shattering. He had considered using tinplate but the quality of the available material was poor.

FIGURE 12.20 Cans are a common example of corrosion prevention found in your kitchen cupboard.

At about the same time, Peter Durand patented the first cans for stopping food from becoming rancid. Again, the food was heated and this seemed to effectively preserve the food. Cans were easier to produce at that time than the thick glass used in Appert's method. These 'tins' were made of steel coated in tin to prevent corrosion and the lids were originally soldered on with lead to provide a good seal. Surprisingly, the can opener was not invented until 50 years later, so it would have been interesting observing the various methods used to open the cans!

It was not long before lead was no longer used because it was found to be toxic. Today, cans are less likely to be coated in tin, which is not readily available. Today's cans are made from alloys and are lined with an epoxy coating. This plastic lining is sometimes bisphenol-A (BPA), which is a controversial compound. Many studies are investigating the safety of this material as well as the use of less contentious alternatives. The development of cans has made a variety of different foods available to us whenever we want them, and the ring top eliminates the need for finding a can opener in the bottom of a drawer.

elog-1844

EXPERIMENT 12.3 online only

Corrosion of iron

Aim

To investigate the corrosion of iron

elog-1845

EXPERIMENT 12.4 online only

Minimising corrosion (student design)

Aim

To design and perform an experiment to investigate the effectiveness of different methods used to minimise metal corrosion

12.5 Activities

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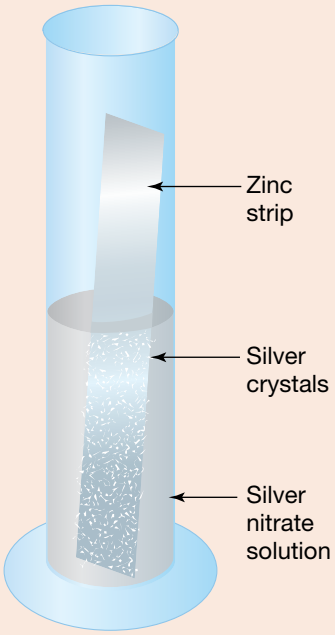
12.5 Quick quiz on | **12.5 Exercise** | **12.5 Exam questions**

12.5 Exercise

1. When zinc metal is dipped into a solution of silver nitrate, it forms a coating of silver, as shown in the provided diagram.
   a. Use this reaction to draw a diagram of a galvanic cell, using KNO_3 in the salt bridge and zinc sulfate as one of the electrolytes. Complete the following steps in constructing your diagram.
      i. Draw the two half-cells.
      ii. Label the flow of:
         - electrons in the wire
         - anions in the salt bridge
         - cations in the salt bridge.
      iii. Label the anode and the cathode.
   b. Write half-equations for the oxidation and reduction reactions.
   c. Write the overall cell reaction.



2. A galvanic cell was set up by combining half-cells containing zinc and magnesium electrodes dipped into the appropriate sulfate solutions. A conducting wire and a salt bridge completed the circuit. After three hours, the two electrodes were removed and weighed. The mass of the Zn electrode had increased, while the mass of the Mg electrode had decreased.
   a. Draw this galvanic cell, clearly indicating the following.
      i. The anode and the cathode
      ii. The ions present in the half cells
      iii. The electrolyte in the salt bridge
      iv. The direction of the flow of electrons
   b. Write half-equations for the anode reaction and the cathode reaction.
   c. Write half-equations for the oxidation reaction and the reduction reaction.
   d. Write the overall cell reaction.
3. Answer the following for an electrochemical cell.
   a. Will the species that competes most strongly for electrons be oxidised or reduced?
   b. Will the half-cell containing this species be in the anode compartment or the cathode compartment?
4. a. What is corrosion?
   b. What type of reaction is involved?
5. Explain the difference between dry corrosion and wet corrosion. Give an example of each.
6. What is the difference between rusting and corrosion?
7. What conditions are required for rusting to occur? Describe the process of rusting, including chemical equations.
8. Describe an experiment to test if iron nails rust more quickly in tap water or salty water.
9. If iron is so susceptible to corrosion, why is it used so extensively?
10. What sort of corrosion protection would you observe in the following examples?
    a. A bike chain
    b. A metal dish-draining rack
    c. A metal roof
    d. A ship's hull

12.5 Exam questions

Question 1 (1 mark)

MC The underground pipe used to deliver gas to some homes passes through damp marshland. Iron pipes are particularly susceptible to corrosion in the environment. What does the overall corrosion reaction involve?

A. The reduction of the iron
B. The oxidation of the iron to Fe^{2+} ions
C. The oxidation of the oxygen to OH^- ions
D. The oxidation of the iron with acid from dissolved carbon dioxide

Question 2 (6 marks)

Consider a galvanic cell containing a copper/copper sulfate half-cell and a lead/lead sulfate half-cell. Potassium nitrate forms the salt bridge.

a. Draw the galvanic cell and use the activity series to predict the following.
- **i.** The substance that is oxidised and the one that is reduced **(1 mark)**
- **ii.** The anode and cathode **(1 mark)**
- **iii.** The direction of electron flow **(1 mark)**

b. Predict the half-cell reactions. **(2 marks)**
c. Predict the overall redox reaction. **(1 mark)**

Question 3 (4 marks)

The conjugate pairs of a half-cell are written in the following way for convenience: $Zn^{2+}(aq)/Zn(s)$. When each of the following half-cells are joined to make a galvanic cell, which is the anode in each?

a. $Pb^{2+}(aq)/Pb(s)$ and $Ag^{+}(aq)/Ag(s)$ **(1 mark)**
b. $Fe^{2+}(aq)/Fe(s)$ and $Mg^{2+}(aq)/Mg(s)$ **(1 mark)**
c. $Pb^{2+}(aq)/Pb(s)$ and $Cu^{2+}(aq)/Cu(s)$ **(1 mark)**
d. $Zn^{2+}(aq)/Zn(s)$ and $Cu^{2+}(aq)/Cu(s)$ **(1 mark)**

Question 4 (8 marks)

Experts estimate that over 20 per cent of the iron produced in the world is used to replace that lost by corrosion.

a. List the substances needed for the rusting of iron to occur. **(2 marks)**
b. Name three other substances that speed up the rusting of iron. **(3 marks)**
c. Explain why iron objects are less likely to rust in the following environments.
- **i.** In a desert **(1 mark)**
- **ii.** In outer space **(1 mark)**
- **iii.** On the ocean floor **(1 mark)**

Question 5 (3 marks)

An iron sculpture was produced to commemorate the anniversary of the founding of a small town. To stop it rusting, the mayor attached it by a wire to a block of zinc that was stored underground out of sight.

a. Explain how the mayor's action would prevent corrosion. **(1 mark)**
b. What would happen if a block of lead was attached to the iron rather than the block of zinc? **(1 mark)**
c. List two other ways in which corrosion of the statue could be prevented. **(1 mark)**

More exam questions are available in your learnON title.

12.6 Review

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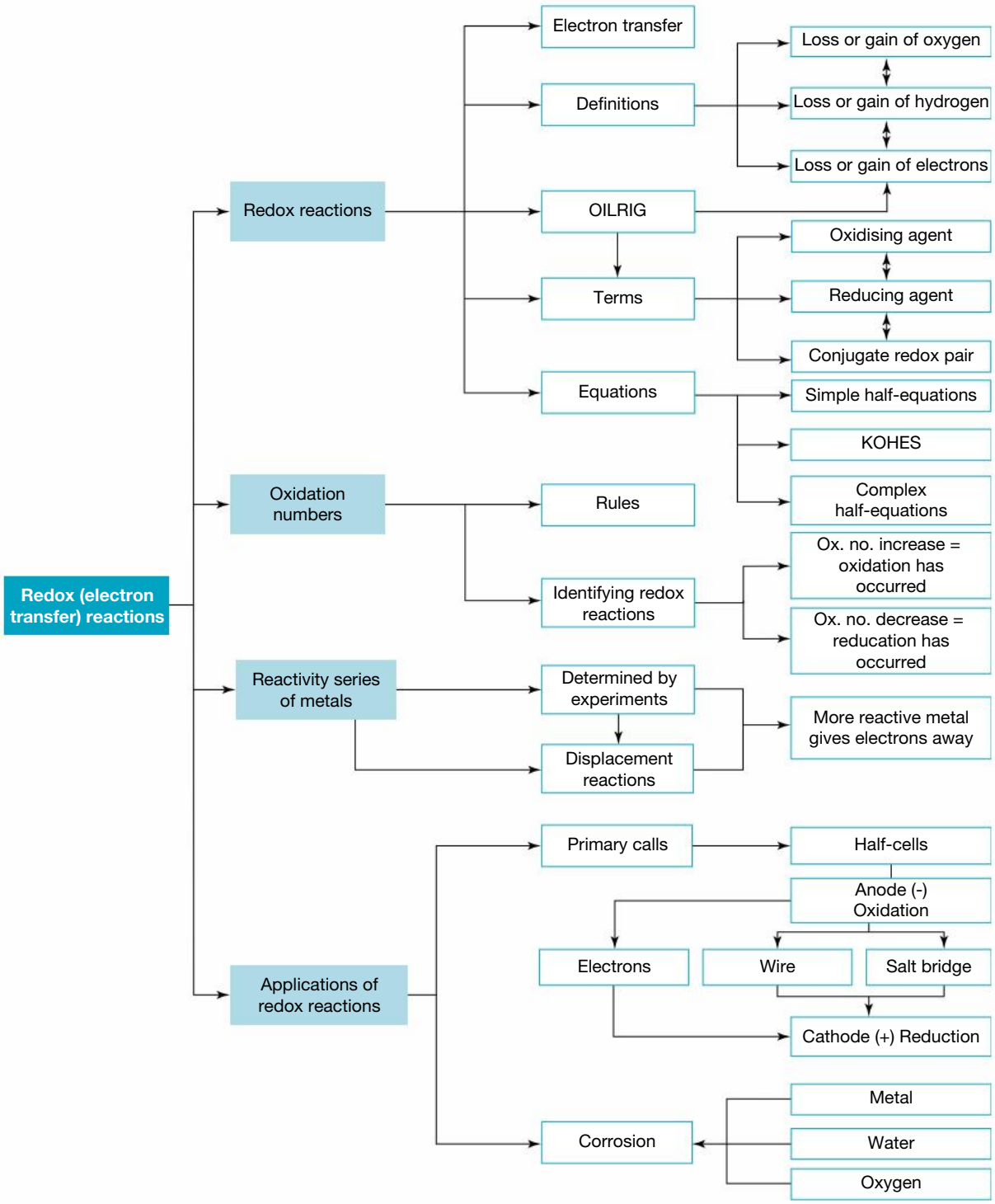
12.6.1 Topic summary



12.6.2 Key ideas summary

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12.6.3 Key terms glossary

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on Resources

Solutions	Solutions — Topic 12 (sol-0811)
Practical investigation eLogbook	Practical investigation logbook — Topic 12 (elog-1617)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 12 (doc-37089) Key idea summary — Topic 12 (doc-37090)
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12.6 Activities

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12.6 Review questions

1. **a.** Define oxidation and reduction reactions in terms of the transfer of oxygen.
   b. Define oxidation and reduction reactions in terms of the transfer of electrons.
   c. What is an oxidising agent?
2. Use the electron transfer approach to identify the oxidising agent, reducing agent, direction of electron transfer, and which species are oxidised and which ones are reduced in the following equations.
   a. $Sn^{2+}(aq) + 2Ce^{4+}(aq) \rightarrow Sn^{4+}(aq) + 2Ce^{3+}(aq)$
   b. $Cd(s) + Pb^{2+}(aq) \rightarrow Cd^{2+}(aq) + Pb(s)$
   c. $2Al(s) + 3Cl_2(g) \rightarrow 2Al^{3+}(aq) + 6Cl^{-}(aq)$
   d. $Mg(s) + S(s) \rightarrow Mg^{2+}(s) + S^{2-}(s)$
   e. $2Na(s) + 2H_2O(l) \rightarrow 2Na^{+}(aq) + OH^{-}(aq) + H_2(g)$
3. Which of the following reactions is more likely to happen if the reactants are mixed?

$$Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$$
$$Pb(s) + Zn^{2+}(aq) \rightarrow Pb^{2+}(aq) + Zn(s)$$

4. Develop overall redox equations by first writing the oxidation and reduction equations for the following reactions. Label the oxidising agent and reducing agent in each case.
   a. Zinc solid is placed in a copper sulfate solution.
   b. An iron nail is placed in a solution of tin(II) nitrate.
   c. A small piece of lead is placed in silver nitrate solution.
   d. A small piece of aluminium is reacted with zinc chloride solution.
5. Tin displaces copper metal from $CuSO_4$ solution and cadmium displaces tin metal from $SnCl_2$ solution. The three metals and their cations are as follows: Sn^{2+}/Sn, Cu^{2+}/Cu, Cd^{2+}/Cd.
   a. Ignoring spectator ions, write an ionic equation for the two displacement reactions described.
   b. Use the information given to deduce the relative order of metals from most reactive to least reactive.
   c. State whether you would expect the following reactions to occur.
   i. $Sn^{2+}(aq) + Cu(s) \rightarrow Sn(s) + Cu^{2+}(aq)$
   ii. $Cd^{2+}(aq) + Cu(s) \rightarrow Cd(s) + Cu^{2+}(aq)$
   iii. $Sn(s) + Cd^{2+}(aq) \rightarrow Sn^{2+}(aq) + Cd(s)$
   iv. $Cd(s) + Cu^{2+}(aq) \rightarrow Cd^{2+}(aq) + Cu(s)$

6. Six different strips of metal, *A* to *F*, were placed in solutions of their metal nitrates to see if any reaction took place. The results are shown in the following table.

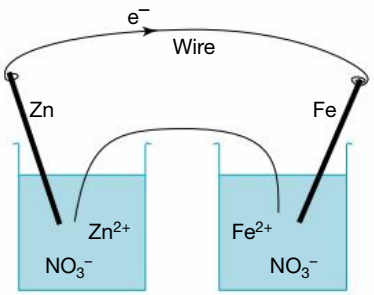
Metal	*A* nitrate	*B* nitrate	*C* nitrate	*D* nitrate	*E* nitrate	*F* nitrate
A	—	✓	✓	✓	×	✓
B	×	—	×	×	×	×
C	×	✓	—	✓	×	×
D	×	✓	×	—	×	×
E	✓	✓	✓	✓	—	✓
F	×	✓	✓	✓	×	—

✓ = metal displaced; × = no reaction

   a. Deduce the order of reactivity of the metals *A* to *F* and list them in order, beginning with the most reactive.
   b. Identify the six metals.

7. Briefly explain the meaning of the following terms.
   a. Galvanic cell
   b. Electrode
   c. Electrolyte
   d. Anode
   e. Cathode

8. Consider the reaction occurring in the diagram shown, and complete the following.
   a. State the anode reaction.
   b. State the cathode reaction.
   c. Find the overall cell reaction.



9. Explain, using equations, why the corrosion of iron to form rust is a redox process.
10. Explain what would happen if the following coatings of a steel object are damaged.
    a. Plastic b. Paint c. Tin d. Zinc

12.6 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC Oxidation is best described as a process that involves what?

A. The gain of protons
B. The loss of oxygen
C. The gain of electrons
D. The loss of electrons

Question 2

MC The process of black and white film photography depends upon a reaction initiated by light and represented by the following equation.

$$2Ag^+(aq) + 2Br^-(aq) \rightarrow 2Ag(s) + Br_2(aq)$$

Name the species that is oxidised in this reaction.

A. $Ag^+(aq)$ **B.** $Br^-(aq)$ **C.** $Ag(s)$ **D.** $Br_2(aq)$

Question 3

MC For the reaction $Al \rightarrow Al^{3+} + 3e^-$, which of the following is *true*?

A. Aluminium loses electrons from its nucleus.
B. Al is reduced to Al^{3+}.
C. The electrons remain free in aqueous solution.
D. This reaction can occur only if another reaction is occurring simultaneously and using electrons.

Question 4

MC Consider the following half-equation.

$$2Cl^-(aq) \rightarrow Cl_2(aq) + 2e^-$$

What is this an example of?

A. An acid–base reaction
B. A precipitation reaction
C. A reduction reaction
D. An oxidation reaction

Question 5

MC A small piece of silver was placed in a solution containing both magnesium nitrate and copper(II) nitrate.

Which one of the following will occur?

A. No reaction occurs.
B. The silver dissolves and only copper is precipitated.
C. The silver dissolves and only magnesium is precipitated.
D. A mixture of copper and magnesium forms on the silver.

Question 6

MC What happens if a piece of shiny aluminium metal is placed in a $CuSO_4$ solution?

A. Sulfur dioxide gas is liberated.
B. No reaction occurs.
C. Copper oxide is formed on the aluminium metal.
D. Copper metal is deposited on the aluminium metal.

Question 7

MC What would happen if you tried to store $FeSO_4$ solution in a container made of Cu metal?

A. The $FeSO_4$ solution would be stored quite safely.
B. The copper of the container would dissolve and Fe metal would form.
C. The copper container would dissolve and Fe^{2+} ions would form.
D. The copper container would dissolve and O_2 gas would form.

Question 8

MC An electrochemical cell will function only when there is a complete circuit for electrical flow. For reactions to occur that produce an electric current, what is it necessary to have?

A. An external and internal circuit for the flow of ions
B. An external circuit for electron flow and an internal circuit for ion flow
C. An external circuit for ion flow and an internal circuit for electron flow
D. An internal and external circuit for the flow of electrons

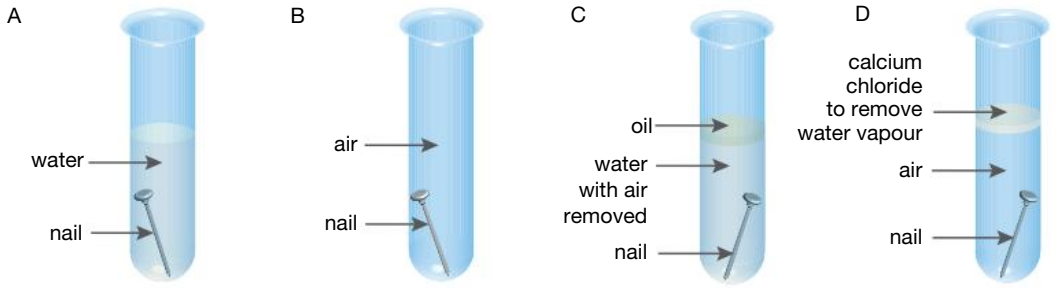
Question 9

MC In any electrochemical cell, the cathode is the electrode:

A. that is closest to the outside of the cell.
B. at which electrons are liberated by some species.
C. at which reduction occurs.
D. at which hydrogen is liberated.

Question 10

MC The following diagrams represent four tests in an experiment designed to find the conditions necessary for a steel nail to rust.



Which diagram represents the most favourable conditions for rust to appear on the nail?

A. A **B.** B **C.** C **D.** D

Section B — Short answer questions

Question 11 (4 marks)

State which of the following are redox reactions:

a. $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$ **(1 mark)**

b. $HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$ **(1 mark)**

c. $Cu^{2+}(aq) + Mg(s) \rightarrow Cu(s) + Mg^{2+}(aq)$ **(1 mark)**

d. $2H_2O(l) + 2F_2(g) \rightarrow 4HF(aq) + O_2(g)$ **(1 mark)**

Question 12 (15 marks)

Balance the following equations using the half-equation method. Identify the oxidation reaction, the reduction reaction and the balanced equation.

a. $Fe_2O_3(s) + CO(g) \rightarrow Fe(s) + CO_2(g)$ **(3 marks)**

b. $P_4(s) + HNO_3(aq) + H_2O(l) \rightarrow H_3PO_4(aq) + NO(g)$ **(3 marks)**

c. $Cr_2O_7^{2-}(aq) + H_2C_2O_4(aq) \rightarrow Cr^{3+}(aq) + CO_2(g)$ **(3 marks)**

d. $Cu(s) + HNO_3(aq) \rightarrow Cu^{2+}(aq) + NO_2(g)$ **(3 marks)**

e. $PbS(s) + H_2O_2(l) \rightarrow PbSO_4(s) + H_2O(l)$ **(3 marks)**

Question 13 (6 marks)

A clean piece of tin metal is placed into a solution of lead nitrate and left for a period of time. It is then observed that the tin has become coated with a layer of dull lead metal.

a. Write the half-equation for the oxidation reaction that has occurred. **(1 mark)**

b. Write the half-equation for the reduction reaction that has occurred. **(1 mark)**

c. State the oxidising agent for this reaction. **(1 mark)**

d. State the reducing agent for this reaction. **(1 mark)**

e. If the tin is removed, cleaned of lead, and then placed into a solution of magnesium chloride, would a further redox reaction occur? Explain. **(2 marks)**

Question 14 (5 marks)

Four hypothetical metallic elements, *A*, *B*, *C* and *D*, form soluble nitrates with the formulas ANO_3, $B(NO_3)_2$, CNO_3 and $D(NO_3)_3$.

Strips of each of the four metals were immersed in solutions of the other metal nitrates and the following observations were recorded.

- Metal *B* underwent reaction in all solutions.
- Metal *A* reacted only with CNO_3.

a. State the formulas of the metal cations in the soluble nitrates. **(2 marks)**

b. Deduce the position of *D* in the series by listing the three metals *A*, *B* and *C* in decreasing order, with respect to their relative reactivities. **(1 mark)**

c. List the four metals in order from most reactive to least reactive. **(1 mark)**

d. Which metal cation is the weakest oxidising agent? **(1 mark)**

Question 15 (2 marks)

Choose the statements that are correct. Justify your response.

- Statement I: The corrosion of iron is enhanced in the presence of salt water.
- Statement II: The corrosion of iron is minimised in the absence of water.
- Statement III: The corrosion of iron is prevented by electrical contact with a less reactive metal.

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AREA OF STUDY 1 How do chemicals interact with water?

OUTCOME 1

Explain the properties of water in terms of structure and bonding, and experimentally investigate and analyse applications of acid–base and redox reactions in society.

PRACTICE EXAMINATION

STRUCTURE OF PRACTICE EXAMINATION		
Section	**Number of questions**	**Number of marks**
A	20	20
B	8	30
	Total	**50**

Duration: 50 minutes

Information:

- This practice examination consists of two parts; you must answer all question sections.
- Pens, pencils, highlighters, erasers, rulers and a scientific calculator are permitted.
- You may use the VCE Chemistry Data Book for this task.

on Resources

Weblink VCE Chemistry Data Book

SECTION A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

1. Which of the following explains why water exists in all three states at the Earth's surface?

A. Water covers about 70 per cent of its surface.

B. Both melting and evaporation can occur within the range of temperatures experienced.

C. Its boiling point is 100 °C.

D. It has covalent bonding between its molecules.

2. New bonds are formed when water freezes to produce ice. What are these bonds between the molecules predominantly?

A. Hydrogen bonds and dispersion forces

B. Covalent bonds and hydrogen bonds

C. Ion–dipole bonds and dispersion forces

D. Ionic bonds and ion–dipole bonds

3. The table provided shows the boiling points of group 16 hydrides.

Hydride	Boiling point (°C)
Water	100
Hydrogen sulfide	–61
Hydrogen selenide	–42
Hydrogen telluride	–2

What is the trend for boiling points going down in this table explained by?

A. The increasing strength of hydrogen bonding between molecules as you go down the group

B. The presence of hydrogen bonding between molecules of water and between molecules of hydrogen telluride, resulting these hydrides having the two highest boiling points

C. The dispersion forces between water molecules being significantly stronger than between molecules of the other hydrides

D. Hydrogen bonding and dispersion forces existing between water molecules, but only dispersion forces existing between molecules of the other hydrides

4. The specific heat capacity of water is a quantitative measurement of what?

A. How easy or hard it is to evaporate water

B. The ease or difficulty of increasing the temperature of a sample of water

C. The heat evolved when a sample of water vapour condenses into liquid water

D. The difference between the melting and boiling points of water

5. Water makes up the largest proportion of sweat, which the body releases as a cooling mechanism. Which of the following reasons can explain this cooling effect?

A. The density of water decreases as it gets warmer.

B. Water is a good solvent for many of the ionic substances found in the body.

C. The electrical conductivity of water increases as it dissolves more ions.

D. Water has a particularly high latent heat of vaporisation for such a small molecule.

6. In which of the following reactions is the bolded species acting as a Brønsted–Lowry acid?

A. $\mathbf{NH_3}(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

B. $HCOO^-(aq) + \mathbf{H_2O}(l) \rightleftharpoons HCOOH(aq) + OH^-(aq)$

C. $H_2O(l) + \mathbf{F^-}(aq) \rightleftharpoons OH^-(aq) + HF(aq)$

D. $2\mathbf{HCl}(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)$

7. Which of the following are *not* acid–base conjugate pairs?

A. HCO_3^- and CO_3^{2-}

B. NH_4^+ and NH_3

C. H_3O^+ and OH^-

D. HBr and Br^-

8. The following table lists some chemical species that have been classified according to their acid–base properties. Which classifications are correct?

	Acid	Base	Amphiprotic
I.	H_3PO_4	PO_4^{3-}	HPO_4^{2-}
II.	HCl	$Ca(OH)_2$	H_2O
III.	CH_4	NH_3	NH_4^+
IV.	HCO_3^-	$HClO_4$	CH_3COOH

A. Classifications (I) and (II) only

B. Classifications (I) and (III) only

C. Classifications (II) and (III) only

D. Classifications (II) and (IV) only

9. When nitric acid reacts with sodium hydrogen carbonate, what is the salt formed in the reaction?

A. Sodium chloride

B. Sodium carbonate

C. Sodium nitrate

D. Sodium sulfate

10. Antacids contain one or more active ingredients for the purpose of neutralising stomach acid (HCl). Which of the following would *not* be suitable as an active ingredient in an antacid tablet?

A. $Al(OH)_3$

B. $MgCO_3$

C. $CaCl_2$

D. $NaHCO_3$

11. Which of the following lists the solutions from lowest to highest pH?

A. 1.0 mol L^{-1} NaOH, 1.0 mol L^{-1} CH_3COOH, 1.0 mol L^{-1} HCl, 2.0 mol L^{-1} HCl

B. 1.0 mol L^{-1} CH_3COOH, 1.0 mol L^{-1} NaOH, 1.0 mol L^{-1} HCl, 2.0 mol L^{-1} HCl

C. 1.0 mol L^{-1} HCl, 2.0 mol L^{-1} HCl, 1.0 mol L^{-1} CH_3COOH, 1.0 mol L^{-1} NaOH

D. 2.0 mol L^{-1} HCl, 1.0 molL^{-1} HCl, 1.0 mol L^{-1} CH_3COOH, 1.0 mol L^{-1} NaOH

12. Two acidic solutions were compared. One was a solution of sulfuric acid, while the other was a solution of a carboxylic acid, ethanoic acid. Both solutions were found to have the same hydronium, H_3O^+, concentration. As a result, what is also true?

A. Both acids have the same molar concentration.

B. Both acids have an equal tendency to react with water.

C. Both solutions have the same pH.

D. Both solutions have most of their acids ionised.

13. A class was given the task of determining the pH of a solution using a variety of methods. These included the following.

I. Using universal indicator and comparing the colour obtained against a colour chart

II. Using a pH meter that was subsequently found to be improperly calibrated

III. Using a correctly calibrated pH meter

Which method would be most likely to produce results that were precise but not accurate?

A. I only

B. II only

C. III only

D. I and III only

14. Is natural rainfall acidic?

A. Yes, but only slightly due to its reaction with atmospheric carbon dioxide.

B. No. Water is neutral and has a pH of 7.

C. No. It is only acidic in regions with a large amount of atmospheric pollution.

D. Yes, but only slightly as evidenced by its pH of 7.05.

15. The following four equations represent four different reactions.

I. $HI(aq) + KOH(aq) \rightarrow KI(aq) + H_2O(l)$

II. $Mg(s) + S(l) \rightarrow MgS(s)$

III. $H_2O(l) + HCl(aq) \rightarrow H_3O^+(aq) + Cl^-(aq)$

IV. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

Which of the following statements is true?

A. Reactions (I) and (IV) are both acid–base reactions.

B. Reactions (II) and (IV) are both redox reactions.

C. Reaction (I) is redox and reaction (III) is an acid–base reaction.

D. Reaction (IV) is neither acid–base nor redox.

16. For the reaction

$$Pb^{2+}(aq) + Cd(s) \rightarrow Cd^{2+}(aq) + Pb(s)$$

which of the following statements is correct?

A. Cd(s) is oxidised and $Pb^{2+}(aq)$ is the reducing agent.
B. Cd(s) is oxidised and $Pb^{2+}(aq)$ is the oxidising agent.
C. Cd(s) is reduced and Pb(aq) is the oxidising agent.
D. Cd(s) is reduced and $Pb^{2+}(aq)$ is the reducing agent.

17. If an element is described as an oxidising agent, what does it do?

A. Accepts protons during a reaction
B. Is reduced when it accepts electrons from an reducing agent
C. Donates electrons to a reducing agent
D. Gains electrons from a reducing agent and is oxidised

18. Which of the following is a balanced *reduction* half-equation?

A. $2H^+(aq) + e^- \rightarrow H_2(g)$
B. $Ni(s) \rightarrow Ni^{2+}(aq) + e^-$
C. $Cl_2(g) \rightarrow 2Cl^-(aq) + 2e^-$
D. $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

19. Several small pieces of zinc, Zn, are added to a solution of cobalt(II) sulfate, which is pink in colour. The equation for this reaction is

$$Zn(s) + CoSO_4(aq) \rightarrow ZnSO_4(aq) + Co(s).$$

Which of the following observations would be made?

A. A deposit of cobalt metal would form and the zinc pieces would dissolve.
B. The zinc pieces would dissolve and the cobalt(II) sulfate solution would become a stronger pink colour.
C. Cobalt metal would dissolve and the zinc pieces would increase in size.
D. A precipitate of zinc sulfate would form and the cobalt(II) sulfate solution would become a paler pink colour.

20. Which of the following methods will *not* minimise the corrosion of iron?

A. Coating the surface with an impenetrable layer of plastic or enamel
B. Forming an alloy of the metal with non-corrosive metals like chromium
C. Placing a less reactive metal like nickel in electrical contact with the iron
D. Applying a light coating of zinc to the surface

SECTION B — Short answer questions

Question 21 (4 marks)

a. The density of water varies with temperature. It is at its highest at 4 °C. At temperatures below 0 °C, however, water has a lower density. Explain why this is the case. **(1 mark)**

b. The density of water at 25 °C is 0.997 g mL^{-1}. Use this value to calculate the mass (in grams) of 10.0 L of water. **(1 mark)**

c. Given that the specific heat capacity of water is 4.18 J g^{-1} $°C^{-1}$, what amount of heat energy in Joules would be required to heat 10.0 L of water from 25 °C to 100 °C ? **(1 mark)**

d. What will happen to the density of water as its temperature rises from 25 °C? Give reasons for your answer. **(1 mark)**

Question 22 (5 marks)

Write equations with water to demonstrate the following.

a. Nitric acid, HNO_3, is a strong monoprotic acid. **(1 mark)**

b. H_2S is a weak diprotic acid. **(2 marks)**

c. The hydrogen sulfite ion, HSO_3^-, is an amphiprotic substance. **(2 marks)**

Question 23 (2 marks)

Explain the following chemical properties.

a. Methanoic acid, HCOOH, is a molecular substance, yet when it is dissolved in water, the resulting solution is able to conduct electricity. **(1 mark)**

b. Citric acid is found in many foods and is able to be consumed, yet spilling hydrochloric acid on your skin would result in a chemical burn. **(1 mark)**

Question 24 (3 marks)

Write balanced equations as directed for the following reactions. Make sure that symbols of state are included.

a. The full equation for the reaction between calcium hydroxide solution, $Ca(OH)_2$, and dilute nitric acid, HNO_3. **(1 mark)**

b. The full equation for the reaction between magnesium metal, Mg, and dilute hydrochloric acid, HCl. **(1 mark)**

c. The ionic equation for the reaction between zinc metal, Zn, and dilute nitric acid, HNO_3. **(1 mark)**

Question 25 (3 marks)

a. A solution of hydrochloric acid, HCl, was prepared by diluting 0.50 mol L^{-1} hydrochloric acid stock solution.

i. Calculate the pH of 0.50 mol L^{-1} hydrochloric acid. **(1 mark)**

ii. What is the final concentration of hydrochloric acid if 100.0 mL of water is added to 20.0 mL of 0.50 mol L^{-1} hydrochloric acid? **(1 mark)**

b. Determine the pH of a 0.0010 mol L^{-1} solution of the alkali, NaOH. **(1 mark)**

Question 26 (7 marks)

a. Concentrated nitric acid, HNO_3, is able to undergo redox reactions with many metals, including less reactive metals such as copper. This reaction produces copper(II) ions and nitrogen dioxide gas among the products.

i. Derive the two half-equations for this reaction. **(2 marks)**

ii. Label the half-equation that is the reduction reaction. **(1 mark)**

iii. Use the two half-equations to write the overall ionic equation for this reaction. **(1 mark)**

b. Nitrogen dioxide molecules often combine (dimerise) to form dinitrogen tetroxide, N_2O_4. What is the oxidation number of nitrogen in the following?

i. HNO_3 **(1 mark)**

ii. N_2O_4 **(1 mark)**

c. In some reactions with metals, nitric acid produces nitrogen monoxide instead of nitrogen dioxide. Balance the following equation using the half-equation method.

$__ HNO_3(aq) + __V(s) + __H^+(aq) \rightarrow __NO(g) + __V^{2+}(aq) + __H_2O$ **(1 mark)**

Question 27 (3 marks)

The reactivity series of metals is listed from highest to lowest:
K, Na, Ca, Mg, Al, Zn, Fe, Ni, Sn, Pb, Cu, Ag, Au.
The following experiments were set up to investigate this series using displacement reactions with metals.

a. In the first experiment, three iron nails were polished and each was placed into a 1 mol L^{-1} solution of one of the following compounds: nickel(II) sulfate, zinc sulfate and tin(II) sulfate.

i. Predict the products of any potential reaction(s). **(1 mark)**

ii. Write the ionic equation for one predicted reaction. **(1 mark)**

b. In a second experiment, three nails made of a different metal from iron were tested in the same solutions. This time none of the three nails was observed to react. Give a possible metal used to make these nails. **(1 mark)**

Question 28 (3 marks)

Metals may be coated in a number of ways to prevent their corrosion.
Use the reactivities of the metals involved to explain the following.

a. 'Tin plate' is sheet iron coated with a thin layer of tin. If it is scratched to expose the iron underneath, the iron corrodes. **(1 mark)**

b. Galvanised iron is sheet iron coated with a thin layer of zinc. If this is scratched to expose the iron underneath, the iron will still stay protected. **(1 mark)**

c. Although aluminium is relatively high on the activity series, external aluminium structures very rarely show signs of significant corrosion. **(1 mark)**

PRACTICE SCHOOL-ASSESSED COURSEWORK

ASSESSMENT TASK — ANALYSIS AND EVALUATION OF A SOCIO-SCIENTIFIC ISSUE

In this task, you will write a scientific inquiry essay on an issue related to the properties and reactions of water. This essay will relate the properties of water to a real-life issue.

- Pens, pencils, highlighters, erasers, rulers and a scientific calculator are permitted.
- You may use the VCE Chemistry Data Book for this task.

Total time: 55 minutes (5 minutes reading, 10 minutes planning, 40 minutes writing)

Total marks: 50 marks

WATER AND THE ENVIRONMENT

Introduction

Science writing is not always in the form of academic research papers. The general public receives the majority of their information about science through the popular media. Science journalists play an important role in presenting up-to-date scientific information and engaging the public with important scientific issues and developments.

Task

This task is a problem-solving assessment in the form of a science essay. You will discuss the real-life applications of chemical concepts, their related issues, and the effect on our everyday lives. The task is to write an essay based on **one** of the following topics.

- The relationship between acid rain and the environment
- The global issue of ocean acidity
- Metal corrosion effects on our environment

Please use the following assessment tasks as a guide to sequencing your essay. The assessment will be based on the following.

1. Presentation of the issue.
2. Causes of the issue.
3. Use of relevant graphs, reactions, equations and diagrams; for example, if referring to ocean acidity, chemical equation examples should be used to demonstrate this chemical process.
4. Effects of the issue, including discussion of the current state of the issue.
5. Solutions to the issue, including opinions or suggestions as to how the problem could be resolved using chemistry.

on Resources

Digital document U2AOS1 School-assessed coursework (doc-30864)

13 Measuring solubility and concentration

KEY KNOWLEDGE

In this topic you will investigate:

Measuring solubility and concentration

- solution concentration as a measure of the quantity of solute dissolved in a given mass or volume of solution (mol L^{-1}, g L^{-1}, %(m/v), %(v/v), ppm), and including unit conversions
- the use of solubility tables and solubility graphs to predict experimental determination of ionic compound solubility; the effect of temperature on the solubility of a given solid, liquid or gases in water
- the use of precipitation reactions to remove impurities from water.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of VCE Chemistry. Experiments and investigations, supported by a **practical investigation eLogbook** and **teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

EXAM PREPARATION

Access exam-style questions and their video solutions in every lesson, to ensure you are ready.

13.1 Overview

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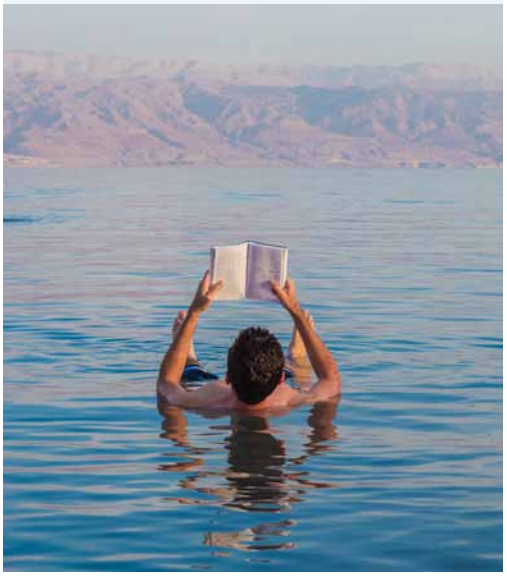
13.1.1 Introduction

The Dead Sea is a salt lake in lower Israel. It is famous for its high concentration of salt. Measurements show that the Dead Sea is about ten times saltier than the ocean, which means that people can float in it very easily. It is so salty that during the hot summers, the evaporation of water in certain parts causes extensive deposits of salt crystals to form.

When such measurements are made, what units are they expressed in? How can solubility data be analysed and processed? Why do the salt crystals form? This topic will examine these questions and build further upon some of the ideas from topic 4. The process by which dissolving occurs will be re-examined and extended to molecular substances as well. Solubility curves will be introduced as graphs that display solubility data in an easy-to-understand format. Factors that influence the dissolving process will also be examined. A range of units whereby solubility can be expressed in quantitative terms will be discussed. This range of units is necessary due to the large number of contexts in which it is desired to know the concentrations of dissolved substances. A special unit of interest to chemists (molarity) will also be introduced.

FIGURE 13.1 The extremely high salt concentration in the Dead Sea means the water is denser than a human.



LEARNING SEQUENCE

13.1 Overview 514
13.2 Measuring solution concentration 515
13.3 Factors that influence solubility 526
13.4 Solubility graphs 536
13.5 Review 550

on Resources

Solutions	Solutions — Topic 13 (sol-0812)
Practical investigation eLogbook	Practical investigation logbook — Topic 13 (elog-1618)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 13 (doc-37091) Key idea summary — Topic 13 (doc-37092)
Exam question booklet	Exam question booklet — Topic 13 (eqb-0094)

13.2 Measuring solution concentration

KEY KNOWLEDGE

- Solution concentration as a measure of the quantity of solute dissolved in a given mass or volume of solution (mol L^{-1}, g L^{-1}, %(m/v), %(v/v), ppm), and including unit conversions

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13.2.1 Defining concentration

Concentration is usually defined in terms of the amount of **solute** per volume of **solvent**. The greater the quantity of solute added to the solvent, the more concentrated the **solution** becomes. For example, the more sugar you dissolve in your cup of coffee, the higher the concentration of sugar in the coffee solution. When only a very small amount of solute is present, the solution may be called **dilute** (that is, its concentration is relatively low).

FIGURE 13.2 Solutions with decreasing concentration from left to right. The lightest coloured flask has $\frac{1}{16}$th of the concentration of the darkest flask.

Often, we use the terms 'concentrated' and 'dilute' comparatively. A concentrated solution of hydrochloric acid, for example, contains 37 g of HCl in 100 g of solution. If the same volume of the solution contained only 10 g of HCl, it would be called a dilute solution of hydrochloric acid. The unit of concentration depends on the units for the quantity of solution and the quantity of solute.

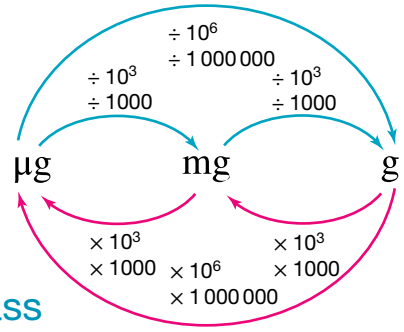
A concentrated solution may be diluted by adding more solvent.

The concentration of a solution is a measure of how much solute has been dissolved in a certain volume of solvent. It is often expressed as a ratio of the solvent amount unit to the volume unit. A number of different units are typically used.

Common units of concentration include:

- grams per litre (g L^{-1})
- milligrams per litre (mg L^{-1})
- parts per million (ppm)
- %(m/m), %(m/v), %(v/v)

FIGURE 13.3 How to convert between different units of mass



13.2.2 Concentration in grams per volume or mass

The concentration of a solution can be represented in various units, and the conversion of those units becomes very important.

- Mass is measured in grams (g), milligrams (mg) and micrograms (µg).
- Volumes of liquids are measured in litres (L), millilitres (mL) and kilolitres (kL), where 1000 mL = 1 L. Remember that 1000 L is 1 kilolitre (kL), which is also equal to 1 cubic metre (1 m^3).

concentration the amount of solute that is dissolved in a known volume of solvent

solute dissolved substance in a solution

solvent substance in which the solute is dissolved

solution solute dissolved in a solvent

dilute not concentrated; has a small quantity of solute in a known volume of solution

Mass and volume conversions

Mass conversions	*Volume conversions*
1 g = 1000 mg = 1 000 000 µg	1 mL = 10^{-3} L
1 g = 10^3 mg = 10^6 µg	1 L = 1000 mL = 10^3 mL
1 µg = 0.001 mg = 0.000 001 g	1000 L = 1 000 000 mL = 1 kL
1 µg = 10^{-3} mg = 10^{-6} g	10^3 L = 10^6 mL = 1 kL

Concentration in grams per L (g L^{-1}; mg L^{-1})

Concentration units can be expressed in many ways. A common form is the ratio of the units of solute to the units of volume (for example, g L^{-1}, mg L^{-1}).

Concentration by mass

$$\text{Concentration } (c) \text{ in grams per litre } \left(\text{g L}^{-1}\right) = \frac{\text{mass of solute in grams}}{\text{volume } (V) \text{ of solution in litres}}$$

$$\text{Concentration } (c) \text{ in milligrams per litre } \left(\text{m g L}^{-1}\right) = \frac{\text{mass of solute in milligrams}}{\text{volume } (V) \text{ of solution in litres}}$$

When the concentration unit mg L^{-1} is required, the mass in grams is converted to mass in milligrams by multiplying by 1000. For smaller solute amounts, mg L^{-1} is usually more convenient to use.

tlvd-0585

SAMPLE PROBLEM 1 Finding the concentration of a solution using mass and volume

A fence-post preservative solution is prepared by dissolving 4.00 g of zinc chloride in enough water to make 2250 mL of solution. Find the concentration (c) of this solution in g L^{-1} and then in mg L^{-1}.

THINK

1. Compare units given to units required. The volume of the solution is given in mL, while the concentration is required in g L^{-1}. We need to convert from mL to L.

2. Use the equation for concentration in g L^{-1}.
$c = \dfrac{\text{mass of solute in grams}}{\text{volume of solution in litres}}$

3. The result from the previous step is in g L^{-1} and needs to be converted to mg L^{-1}.
Remember, 1 g = 1000 mg L^{-1}.

WRITE

Convert 2250 mL to a volume in litres.
Since 1000 mL = 1 L:

$V = \dfrac{2250}{1000} = 2.250 \text{ L}$

$c = \dfrac{\text{mass of solute in grams}}{\text{volume of solution in litres}}$
$= \dfrac{4.00}{2.250}$
$= 1.78 \text{ g L}^{-1}$

$c = 1.78 \times 1000$
$= 1780 \text{ mg L}^{-1}$

PRACTICE PROBLEM 1

Find the mass of sodium bromide required to prepare 50 mL of a 0.40 g L^{-1} solution.

tlvd-0586

SAMPLE PROBLEM 2 Finding the volume of a solution using mass and concentration

What volume of solution is required to dissolve 125 mg of lithium chloride to give a concentration of 0.0500 g L^{-1}?

THINK

1. Compare units given to units required. The mass of the solution is given in mg, while the concentration is required in g L^{-1}. We need to convert from mg to g.

2. Use the equation for volume of solution (L).
$$V = \frac{\text{mass of solute (g)}}{\text{concentration } (\text{g L}^{-1})}$$
Record answer to 3 significant figures.

WRITE

Since 1000 mg = 1 g

$$\text{Mass of solute} = \frac{125 \text{ mg}}{1000}$$
$$= 0.125 \text{ g}$$

$$V = \frac{\text{mass of solute (g)}}{\text{concentration } (\text{g L}^{-1})}$$
$$= \frac{0.125}{0.0500}$$
$$= 2.50 \text{ L}$$

PRACTICE PROBLEM 2

What volume of solution is required to dissolve 425 mg of sodium chloride to give a concentration of 0.100 g L^{-1}?

Concentration in ppm

When very small quantities of solute are dissolved to form a solution, the concentration can be measured in **parts per million (ppm)**. This can be written as mg L^{-1} or µg g^{-1}. The concentration of chemicals in the environment and trace elements in the soil are often expressed in ppm. This is a method of comparing the amount of solute, measured in millionths of a gram (micrograms or µg), to the mass of solution, measured in grams. So this is really a mass per mass ratio.

Concentration in ppm

$$\text{Parts per million (ppm)} = \frac{\text{mass of solute in micrograms}}{\text{mass of solution in grams}} = \frac{\text{mass of solute in milligrams}}{\text{volume of solution in litres}}$$

Our drinking water is treated with chlorine to kill bacteria. Too much chlorine is dangerous and too little does not kill the bacteria. A safe amount is about 1 part per million (ppm), or 1 microgram of chlorine in each gram of solution. (A microgram, µg, is 10^{-6} g, or 0.000 001 g.)

parts per million (ppm) number of a particular component present within one million parts

tlvd-0587

SAMPLE PROBLEM 3 Calculating concentration in g L^{-1} and ppm

a. If 10 g of chlorine gas is dissolved in every 2 500 000 L of water, express the concentration of the chlorine water in:

i. g L^{-1}

ii. ppm (µg g^{-1}).

(*Note:* 1 mL water = 1 g.)

b. Would this water be fit to drink?

THINK

a. Compare units given to units required.

i. Use the equation for concentration in g L^{-1}.

$$c = \frac{\text{mass of solute in grams}}{\text{volume of solution in litres}}$$

ii. Use the equation for concentration in ppm to find the mass of the solute and the mass of the solution.

$$c = \frac{\text{mass of solute in micrograms}}{\text{mass of solution in grams}}$$

b. Recall that a safe amount of chlorine in water to kill bacteria is about 1 part per million (ppm), or 1 microgram of chlorine in each gram of solution. Determine whether water is fit to drink.

WRITE

$V = 2.5 \times 10^6$ L

$$c = \frac{\text{mass of solute in grams}}{\text{volume of solution in litres}}$$

$$= \frac{10\text{g}}{2.5 \times 10^6}$$

$$= 4.0 \times 10^{-6}\text{g L}^{-1}$$

Mass of solute:

Convert g to µg; mass of solute $= 10 \times 10^6$

$= 1 \times 10^7$ µg

Mass of solution:

Convert L to g:

mass of solution $= 2.5 \times 10^6 \times 10^3$

$= 2.5 \times 10^9$ g

Calculate the concentration in ppm.

$$c = \frac{1 \times 10^7}{2.5 \times 10^9}$$

$= 4.0 \times 10^{-3}$ppm

The water is unfit to drink because insufficient chlorine is present to kill the bacteria. (*Note:* when the solvent is water, mg L^{-1} also equals ppm.)

PRACTICE PROBLEM 3

A mass of 15 g of a substance is dissolved in 100 m^3 of water. Calculate its concentration in g L^{-1} and ppm.

13.2.3 Concentration in percentage by mass or volume

Concentration in percentage by mass, %(m/m)

Solutions of concentrated acids and some household cleaners often carry a label showing **%(m/m)**; for example, sulfuric acid, 98%(m/m). The unit %(m/m) expresses the concentration of the solute as a percentage of the mass of solution. So, 98%(m/m) sulfuric acid means there is 98 grams of sulfuric acid solute per 100 grams of solution. We can calculate this mass percentage using the following equation.

Concentration in percentage by mass, %(m/m)

$$\text{Concentration in \% (m/m)} = \frac{\text{mass of solute in g}}{\text{mass of solution in g}} \times \frac{100}{1}$$

tlvd-0589

SAMPLE PROBLEM 4 Finding the concentration in %(m/m)

Find the concentration in %(m/m) obtained when 18.5 g of hydrochloric acid is dissolved in 50.0 g of solution.

THINK

1. Compare units given to units required for the formula for %(m/m).

   $$\%(\text{m/m}) = \frac{\text{mass of solute in g}}{\text{mass of solution in g}} \times \frac{100}{1}$$

   No need to convert units.
2. Calculate %(m/m) by applying the formula for %(m/m).

WRITE

18.5 g HCl and 50.0 g solution.

$$\begin{aligned}\text{\% (m/m)} &= \frac{\text{mass of solute in g}}{\text{mass of solution in g}} \times \frac{100}{1} \\ &= \frac{18.5}{50.0} \times \frac{100}{1} \\ &= 37.0\%\ \text{(m/m)}\end{aligned}$$

PRACTICE PROBLEM 4

Calculate the concentration in %(m/m) when 24.5 g of copper(II) sulfate is dissolved in 50 g of solution.

Concentration in percentage by mass/volume, %(m/v)

An oven cleaner may have on its label 'active ingredient 1%(m/v) sodium hydroxide'. The expression **%(m/v)** expresses the concentration as mass of solute in grams per 100 mL of solution. We can use the following formula to perform calculations.

%(m/m) percentage mass per mass
%(m/v) percentage mass per volume

Concentration in percentage by mass/volume, %(m/v)

$$\text{Concentration \%(m/v)} = \frac{\text{mass of solute in g}}{\text{volume of solution in mL}} \times \frac{100}{1}$$

tlvd-0590

SAMPLE PROBLEM 5 Finding the concentration in %(m/v)

Saline bags are commonly used in hospitals. A saline bag contains 4.30 g of sodium chloride dissolved in 0.500 L of aqueous solution. Calculate the %(m/v) of NaCl in the saline bag.

THINK

1. Compare units given to units required for the formula for %(m/v).

$$\%(\text{m/v}) = \frac{\text{mass of solute in grams}}{\text{volume of solution in mL}} \times \frac{100}{1}$$

Need to convert L to mL.

2. Calculate %(m/v) by applying the formula.

WRITE

$V = 0.500\ \text{L} \times 1000$
$= 500\ \text{mL solution}$
$m = 4.30\ \text{g NaCl}$

$$\%\ (\text{m/v}) = \frac{\text{mass of solute in grams}}{\text{volume of solution in mL}} \times \frac{100}{1}$$
$$= \frac{4.30}{500} \times \frac{100}{1}$$
$$= 0.860\%\ (\text{m/v})$$

PRACTICE PROBLEM 5

Calculate the %(m/v) of 200 mg of potassium chloride in 50.0 mL of solution.

Concentration in percentage by volume, %(v/v)

The percentage by volume, **%(v/v)**, expresses the concentration unit for liquid solutes as volume of liquid solute in mL per 100 mL of solution. For example, a one-litre bottle of liqueur is labelled 40% alc/vol. This means 40 mL of ethanol (alcohol) is present in each 100 mL of liqueur. The following equation is used.

%(v/v) percentage volume per volume

Concentration in percentage by volume, %(v/v)

$$\text{Concentration } \%(\text{v/v}) = \frac{\text{volume of solute in mL}}{\text{volume of solution in mL}} \times \frac{100}{1}$$

tlvd-0591

SAMPLE PROBLEM 6 Finding the concentration in %(v/v)

If a standard glass (0.200 L) of a particular brand of beer contains 9.80 mL of ethanol, calculate the concentration of the beer in %(v/v).

THINK

1. Compare units given to units required for the formula for %(v/v).

$$\%(\text{v/v}) = \frac{\text{volume of solute in mL}}{\text{volume of solution in mL}} \times 100\%$$

We need to convert L to mL.

WRITE

$V = 0.200\ \text{L} \times 1000$
$= 200\ \text{mL solution}$
$V = 9.8\ \text{mL solute}$

2. Calculate %(v/v) by applying the formula.

$$\%(\text{v/v}) = \frac{\text{volume of solute in mL}}{\text{volume of solution in mL}} \times \frac{100}{1}$$
$$= \frac{9.80}{200} \times \frac{100}{1}$$
$$= 4.90\%(\text{v/v})$$

PRACTICE PROBLEM 6

The label of an organic vanilla essence used for cooking shows a 35%(v/v) alcohol content. Tests in the laboratory found that a 25 mL sample of the essence contained 8 mL organic alcohol. Is the label on the bottle correct?

13.2.4 Molar concentration

Concentration can be expressed in a number of ways. Chemists most commonly use molar concentration or **molarity**, which is *the amount of solute, in moles, present in each litre of solution*. The symbol for concentration is *c*. The unit of measurement for molar concentration is moles per litre, which can be written as either M or mol L^{-1}. Therefore, a 1.5 M solution has a concentration of 1.5 mol L^{-1}.

If 1 mol of solute is dissolved in a total volume of 1 L of water, the concentration of the solution is 1 M. If 0.5 mol of solute is dissolved in a total volume of 1 L of water, the concentration of the solution is 0.5 M. If, however, 0.5 mol of solute is dissolved in a total volume of 0.5 L of water, the concentration of the solution is 1 M.

It follows from the provided definition that we can produce the following formula for calculating the number of moles present in a solution. This formula is $n = cV$, where n is the number of moles, c is the concentration (in mol L^{-1}) and V is the volume (in L).

molarity concentration of moles of solute per litre of solution, with units M or mol L^{-1}

Molar concentration

$$\text{Concentration } (c) = \frac{\text{quantity of solute (moles)}}{\text{volume of solution (L)}}$$
$$c = \frac{n}{V} \left(\text{M or mol L}^{-1}\right)$$

The VCE Chemistry Data Book contains the formulas $n = cV$ and $n = \frac{m}{M}$. These formulas can be manipulated to determine the concentration and subsequently the mass of solute.

Determining the mass of solute (*m*)

$$n = c \times V$$
$$m = n \times M$$

Where:
V is volume (litres, L)
c is concentration (M or mol L^{-1})
M is molar mass.

In order to prepare a particular volume of solution of known concentration, the following five steps should be followed:

1. Calculate the number of moles of solute that are needed to obtain the correct concentration of solution for the volume of solvent to be used, according to the formula $n = c \times V$.
2. Calculate the mass of the solute needed, using the formula $m = n \times M$.
3. Partially fill a volumetric flask with water, and add the correct mass of solute.
4. Dissolve the solute.
5. Add water to the required volume.

Determining the volume (V) required for a specific concentration

$$\text{Volume } (V) = \frac{\text{quantity of solute (moles)}}{\text{concentration (moles L}^{-1})} = \frac{n}{c}$$

Note: volume is calculated in litres.

tlvd-0592

SAMPLE PROBLEM 7 Calculating the number of moles needed to obtain a specific concentration

Calculate the number of moles of sodium chloride needed to prepare 500 mL of a 0.0800 mol L^{-1} salt solution.

THINK	WRITE
1. List the known information, and then compare the units given to units required for the formula. $n = c \times V$. Convert mL to L.	$V = 500 \text{ mL} \div 1000$ $= 0.500 \text{ L}$ $c = 0.0800 \text{ mol L}^{-1}$
2. Calculate the number of moles (n) of NaCl needed.	$n = c \times V$ $= 0.0800 \times 0.500$ $= 0.0400 \text{ mol}$

PRACTICE PROBLEM 7

Calculate the concentration (molarity) of copper(II) sulfate, $CuSO_4$, in 1 L of solution containing 200 g $CuSO_4 \cdot 5H_2O$.

tlvd-0593

SAMPLE PROBLEM 8 Calculating the mass needed to obtain a specific concentration

What mass of NaCl would be required to prepare 0.250 L of a solution of 0.0800 mol L^{-1} solution?

THINK	WRITE
1. List the known information and then compare the units given to units required for the formulas that can assist. $c = \frac{n}{V}$ and $n = \frac{m}{M}$. No conversion required.	$V = 0.250 \text{ L}$ $c = 0.0800 \text{ mol L}^{-1}$

2. Calculate moles from concentration.	$c = \frac{n}{V}$ $n = cV$ $= 0.0800 \text{ mol L}^{-1} \times 0.25 \text{ L}$ $= 0.0200 \text{ mol}$
3. Rearrange $n = \frac{m}{M}$ to calculate mass of NaCl.	$n = \frac{m}{M}$ $m = n \times M$ $= 0.0200 \text{ mol} \times (23 + 35.5)$ $= 1.17 \text{ g}$

PRACTICE PROBLEM 8

Calculate the mass of solute in 120 mL of 1.7 M $Mg(NO_3)_2$.

tlvd-0594

SAMPLE PROBLEM 9 Calculating the volume needed to obtain a specific concentration

What volume of water would be required to prepare a solution of 0.0800 mol L^{-1} with 5.00 g of NaCl?

THINK	WRITE
1. List the known information and then compare the units given to units required for the formulas that can assist. $C = \frac{n}{V}$ and $n = \frac{m}{M}$ No conversion required.	$c = 0.0800 \text{ mol L}^{-1}$ $m = 5.00 \text{ g}$
2. Calculate moles from the mass of NaCl given.	$n = \frac{m}{M}$ $= \frac{5.00}{(23 + 35.5)}$ $= 0.0855 \text{ mol}$
3. Use calculated moles to calculate V by rearranging formula $c = \frac{n}{V}$.	$C = \frac{n}{V}$ $V = \frac{n}{c}$ $= \frac{0.0855 \text{ mol}}{0.0800 \text{ mol L}^{-1}}$ $= 1.07 \text{ L}$

PRACTICE PROBLEM 9

A student needs to prepare a solution of 0.100 mol L^{-1} from 7.50 g of $AlCl_3$. What volume of water would they need to use?

13.2.5 Concentration of ions in solution

As has already been seen, many ionic substances are **soluble**. The process of dissolving involves the ionic lattice breaking up and the anions and cations dissociating from each other.

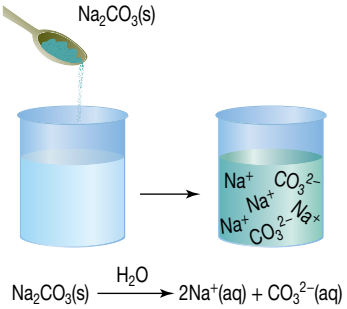
The concentrations of the resulting ions can be calculated. These concentrations are called **ionic concentrations** and they are designated by square brackets.

For example:

$$[Na^+] = 0.5 \text{ mol L}^{-1} \text{ (or M)}$$

means that the ionic concentration of Na^+ ions in solution is 0.5 mol L^{-1} (or M).

FIGURE 13.4 Dissolving sodium carbonate brings about dissociation of the sodium and carbonate ions.



Ionic concentrations

Ionic concentrations are calculated by finding the concentration of the solution, and then multiplying the mole ratio of ions in the dissociation equation.

soluble will dissolve

ionic concentrations concentrations of the anions and cations resulting from the dissolution of an ionic lattice; designated by square brackets

tlvd-0595

SAMPLE PROBLEM 10 Calculating the concentration of a solution and its ions

Sodium carbonate, Na_2CO_3, is often used as an ingredient in washing powders. Na_2CO_3 softens hard water because it readily dissolves in water. If 10.6 g of sodium carbonate is dissolved in 500 mL of water:

a. what is the concentration of the solution?

b. what is the concentration of each ion in the solution?

THINK

WRITE

a. 1. Calculate the number of moles by applying the formula $n = \frac{m}{M}$.

$$n(Na_2CO_3) = \frac{m}{M}$$
$$= \frac{10.6}{106}$$
$$= 0.100 \text{ mol}$$

2. Calculate the concentration by applying the formula $c = \frac{n}{V}$.

$$c(Na_2CO_3) = \frac{n}{V}$$
$$= \frac{0.100}{0.500}$$
$$= 0.200 \text{ mol L}^{-1}$$

b. 1. Write the dissociation equation.

$$Na_2CO_3(s) \xrightarrow{H_2O(l)} 2Na^+(aq) + CO_3^{2-}(aq)$$

2. Look at the equation to observe how many moles of ions are produced.

One mol Na_2CO_3 produces two mol Na^+ and one mol CO_3^{2-}.

3. Since $c = \frac{n}{V}$ and the volume is constant, calculate concentration of each ion.

$$c(Na^+) = 2 \times c(Na_2CO_3)$$
$$= 2 \times 0.200$$
$$= 0.400 \text{ mol L}^{-1}$$
$$c(CO_3^{2-}) = c(Na_2CO_3)$$
$$= 0.200 \times 1$$
$$= 0.200 \text{ mol L}^{-1}$$

PRACTICE PROBLEM 10

In an $Al_2(SO_4)_3$ solution, the concentration of SO_4^{2-} is 0.050 mol L^{-1}. What is the concentration of solute?

elog-1850

EXPERIMENT 13.1 online only

Preparation of a solution of known concentration

Aim

To prepare a solution of known concentration

13.2 Activities

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13.2 Quick quiz on | **13.2 Exercise** | **13.2 Exam questions**

13.2 Exercise

1. Calculate the concentration in g L^{-1} and mg L^{-1} for each of the following solutions.
   a. 0.425 g barium nitrate in 25.0 mL of solution
   b. 26.4 mg sodium bicarbonate in 55.0 mL of solution
2. How many grams of sodium chloride would you need to prepare 475 mL of a solution with a concentration of 0.200 g L^{-1}?
3. If you dissolve 200 mg of potassium chloride in sufficient solution to give a concentration of 50.0 g L^{-1}, what volume of solution is required?
4. Calculate the concentration in %(m/m) for the following solutions.
   a. 38.4 g of NaCl in 150.0 g of solution
   b. 0.25 g benzalkonium chloride (active ingredient in an antibacterial kitchen cleaner) in 50.0 g of solution
   c. 144 mg of sugar in 200 g of solution
5. Calculate the %(m/v) of the following solutions.
   a. 50 kg of sodium bromide in 250 L of solution b. 0.025 g NaCl in 500 mL of solution
6. Ethanoic acid is dissolved in aqueous solution to make a 5.00%(v/v) solution of vinegar.
   a. What volume of ethanoic acid is in 0.750 L of vinegar?
   b. What volume of vinegar contains 25.0 mL of ethanoic acid?
7. Calculate the concentration (molarity) of copper(II) sulfate, $CuSO_4$, in hydrated form in each of the following.
   a. 2.0 L of solution containing 250 g $CuSO_4 \cdot 5H_4O$
   b. 700 mL of solution containing 750 g $CuSO_4 \cdot 5H_2O$
   c. 125 mL of solution containing 50 g $CuSO_4 \cdot 5H_2O$
8. Caustic soda, NaOH, is an ingredient in household drain cleaners. If a 250 mL bottle of cleaner contains 20 g of caustic soda, calculate the following.
   a. The concentration of the solution (in mol L^{-1}) b. The concentration of each ion in the solution
9. Calculate the concentration of an $Mg(NO_3)_2$ solution in which the concentration of nitrate ions is 0.030 M.
10. A laboratory contains a large bottle of hydrochloric acid labelled 11.6 M.
    a. Explain the significance of the 11.6 M.
    b. It is desired to use this to prepare 2 L of 1 M acid solution.
    Clearly explain, showing calculations where appropriate, how this could be done.

13.2 Exam questions

Question 1 (1 mark)

MC What volume of 0.010 M calcium hydroxide solution contains 0.20 mol of hydroxide ions?

A. 0.0020 L **B.** 0.0040 L **C.** 10 L **D.** 20 L

Question 2 (1 mark)

A 375 mL bottle of an alcoholic beverage is found to contain 26.0 mL of alcohol. Express the concentration of alcohol in %(v/v).

Question 3 (1 mark)

A household cleaner is labelled as containing NaOH at a concentration of 12.0 g L^{-1}. Express this concentration in M.

Question 4 (1 mark)

MC A particular cough medicine contains 0.125 g of ammonium chloride (molar mass = 53.5 g mol^{-1}) in each 5 mL. Which of the following concentrations is *not* equivalent to this?

A. 25 g L^{-1} **B.** 25 mg mL^{-1} **C.** 0.25%(m/v) **D.** 0.47 M

Question 5 (2 marks)

1.25 g of silver nitrate is dissolved in 150 mL. Calculate its molar concentration.

More exam questions are available in your learnON title.

13.3 Factors that influence solubility

KEY KNOWLEDGE

- The use of solubility tables to predict experimental determination of ionic compound solubility; the effect of temperature on the solubility of a given solid, liquid or gases in water
- The use of precipitation reactions to remove impurities from water

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13.3.1 Water as a solvent for ionic substances

You will recall from topic 4 that a solvent is a substance in which another substance (the solute) can be dissolved. The most common solvent used today is water. This is due not only to its widespread abundance, but also to the polarity of its molecules that you learned about in topic 2.

The process by which ionic substances dissolve in water was discussed in section 4.4.2. This is due to the interactions between ions in the solid lattice structure and the dipoles that are present in the covalent bonds within the water molecules. You will recall from topic 2 that these dipoles are regions of positive and negative charge that are *slightly* unbalanced, and that they exist due to the electronegativity difference between oxygen and hydrogen atoms.

FIGURE 13.5 Most phosphate ionic compounds are insoluble.

While water will dissolve many ionic substances, in some the strength of attraction within the lattice structure is strong enough to withstand the attempts of the water molecules to disrupt it. Such substances are, therefore, insoluble. Solubility tables such as table 4.4 (in topic 4) can be used to predict whether or not an ionic substance is soluble or insoluble.

Making use of solubility tables to identify ions in solution

Situations often arise where a quick but simple test is required to qualitatively establish whether or not certain ions are present a sample of water. In situations such as these, the use of solubility tables such as table 4.4 can prove very useful. By adding various solutions and seeing whether or not a precipitate forms (and looking at its appearance), useful information may be deduced as to the contents of a water sample.

FIGURE 13.6 Barium sulfate precipitate

For example, suppose that a sample of water is known to contain either chloride ions or sulfate ions as an impurity. If a few drops of barium nitrate solution are added to the water, a precipitate will only be observed if sulfate ions are present. This is due to insoluble barium sulfate being formed. As seen in topic 4, an equation may be written for this as follows:

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

If chloride ions are present, no precipitate will form because barium chloride is soluble.

As an alternative, lead(II) nitrate could be used. This time, both the possibilities form precipitates according to:

$$Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) \text{ and}$$

$$Pb^{2+}(aq) + 2Cl^{-}(aq) \rightarrow PbCl_2(s)$$

If the resulting sample is then heated, the precipitate will dissolve if it is $PbCl_2$, but not if it is $PbSO_4$. The ion responsible may thus be identified.

13.3.2 Using precipitation reactions to remove impurities from water

Removing hardness from water

Water can be described as 'hard' or 'soft'. Hard water is caused by certain dissolved ions such as calcium and magnesium ions, and occasionally iron(II) ions. The presence of these ions causes problems when the water is used in certain situations — for example:

- In applications where hot water is involved: examples include in pipes, boilers and domestic hot water heaters. The presence of calcium and/or magnesium ions causes the build-up of 'scale' over time. This coats surfaces and leads to decreased thermal efficiency and increased corrosion.
- When washing with soap: soap is made from long-chain carboxylic acids that are reacted with sodium hydroxide to form the soap. Stearic acid, $C_{17}H_{35}COOH$, is a typical example of such a carboxylic acid. The equation for its reaction to form soap is

$$C_{17}H_{35}COOH(s) + NaOH(aq) \rightarrow C_{17}H_{35}COONa(aq) + H_2O(l)$$

- If soap is added to hard water, the stearate ions, $C_{17}H_{35}COO^-$, from the soap can react with any calcium or magnesium ions that are present to form an insoluble product called scum. This floats on the surface and coats whatever is being washed. The equations for this process are

$$Ca^{2+}(aq) + 2C_{17}H_{35}COO^-(aq) \rightarrow Ca(C_{17}H_{35}COO)_2(s)$$
and/or
$$Mg^{2+}(aq) + 2C_{17}H_{35}COO^-(aq) \rightarrow Mg(C_{17}H_{35}COO)_2(s)$$

- This process also prevents the soap from lathering, and more and more must be used until all the calcium and magnesium ions are used and a lather can eventually form. Although modern detergents have largely alleviated this problem, in some situations soap use is still preferred.

To overcome problems such as these, hard water must be treated before being piped to the consumer. Precipitation reactions play a vital role in this process.

A common method used is the so-called lime-soda method. Lime is a common name for calcium hydroxide, $Ca(OH)_2$, which is mildly soluble (refer to table 4.4). Soda is a common name for sodium carbonate, Na_2CO_3, and, as expected of a sodium compound, it is soluble in water.

$$Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$$

The hydroxide ions increase the pH of the water and enable the formation of magnesium hydroxide precipitate.

FIGURE 13.7 Calcium and magnesium ions in water cause the build-up of scale on appliances that use hot water.

Calcium ions are removed by the formation of calcium carbonate as a precipitate.

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$$

The amounts of these chemicals added are carefully controlled to enable maximum removal of calcium and magnesium ions while also ensuring that they themselves do not become a problem if left in excess. These precipitates are then removed by filtering. Other chemicals, called coagulants, are sometimes also added to make this filtration easier.

On a domestic scale, hardness in water can be removed through the use of special filters containing ion-exchange resins. These operate by a different principle and do not involve precipitation reactions.

Removing turbidity from water

Turbidity refers to the cloudiness in a sample of water. Besides looking unattractive, it can also be the cause of odours in the water. In most large-scale supplies, turbidity is caused by very fine silt and clay particles that are too small to settle under gravity. These particles can also have bacteria and other undesirable chemicals adsorbed to their surfaces.

Aluminium sulfate, $Al_2(SO_4)_3$, also known as alum, is commonly used to remove turbidity.

Alum dissolves in water to form aluminium hydroxide. This forms as a gelatinous precipitate and, as it settles, physically traps many of the fine silt particles, thus removing them from the water.

FIGURE 13.8 Water treatment plants contain large tanks that allow precipitates and other solid material to settle.

The water is allowed to settle to remove the bulk of this precipitate and then filtered to remove any remaining solids.

The reaction for the dissolving of aluminium sulfate is interesting. After dissociating, the aluminium ions then react with water in a type of acid–base reaction. This is a slightly different type of acid–base reaction to those studied in topic 11 and is outside the scope of VCE Chemistry. The overall equation for the dissolving of aluminium sulfate and its subsequent reaction with water is as follows.

$$Al_2(SO_4)_3(s) + 6H_2O(l) \rightarrow 2Al(OH)_3(s) + 6H^+(aq) + 3SO_4^{2-}(aq)$$

Removing phosphate ions from water

Phosphate ions are undesirable in bodies of water because they can lead to eutrophication (refer to topic 10). The major sources of such pollution are the use of older style detergents and agricultural run-off.

A number of chemicals can be used to remove these ions by precipitation. Alum (see the preceding section) can also be used for this purpose, where it is able to form aluminium precipitates with various forms of phosphate ions. It also has the advantage of trapping the aluminium phosphate precipitate in the aluminium hydroxide precipitate formed, thus aiding the filtering process that follows. The relevant equations for this are as follows.

$$Al^{3+}(aq) + PO_4^{3-}(aq) \rightarrow AlPO_4(s)$$
$$Al^{3+}(aq) + HPO_4^{2-}(aq) \rightarrow AlPO_4(s) + H^+(aq)$$
$$Al^{3+}(aq) + H_2PO_4^-(aq) \rightarrow AlPO_4(s) + 2H^+(aq)$$

Treating wastewater to remove heavy metals

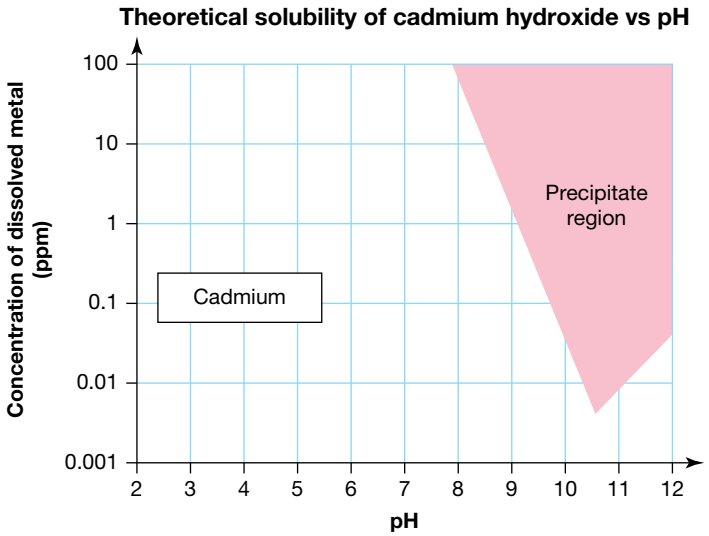
Many industrial processes produce wastewater that contains metal ions that are detrimental to health. Reducing these to an environmentally acceptable level before such water can be discharged is, therefore, necessary. In this context, the term ‘heavy metal’ is usually taken to mean ions of metals such as cadmium, mercury, silver, tin and lead, but it can also include nickel, copper, zinc and chromium.

The removal mechanism is similar to that already described for calcium ions in hard water. The offending ion is precipitated by adding sodium hydroxide to form the relevant insoluble hydroxide. For example, cadmium ions are removed according to the following.

$$Cd^{2+}(aq) + 2OH^-(aq) \rightarrow Cd(OH)_2(s)$$

Studies have shown that the optimal pH range for this process depends on the concentration of ions present. This optimal range is shown in figure 13.9 for cadmium.

FIGURE 13.9 The removal of ions by hydroxide precipitation depends on the concentration and the pH.



Once again, other chemicals are often added as coagulants to assist in the filtration process that follows to remove the precipitate. Because the use of sodium hydroxide increases the pH of the water, it is then necessary to reduce this in some way prior to ultimate discharge.

13.3.3 Water as a solvent for molecular substances

While ionic substances are all solids at room temperature, **molecular** compounds may be solids, liquids or gases at these temperatures. Molecular compounds are covalently bonded and may or may not dissolve in water. As seen in topic 2, the interactions between **covalent** molecules include **dispersion forces**, **dipole–dipole interactions** and **hydrogen bonding**.

molecular group of atoms bonded together covalently

covalent molecules which share electrons between nuclei to form the bond that holds them together in a molecule

dispersion force the bond between adjacent molecules formed by instantaneous dipoles; this weak non-directional bonding is also known as van der Waals force

dipole–dipole interactions weak bonding caused by the positive end of one dipole attracting the negative end of another dipole

hydrogen bonding the bond between a hydrogen atom covalently bonded to an atom of F, O or N and another molecule that also contains an atom of H, F, O or N

polar molecules that have an uneven distribution of electrons due to differences in electronegativity of the atoms

ionisation a process in which ions are formed

Water molecules are attracted to each other by comparatively strong hydrogen bonds. Molecules that do dissolve in water are highly **polar**, and are described as dipoles. These polar molecules will interact with polar water molecules and dissolve, either because they form ions in water or because they form hydrogen bonds with water.

Molecular compounds that ionise in water

Pure hydrogen chloride, HCl, is a gas consisting of covalently bonded H–Cl molecules, which are highly polar due to the large electronegativity difference between hydrogen and chlorine atoms. Yet, when dissolved in water, HCl undergoes almost complete ionisation to form hydrogen and chloride ions. The hydrogen ions combine with water to form hydronium ions, H_3O^+, and the aqueous solution is called hydrochloric acid. This happens because attraction of the polar water molecules increases the polarity of the H–Cl bond and the slightly negative end of the water molecule attracts the hydrogen atom in HCl, pulling it away from the hydrogen chloride molecule. This hydrogen atom now becomes covalently bonded to the water molecule. Both of these ions are hydrated and form ion–dipole interactions in a similar way to the ions described in topic 4. **Ionisation** is the process by which polar molecular compounds form ions when dissolved in water.

$$HCl(l) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

Ionisation

Highly polar molecular substances ionise in water when they dissolve.

FIGURE 13.10 HCl dissolving in water

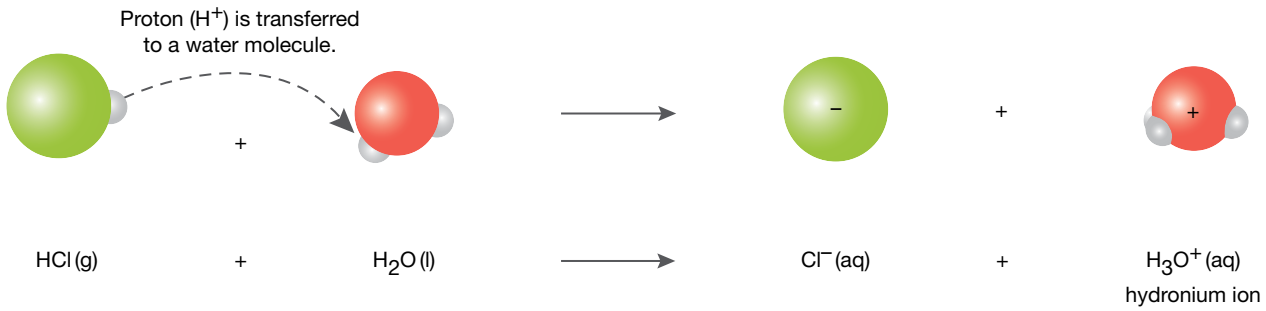
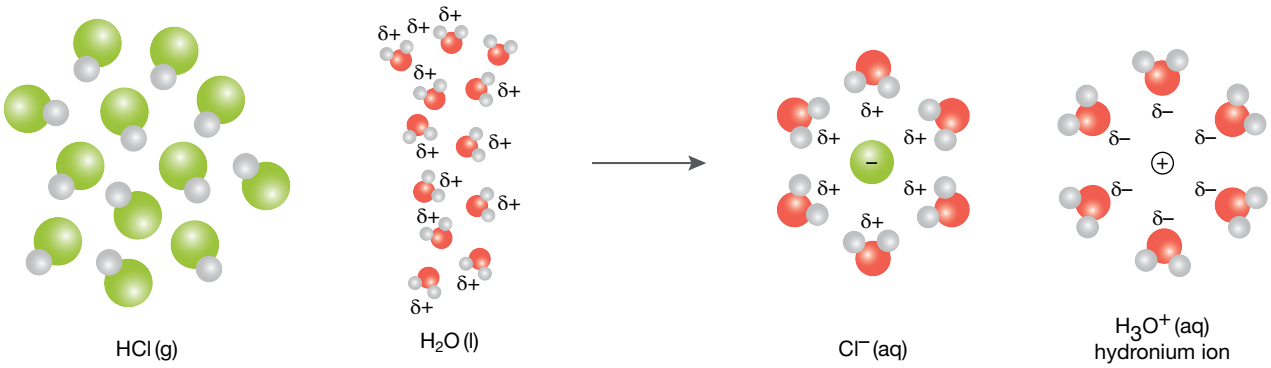


FIGURE 13.11 Hydrated ions with ion–dipole attractions



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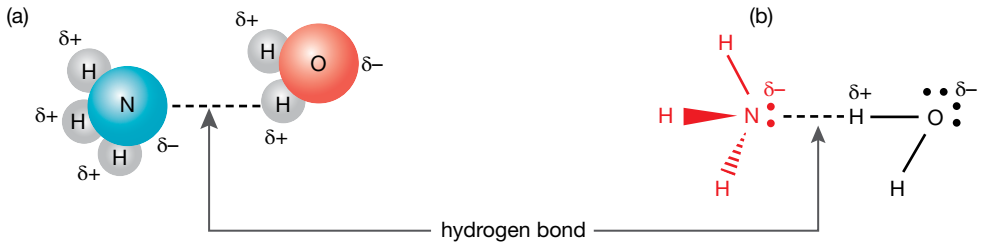
Weblink HCl dissolving in water

Molecular compounds that form hydrogen bonds with water

Other polar molecules, although they do not react with water as outlined in the previous section, may still form hydrogen bonds with water and are soluble. These molecules can be identified because they have a hydrogen atom bonded to one of fluorine, oxygen or nitrogen. Consider ammonia, NH_3: the covalent bonds between the nitrogen atom and hydrogen atoms in the molecule are highly polar due to the electronegativity of nitrogen being considerably greater than that of hydrogen. Water can form a hydrogen bond with nitrogen in ammonia, $NH_3(g) \rightarrow NH_3(aq)$, as shown in figure 13.12.

FIGURE 13.12 Hydrogen bond between ammonia and water molecules shown as (a) a space-filling model and (b) Lewis (electron dot) structure

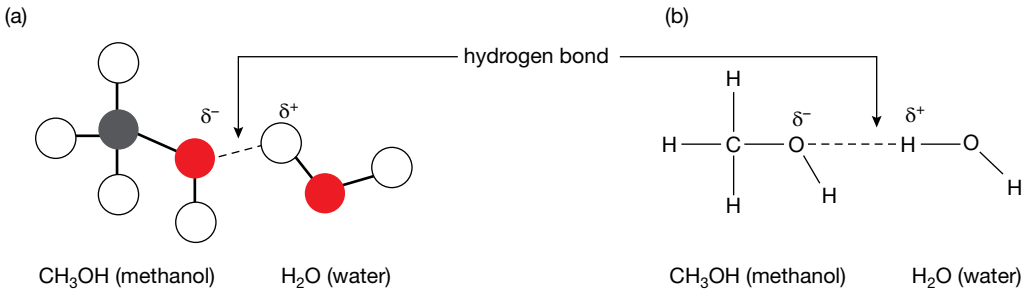


Other molecules that are soluble in water include small alcohol molecules. In the simplest alcohol, methanol, CH_3OH, the covalent bond between oxygen and hydrogen in the hydroxy (–OH) group is highly polar. Water can form a hydrogen bond with methanol, as shown in figure 13.13. Hydrogen bonding does not occur between the hydrogen atoms that are bonded to the carbon atoms because the bonds are not very polar — this is due to a relatively small electronegativity difference between carbon and hydrogen.

$$CH_3OH(l) \xrightarrow{H_2O(l)} CH_3OH(aq)$$

Interestingly, when water and alcohol molecules interact, they take up less volume due to the rearrangement of the molecules. The volume of the combined liquids is less than the sum of the individual volumes.

FIGURE 13.13 Hydrogen bonding between methanol and water molecules shown as (a) a ball-and-stick model and (b) structural formula



Molecular compounds with hydrogen bonds

Molecular compounds that contain hydrogen bonded to fluorine, oxygen or nitrogen dissolve by forming hydrogen bonds with water.

Other organic compounds that have a high proportion of very electronegative atoms such as N or O or, preferably, a number of N–H or O–H bonds will also form hydrogen bonds with water and, therefore, be soluble. For example, simple sugars have many hydroxy groups (–OH) and are, therefore, soluble.

13.3.4 Factors that affect solubility

The solubility of a substance refers to the amount that can dissolve in a given volume of solvent. In topic 4, a qualitative way of describing this was introduced. Substances were described as either insoluble, sparingly soluble, moderately soluble or soluble. Such descriptions can be placed on a more quantitative basis by *measuring* the amount of solute that dissolves. We have already seen that the nature of the solute–solvent combination plays an important role in this process. Whether or not a substance dissolves in water ultimately comes down to how the solute particles interact with the polar water molecules. Where the solute particles are ions or polar covalent molecules, dissolving may take place. However, if the attraction between the ions in an ionic solute is too strong, or if the molecules of a covalent solute are non-polar, dissolving will not be possible.

Why substances dissolve

Whether or not a substance dissolves depends on the interactions between the solute and solvent particles.

The effect of temperature

The temperature of the solvent can have a significant effect on the amount of solute that can be dissolved. As a general rule, for solid solutes, the higher the temperature, the greater the amount of solute that can be dissolved. In other words, *solubility increases with temperature*. For some substances this difference is only slight. Others, however, show a big difference. Lead (II) chloride, $PbCl_2$, for example, is often described as insoluble in cold water but soluble in hot water.

The model already presented for the dissolving process explains this trend. The higher the temperature, the more energy (on average) the solvent particles possess. This makes it easier for the solid solute structure to be disrupted but, also importantly, makes it easier to keep the particles apart once they have dissolved. However, for all solvents, a point will come where so many particles are in solution and moving around that the solvent particles will not have enough energy to keep any further particles apart. This point then represents a stage at which a maximum amount of dissolved solute has been reached. It then follows that, the more energy the solvent particles have, the more solute is able to be dissolved before this point is reached. This is discussed further later in this topic.

Agitation — a misconception

We stir our tea or coffee to help dissolve sugar. A common misconception that follows from this is that stirring (agitation) causes more sugar to dissolve. This is *not* true. Stirring merely increases the rate of dissolving. It does not increase the *amount* of sugar that can be dissolved.

Another factor that increases the rate of dissolving but not the amount is the *surface area* of the solute. Finely ground sugar will dissolve faster than sugar crystals but, for a certain temperature, the amount that can eventually be dissolved will remain the same.

FIGURE 13.14 Surface area and temperature affect the rate at which sugar dissolves.

13.3.5 Solubility of liquids and gases

Substances that are liquids and gases at normal temperature are all covalent molecular substances. The processes by which these dissolve in water were discussed in section 13.3.3.

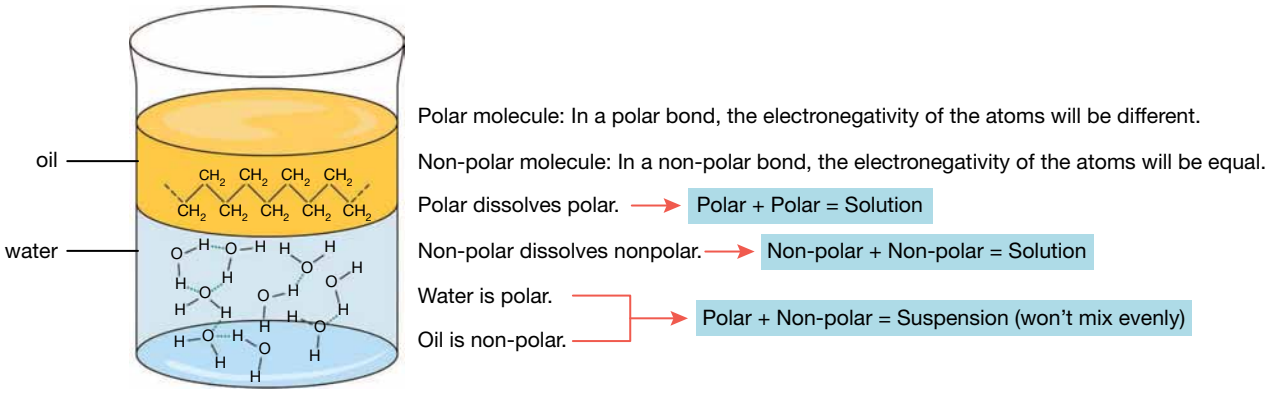
- Liquids that dissolve in water (such as ethanol) are polar molecules and are said to be *miscible*.
- Liquids that do not dissolve (such as oil in water) are non-polar molecules and will form layers if left to settle. Such liquids are termed *immiscible*.

FIGURE 13.15 Oil is immiscible in water.

As described in topic 5, chemists use the like dissolves like rule, where polar solutes tend to dissolve in polar solvents, and non-polar solutes tend to dissolve in non-polar solvents.

With gases, the situation is the same. That is, polar gases such as ammonia are more soluble in water than non-polar gases such as oxygen. It should also be noted that gases usually behave in the opposite way to solids as far as temperature is concerned — that is, as the temperature increases, the solubility of a gas decreases. With gases, however, another factor affects solubility. This is pressure. The solubility of a gas in water increases with increasing pressure. This is seen in reverse when the top is removed from a bottle of soft drink. The resulting decrease in pressure means that the dissolved gas (carbon dioxide) can no longer stay dissolved to the same extent. Bubbles, therefore, form as this excess gas comes out of solution.

FIGURE 13.16 Oil molecules are non-polar and will not dissolve in water.



The decreasing solubility of gases with increasing temperature has the potential to become a significant problem for the Earth. As atmospheric temperatures increase due to higher levels of greenhouse gases (such as carbon dioxide), the upper layers of the oceans also increase in temperature. This leads to reduced amounts of gases such as oxygen and carbon dioxide being dissolved in the water and significant amounts of these gases being returned to the atmosphere. This would further increase the level of carbon dioxide in the atmosphere and would seriously hamper respiration and photosynthesis for marine organisms due to the lower amounts of dissolved oxygen and carbon dioxide that remain. The solubility of gases is described in more detail in topic 15.

EXTENSION: Using greener solvents

Although water is a very useful solvent, in many situations in the laboratory and in industry it is not suitable. The most obvious situation is where a non-polar solute is desired to be dissolved. We are all familiar with petrol and may know that it is a good solvent for dissolving grease and oil. Petrol is primarily a mixture of alkanes (see topic 7) so it is not surprising that liquid alkanes such as hexane and cyclohexane are sometimes chosen as solvents in this situation. Sometimes, a degree of polarity is required but not as great as in water alone. In these cases, other solvents may be chosen. Sometimes solvents may even be blended to achieve a desired outcome.

FIGURE 13.17 Halogens (Cl_2, Br_2, I_2) dissolved in non-polar cyclohexane (top of each test tube), demonstrating immiscibility with water (bottom of each test tube)

Common alternative solvents to water include, but are by no means limited to:

- tetrachloroethene, C_2Cl_4
- methanol, CH_3OH
- ethanol, CH_3CH_2OH
- hexane and cyclohexane, C_6H_{14} and C_6H_{12}
- propan-2-one, also known as acetone, $(CH_3)_2CO$
- ethyl ethanoate, also known as ethyl acetate, $CH_3COOCH_2CH_3$
- methyl benzene, also known as toluene, $C_6H_5CH_3$
- xylene, a hydrocarbon mixture consisting mainly of isomers of dimethylbenzene, C_8H_{10}.

With the advent of green chemistry has come a push to develop more user and environmentally friendly solvents because many of the alternative solvents used have health and environmental drawbacks. Examples in daily life include solvent use in dry cleaning.

To learn more about the use of green chemistry solvents, download the digital document.

on Resources

Digital document EXTENSION: Using greener solvents (doc-37911)

13.3.6 Summary of factors that influence solubility

When everything is considered from this topic and from topic 4, the factors that influence solubility can be summarised as follows:

- The nature of the particles in the solute: these can be either ions, polar molecules or non-polar molecules.
- The nature of the solvent: if the solvent consists of polar molecules (for example, water) it is more likely to be able to dissolve ionic solids or substances that contain polar molecules themselves. If the solvent molecules are non-polar, the solvent will be more likely to dissolve non-polar substances. Chemists refer to this as the *like dissolves like* rule.

- The attraction between the solute particles: if this attraction is too strong, the solvent may not be able to disrupt it and affect dissolution.
- The temperature: in solids, as temperature increases, the solvent particles move faster and have more energy. This makes it easier for them to overcome the attractive forces between the solute particles.
- For gases, both the pressure and temperature: the higher the pressure, the more gas will dissolve. The higher the temperature, the less gas dissolves.

13.3 Activities

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13.3 Quick quiz on | **13.3 Exercise** | **13.3 Exam questions**

13.3 Exercise

1. Explain why temperature is one of the main factors that affects the solubility of a solute in water.
2. Why are substances such as oil insoluble in water?
3. Describe the solubility of barium sulfate, $BaSO_4$, at the following temperatures, using accepted terminology. Explain your reasoning for each part.
   a. 20 °C
   b. 50 °C
   c. 75 °C
4. Using table 4.4 from topic 4, describe how a sample of lead(II) sulfate that is contaminated with a small amount of lead(II) chloride may be purified.
5. Name the interactions that occur when each of the following types of compounds dissolve in water.
   a. Ionic compound
   b. Polar molecular compound such as HCl
   c. Molecular compound that contains H bonded to F, O or N

13.3 Exam questions

Question 1 (6 marks)

Answer the following.

a. Draw a labelled diagram showing potassium chloride before and after being dissolved in water. **(3 marks)**
b. Label the types of interactions present in water and potassium chloride both before and after dissolving. **(3 marks)**

Question 2 (1 mark)

MC Which of the following lists contains substances that can all hydrogen bond with water?

A. Methane, methanol, methanoic acid
B. HF, HCl, HBr
C. NH_3, NaCl, $NaNO_3$
D. CH_3OH, C_2H_5OH, C_3H_7OH

Question 3 (1 mark)

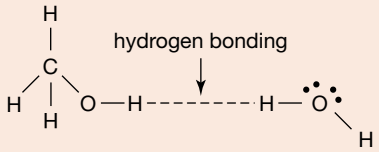
MC Which of the following would not display ion–dipole interactions in water?

A. $C_6H_{12}O_6$ B. $Zn(NO_3)_2$ C. LiCl D. $(NH_4)_2SO_4$

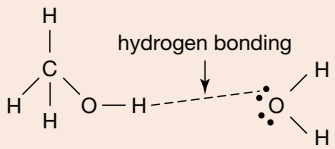
Question 4 (1 mark)

MC Which of the following diagrams correctly shows the bonding between methanol and water?

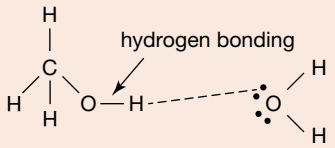
A.



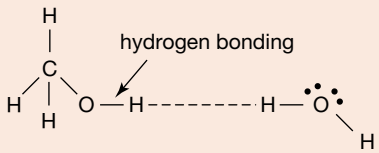
B.



C.



D.



Question 5 (1 mark)

MC Which of the following substances is the least soluble in water?

A. Propan-1-ol
B. Propan-2-ol
C. Propane
D. Propanoic acid

More exam questions are available in your learnON title.

13.4 Solubility graphs

KEY KNOWLEDGE

- The use of solubility graphs to predict experimental determination of ionic compound solubility

Source: Adapted from VCE Chemistry Study Design (2023–2027) extracts © VCAA; reproduced by permission.

13.4.1 Solution terminology

What do soft drinks, coffee, tea, soup and cordial have in common? They are all mixtures and they all contain water. Some of them are heterogeneous mixtures (known simply as 'mixtures') and others are homogeneous mixtures called solutions. For example, when sodium chloride is dissolved in water, sodium chloride is the solute and water is the solvent. The end result is called the solution.

The solvent properties of water make it an excellent medium for chemical reactions. Water samples containing dissolved substances are called **aqueous solutions**.

aqueous solutions mixtures in which substances are dissolved in water

Most salts are soluble in water to some extent (see examples of the solubility of various salts in table 13.1). Solubility is influenced by the temperature of the solution. Most salts become more soluble at high temperatures, but a few become less soluble

as the temperature increases. Salts with a solubility greater than 0.1 M of water are classed as soluble and salts with a solubility less than 0.1 M of water are classed as insoluble. M refers to a unit of concentration and was introduced in section 13.2.4.

TABLE 13.1 The solubility of various salts

Salt	Formula	Solubility (g per 100 g water at 20 °C)
Barium chloride	$BaCl_2$	36.0
Barium sulfate	$BaSO_4$	0.000 24
Calcium chloride	$CaCl_2$	74.0
Calcium sulfate	$CaSO_4$	0.21
Copper(II) sulfate	$CuSO_4$	20.5
Copper(II) sulfide	CuS	0.000 03
Lead(II) sulfate	$PbSO_4$	0.004
Potassium chlorate	$KClO_3$	7.3
Potassium nitrite	KNO_2	300.0
Silver chloride	AgCl	0.000 000 1
Silver nitrate	$AgNO_3$	217.0
Sodium chloride	NaCl	36.0
Sodium nitrate	$NaNO_3$	87.0

- A **saturated solution** contains the maximum amount of solute for the volume of solution at a particular temperature.
- An **unsaturated solution** contains less than the maximum amount of solute that could be added to the solvent.
- A **supersaturated solution** is an unstable solution that can be prepared by slowly cooling a saturated solution. When this is done, the solution contains more solute than it should at that temperature. The addition of a small crystal seed or even some dust causes the excess solute to crystallise. Increasing the temperature of a supersaturated solution will increase the solubility, leading to more of the solid to be dissolved and to obtain a saturated solution. Alternatively, more water could be added, so more solid can dissolve.

13.4.2 Displaying solubility data

Depending on the amount of solute present, solutions may be described as saturated, unsaturated or supersaturated.

The solubility of a substance in water depends on temperature. Most solids are more soluble as temperature increases. The solubility of soluble substances can be easily measured by experiment. A convenient unit used for this purpose is g/100 g of water. When this is done across a range of temperatures, the result may be drawn as a graph of solubility versus temperature. Such graphs are called **solubility curves**. These find extensive use in medical, pharmaceutical and industrial applications where compounds need to be isolated and purified.

In this context, 100 mL of water is the same as 100 g water. Several examples of solubility curves are shown in figure 13.18. Most substances in these examples increase their solubility as the water temperature increases, except for some gases.

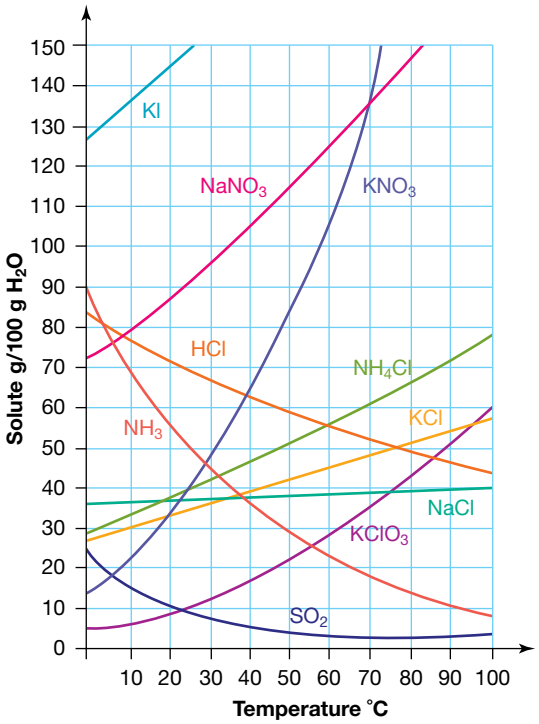
saturated solution solution in which the maximum amount of solute has been dissolved

unsaturated solution solution containing less than the maximum amount of solute that can dissolve

supersaturated solution solution in which a greater amount of solute is dissolved at a particular temperature than is predicted by a solubility curve temperature

solubility curves graphs of grams of a substance that will dissolve in 100 g of solvent at a particular temperature

FIGURE 13.18 Solubility curves for several substances



Important characteristics for interpreting solubility curves include the following:

- All the points *on* the solubility curve represent the maximum amount of solute that can be dissolved in 100 g water at the specific temperature. Therefore, any point on the curve represents a *saturated solution*.
- Points *above* the curve represent the amount of substance in a *supersaturated solution* (see section 13.4.1).
- Points *below* the curve represent the amount of substance completely dissolved, in an *unsaturated solution*.
- Values determined from solubility curves are only considered reliable if they fall within the range of the data used to create the curve. Extrapolating values outside this range is unreliable, because no evidence is provided that the trend will continue in the same way.

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SAMPLE PROBLEM 11 Investigating solubility curves

The solubility of a white crystalline substance known to be a type of sugar was determined over a range of temperatures. The results are shown in the following table.

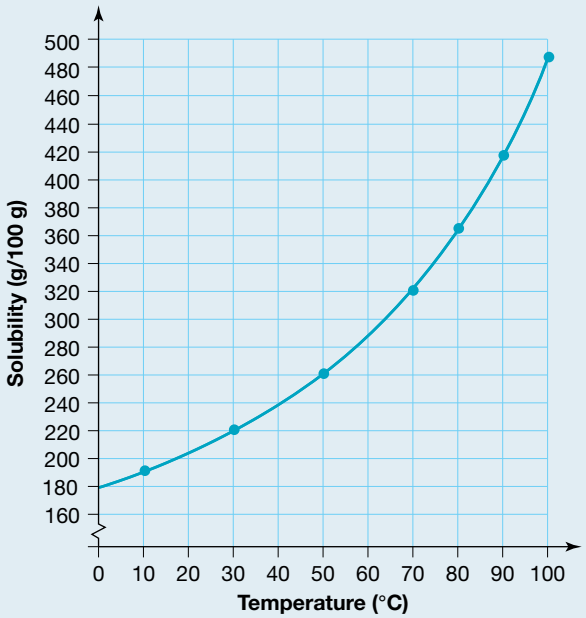
Temperature (°C)	10	30	50	70	80	90
Solubility (g/100 g)	190	220	260	320	362	415

a. Plot the solubility curve for this sugar.
b. What mass of sugar dissolves in 100 g of water at 75 °C?
c. What mass of sugar dissolves in 250 g of water at 20 °C?
d. What mass of water is required to dissolve 110 g of sugar at 50 °C?

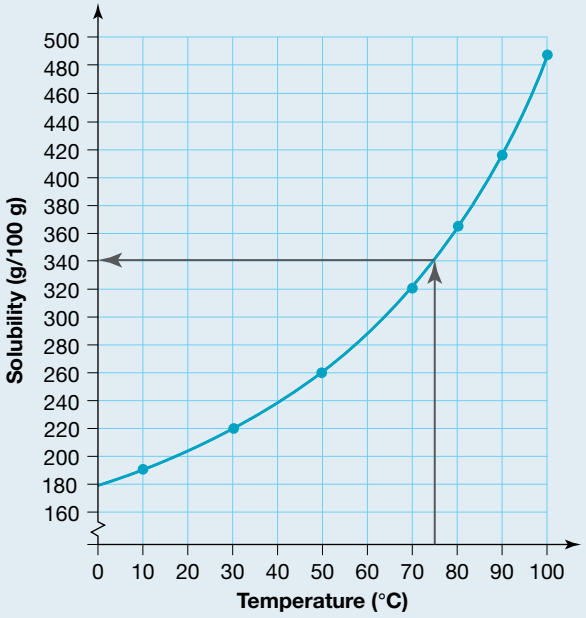
THINK

a. Read values from the table and create a graph using appropriate scales, with the *x*-axis showing Temperature °C units and the *y*-axis showing g/100 g units. Plot the graph.

WRITE

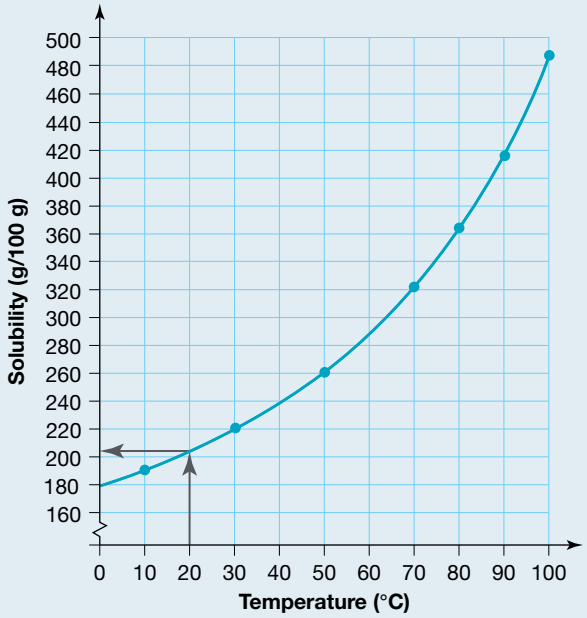


b. Starting from 75 °C on the *x*-axis, draw a line to the curve, and then from the curve to the *y*-axis, to determine mass of sugar dissolved in 100 g water. Read the answer off the graph.



At 75 °C, solubility = 340 g/100 g (from graph).
Therefore at 75 °C, 340 g sugar dissolves in 100 g of water.

c. 1. Starting from 20 °C on the *x*-axis, draw a line to the curve, and then from the curve to the *y*-axis, to determine mass of sugar dissolved in 100 g water. Read the answer off the graph.



At 20 °C, solubility = 204 g/100 g.

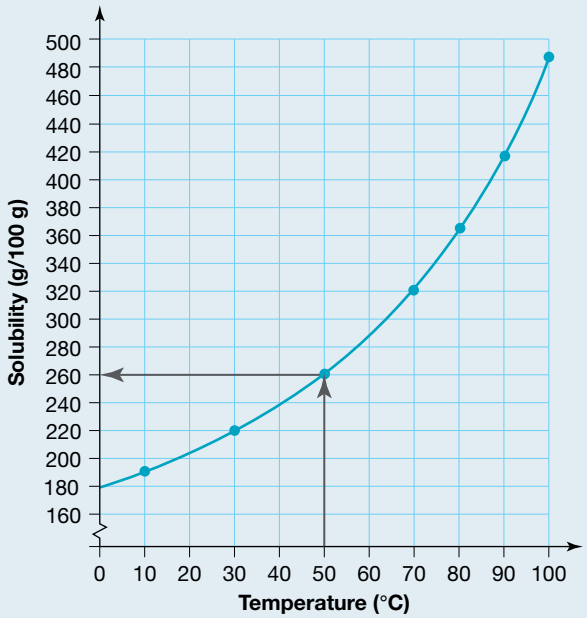
2. Determine the amount of sugar in 250 g of water.

For 250 g water:

$$204 \times \frac{250}{100} = 510 \text{ g}$$

Therefore, at 20 °C, 510 g sugar dissolves in 250 g water.

d. 1. Starting from 50 °C on the x-axis, draw a line to the curve, and then from the curve to the y-axis, to determine mass of sugar that can be dissolved in 100 g water.



2. The actual mass of sugar required to be dissolved (110 g) at 50 °C is less than the maximum that can be dissolved (260 g). To determine how much water is required to dissolve 110 g sugar at 50 °C, ratios can be used.

At 50 °C, solubility = 260 g/100 g.

Therefore, $100 \times \frac{110}{260} = 42.3$ g of water is required at this temperature.

PRACTICE PROBLEM 11

The solubility of NaCl was determined over a range of temperatures. The results are shown in the following table.

Temperature (°C)	0	10	20	30	40	50
Solubility (g/100 g)	35.7	35.8	36.0	36.3	36.6	37.0

a. Plot the solubility curve for NaCl.
b. What mass of NaCl dissolves in 100 g water at 35 °C?
c. What mass of NaCl dissolves in 150 g water at 10 °C?

tlvd-3485

SAMPLE PROBLEM 12 Obtaining unsaturated solutions

Discuss two analytical methods you can use under standard laboratory conditions (SLC) to obtain an unsaturated solution of $NaNO_3$ from 1.53 mols of $NaNO_3$/100 g (use figure 13.18).

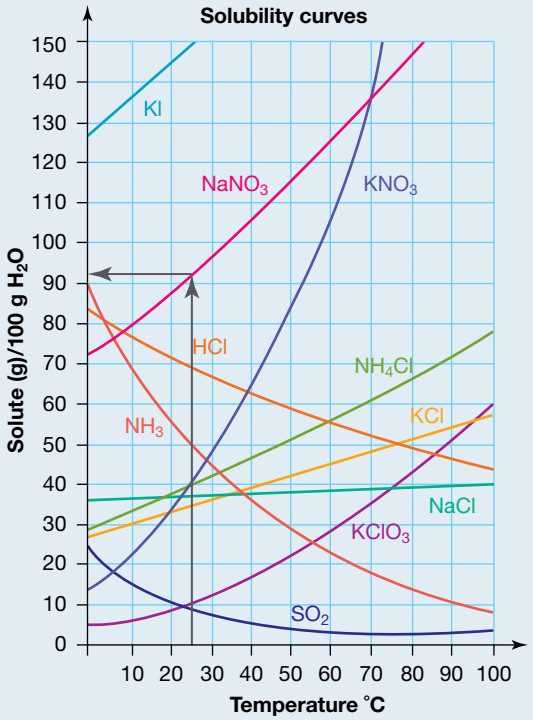
THINK

1. As solubility curves are shown in g/100 g, the mass of $NaNO_3$ in the 1.53 mols of $NaNO_3$ must be determined.
   Recall the formula $n = \frac{m}{M}$ and calculate the mass of $NaNO_3$ per 100 g H_2O.

2. Use figure 13.18 to determine solubility of $NaNO_3$ at 25 °C (laboratory conditions) and identify the state of the solution.
   From 25 °C on the *x*-axis, draw a line until it intersects with the solubility curve for $NaNO_3$. From that point draw a horizontal line to intersect with the *y*-axis. This will determine the amount of grams solute per 100 g H_2O.

WRITE

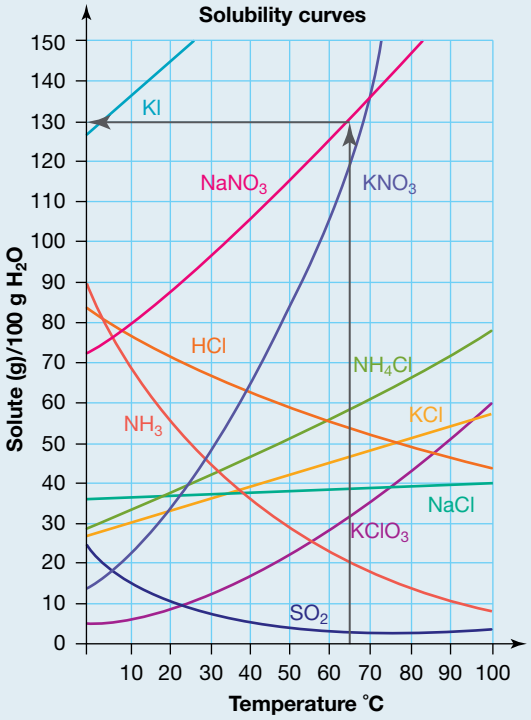
$$n = \frac{m}{M}$$
$$m = n \times M$$
$$= 1.53 \times (23 + 14 + 3 \times 16)$$
$$= 1.53 \times 85$$
$$= 130\text{g } NaNO_3/100\text{g } H_2O$$



3. Identify the state of the solution.

At 25 °C, solubility of $NaNO_3$ is around 93 g $NaNO_3$/100 g H_2O.
Therefore, a sample of 130 g $NaNO_3$/100 g H_2O will form a supersaturated solution.

4. Answer the question. Since the solubility graphs represent mass of solute with per 100 g H_2O versus temperature, think of changes to either the temperature or the volume of water to obtain an unsaturated solution. (In this case, the concentration was not required to be kept constant.)
   a. Option 1: consider the effect of a change in temperature on the solubility curve. To obtain an unsaturated solution of $NaNO_3$, the temperature would need to be increased to at least 65 °C so all $NaNO_3$ can dissolve.
   b. Option 2: consider the effect of a change of water volume while maintaining temperature. Determine how much water is required to be added to the sample to obtain an unsaturated solution.



Option 1: The $NaNO_3$ sample would need to be heated to at least 65 °C.
Option 2: Since solubility of $NaNO_3$ at 25 °C is 93 g/100 g H_2O, 130 g $NaNO_3$ would need:
$130 \times \frac{100}{93} = 139.78 = 140$ g water to become completely dissolved.
So, at least 40 g of H_2O must be added to the supersaturated solution.

PRACTICE PROBLEM 12

Using figure 13.18 and working at SLC, discuss a method to obtain an unsaturated solution of KCl from 0.700 mols of KCl/100 g.

13.4.3 Crystallisation

Crystallisation is a technique that is used to isolate substances and is based on the difference in solubility between two temperatures. As a hot solution is cooled, it eventually reaches a temperature at which it is saturated. Further cooling results in it becoming supersaturated — an unstable situation. Crystals then precipitate out of the solution and, the more it is cooled, the more crystals are formed. These crystals can then be collected easily by filtration. This technique can also be used to purify substances. Solubility curves are an important predictive tool for both these processes.

Sample problem 13 illustrates this technique.

Evaporative crystallisation

An example of natural crystallisation through evaporation takes place around hypersaline seas situated in a dry and warm climate. Hypersaline seas contain salt concentrations that are much greater than normal seawater. The Dead Sea is known to display salt bodies growing from the water during certain periods of the year (see figure 13.19 and the opening image for this topic). The salinity of the Dead Sea is 33.7 g/100 g water.

FIGURE 13.19 Salt formation at the Dead Sea

As seen in table 13.2, solubility of NaCl is only minimally affected by temperature increase. Climate characteristics around the Dead Sea (such as average temperatures between 32 °C and 40 °C for half of the year, very low humidity and a low 50 mm annual rainfall) lead to water evaporation to a point where crystallisation occurs. The crystals dissolve as soon as the water is replenished during the wetter months.

TABLE 13.2 Solubility of $AgNO_3$ and NaCl salts at various temperatures

$AgNO_3$		NaCl	
Temperature (°C)	Solubility (g of salt per 100 g of water)	Temperature (°C)	Solubility (g of salt per 100 g of water)
0	122	0	35.7
10	170	10	35.8
20	222	20	36.0
30	300	30	36.3
40	376	40	36.6
50	455	50	37.0

Reduced-pressure crystallisation

Crystallisation is used extensively in industry. However, because heating, especially on a large scale, is a major cost, a modification of the process described previously is often employed. Also, some substances, especially pharmaceuticals, can be damaged if the temperature is too high. If pressure is reduced, water boils at a lower temperature. It also evaporates faster under such conditions. By reducing the pressure, the water may, therefore, be 'sucked' out of solution, making the remaining solution more and more concentrated. Once the solution becomes saturated, removing further water from the solution results in the formation of crystals. Using reduced pressure to achieve crystallisation, rather than heating and then cooling, is more cost effective in some situations.

EXTENSION: Purifying by recrystallisation

Often when substances are produced, they are contaminated by impurities that remain from the manufacturing process. It then becomes necessary to remove these impurities before the substance can be further used or on-sold. Purification by recrystallisation is a technique that is commonly used to achieve this. This technique depends upon:

- the fact that the compound and the impurity often have differences in solubility
- the fact that temperature affects solubility.

The technique involves dissolving the impure product in the *minimum amount* of hot solvent and then cooling it to a lower temperature.

The higher temperature is chosen so that the minimum amount of water can be used to affect dissolution of the impure product. The lower temperature is chosen so that crystals of the desired product can form and the impurity stays in solution (and, therefore, does not form crystals itself).

To see an example of how this works, download the digital document.

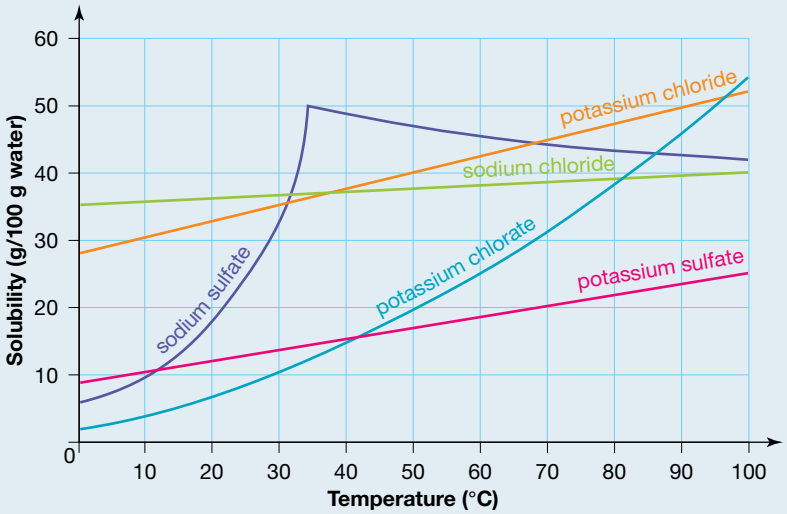
on Resources

Digital document EXTENSION: Purifying by recrystallisation (doc-37912)

tlvd-3486

SAMPLE PROBLEM 13 Determining the mass of crystals formed

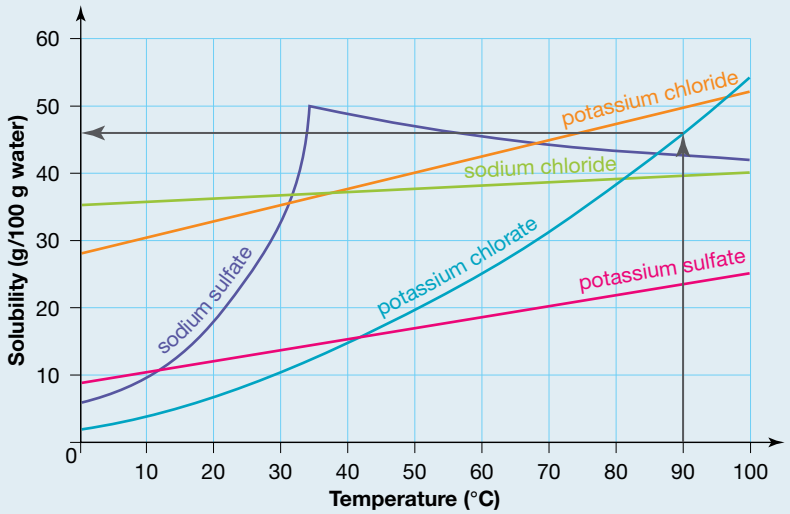
A solution of potassium chlorate, $KClO_3$, contains 40 g dissolved in 100 g of water at 90 °C. It is cooled to 20 °C. What mass of crystals form? Use the following figure to determine your answer.



THINK

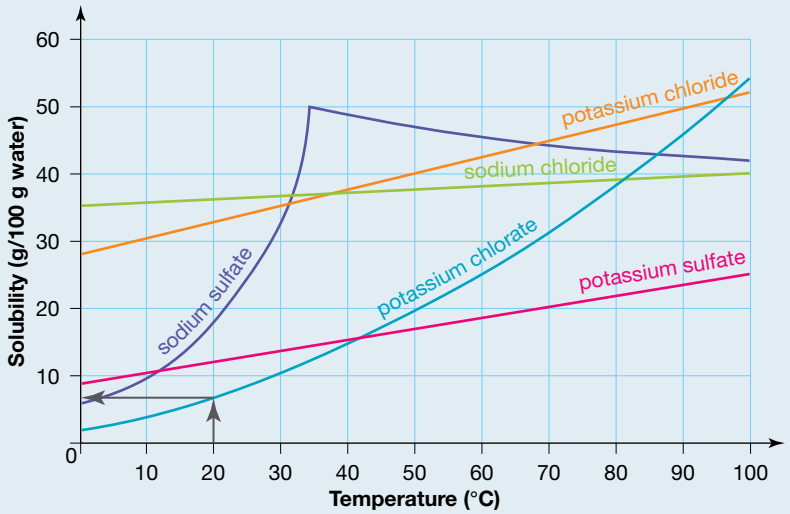
1. Identify the curve for potassium chlorate. From 90 °C on the *x*-axis, draw a vertical line to intersect the curve. Then draw a horizontal line to the *y*-axis to determine the solubility of $KClO_3$ per 100 g H_2O. Identify the state of solution.

WRITE



At 90 °C, solubility equals 46 g/100 g. Therefore, all the 40 g of $KClO_3$ is dissolved, and the solution is unsaturated.

2. Repeat the same steps for 20 °C and determine the state of the solution.



At 20 °C, solubility equals 6 g/100 g.
Therefore, only 6 g can remain dissolved, and the solution is supersaturated.

3. Answer the question. Recall the mass of $KClO_3$ at the start and calculate the difference with maximum mass that can be dissolved at 20 °C.

The mass of crystals formed
$m(KClO_3) = 40 - 6$
$= 34\text{g}.$

PRACTICE PROBLEM 13

A mixture is known to contain 46 g of potassium chloride and 12 g of potassium sulfate. The mixture is dissolved in 100 g of water at 80 °C. Use the solubility curve in sample problem 13 to complete the following.

a. Do both these salts dissolve at this temperature?
b. The solution is cooled to 20 °C. Describe what you would expect to happen.
c. If you wish to obtain a pure sample of potassium chloride, why would it be inadvisable to cool the solution below 20 °C?

elog-1851

tlvd-0639

EXPERIMENT 13.2 online only

Determine a solubility curve

Aim

To obtain the solubility curve for potassium chloride, KCl

CASE STUDY: Detoxifying cycad seeds and First Nations Australian knowledge

Aboriginal and Torres Strait Islander communities have used cycad plants as a food source for a long time. Cycad plants are suited to this purpose because they are rich in carbohydrates, are drought resistant and tolerate poor quality soils, which are found in many parts of Australia. However, cycad seeds also contain a powerful toxin called cycasin. This can cause serious damage to the nervous system as well as cancer and even death. Treating the seeds to remove this toxin is, therefore, necessary before they are consumed. Aboriginal and Torres Strait Islander communities have developed various processes to remove cycasin from cycad seeds, making the seeds edible.

FIGURE 13.20 The seeds of the cycad plant are a rich food source but must be treated before they are consumed.

Two of the ways that this is done rely on solubility differences between the seed nutrients and the cycasin. Cycasin is water soluble, whereas the bulk of the seed containing the nutrients is not. Water can, therefore, be used to dissolve out the toxin. However, under normal circumstances this is a very slow process and steps must be taken to speed it up. This can be achieved by roasting and then pounding the seeds to break them up. The contact area between the cycasin molecules and the water molecules is thus increased and the attractions between them allow the cycasin to be dissolved more quickly. Running water is also used where possible to make the process more effective. If the seeds are left in still water for long enough, a saturated cycasin solution will be formed and no further cycasin will be able to dissolve. If the water is continually replaced (as in running water), the cycasin can still dissolve but a saturated solution will never be formed.

This process is thus an illustration of how solubility differences can be used to purify a substance (the seeds) and how increasing the surface area can assist the dissolving process. The process is further enhanced by avoiding the formation of saturated solutions.

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13.4 Quick quiz on | 13.4 Exercise | 13.4 Exam questions

13.4 Exercise

1. Using the solubility values for sugar provided in the table, determine the following.

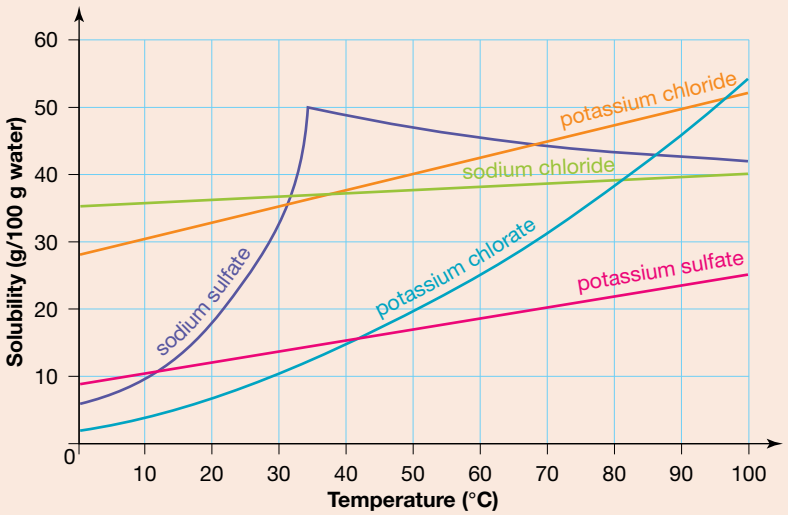
Temperature (°C)	10	30	50	70	80	90
Solubility (g/100 g)	190	220	260	320	362	415

   a. What type of solution will 260 g of sugar form at 55 °C?
   b. What mass of sugar dissolves in 150 g water at 50 °C?
   c. What mass of water is required to dissolve 245 g of sugar at 20 °C?
   d. What type of solution would form from 250 g of sugar/100 g water at 20 °C?
2. Use the information in table 13.2 to draw solubility curves for silver nitrate, $AgNO_3$, and sodium chloride, NaCl, on the same set of axes.

Use figure 13.18 to answer questions 3 to 6.

3. 60 mL of a saturated solution of NH_4Cl contains 25 g of solute. What temperature is it at?
4. What mass of water would be needed to completely dissolve 300 g of $NaNO_3$ at 60 °C?
5. At what temperature is the solubility of HCl 27 g/50 g water?
6. How many mols of KNO_3 could dissolve in 40 mL of water at 35 °C?

Use the following solubility curves to answer questions 7 to 10.

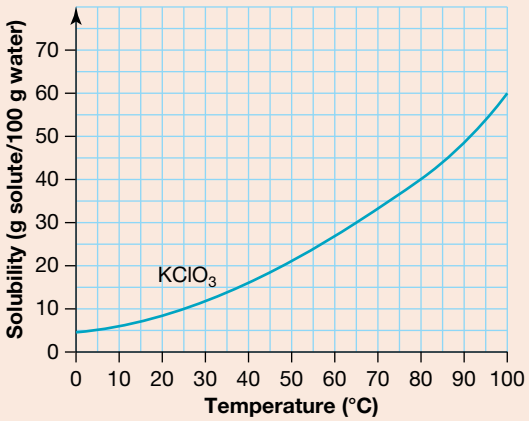


7. Determine the amount of potassium chlorate that dissolves in 100 g of water at 35 °C.
8. What mass of potassium chloride dissolves in 50 g of water at 80 °C?
9. Which salt is most soluble at 40 °C?

10. Daisy and her partner Liu Sin have just completed an experiment involving the solubility of potassium chlorate, $KClO_3$.
In this experiment, they added 31 g of potassium chlorate to 150 mL of water at 90 °C.
a Did all the potassium chlorate dissolve? Explain.
b The solution formed was heated to 95 °C and then cooled to 20 °C.
At what temperature did it become saturated?
c Once the solution reached 20 °C, it was immediately filtered and the residue dried.
What mass of potassium chlorate crystals would have been obtained?
d The filtrate was then gently evaporated to dryness.
Calculate the mass of potassium chlorate that was obtained in this step.
e Calculate the moles of potassium chlorate used in this experiment.

13.4 Exam questions

Use the following graph to answer questions 1 *and* 2.



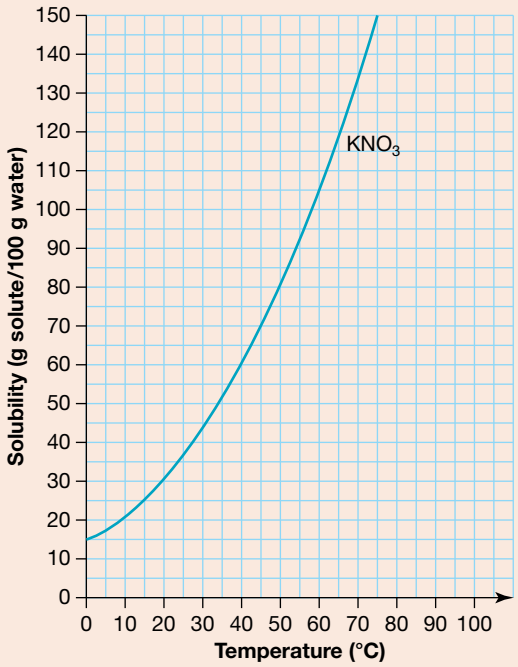
Question 1 (1 mark)

Use the solubility curve to find the solubility of potassium chlorate, $KClO_3$, at 50 °C.

Question 2 (1 mark)

MC If 35 g of $KClO_3$ is added to a beaker containing 50 g of water and then heated, with stirring, to 65 °C, what would happen?
A. All the crystals would dissolve and the solution would be unsaturated.
B. All the crystals would dissolve and the solution would be saturated.
C. 5 g of crystals would remain undissolved.
D. 20 g of crystals would remain undissolved.

Use the following graph to answer questions 3, 4 *and* 5.



Question 3 (1 mark)

MC If 100 g of KNO_3 is stirred into 250 g of water, at what minimum temperature would all the solute dissolve?

A. 27 °C
B. 58 °C
C. 145 °C
D. 150 °C

Question 4 (1 mark)

A saturated solution of KNO_3 was prepared in 200 g of water at 60 °C. If the solution was cooled to 25 °C, what mass of crystals would be produced?

Question 5 (2 marks)

a. Calculate the minimum volume of water at 60 °C that is required to dissolve 9.123 g of potassium nitrate. **(1 mark)**

b. If the solution formed in part **a** is cooled to 15 °C, calculate the mass of potassium nitrate that will crystallise. **(1 mark)**

More exam questions are available in your learnON title.

13.5 Review

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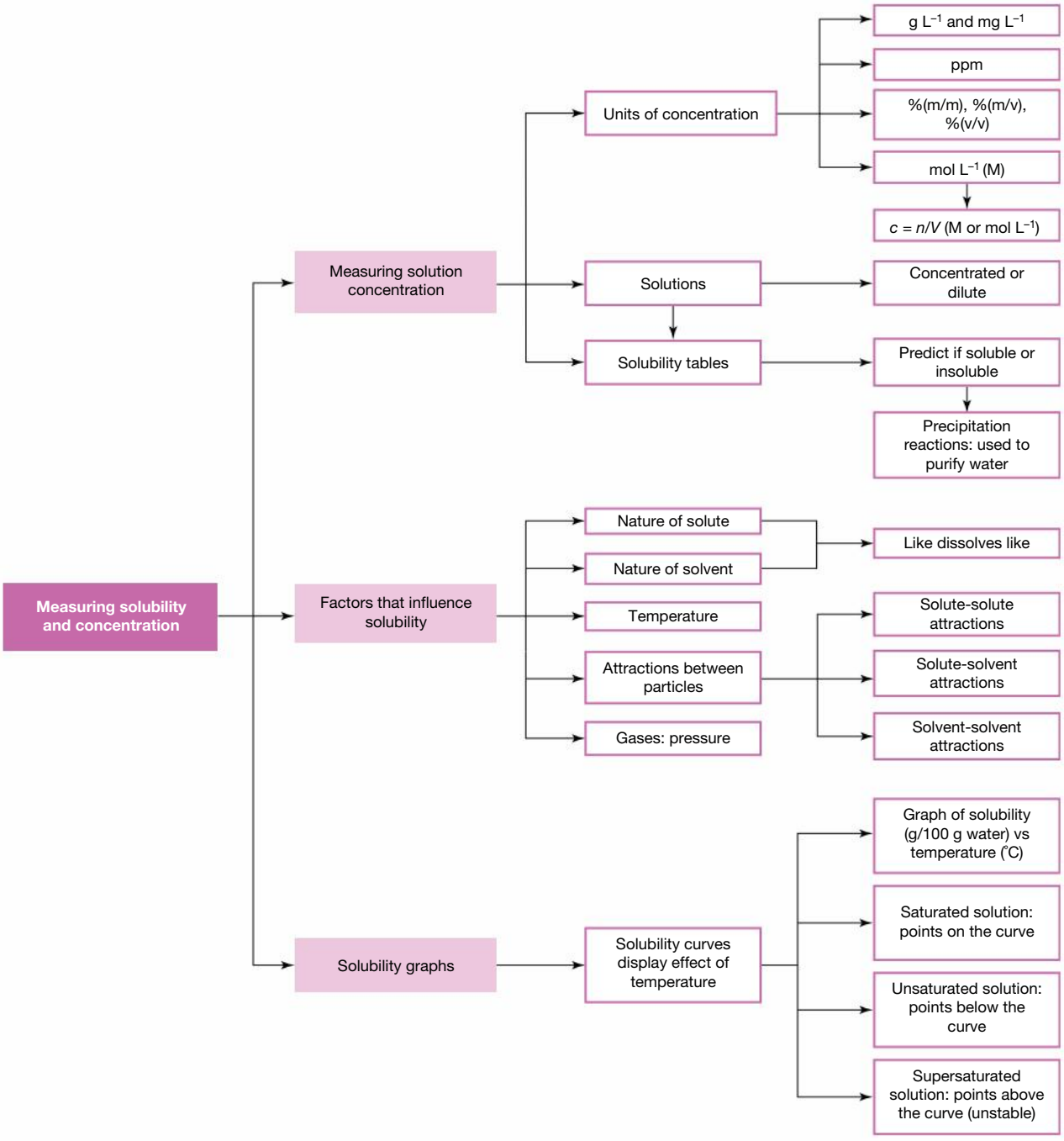
13.5.1 Topic summary



13.5.2 Key ideas summary

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13.5.3 Key terms glossary

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Solutions	Solutions — Topic 13 (sol-0812)
Practical investigation eLogbook	Practical investigation logbook — Topic 13 (elog-1618)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 13 (doc-37091) Key idea summary — Topic 13 (doc-37092)
Exam question booklet	Exam question booklet — Topic 13 (eqb-0094)

13.5 Activities

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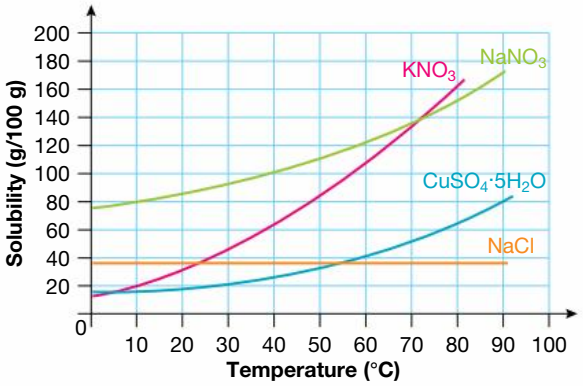
13.5 Review questions

1. An amount of 0.2 g of solute is dissolved in 500 L (500 kg) of aqueous solution. Express this in the following units.
   a. g L^{-1}
   b. ppm (μg g^{-1})
2. Consider the solubility curves shown and answer the questions.
   a. At what temperature is the solubility of $CuSO_4 \cdot 5H_2O$
      i. 30 g/100 g?
      ii. 40 g/100 g?
      iii. 20 g/100 g?
   b. At what temperature does 60 g of KNO_3 dissolve in 50 g of water?
   c. What is the solubility of the following solids at 35°C?
      i. NaCl
      ii. $NaNO_3$
      iii. KNO_3



3. Distinguish between the terms ‘unsaturated’, ‘saturated’ and ‘supersaturated’.
4. A solution of potassium chloride is 8.5%(m/m). An experiment requires 3.4 g of KCl. How many grams of the solution do you need for this experiment?
5. A chemist dissolves 3.5 mL of ethanoic acid in water to give a total of 120 mL. What is the concentration in %(v/v)?

6. Calculate the mass of solute in each of the following solutions.
   a. 300 mL of 1.5 M NaOH
   b. 250 mL of 2.0 M H_2SO_4
   c. 17.5 L of 1.5 M Na_2CO_3
   d. 200 mL of 2.5 M $CuSO_4 \cdot 5H_2O$
7. Calculate the mass of solute needed to make each of the following quantities of solution.
   a. 250 mL of a 1.5 M sodium bromide, NaBr, solution
   b. 500 mL of a 1.75 M magnesium sulfate, $MgSO_4$, solution
   c. 1.50 L of a 0.575 M magnesium nitrate, $Mg(NO_3)_2$, solution
   d. 100 mL of a 0.850 M sodium phosphate, Na_3PO_4, solution
   e. 2.00 L of a 0.00500 M potassium carbonate, K_2CO_3, solution
   f. 40.0 mL of a 2.30 M aluminium chloride, $AlCl_3$, solution
8. How many grams of CH_3COONa are obtained when 400.0 mL of a 0.500 M solution of CH_3COONa is evaporated to dryness?
9. Aluminium chloride, $AlCl_3$, is found in antiperspirants. A student wishes to do some tests on antiperspirant, so a jar of it is dissolved into a beaker containing 700 mL of water. The label on the jar states that it contains 200 g of aluminium chloride. Calculate the following.
   a. The concentration of the solution
   b. The concentration of each ion in the solution
10. In an $Fe_2(SO_4)_3$ solution, the concentration of SO_4^{2-} is 0.25 M. What is the concentration of solute?

13.5 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC Which of the following equations represents $MgCl_2$ dissolving in water?

A. $MgCl_2(s) \xrightarrow{H_2O(l)} Mg(s) + Cl_2(g)$

B. $MgCl_2(s) \xrightarrow{H_2O(l)} Mg^{2+}(aq) + Cl^-(aq)$

C. $MgCl_2(s) \xrightarrow{H_2O(l)} Mg(aq) + Cl_2(aq)$

D. $MgCl_2(s) \xrightarrow{H_2O(l)} Mg^{2+}(aq) + 2Cl^-(aq)$

Question 2

MC Which of the following lists contains only substances that are soluble in water?

A. Potassium iodide, sodium iodide, silver iodide
B. Silver nitrate, silver chloride, silver bromide
C. Sodium carbonate, ammonium carbonate, calcium carbonate
D. Copper(II) nitrate, iron(II) nitrate, iron(III) nitrate

Question 3

MC The unit 'parts per million' or ppm is used for very dilute solutions. Which of the following is *not* equivalent to 1 ppm for a dilute aqueous solution?

A. 1 kg/10^6 g
B. 1 mg L^{-1}
C. 1 mg kg^{-1}
D. 1 µg g^{-1}

Question 4

MC The following table shows solubility data for potassium fluoride.

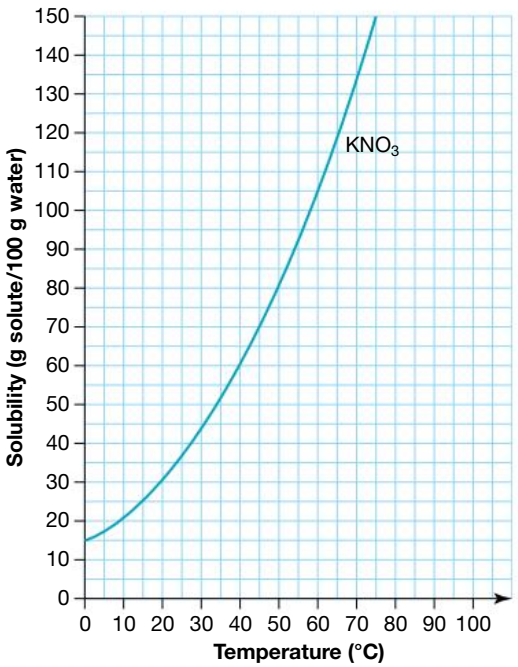
Temperature (°C)	0	20	40	60	80
Solubility (g per 100 g H_2O)	48	95	138	142	150

35 g of potassium fluoride is added to 20 g of water at 50 °C. The resultant mixture

A. is unsaturated.
B. is at saturation point.
C. contains about 2 g of undissolved solid.
D. contains about 7 g of undissolved solid.

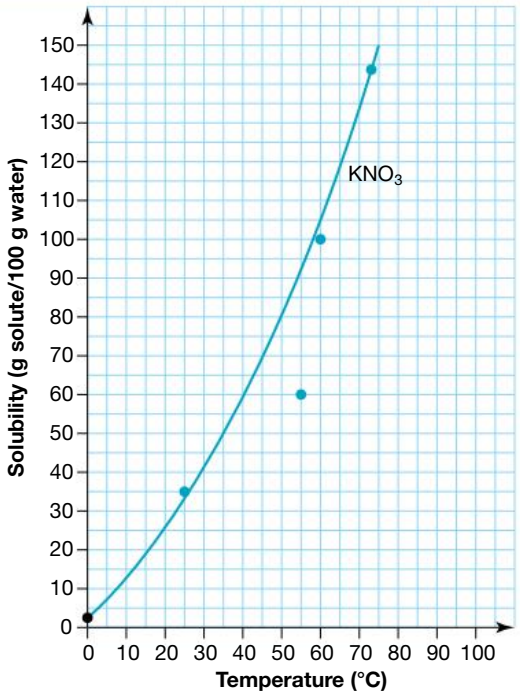
Question 5

MC According to the solubility graph shown, at what temperature would 40 g of KNO_3 dissolve fully in 50 g of water?



A. 25 °C **B.** 27 °C **C.** 50 °C **D.** 60 °C

Question 6



MC A student carried out an experiment to determine the solubility curve for potassium nitrate.

The expected results are shown here marked with a solid line. The student's data points are shown.

What may account for the outlier at 55 °C?

A. The student recorded 30 g instead of the 20 g of water actually used in the experiment.
B. The student forgot to subtract the mass of the test tube when weighing the KNO_3.
C. The student allowed the solution to cool too much before reading the thermometer.
D. The student used a larger test tube than other students.

Question 7

MC Which of the following contains the highest mass of solute?

A. 100 g of 2.2%(w/w) NaCl solution
B. 200 mL of 11 g L^{-1} KCl solution
C. 50 mL of 1.0 M NaCl solution
D. 50 mL of 1.0 M KCl solution

Question 8

MC What is the concentration of a solution prepared by dissolving 1.540 g of sodium carbonate in 250.0 mL volumetric flask?

A. 6.16×10^{-3} M
B. 3.633×10^{-3} M
C. 0.05811 M
D. 0.0742 M

Question 9

MC Which of the following solutions contains the largest amount (in mol) of copper(II) sulfate?

A. 200 mL of 0.050 M solution
B. 18.0 mL of 1.60 M solution
C. 0.500 L of 0.0300 M solution
D. 250 mL of 25.0 g L^{-1} solution

Question 10

MC What mass of ammonium sulfate would be needed to prepare 250.0 mL of a 0.100 M solution of this compound?

A. 0.0250 g **B.** 1.25 g **C.** 2.85 g **D.** 3.30 g

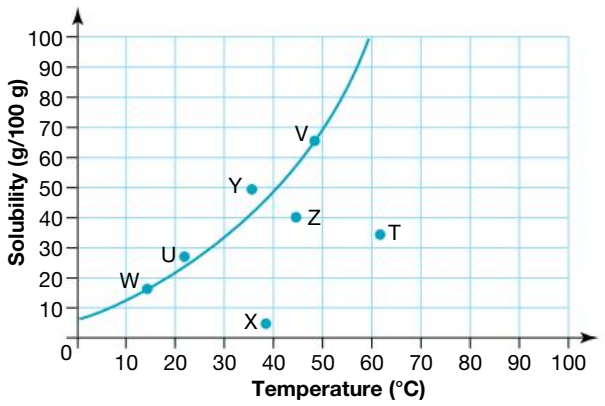
Section B — Short answer questions

Question 11 (1 mark)

What volume of 0.110 M NaCl would need to be evaporated to produce 5.00 g of NaCl?

Question 12 (3 marks)

Various solutions of potassium nitrate are represented by the letters T, U, V, W, X, Y and Z on the solubility curve shown in the figure. Which letters represent the following?



a. Saturated solutions **(1 mark)**
b. Unsaturated solutions **(1 mark)**
c. Supersaturated solutions **(1 mark)**

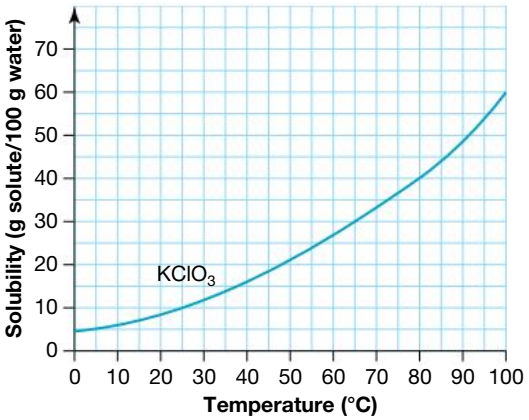
Question 13 (6 marks)

Calculate the concentration of the following aqueous solutions in (i) g L^{-1} and (ii) mg L^{-1} (ppm).

a. 3 g of solute in 100 mL solution **(2 marks)**
b. 0.004 g in 800 mL **(2 marks)**
c. 500 mg in 200 mL **(2 marks)**

Question 14 (6 marks)

The solubility curve for potassium chlorate, $KClO_3$, is shown.



a. What is the solubility of potassium chlorate at 40 °C? **(1 mark)**

b. At what temperature would 30 g of potassium chlorate dissolve in 100 g of water? **(1 mark)**

c. What mass of potassium chlorate would produce 250 mL of a saturated solution at 80 °C? **(1 mark)**

d. If a saturated solution in 50 mL of water at 90 °C is cooled to 25 °C, what mass of crystals would be produced? **(2 marks)**

e. Write the equation showing the dissociation of potassium chlorate in water. **(1 mark)**

Question 15 (5 marks)

It has been proposed that gold could be obtained from sea water. The concentration of gold varies from 0.1 to 2.0 mg per tonne.

a. Assuming that 1 tonne is equal to 1000 L of sea water, calculate the concentration of gold in parts per million. **(2 marks)**

b. Assuming that sea water contains 1.0 mg of gold per tonne, calculate the volume of sea water that would need to be processed to produce a one-kilogram gold ingot. **(2 marks)**

c. Why do you think this proposal would be uneconomical? **(1 mark)**

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14 Analysis for acids and bases

KEY KNOWLEDGE

In this topic you will investigate:

Analysis for acids and bases

- Volume–volume stoichiometry (solutions only) and application of volumetric analysis including the use of indicators, calculations related to the preparation of standard solutions, dilution of solutions, and use of acid–base titrations (excluding back titrations) to determine the concentration of an acid or a base in a water sample.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of VCE Chemistry. Experiments and investigations, supported by a **practical investigation eLogbook** and **teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

EXAM PREPARATION

Access exam-style questions and their video solutions in every lesson, to ensure you are ready.

14.1 Overview

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14.1.1 Introduction

To determine the concentration of an acid or a base in solution, a technique called volumetric analysis is often used. One of the pieces of equipment used is a burette, which delivers an accurate amount of solution.

As you will discover, completing a successful titration requires careful cleaning and delicate use of specialised equipment. While professional laboratories now use automated equipment to carry out titrations, learning how to complete a titration by hand is an important skill for all chemistry students. Such is the skill required, there are even titration competitions! Titrations are used in a range of industries, including monitoring blood glucose levels, in urine tests such as for pregnancy, and to determine the levels of salt, sugar, vitamins and fats in foods.

This topic examines the sources of acids and bases in water, their influence on the pH change of the solutions, and how stoichiometry can be applied to reactions in solutions.

The use of stoichiometry in solutions is known as volumetric analysis, and applications explored in this topic include preparations of standard solutions, calculations in diluted solutions and the use of acid–base titrations to determine the concentration of an acid or a base in a water sample.

FIGURE 14.1 A burette used for titration

LEARNING SEQUENCE

14.1 Overview 558
14.2 Solution stoichiometry (volume–volume stoichiometry) 559
14.3 Acid–base titrations 568
14.4 Review 581

on Resources

Solutions	Solutions — Topic 14 (sol-0813)
Practical investigation eLogbook	Practical investigation logbook — Topic 14 (elog-1619)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 14 (doc-37093) Key ideas summary — Topic 14 (doc-37094)
Exam question booklet	Exam question booklet — Topic 14 (eqb-0095)

14.2 Solution stoichiometry (volume–volume stoichiometry)

KEY KNOWLEDGE

- Volume-volume stoichiometry (solutions only) and application of volumetric analysis including the use of indicators, calculations related to the preparation of standard solutions, dilution of solutions to determine the concentration of an acid or a base in a water sample

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14.2.1 Stoichiometry

Stoichiometry is the term used for calculations based on chemical equations. It is the study and prediction of relative amounts (or ratios) of substances involved in chemical reactions. The word 'stoichiometry' comes from two Greek words meaning 'element' and 'measure'. Knowledge of stoichiometry is essential whenever quantitative information about a chemical reaction is required.

CASE STUDY: Using stoichiometry to design airbags

How is chemistry involved in the operation of an automotive airbag? Before being inflated, airbags contain a small amount of solid sodium azide, NaN_3. On impact with another vehicle or object, an electric spark is released, acting as a trigger. Very quickly, the sodium azide in the airbag decomposes to form solid sodium, Na, and nitrogen gas, N_2, according to the chemical equation:

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$$

The sodium then reacts with water vapour in air. The nitrogen gas produced in the reaction inflates the airbag in time to absorb much of the impact energy of the crash, which may otherwise cause injury to the driver. The sodium azide reaction must produce just the right amount of nitrogen gas to inflate the bag to the correct pressure. How do designers know how much sodium azide to use so that the airbag is correctly inflated? They use stoichiometric calculations.

Solving stoichiometric problems is particularly important in fields such as food chemistry, drug chemistry, forensic science and in any industry where chemicals are manufactured. Stoichiometry may be thought of as the means of obtaining a recipe for a chemical reaction.

Mole ratios

Chemical equations show the relative number of reactants participating in a chemical reaction. They can also show the relative number of moles, the **mole ratio**, of each substance involved.

Chemically speaking, if we wish to say, 'One molecule of nitrogen gas, N_2, reacts with three molecules of hydrogen gas, $3H_2$, to form two molecules of ammonia gas, $2NH_3$', we simply write:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

However, it would be impossible and impractical to make two ammonia molecules using one molecule of nitrogen gas and three molecules of hydrogen gas. Chemists work with larger quantities called moles (symbol mol — see topic 6). The coefficients in a balanced chemical equation tell us the relative numbers of moles of reactants and products taking part in the reaction.

stoichiometry the amount and ratio of reactants and products in a chemical reaction

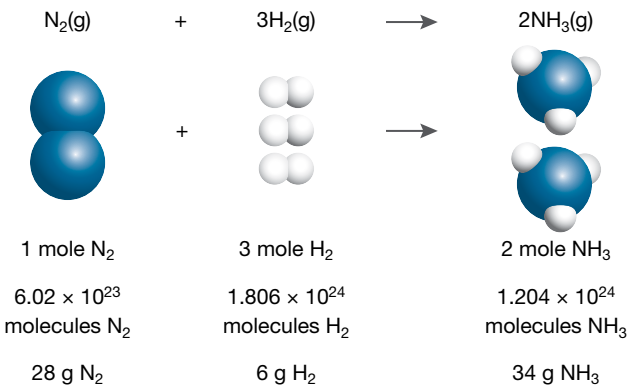
mole ratio represents the relative number of moles of each substance involved in the reaction

For example, one mole of nitrogen gas reacts with three moles of hydrogen gas to produce two moles of ammonia gas. The equation actually indicates the mole ratio of reactants and products, and the ratios never change. If one mole of nitrogen reacts with three moles of hydrogen, two moles of nitrogen react with six moles of hydrogen, 0.5 mole of nitrogen react with 1.5 moles of hydrogen and so on.

Assuming that plenty of nitrogen is available, how much ammonia is produced by three moles of hydrogen? The equation tells us: two moles of ammonia. Similarly, six moles of hydrogen produce four moles of ammonia.

An equation also tells us about the ratio between other quantities of the reactants and the products, including the number of particles and the masses of the reactants and products (see figure 14.2).

FIGURE 14.2 The balanced chemical equation for the formation of ammonia provides information about the ratios and abundances of products and reactants.



What a chemical equation *does not* tell us

An equation conveys no information about the rate of a reaction. It may be fast, as in the explosive oxidation of hydrogen to form water, or it may be slow, as in the oxidation or rusting of iron.

An equation does not tell us whether a reaction requires heat or gives off heat, or what temperature or pressure is needed.

Finally, an equation gives no details as to how the individual atoms or molecules are transformed from reactants to products. This information is essential for full understanding of how a chemical reaction takes place.

tlvd-0596

SAMPLE PROBLEM 1 Determining the number of moles of oxygen needed to react with methanol

Methanol, CH_3OH, is used as a fuel in some racing cars. The equation for the combustion of methanol in the car's engine is:

$$2CH_3OH\,(l) + 3O_2\,(g) \rightarrow 2CO_2\,(g) + 4H_2O\,(g)$$

How many moles of oxygen are needed to react with four moles of methanol?

THINK

From the coefficients in the reaction, we can see that two moles of methanol will react with three moles of oxygen.

WRITE

The mole ratio of methanol : oxygen is 2 : 3. Therefore, four moles of methanol will require six moles of oxygen.

PRACTICE PROBLEM 1

Using the reaction from sample problem 1, $2CH_3OH\,(l) + 3O_2\,(g) \rightarrow 2CO_2\,(g) + 4H_2O\,(g)$, how many moles of water are formed when one mole of oxygen reacts completely with methanol?

14.2.2 Volumetric analysis

Acids and bases are usually analysed using **volumetric analysis**. As the name suggests, volumetric analysis involves measuring volumes of solutions. If one of the solutions is of accurately known concentration, the concentration of a second solution may be derived using stoichiometry. The more accurately volumes and known concentrations are determined, the more accurate is the derived concentration. Great care, therefore, is taken in how this is done.

Although volumetric analysis is the technique of choice for many acid–base determinations, it can also be applied to other types of reactions that occur in solution. For example, many redox reactions are analysed using this method.

Before we look at the specifics of how volumetric analysis is performed, we first need to examine how stoichiometry can be applied to reactions in solution.

Solution stoichiometry

When some solutions are mixed, chemical reactions may occur. Solution stoichiometry, sometimes also called 'concentration–concentration' stoichiometry, involves reactions in solution such as precipitation reactions and the neutralisation of an acid and a base to form water and a salt. Solution stoichiometry determines the number of moles of solute in solutions that may react or that may be produced.

Solution stoichiometry formula

$$n = cV$$

Where n = number of moles (mol), c = concentration (mol L^{-1}) and V = volume (L).

Solution stoichiometry problems can be solved in four steps

1. Write a balanced chemical equation for the reaction, identifying the known (given) and unknown (required) quantities of substance.
2. Calculate the number of moles of the known quantity of substance present, using the formula $n = cV$.
3. From the equation, find the molar ratio of known to unknown quantities in the reaction and use it to calculate the number of moles of the required substance. Remember to always set up your mole ratio as your *unknown* coefficient divided by *known* coefficient; this way when you multiply by your known mole, this cancels and you can determine your unknown mole.
4. Calculate the quantity (that is, concentration) of the required substance.

Solution stoichiometry also uses the same steps as for mass–mass stoichiometry (see topic 17). The difference is that the formula $n = cV$ is used for both the given and the required substances in solution stoichiometry.

volumetric analysis
measurement of volume of one solution of accurately known concentration to determine the concentration of a second solution using stoichiometry

tlvd-0608

SAMPLE PROBLEM 2 Determining the concentration using solution stoichiometry

A solution of cloudy ammonia is analysed for its ammonium hydroxide concentration. A 20.00-mL sample is neutralised with 30.00 mL of 1.10 mol L^{-1} hydrochloric acid. Find the molarity or concentration of the ammonium hydroxide in the cloudy ammonia.

THINK	WRITE
1. Write the equation, identifying the known and unknown quantities of substance. Convert V units from mL to L.	$NH_4OH(aq) + HCl(aq) \rightarrow NH_4Cl(aq) + H_2O(l)$ $c(NH_4OH) = ?\ mol\ L^{-1}$ $c(HCl) = 1.10\ mol\ L^{-1}$ $V(NH_4OH) = 0.0200\ L$ $V(HCl) = 0.0300\ L$
2. Calculate the number of moles of the known quantity of HCl by recalling the formula $n = cV$.	$n(HCl) = c(HCl) \times V(HCl)$ $= 1.10 \times 0.030\ 00$ $= 0.0330\ mol$
3. Find the molar ratio from the equation and use it to calculate the number of moles of the required substance.	$n(NH_4OH) : n(HCl) = 1 : 1$ $n(NH_4OH) = n(HCl)$ So, $n(NH_4OH) = 0.0330\ mol$
4. Calculate the concentration of the required substance.	$c(NH_4OH) = \frac{n}{V}$ $= \frac{0.0330}{0.02000}$ $= 1.65\ mol\ L^{-1}$

PRACTICE PROBLEM 2

A solution of 25.0 mL of 1.00 mol L^{-1} H_2SO_4 was neutralised with 0.250 mol L^{-1} NaOH. What volume of NaOH was required?

14.2.3 Dilution

When a solution is diluted by the addition of more solvent (for example, water), the concentration of that solution changes. So, as its volume is increasing, the concentration is decreasing. Volume and concentration are inversely proportional to one another.

If we think of the number of solute particles, nothing has been added or taken away, so the number of moles (the amount) of solute remains the same. The addition of water to a concentrated solution does not alter the number of moles or the mass of the solute in that solution.

Number of moles before and after dilution

If n_1 represents the number of moles of the initial or concentrated solution and n_2 represents the number of moles of the final or dilute solution, we can say that:

$$n_1 = n_2$$

Using the equation $n = cV$, we can write equations for n_1 and n_2 as follows:

$$n_1 = c_1V_1 \text{ and } n_2 = c_2V_2$$

Since the values of n_1 and n_2 are equal, these equations can be combined.

The relationship between concentration and volume before and after dilution

The relationship between the concentration (c) and volume (V) of the initial (or concentrated) solution and the final (or dilute) solution is:

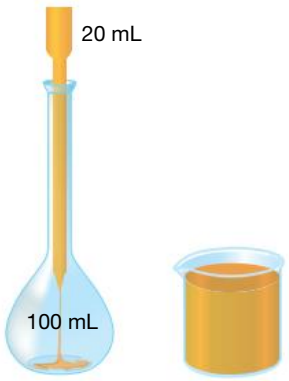
$$c_1V_1 = c_2V_2$$

Any unit can be used for volume, as long as the same unit is used in both cases.

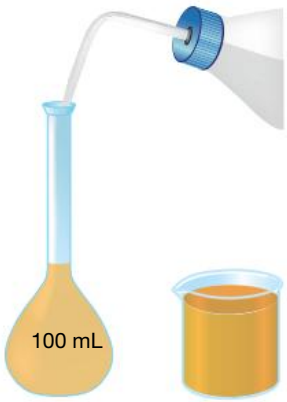
The dilution procedure is demonstrated in figure 14.3.

FIGURE 14.3 The dilution of a 0.20 M hydrochloric acid HCl, solution to 0.04 M

Step 1
Measure 20 mL of 0.20 M $K_2Cr_2O_7$ into a pipette. Slowly add the $K_2Cr_2O_7$ to a calibrated 100 mL flask.



Step 2
Rinse the pipette with water and also pour slowly into the calibrated 100 mL flask.



Step 3
Stopper the flask and mix thoroughly, and then add more H_2O to the 100 mL mark. The original 0.200 M $K_2Cr_2O_7$ has been diluted by a factor of 5, so the concentration decreases by a factor of 5. You now have 100 mL of 0.0400 M $K_2Cr_2O_7$.

tlvd-3487

SAMPLE PROBLEM 3 Determining volumes of stock solutions

What volume of 10 mol L^{-1} stock solution of hydrochloric acid is needed to prepare 250 mL of a 2.0-mol L^{-1} hydrochloric acid solution?

THINK

1. Let c_1 and V_1 represent the quantities of concentrated solution and c_2 and V_2 represent the quantities of dilute solution.
   Convert volume units from mL to L.
2. Calculate the quantity of 10 mol L^{-1} HCL needed to produce the required dilution.

WRITE

$V_1 = ?\,L$ $\quad V_2 = 0.250\,L$

$c_1 = 10.0\,mol\,L^{-1}$ $\quad c_2 = 2.0\,mol\,L^{-1}$

$$c_1V_1 = c_2V_2$$
$$10.0 \times V_1 = 2.0 \times 0.250$$
$$V_1 = \frac{2.0 \times 0.250}{10.0}$$
$$= 0.050\,L$$
$$= 50\,mL$$

PRACTICE PROBLEM 3

From a stock solution of nitric acid, a volume of 20 mL was used to prepare 250 mL solution of 0.80 mol L^{-1}. What is the concentration of nitric acid in the stock solution?

14.2.4 Standard solutions

A **standard solution** is a solution with an accurately known concentration. Two methods can usually be used to determine a solution's concentration accurately.

Method 1: by reacting it with another solution with a concentration that is accurately known.

Method 2: by taking a substance called a **primary standard** and dissolving it in a known volume of water. Primary standards are pure substances that satisfy several criteria.

To qualify as a primary standard, a substance must have a number of the following properties:

- It must be a solid with a high state of purity and soluble in the solvent used.
- It must have an accurately known formula — that is, its degree of hydration must not change.
- It must be stable. In other words, its composition or formula must not change over time. Changes can happen, for example, as a result of storage or reaction with the atmosphere.
- It should be cheap and readily available.
- It should have a relatively high molar mass so weighing errors are minimised.

Note that this means that not all substances are suitable for use as primary standards.

A **secondary standard** can be prepared by standardising it (measuring its concentration) against a primary standard.

Substances not suitable as primary standards

Not all substances are suitable for use as primary standards, for the following reasons:

- Despite being available as a commercial liquid concentrate, their exact concentration is unknown (for example, hydrochloric acid, HCl, and sulfuric acid, H_2SO_4), and they can produce fumes, increasing the dangers involved in trying to measure an accurate volume.
- Some solids (of known chemical formula) can readily absorb water vapor, H_2O, changing their chemical formula (that is, most alkali hydroxides such as sodium hydroxide, NaOH).
- Can readily absorb carbon dioxide, CO_2, from the atmosphere (that is, most alkali hydroxides).

standard solution solution that has a precisely known concentration

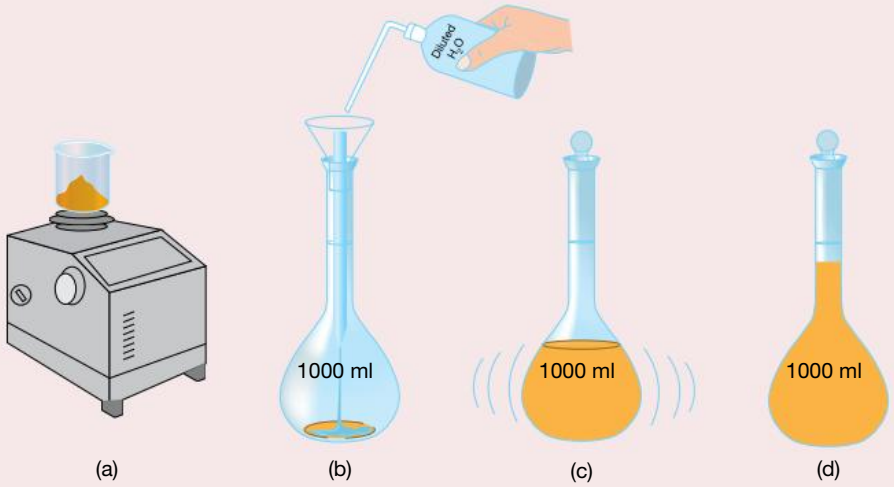
primary standard substance used in volumetric analysis that is of such high purity and stability that it can be used to prepare a solution of accurately known concentration

secondary standard a standard prepared for a specific volumetric analysis that has been standardised against a primary standard

Preparing a standard solution using a primary standard such as anhydrous sodium carbonate, Na_2CO_3

1. Weigh out the required mass of solute in a special weighing bottle or small beaker (figure 14.4).
2. Transfer this to the volumetric flask using a dry funnel and wash the weighing bottle and funnel a number of times, each time adding the rinsings to the flask.
3. Add distilled water to approximately half the required volume, and then swirl to dissolve the solute.
4. Add further distilled water until the bottom of the meniscus is level with the etched line on the neck of the flask. Fully invert at least 15 times to ensure that the concentration of the solution is uniform.

FIGURE 14.4 Steps in preparing a standard solution



tlvd-3488

SAMPLE PROBLEM 4 Calculating the concentration of a standard solution

Calculate the concentration of a standard solution of potassium dichromate, $K_2Cr_2O_7$, which was prepared from 30.0 g of $K_2Cr_2O_7$ in a 250-mL volumetric flask.

THINK

1. List all required formulas and convert given units to required units.

WRITE

$$c = \frac{n}{V}$$
$$n = \frac{m}{M}$$
$$V = \frac{250\text{ mL}}{1000} = 0.25\text{ L}$$

2. Calculate moles of $K_2Cr_2O_7$.

$$n(K_2Cr_2O_7) = \frac{m}{M}$$
$$= \frac{30}{(2\times 39.1 + 2\times 52 + 7\times 16)}$$
$$= 0.102\text{ mol}$$

3. Calculate concentration.

$$c = \frac{n}{V}$$
$$= \frac{0.102}{0.250}$$
$$= 0.408\text{ mol L}^{-1}$$

PRACTICE PROBLEM 4

A standard solution of 0.500 mol L^{-1} Na_2CO_3 is required. What mass of anhydrous sodium carbonate, Na_2CO_3, should be used to prepare the solution in a 250 mL volumetric flask?

Example of preparation of a secondary standard solution

A primary standard solution of hydrochloric acid is difficult to prepare for reasons already outlined. To accurately determine hydrochloric acid's concentration, it is usually standardised by titration against a known concentration of a suitable base. A standard solution of anhydrous sodium carbonate is sometimes used as the base because it is stable and does not deteriorate over time.

tlvd-0611

SAMPLE PROBLEM 5 Determining the molarity of a solution from titrations

1.33 g of anhydrous sodium carbonate is placed in a 250.0-mL (0.25 L) volumetric flask and water is added to the required mark.

This standard solution is poured into the burette and titrated against a solution of hydrochloric acid; its accurate concentration is to be determined. It was found that 23.50 mL of Na_2CO_3 was required to completely neutralise 20.00 mL of the HCl.

Calculate the concentrations of the following.

a. The standard solution of anhydrous Na_2CO_3

b. The HCl solution

THINK

a. To calculate the concentration of the standard solution use

$n = \frac{m}{M}$ and $c = \frac{n}{V}$.

WRITE

$$n(Na_2CO_3) = \frac{m}{M}$$
$$= \frac{1.33}{106.0}$$
$$= 0.01\ 255 \text{ mol}$$
$$c(Na_2CO_3) = \frac{n}{V}$$
$$= \frac{0.01\ 255}{0.2500}$$
$$= 0.0502 \text{ mol L}^{-1}$$

b. 1. Write the equation for the reaction and use it to find the molar ratio between Na_2CO_3 and HCl.

$Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g) Na_2CO_3 : HCl$
$1 : 2$

2. Calculate the concentration of the solution.

$$n(Na_2CO_3) = cV$$
$$= 0.0502 \times 0.023\ 50$$
$$= 1.18 \times 10^{-3}$$
$$n(HCl) = 2 \times 1.18 \times 10^{-3}$$
$$= 2.36 \times 10^{-3} \text{ mol}$$
$$c(HCl) = \frac{n}{V}$$
$$= \frac{2.36 \times 10^{-3}}{0.020\ 00}$$
$$= 0.118 \text{ mol L}^{-1}$$

Therefore, the hydrochloric acid has a molarity of 0.118 mol L^{-1}.

PRACTICE PROBLEM 5

A solution of previously standardised sodium hydroxide was used to accurately determine the concentration of a sulfuric acid solution. On average, 18.56 mL of sulfuric acid was required to neutralise 20.00 mL of 0.0927 mol L^{-1} sodium hydroxide solution. Calculate the molarity of the sulfuric acid solution.

14.2 Activities

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14.2 Quick quiz on | 14.2 Exercise | 14.2 Exam questions

14.2 Exercise

1. In an experiment to determine the concentration of a solution of hydrochloric acid, 15.5 mL of the HCl was neutralised by 25.0 mL of a 0.055 mol L^{-1} solution of Na_2CO_3 in a conical flask.
   a. Write a balanced equation for the reaction.
   b. Determine the concentration of the hydrochloric acid.
2. a. What volume of a solution of 0.200 mol L^{-1} potassium hydroxide, KOH, exactly neutralises 20.0 mL of a solution of the following.
      i. 0.300 M hydrochloric acid, HCl
      ii. 0.300 M sulfuric acid, H_2SO_4
      iii. 0.300 M phosphoric acid, H_3PO_4
   b. What conclusions or trends in data can you draw from your answer to part a?
3. The salinity of sea water can be measured by adding silver nitrate solution until all the chloride ions have been precipitated as silver chloride, according to the equation:

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

   A 20.00 mL sample of sea water required 22.40 mL of 0.500 mol L^{-1} silver nitrate solution to precipitate all the chloride ions from solution. Calculate the molarity (concentration) of the sodium chloride in the sample.
4. Calculate the volume of stock solution needed to prepare 250 mL of a 2.15 M solution of hydrochloric acid (hydrochloric acid stock solution = 12 mol L^{-1}).
5. Calculate the concentration (molarity) of 2.0 L of 2.0 mol L^{-1} HCl solution after dilution with 500 mL of water.
6. If 30 mL of stock solution of initial 2.0 mol L^{-1} was diluted to 0.500 L, what is its new concentration?
7. How much water is required to dilute a solution of 25 mL of 3.00 mol L^{-1} to 0.75 mol L^{-1}?
8. If 170 mL of water was used to prepare a solution of 250 mL of 0.600 mol L^{-1}, what was the original concentration?
9. a. Sodium oxalate, $Na_2C_2O_4$, can be used to make a standard solution. Describe how this is done.
   b. If 3.183 g of sodium oxalate is dissolved in water in a 250.0 mL volumetric flask, find the concentration of the solution.
10. 1.461 g of anhydrous sodium carbonate, Na_2CO_3, is dissolved in water in a 250 mL volumetric flask.
    a. Why can this be called a standard solution?
    b. Calculate the concentration of the solution.

14.2 Exam questions

Question 1 (1 mark)

MC A solution containing 0.001 mol of HCl reacts exactly with which of the following?

A. 1 mL of 0.1 mol L^{-1} NaOH
B. 1 mL of 0.01 mol L^{-1} NaOH
C. 10 mL of 0.1 mol L^{-1} NaOH
D. 10 mL of 0.01 mol L^{-1} NaOH

Question 2 (4 marks)

A student was required to find the concentration of a solution of sulfuric acid.

In order to carry out a titration, the student first prepared a standard solution of anhydrous sodium carbonate by dissolving 2.390 g in 100.0 mL of water using a volumetric flask.

Once prepared and thoroughly mixed, 20.00 mL aliquots of the sodium carbonate solution were pipetted into conical flasks and titrated with the acid. The average titre was found to be 17.58 mL.

a. Write the equation for the reaction between sulfuric acid and sodium carbonate. **(1 mark)**
b. Calculate the concentration of the standard sodium carbonate solution. **(1 mark)**
c. Calculate the concentration of the sulfuric acid solution. **(2 marks)**

Question 3 (2 marks)

What volume of water is needed to produce a 2.00 mol L^{-1} Na_2CO_3 solution if 2.35 g of Na_2CO_3 is present?

Question 4 (3 marks)

Calculate the concentration (molarity) of the following.

a. 100 mL of 0.1 mol L^{-1} solution after dilution with 10 mL water **(1 mark)**
b. 16 mL of 0.20 mol L^{-1} solution after dilution with 4.0 mL water **(1 mark)**
c. 750 mL of 1.50 mol L^{-1} solution after dilution with 150 mL water **(1 mark)**

Question 5 (3 marks)

An iron ore sample may be analysed by preparing an iron(II) ion solution of the ore, which is then reacted with acidic permanganate solution as in the given equation. If 15.0 mL of 0.0100 mol L^{-1} MnO_4^- solution is required to completely react with 10.0 mL of the Fe^{2+} solution, determine the concentration of the iron(II) ion solution.

$$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$$

More exam questions are available in your learnON title.

14.3 Acid–base titrations

KEY KNOWLEDGE

- Volume-volume stoichiometry (solutions only) and application of volumetric analysis including the use of indicators, calculations related to the preparation of standard solutions, dilution of solutions, and use of acid–base titrations (excluding back titrations) to determine the concentration of an acid or a base in a water sample

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Acid–base **titration** is an experimental method that uses specialised equipment to determine the concentration of a solution, based on neutralisation reactions.

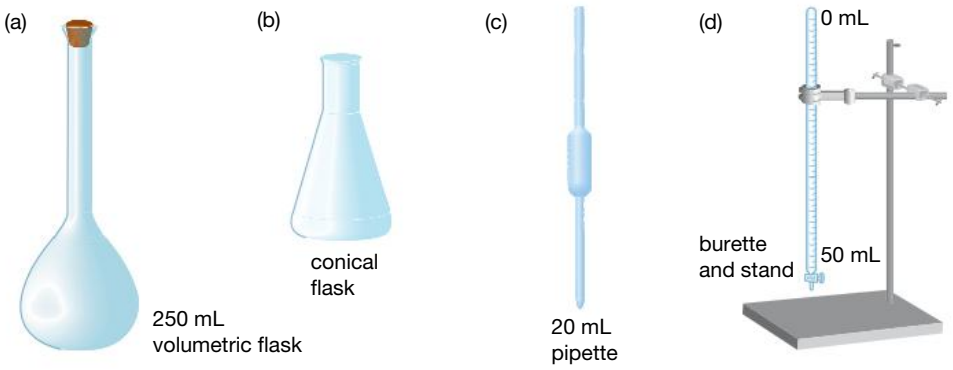
titration process used to determine the concentration of a substance by using a pipette to deliver one substance and a burette to deliver another substance until they have reacted in exactly their mole ratio

Acid–base titration is a type of volumetric analysis where the unknown concentration of a solution called the **analyte** is determined by reacting it with a solution of known concentration called the **titrant**. The solution with a concentration that is known accurately is called the standard solution, explained in section 14.2.4.

14.3.1 Titration apparatus and method

Successful volumetric analysis relies on accurate measurement of volumes and often, but not always, accurate knowledge of one of the initial concentrations, using very accurate volume-measuring equipment. Keeping accuracy in mind, the correct operating technique must be applied, as outlined in this section.

FIGURE 14.5 Basic equipment required in a volumetric analysis



The volumetric flask

Purpose: Volumetric flasks, also called standard flasks, are used either when making up a standard solution or when an accurate dilution of an existing standard solution is required.

Features: Broad and bulbus at the bottom with an etched calibrated mark of known volume in the narrow neck of the flask.

Preparation for titration: Rinse the volumetric flask first with distilled water and follow the steps as per figure 14.4 to prepare the secondary standard.

Be aware: An important difference in filling a volumetric flask is that it is always filled so the bottom of the meniscus is on the calibration mark and the volume of the solution is never higher than the exact volume.

The pipette

Purpose: A volumetric pipette is used to deliver analyte as a known, fixed volume called an **aliquot** to the conical flask for a titration.

Features: Each pipette has an etched calibrated mark above the bulge section. (Drawing the liquid up above this mark and then allowing the bottom of the **meniscus** to drain down to the mark ensures the pipette delivers an accurate volume.)

analyte the solution of unknown concentration to go into the conical flask and be determined by titration

titrant the standard solution of known concentration to go into the burette

aliquot known, fixed volume delivered by a pipette

meniscus the curved liquid surface in a tube caused by the hydrostatic properties of fluid with the walls of the glassware

Preparation and procedure for titration:

- Rinse the pipette first with distilled water, and second with a small portion of the analyte. (This ensures the analyte is neither contaminated nor diluted, which would ultimately decrease accuracy.)
- Fill the pipette to the calibrated mark using an appropriate a pipette filler (such as a bulb attachment). *Never use your mouth.*
- By holding the tip of the pipette against the side of the conical flask, transfer the analyte aliquot by allowing the solution to drain. (*Note:* a small amount of solution will remain in the pipette. This is normal, and calibration of the pipette takes this small volume into account to ensure accuracy remains.)

Be aware:

- When filling the pipette, *always* keep the tip of the pipette below the level of solution — otherwise, you will get air into your pipette, causing the solution to rush up into the pipette filler where contamination may result.
- *Always* read from the bottom of the meniscus level because this is where pipettes are calibrated to, but also *always* read it from eye level. A parallax error will result if your eyes are not level to the meniscus.

FIGURE 14.6 A volumetric pipette

The conical flask

Purpose: Conical flasks have an aliquot of analyte and a few drops of pH indicator. During the titration, the burette tap is open, allowing the titrant to flow into the flask and causing neutralisation between the acid and the base. This reaction will occur until the **end point** is determined by the colour change of the indicator. The end point is very close to the **equivalence point**, which equates to the reactants (the acid and the base) being present in their molar equivalent ratios, and thus having reached neutralisation.

Features: The flask itself is narrow at the neck before flaring out towards the bottom. Its shape is perfect for swirling the liquid contents.

Preparation for titration:

- Rinse the conical flask with distilled water in preparation for the aliquot of analyte. (This ensures the aliquot is not contaminated and the number of mole is not changed.)
- Add several drops of a suitable pH indicator to the solution, depending on your specific neutralisation reaction (see section 14.3.2).

Be aware: If the titrant produces any splashes on the side of the conical flask, ensure you use distilled water to rinse the side of the flask. This will ensure accuracy of your titre.

FIGURE 14.7 A conical flask

end point the point at which the indicator just changes colour or the pH curve becomes vertical

equivalence point point at which two reactants have reacted in their stoichiometric molar ratios in a titration as determined by a balanced equation

The burette

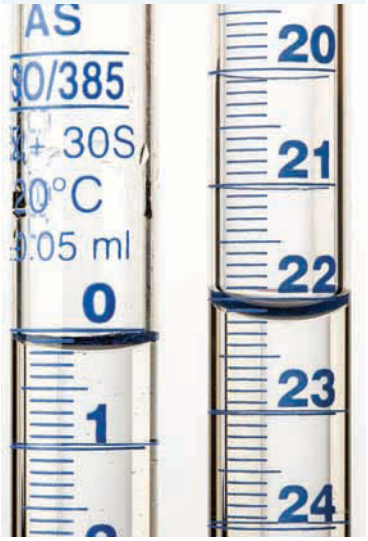
Purpose: To deliver a variable volume of titrant called a **titre** to the conical flask to neutralise the analyte.

Features: The burette is marked with an accurate scale, calibrated to tenths of a millilitre (0.1 mL).

Preparation and procedure for titration:

- Rinse the burette first with distilled water, and second with a small portion of the titrant. (This ensures the titrant is neither contaminated nor diluted, which would ultimately decrease accuracy.)
- Check the burette tap does not leak, and the tip is not blocked by a build-up of sediment.
- Clamp the burette vertically to the stand.
- To fill the burette, place a small funnel on top of the burette to prevent spillage, and fill with titrant. It does not matter what level you fill the burette (for example, it does not have to be exactly on the 0.00 mL mark), provided it is within the marked scale region. Remove the funnel before you start your titration.
- Open the tap to drain some titrant through the tip and close the tap, ensuringno air bubbles remain in the tap.
- When reading the measurement, common practice is to interpolate between the calibrated marks to obtain a reading to a hundredth of a millilitre (that is, two decimal places — for example, 0.11 mL).

FIGURE 14.8 Read from the bottom of the meniscus when reading from a burette.



- To control the process of titration, adopt the 'left-hand turn — right-hand swirl' technique. This involves wrapping the fingers and thumb of the left hand around the burette tip and tap, and using them to open and close the tap. The right hand then swirls the flask as the titrant and analyte are mixed.
- A white tile placed under your conical flask will assist with determining the colour change of the indicator.
- When you get close to the expected end point (colour change) of the reaction, reduce the flow of the burette to drop-sized amounts until the analyte shows a permanent colour change

Be aware:

- *Always* read from the bottom of the meniscus level and *always* read at eye level. Not reading at eye level will cause a parallax error.
  (Note the numerals increase from 0.00 mL at top of the burette to 50.00 mL at the bottom of the burette, so be careful when taking your measurement — see figure 14.8.)
- You must remove your filling funnel. Forgetting to do so may interfere with the accuracy of your measurements, because the funnel may drip some drops of solution.
- Slowly add the titrant. If you add too much titrant and overshoot the end point, the indicator colour change may be too much, causing an error of accuracy with a concentration less that what it is.
- To minimise errors from reading the titre, **repeated measurements** of the titration are required until **concordant titres** are obtained. At least three concordant titres (volumes with readings within ± 0.10 mL of each other) are required to calculate the **average titre**.

titre volume of liquid measured by a burette during titration

repeated measurements precision between the results of successive measurements of the same quantity being measured, carried out under the same conditions of measurement

concordant titres results of repeated titrations showing precision to be within ± 0.10 mL of each other

average titre titre determined by averaging the results of several titrations

14.3.2 Indicators

Acid–base indicators change colour over a narrow pH range (the end point), allowing us to stop a titration close to the equivalence point if the indicators have been chosen carefully and specific to the neutralisation reaction taking place.

As we have seen in subtopic 11.5, a range of indicators exist so it is important to choose an indicator with an end point as close to the equivalence point for your titration reaction. Remember the pH is not necessarily 7 at the equivalence point. If we remember that a conjugate acid and base are always formed in any acid–base reaction (refer to section 11.2.2), it is possible that these new acids and bases have an effect on the pH. Knowledge of this pH for a particular titration means that an indicator can be chosen that changes colour at (or close to) the correct pH of the endpoint. Table 14.1 shows the details of some common indicators, and this is visually shown in figure 14.10.

Table 14.1 shows that, if we know the pH at the equivalence point of a titration is 3.2, a suitable choice of indicator might be methyl orange. On the other hand, phenolphthalein would not be suitable. Several factors may also influence the choice of indicator, such as how *quickly* the pH changes at the equivalence point.

FIGURE 14.9 Acid–base titration using phenolphthalein pH indicator

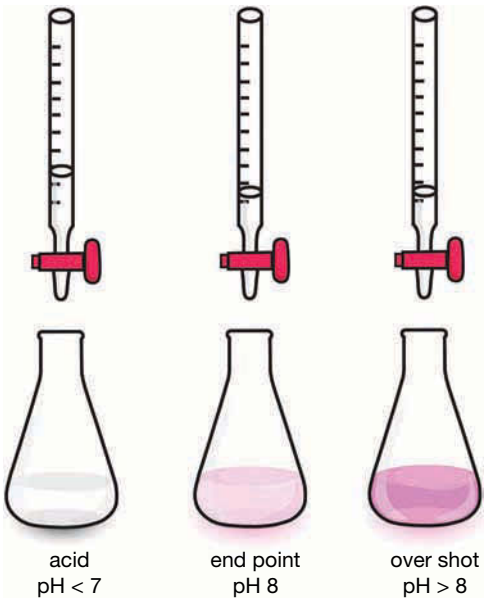


FIGURE 14.10 pH chart for some common acid–base indicators

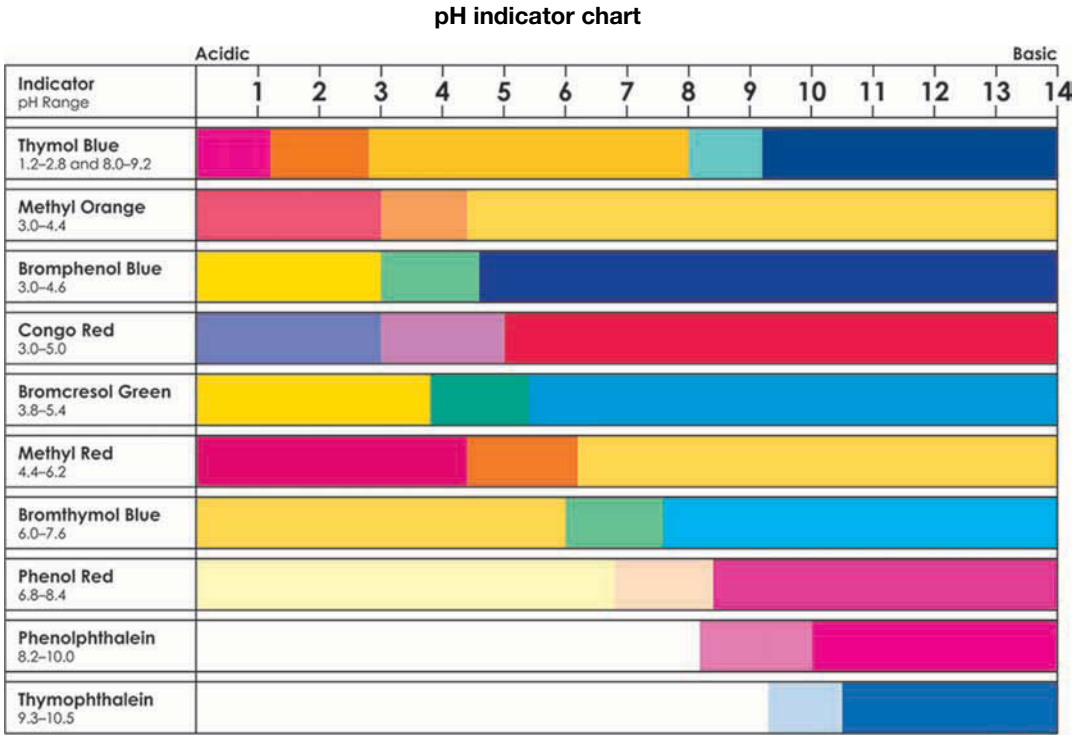


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TABLE 14.1 Some common acid–base indicators

Indicator	Colour at lower pH	pH range for colour change	Colour at higher pH
Methyl orange	Red	3.1–4.4	Yellow
Methyl red	Red	4.2–6.3	Yellow
Litmus	Red	5.0–9.0	Blue
Bromothymol blue	Yellow	6.0–7.6	Blue
Phenolphthalein	Colourless	8.3–10.0	Crimson

Determining the equivalence point without using indicators

Sometimes, the equivalence point of a titration may be gauged in other ways that do not involve end points and indicators. Some of these are by:

- using a pH meter and stopping at the correct pH.
- doing a potentiometric titration. This method uses two electrodes and measures the changes in voltage produced between them.
- measuring conductivity. If the product of an acid–base reaction is a precipitate, the conductivity in the titration flask falls as ions are removed in the precipitate (remember that the other product of an acid–base reaction is water). As soon as all precipitation has finished, the conductivity increases again as excess ions are now added from the burette.

FIGURE 14.11 Titrations can be performed automatically in many modern laboratories.

14.3.3 Sources of errors

In volumetric analysis, it is important to understand the steps in a procedure and predict the effect on the final calculated result if the steps are not performed correctly. Table 14.2 shows some general areas in which mistakes could be made and what effects these mistakes can have.

Other situations can also give rise to mistakes when the properties of the chemicals involved are misunderstood. For example, a common method of standardising hydrochloric acid solution is to react it with a standard sodium carbonate solution. Solid sodium carbonate needs to be anhydrous to qualify as a primary standard (see section 14.2.4). However, if it is left exposed to the atmosphere for too long, it absorbs moisture and begins to rehydrate. Should this then be used to make up a standard solution, not all the mass weighed out would be sodium carbonate, due to the presence of water molecules. The resulting solution of sodium carbonate is, therefore, lower in concentration, and the effect on the calculated concentration of the hydrochloric acid would be an overestimation.

TABLE 14.2 Effect on the calculated result of some possible mistakes during volumetric analysis

Situation	Effect if substance under analysis is in the . . .		Comments
	burette	titration flask	
Rinsing water left in burette	Underestimated	Overestimated	The burette solution is diluted with water, so more is used.
Rinsing water left in pipette	Overestimated	Underestimated	The solution aliquot in the titration flask is diluted.
Indicator chosen changes colour too soon	Overestimated	Underestimated	The choice of indicator can be critical.
Water in conical flask	No effect	No effect	All necessary measurements are made before the chemicals are mixed with this water.

Uncertainties of equipment

Experiments involving glassware and other equipment will always be the source of some uncertainties of the measurements, which lead to errors associated with the equipment.

The glassware and scales used in volumetric analysis typically present the following uncertainties:

- electronic scales with uncertainty of ± 0.0001 g or ± 0.001 g
- 50 mL burettes with ± 0.02 mL for each reading
- 20 mL pipettes with ± 0.03 mL
- 250 mL volumetric flask with ± 0.1 to 0.3 mL (depending on the volume and the grade of the flask)
- 100 mL graduated beaker ± 0.5 mL.

elog-1855

tlvd-0761

EXPERIMENT 14.1 online only

Standardisation of hydrochloric acid

Aim

To determine the accurate concentration of a solution of hydrochloric acid

14.3.4 Titration curves

As a titration is performed, a **titration curve** can be constructed to plot the pH change versus the volume of the titre added during titration.

- The titration curve shows the pH change over time, as the equivalence point is reached or exceeded.
- Titration curves have different shapes depending on the strength of the acids and bases used in the titration, because each combination will have a different equivalence point. Therefore, identifying the strength of the acids and bases used in the titration will assist in choosing the correct indicator that will be suitable for the expected equivalence point.

titration curve results of a titration showing the pH change versus the volume of the titre added during titration

Figure 14.12 highlights a titration of a strong base, NaOH, titrated with a strong acid, HCl, and figure 14.13 highlights a strong acid, HCl, titrated with a strong base, NaOH.

FIGURE 14.12 Titration curve of a strong base (NaOH) titrated with a strong acid (HCl)

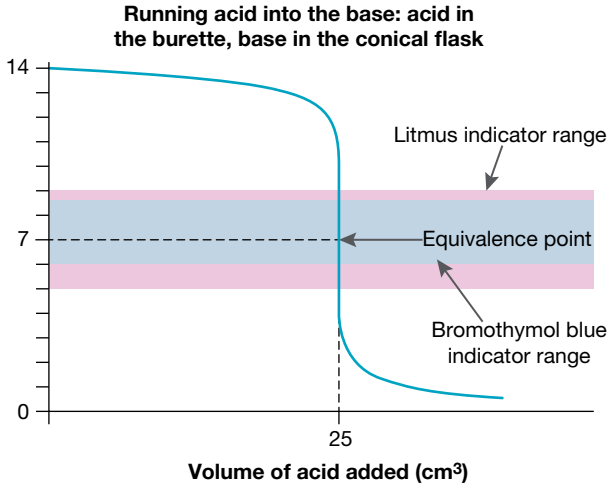
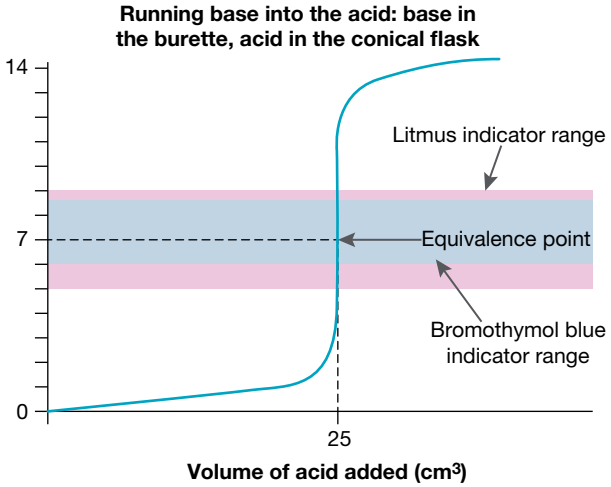


FIGURE 14.13 Titration curves of a strong acid (HCl) titrated with a strong base (NaOH)



In both cases, the equivalence point is around a pH = 7; therefore, an appropriate indicator to capture the pH change would be any indicator with a colour change between pH 4 and pH 8. According to table 14.1, the best indicators for this example would be either bromothymol blue (colour change between 6.0 and 7.6) or litmus (colour change between 5.0 and 9.0).

TIP: Choosing indicators

Always choose an indicator with an end-point colour change range that occurs as close as possible to the equivalence point of the neutralisation reaction of the titration, for the least volume of titrant added.

on Resources

Weblink Titration simulator

14.3.5 Calculations in acid–base titrations

Before starting titration calculations, identify the standard solution information provided. Sometimes calculations for standard solution also need to be employed.

Calculation steps required in a titration

1. Write the balanced chemical equation.
2. Determine the average titre (concordant volumes only).
3. Calculate moles of standard solution (titration) from the average titre.
4. Use molar ratio from equation to derive the moles of the unknown solution (analyte).
5. Calculate the concentration of the unknown solution from the calculated moles to the correct number of significant figures.

tlvd-0612

SAMPLE PROBLEM 6 Performing titration calculations using numerous aliquots

Calculate the concentration of the NaOH solution if titres of 24.45 mL, 24.78 mL, 24.69 mL and 24.65 mL of 0.300 mol L^{-1} HCl were used to titrate 20.0 mL of NaOH solution.

THINK	WRITE
1. Write the balanced chemical equation.	$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$
2. Determine the average titre (concordant volumes only) for the standard solution, and convert volume units from mL to L.	$V = \frac{24.78 + 24.69 + 24.65}{3}$ $= 24.71 \text{ mL}$ $= \frac{24.71 \text{ mL}}{1000}$ $= 0.02471 \text{ L}$
3. Calculate moles of standard solution from the average titre by applying the formula $c = \frac{n}{V}$.	$n = cV$ $n(HCl) = 0.300 \times 0.02471$ $= 0.00741 \text{ mol}$
4. Use molar ratio from equation to derive the moles of the unknown substance.	$n(NaOH) : n(HCl) = 1 : 1$ Thus, $n(NaOH) = 0.00741 \text{ mol}$
5. Calculate the concentration of NaOH solution from the calculated moles. **TIP** Remember to give your answer to the correct number of significant figures.	$c(NaOH) = \frac{n}{V}$ $= \frac{0.00741}{0.0200}$ $= 0.371 \text{ mol } L^{-1}$

PRACTICE PROBLEM 6

A standard solution of 0.200 mol L^{-1} K_2CO_3 was used to titrate a 23.50-mL aliquot of HCl. The titres used were 20.03 mL, 19. 95 mL, 20.01 mL and 19.98 mL. Calculate the concentration of the aliquot.

14.3.6 Titrations with dilutions

Sometimes when performing titrations with concentrated solutions, we use large volumes for titres. To minimise the volume used, we can dilute the aliquot to a workable volume.

When dilution is undertaken, we need to remember to account for it in calculations.

tlvd-0613

SAMPLE PROBLEM 7 Determining the concentration of dilution solutions using titrations

A commercial bathroom cleaner contains NaOH as its active ingredient. A 40.0-mL volume of bathroom cleaner was diluted to 250 mL in a volumetric flask.
An aliquot of 20.0 mL of the diluted cleaner was then titrated with 30.0 mL 1.00 mol L^{-1} HCl standard solution. What was the concentration of NaOH in the bathroom cleaner?

THINK	WRITE
1. Write the balanced equation.	$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$
2. Calculate the moles of standard solution.	$n(HCl) = cV$ $= 1.00 \times 0.0300$ $= 0.0300$ mol
3. Use molar ratio to calculate moles of diluted aliquot.	$n(HCl) : n(NaOH) = 1 : 1$ $n(NaOH\ diluted) = 0.0300$ mol
4. Calculate the concentration of the diluted aliquot.	$c(NaOH\ diluted) = \frac{n}{V}$ $= \frac{0.030}{0.0200}$ $= 1.50\ \text{mol}\ L^{-1}$
5. Multiply by the dilution factor to calculate the NaOH concentration in the bathroom cleaner.	Dilution factor: $\frac{250}{40.0} = 6.25$ $c(NaOH\ in\ cleaner) = c(NaOH\ diluted) \times \text{dilution factor}$ $= 1.50 \times 6.25$ $= 9.38\ \text{mol}\ L^{-1}$
Alternatively, use the relationship between the concentration (c) and volume (V) of the initial (or concentrated) solution and the final (or dilute) solution, $c_1V_1 = c_2V_2$. *Note:* this relationship only works if the mole ratio in the equation is 1 : 1.	c_1 (concentrated) = ? V_1 (concentrated) = 40.0 mL c_2 (diluted) = 1.50 mol L^{-1} V_2 (diluted) = 250 mL $c_1V_1 = c_2V_2$ $c_1 = \frac{c_2V_2}{V_1}$ $= 1.50 \times \frac{250}{40}$ $= 9.38\ \text{mol}\ L^{-1}$

PRACTICE PROBLEM 7

A commercial bathroom cleaner contains NaOH as its active ingredient. A 50.0-mL volume of bathroom cleaner was diluted to 250 mL in a volumetric flask.

An aliquot of 15.0 mL of the diluted cleaner was then titrated with 25.0 mL 0.800 mol L^{-1} HCl standard solution. What was the concentration of NaOH in the bathroom cleaner?

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tlvd-0642

EXPERIMENT 14.2 online only

Determining the CO_2 content of different brands of soft drinks

Aim

To determine the CO_2 content of different brands of soft drinks

on Resources

Weblink RSC Titration screen experiment

14.3 Activities

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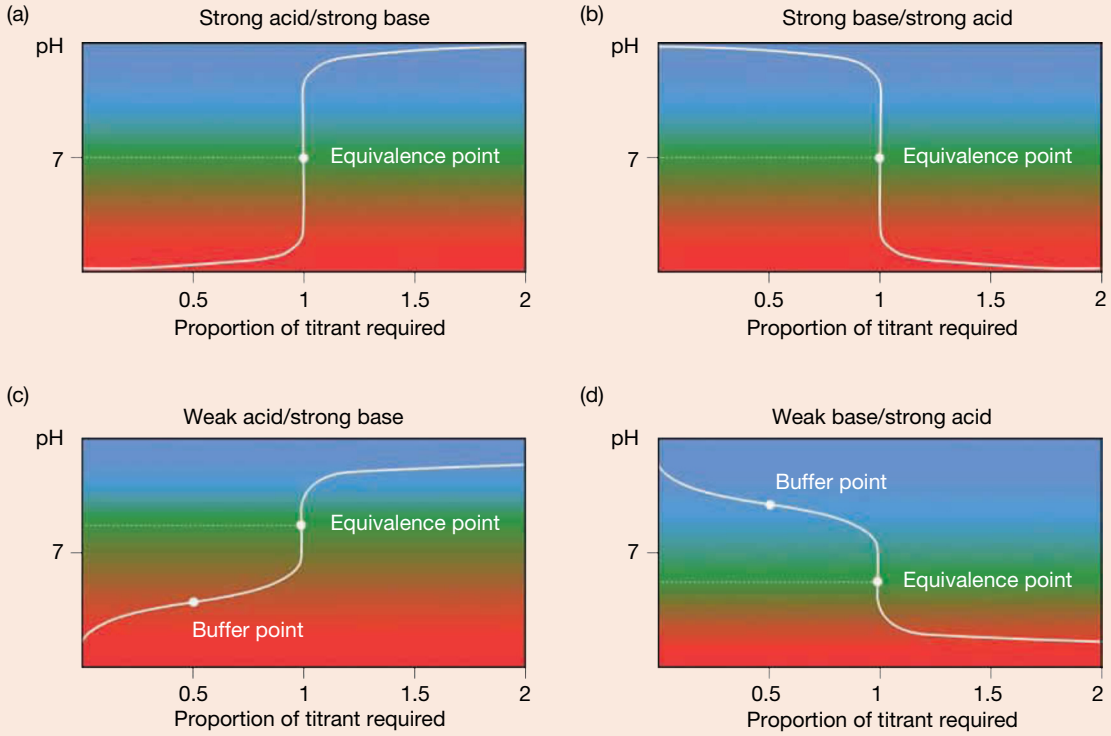
14.3 Exercise

1. MC A student needs to choose an indicator for an acid–base reaction. Which of the following factors is not important in the decision of which indicator to choose?
   - **A.** The indicator needs to be soluble and to not interfere with the reaction.
   - **B.** The indicator needs to have a known concentration.
   - **C.** The indicator needs to have only two different colours so that the end point is more easily determined.
   - **D.** The indicator must change colour around the equivalence point of the reaction.
2. Thymol blue (abbreviated as HTb) is an indicator that, despite its name, changes colour from red to yellow according to the equation:

$$\underset{\text{red}}{HTb(aq)} + H_2O(1) \rightleftharpoons \underset{\text{yellow}}{Tb^-(aq)} + H_3O^+(aq)$$

   What colour is the indicator when the concentration of HTb is approximately equal to the concentration of Tb^-?
3. 1.461 g of anhydrous sodium carbonate, Na_2CO_3, is dissolved in water in a 250 mL volumetric flask. 20.00 mL aliquots of the solution were titrated with nitric acid, HNO_3, until three concordant titres are achieved. The average concordant titre was found to be 22.17 mL.
   - **a.** Write the equation for the reaction of sodium carbonate with nitric acid.
   - **b.** Find the concentration of the nitric acid.

4. Describe the difference between equivalence point and end point.
5. What indicator would be suitable to be used in a titration where equivalence point is at pH 4.1?
6. What indicators would be most suitable for the following titration combinations given that the general titration curves are known?



7. What volume of 0.80 mol L^{-1} HCl will be required to titrate a 20.0 mL solution of 0.60 mol L^{-1} KOH to the equivalence point?
8. A factory is suspected of polluting a stream with hydrochloric acid from its waste water. To gather evidence, a sample of stream water was collected downstream from the discharge point. Back in the laboratory, a 20.00 mL sample of this water was titrated against a solution of sodium carbonate that had been accurately diluted to 0.0010 mol L^{-1}; 17.52 mL was required.
   Assuming that the only acid in the sample was hydrochloric acid:
   a. calculate the molarity of hydrochloric acid in the sample
   b. calculate the concentration of the hydrochloric acid in g L^{-1}, mg L^{-1} and ppm.
9. Ten 1 L water samples were collected from different locations in a lake and then mixed together. One litre of this mixture was then titrated with 0.050 mol L^{-1} sulfuric acid until the pH equalled 4.5. 15 mL of acid was required.
   Calculate the total alkalinity of this water as mg L^{-1} calcium carbonate equivalent.
10. A hydrochloric acid solution of unknown concentration is 'standardised' by titrating it against a standard solution of 0.102 M sodium carbonate solution. 25.00 mL aliquots of the base required an average titre of 19.46 mL. Calculate the concentration of the hydrochloric acid to the correct number of significant figures.

14.3 Exam questions

Question 1 (1 mark)

MC A student carried out an acid–base titration to find the concentration of a nitric acid solution. A standard solution of sodium carbonate was prepared in a volumetric flask. 20.00 mL aliquots were drawn from the standard solution and poured into a conical flask for each titration.

With what liquid or solution should each of the following pieces of apparatus be rinsed?

Apparatus	Pipette	Burette	Volumetric flask	Conical flask
A	Water	Water	Water	Water
B	Nitric acid	Sodium carbonate	Sodium carbonate	Water
C	Sodium carbonate	Nitric acid	Water	Water
D	Sodium carbonate	Nitric acid	Sodium carbonate	Nitric acid

A. A
B. B
C. C
D. D

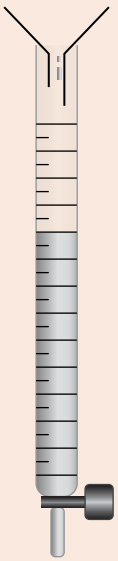
Question 2 (1 mark)

MC What is the piece of equipment shown and what does it do?

A. A pipette and it delivers a volume called a titre.
B. A burette and it delivers a volume called a titre.
C. A pipette and it delivers a volume called an aliquot.
D. A burette and it delivers a volume called an aliquot.

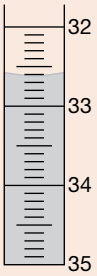
Question 3 (1 mark)

A student left the funnel in the top of the burette while titrating, as shown in the diagram. Why is this a technical error?



Question 4 (1 mark)

A section of a burette is shown.



What is the reading of the level of the meniscus?

Question 5 (12 marks)

One of the uses of hydrochloric acid is to control the pH of water supplies and swimming pools when they become too basic.

a. A chemist was required to find the accurate concentration of a solution of hydrochloric acid that was estimated to be near 0.15 M in concentration.

i. What is the pH of a hydrochloric acid solution of 0.15 M? **(1 mark)**

ii. The chemist decided to prepare a standard solution of sodium carbonate to use in a volumetric analysis of the HCl solution. **(1 mark)**

What is meant by the term 'standard solution'?

iii. The chemist decided to prepare a standard solution of sodium carbonate to use in a volumetric analysis of the HCl solution. **(2 marks)**

In order to prepare 250 mL of a 0.100 M solution of sodium carbonate, what mass of anhydrous sodium carbonate must be used?

iv. Write the full equation for the reaction between hydrochloric acid and sodium carbonate. **(1 mark)**

The chemist used 20.00 mL aliquots of the sodium carbonate solution and performed four titrations as shown in the provided table.

Trial	1	2	3	4
Titre (mL)	28.39	27.95	27.98	27.99

v. What should the chemist have used to rinse the pipette? **(1 mark)**

vi. Calculate the average titre of the acid solution and use it to determine the accurate concentration of the hydrochloric acid solution. **(4 marks)**

b. On a different occasion, the chemist needed to change the pH of a tank of waste water from pH 9 to pH 7 by adding pure water.

i. What is the concentration of hydroxide ions in a solution with pH 9? **(1 mark)**

ii. What volume of pure water would need to be added to 5 L of waste water in order to reduce its pH from 9 to 7? **(1 mark)**

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14.4 Review

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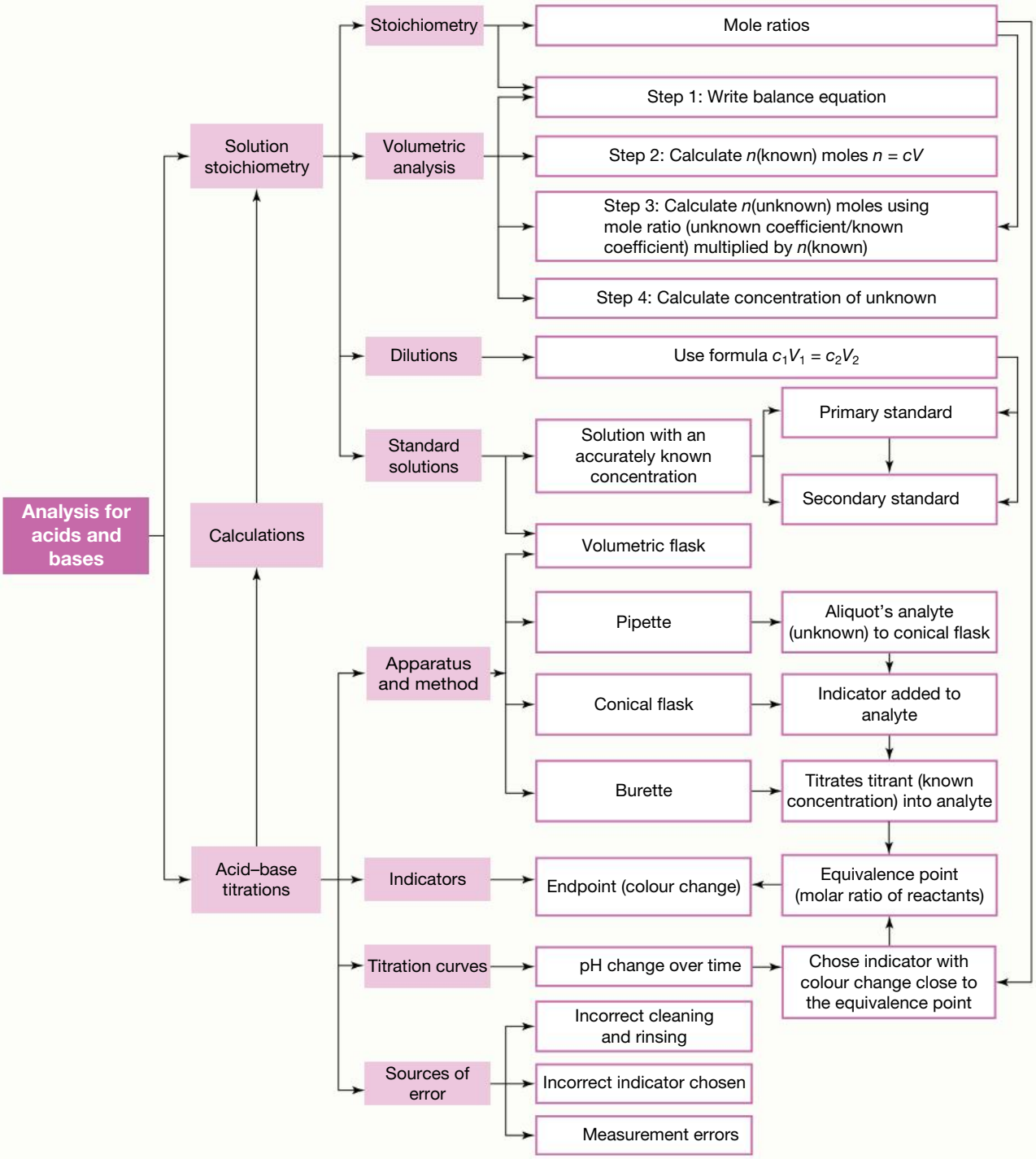
14.4.1 Topic summary



14.4.2 Key ideas summary

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14.4.3 Key terms glossary

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on Resources

Solutions	Solutions — Topic 14 (sol-0813)
Practical investigation eLogbook	Practical investigation logbook — Topic 14 (elog-1619)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 14 (doc-37093) Key idea summary — Topic 14 (doc-37094)
Exam question booklet	Exam question booklet — Topic 14 (eqb-0095)

14.4 Activities

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14.4 Review questions

1. Define the following terms.
   a. Aliquot
   b. Concordant titres
   c. Equivalence point
2. Calculate the volume of stock solution needed to prepare the following.
   a. 500 mL of a 0.750 mol L^{-1} solution of sulfuric acid (sulfuric acid stock solution = 18 mol L^{-1})
   b. 200 mL of a 2.5 mol L^{-1} solution of ammonium hydroxide (ammonium hydroxide stock solution = 15 mol L^{-1})
   c. 350 mL of a 0.150 mol L^{-1} solution of ethanoic acid (ethanoic acid stock solution = 17 mol L^{-1})
3. What volume of 0.250 mol L^{-1} HNO_3 reacts completely with 280 mL of 0.200 mol L^{-1} H_2S according to the reaction:

$$2HNO_3(aq) + 3H_2S(aq) \rightarrow 3S(s) + 4H_2O(l) + 2NO(g)$$

4. What volume of 0.460 mol L^{-1} H_2SO_4 is required to neutralise 24.00 mL of 0.620 mol L^{-1} NaOH?
5. What volume of water must be added to
   a. 100 mL of 15 mol L^{-1} H_2SO_4 to dilute it to 5.0 mol L^{-1}?
   b. 130 mL of 3.50 mol L^{-1} HNO_3 to dilute it to 1.00 mol L^{-1}?
   c. 170 mL of 2.60 mol L^{-1} HCL to dilute it to 0.250 mol L^{-1}?
6. A student is required to make up 500 mL of 0.100 mol L^{-1} $H_2SO_4(aq)$ by diluting a 0.500 mol L^{-1} solution of the acid. What volume of 0.500 mol L^{-1} $H_2SO_4(aq)$ is required?
7. A 3.50 g mass of $MgCl_2$ is dissolved in 200 mL of water.
   a. What is the concentration of the solution?
   b. If 150 mL of water is added to the solution, what is the concentration of the new, diluted solution?
8. 50 mL of 1.2 mol L^{-1} sulfuric acid was added to 30 mL of sodium hydroxide in order to neutralise it. Find the molarity of the sodium hydroxide.

9. What volume of a 2.30 mol L^{-1} solution of Na_3PO_4 is required to react completely with 560 mL of a 3.25 mol L^{-1} solution of $Ca(ClO_3)_2$, according to the reaction:

$$3Ca(ClO_3)_2(aq) + 2Na_3PO_4(aq) \rightarrow Ca_3(PO_4)_2(aq) + 6NaClO_3(aq)$$

10. A sample of waste water known to contain sodium hydroxide as the only base was analysed prior to discharge into a river. It was found that a 20.00 mL sample required 8.19 mL of 0.0100 mol L^{-1} hydrochloric acid solution for neutralisation.
    a. Calculate the molar concentration of sodium hydroxide in the water sample.
    b. Express the answer from part **a** in g L^{-1}.

14.4 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC A student dissolved 2.00 g of sodium hydroxide in 1.00 L to produce solution A. The student then diluted 100 mL of solution A with water to produce 1.00 L of solution B. A 10.0 mL sample of solution B was then diluted to 100 mL with water to produce solution C.

What was the concentration of solution C?

A. 0.500 mol L^{-1}
B. 0.0500 mol L^{-1}
C. 0.005 00 mol L^{-1}
D. 0.000 500 mol L^{-1}

Question 2

MC What volume of 11 mol L^{-1} hydrochloric acid is required to produce 2 L of 1 mol L^{-1} acid when diluted with water?

A. 91 mL **B.** 182 mL **C.** 600 mL **D.** 1838 mL

Question 3

MC Which of the following is *not* a requirement for a substance to be a primary standard?

A. It must have a high state of purity.
B. It must not be a danger to health.
C. It must have an accurately known formula.
D. It must be stable.

Question 4

MC When is the end point reached in an acid–base titration?

A. When the solution is neutral.
B. When the indicator changes to a permanent colour.
C. When equal volumes of reactants have been mixed.
D. When reactants have been mixed in the appropriate stoichiometric ratio.

Question 5

MC When is the equivalence point reached in an acid–base titration?

A. When the solution is neutral.
B. When the indicator changes to a permanent colour.
C. When equal volumes of reactants have been mixed.
D. When reactants have been mixed in the appropriate stoichiometric ratio.

Question 6

MC What mass of potassium hydroxide, KOH, is needed to produce 200 mL of a 0.25 mol L^{-1} solution?

A. 0.050 g **B.** 2.8 g **C.** 50 g **D.** 2.8 kg

Question 7

MC What volume of 0.100 mol L^{-1} H_2SO_4 is required to react completely with 25.0 mL of 0.150 mol L^{-1} NaOH solution?

A. 9.38 mL **B.** 18.8 mL **C.** 37.5 mL **D.** 75.0 mL

Question 8

MC A student obtained the following titres during an analysis: 18.90 mL, 19.02 mL, 18.97 mL, 19.12 mL, 18.98 mL.

What is the average of the concordant titres?

A. 18.95 mL **B.** 18.99 mL **C.** 19.01 mL **D.** 19.07 mL

Question 9

MC What should the volume from a 25 mL pipette and a 250 mL volumetric flask, respectively, be recorded as?

A. 25.00 mL and 250.0 mL
B. 25 mL and 250 mL
C. 25.0 mL and 250.0 mL
D. 25.00 mL and 250.00 mL

Question 10

MC A chemist needed to determine whether an acid was monoprotic, diprotic or triprotic. A 0.100 M solution of the acid was poured into a burette and used to titrate 20.00 mL aliquots of 0.105 M sodium carbonate. If the average titre was 21.02 mL, what was the acid?

A. Monoprotic
B. Diprotic
C. Triprotic
D. Unable to be determined without an equation

Section B — Short answer questions

Question 11 (11 marks)

A company operates a batch process that uses sulfuric acid. At the end of each batch, the leftover acid must be diluted before it is released. The amount of water to be added to enable this dilution is carefully calculated following analysis of the leftover sulfuric acid in the batch liquid.

The terms of the discharge licence state that no more than 1.00 g L^{-1} of sulfuric acid should be in the discharge water.

A 25.00 mL sample of the batch liquid was accurately diluted to 250.0 mL in a standard flask. 25.00 mL of this diluted sample was then titrated with 0.100 mol L^{-1} standardised sodium hydroxide solution using a suitable indicator. An average titre of 18.52 mL was obtained.

a. Write the equation for the reaction between sodium hydroxide and sulfuric acid. **(1 mark)**

b. Calculate the number of moles of sodium hydroxide that were added. **(1 mark)**

c. From your part **b** answer, calculate the number of moles of sulfuric acid in the diluted sample. **(2 marks)**

d. Hence, calculate the molarity of sulfuric acid in the diluted sample. **(1 mark)**

e. Calculate the molarity of sulfuric acid in the undiluted batch liquid. **(2 marks)**

f. Calculate the molarity of the sulfuric acid that is allowed in the discharge water. **(1 mark)**

g. Hence, calculate the number of litres of water that must be added to each litre of batch liquid to meet the terms of the discharge licence. Give your answer to one decimal place. **(3 marks)**

Question 12 (6 marks)

Anhydrous sodium carbonate is often used to make standard solutions of sodium carbonate. This solution may then be used in acid–base titrations to calculate the exact concentration of acidic solutions.

a. Calculate the mass of anhydrous sodium carbonate, Na_2CO_3, that is required to make 500.00 mL of 0.0500 mol L^{-1} solution. **(2 marks)**

b. Exactly 25.00 mL of the solution from part **a** is reacted with hydrochloric acid. The equation for this reaction is:

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

It is found that 23.50 mL of the acid is required for exact neutralisation.
Calculate the molarity of the hydrochloric acid solution. **(4 marks)**

Question 13 (10 marks)

A student is required to find the concentration of a solution containing hydrochloric acid. Before titrating, the student must first prepare a standard solution of sodium carbonate to react with the acid. The sodium carbonate solution was prepared by dissolving 2.773 g of sodium carbonate in 250.0 mL of water using a volumetric flask.

a. Calculate the concentration of the sodium carbonate solution to the correct number of significant figures. **(1 mark)**

b. 20.00 mL aliquots of the sodium carbonate solution were pipetted into conical flasks. Methyl orange indicator was added and the solutions were titrated with the acid. The average titre was 19.58 mL.

i. Write the equation for the reaction between hydrochloric acid and sodium carbonate. **(1 mark)**

ii. Calculate the concentration of the hydrochloric acid. **(3 marks)**

c. Identify the solutions used to rinse the following pieces of glassware.

i. The conical flask used for the reaction between the acid and base **(1 mark)**

ii. The burette **(1 mark)**

iii. The volumetric flask used to prepare the standard solution of sodium carbonate **(1 mark)**

d. Identify the solution used to rinse a volumetric flask, when preparing a solution of citric acid. **(2 marks)**

Question 14 (6 marks)

To ensure accuracy, it is important that the equipment used in volumetric analysis has been correctly rinsed.

Consider a reaction where aliquots of citric acid solution are titrated using a standardised sodium hydroxide solution.

a. Identify the solutions used to rinse the following pieces of glassware.

i. The burette **(1 mark)**

ii. The conical flask (reaction flask) **(1 mark)**

b. If water was used to rinse the pipette at the start of this experiment, answer the following.

i. Explain how would it affect the calculated value of citric acid. **(2 marks)**

ii. For how many titrations would this rinsing error affect the titration results? **(1 mark)**

c. Explain how a volumetric flask is rinsed when preparing a standard solution of citric acid. **(1 mark)**

Question 15 (7 marks)

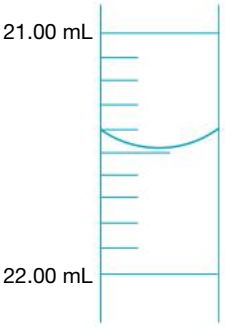
The unknown concentration of a solution of sodium hydroxide, NaOH(aq), can be determined by using a sulfuric acid solution of known concentration. The volumetric analysis can be carried out using methyl orange as an indicator. The results of an analysis are as follows.

Concentration of sulfuric acid, H_2SO_4 = 0.095 60 M

Aliquot of sodium hydroxide, NaOH = 25.00 mL

Initial volume of sulfuric acid, H_2SO_4 = 2.32 mL

A magnified burette scale of the final burette volume of H_2SO_4 is shown in the figure.



a. Write a balanced equation, with states, for the reaction. **(1 mark)**

b. Describe the function of the methyl orange indicator. **(1 mark)**

c. **i.** Read the final volume in the burette from the figure provided. **(1 mark)**

ii. Determine the volume of the titre, in mL, of H_2SO_4. **(1 mark)**

d. Calculate the amount in moles of sulfuric acid used in the titration. **(1 mark)**

e. Calculate the concentration, in M, of the sodium hydroxide solution. **(2 marks)**

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15 Measuring gases

KEY KNOWLEDGE

In this topic you will investigate:

Measuring gases

- CO_2, CH_4 and H_2O as three of the major gases that contribute to the natural and enhanced greenhouse effects due to their ability to absorb infrared radiation
- the definitions of gas pressure and standard laboratory conditions (SLC) at 25 °C and 100 kPa
- calculations using the ideal gas equation ($PV = nRT$), limited to the units kPa, Pa, atm, mL, L,°C, and K (including unit conversions)
- the use of stoichiometry to solve calculations related to chemical reactions involving gases (including moles, mass and volume of gases)
- calculations of the molar volume or molar mass of a gas produced by a chemical reaction.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of VCE Chemistry. Experiments and investigations, supported by a **practical investigation eLogbook** and **teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

EXAM PREPARATION

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15.1 Overview

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15.1.1 Introduction

Gases have properties and exhibit behaviours that are different from liquids and solids. Scientists use the kinetic molecular theory of gases to explain these observed properties and behaviours. The theory consists of five points that allow us to form a mental picture of how particles in a gas would look and behave if we could observe them directly.

The principles of stoichiometry and gas behaviour can be combined to understand the impact that gases have on the environment. Gases such as carbon dioxide, methane, nitrous oxide and ozone contribute to the enhanced greenhouse effect. The 2021 COP26 summit brought parties together to accelerate action towards the goals of the Paris Agreement of 2016.

How can we as individuals lessen our contribution to climate change? In this topic, you will combine your understanding of stoichiometry with the introduced gas laws to calculate how much gas is produced or consumed in a reaction. You will build on this in unit 3, when you will combine your knowledge and skills gained in this topic with thermochemical equations to calculate how much gas is released when different fuels are burned.

FIGURE 15.1 Carbon dioxide is the most significant greenhouse gas due to the massive amounts produced by human activity.

LEARNING SEQUENCE

15.1 Overview 588
15.2 Gases and the enhanced greenhouse effect 589
15.3 Gases at standard laboratory conditions (SLC) 595
15.4 Calculations using the ideal gas equation and stoichiometry 603
15.5 Review 610

on Resources

Solutions	Solutions — Topic 15 (sol-0814)
Practical investigation eLogbook	Practical investigation logbook — Topic 15 (elog-1620)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 15 (doc-37095) Key ideas summary — Topic 15 (doc-37096)
Exam question booklet	Exam question booklet — Topic 15 (eqb-0096)

15.2 Gases and the enhanced greenhouse effect

KEY KNOWLEDGE

- CO_2, CH_4 and H_2O as three of the major gases that contribute to the natural and enhanced greenhouse effects due to their ability to absorb infrared radiation

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15.2.1 Greenhouse gases

Carbon dioxide, CO_2, methane, CH_4, and water vapour, H_2O, are all classified as **greenhouse gases**; that is, they are gases that absorb infrared radiation and, hence, contribute to the **greenhouse effect**. The concentration of water vapour in the atmosphere is not directly linked to human activity, unlike the levels of carbon dioxide and methane. This and the **enhanced greenhouse effect** will be covered in more detail in section 15.2.2.

The human activities that are the major sources of greenhouse gases are summarised in table 15.1.

FIGURE 15.2 Our reliance on vehicles powered by petroleum-based fuels is a contributor to increased CO_2 and CH_4 levels.

greenhouse gases gases that contribute to the greenhouse effect by absorbing infrared radiation

greenhouse effect a natural process that warms the Earth's surface; when the Sun's energy reaches the Earth's atmosphere, some is reflected back to space and the rest is absorbed and re-radiated by greenhouse gases

enhanced greenhouse effect the effect of increasing concentrations of greenhouse gases in the atmosphere as the result of human activity

TABLE 15.1 CO_2 and CH_4 production as a result of human activity

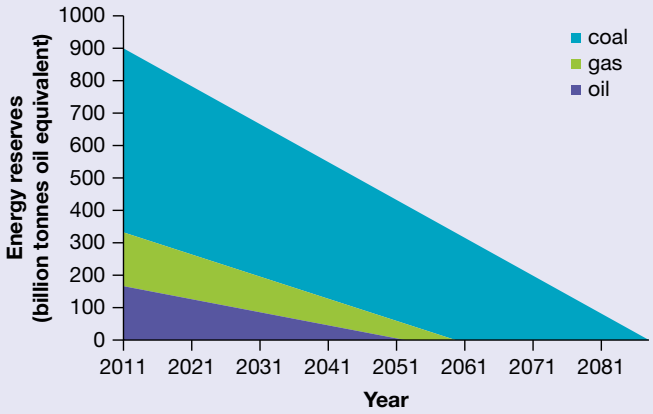
Human activity	Source of CO_2 and CH_4
Electricity and heat production	The combustion of coal, oil and natural gas releases CO_2 into the atmosphere that was previously trapped in the fossil fuels.
Industry	The manufacture of goods and materials, such as cement and steel. Cement production releases vast amounts of CO_2.
Transportation	Transportation mostly relies on the burning of petroleum-based fuels, such as petrol, diesel and aviation fuel. Carbon dioxide is the primary gas emitted in the combustion of these fuels.
Agriculture and land use changes	Deforestation, raising livestock and crops for food. Trees absorb CO_2 from the atmosphere through photosynthesis. Deforestation can, therefore, increase CO_2 levels in the atmosphere. Cows produce large amounts of methane, the most potent greenhouse gas.
Buildings	Operating homes and buildings and managing their wastewater creates CO_2 and CH_4 emissions.

Source: Adapted from www.csiro.au/en/research/environmental-impacts/climate-change/climate-change-qa/sources-of-ghg-gases

BACKGROUND KNOWLEDGE: Australia's energy resources

Australia has a diverse range of energy resources, but the availability, **abundance** and relatively low cost of fossil fuels in Australia mean that they are in high demand — about 95 per cent of Australia's energy requirements are provided by coal, oil and gas. Most of Australia's electricity (75 per cent) is generated by black and brown coal, and 16 per cent is supplied by natural gas. Access to these low-priced fuels is important for our manufacturing industry.

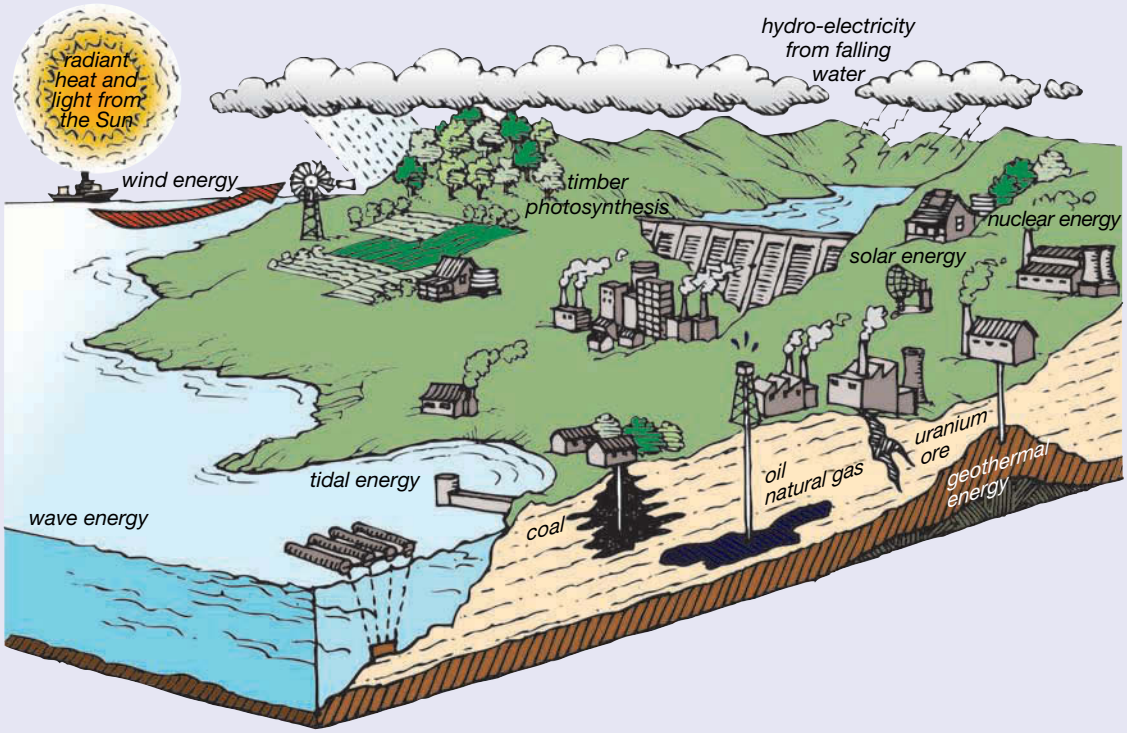
FIGURE 15.3 Projected depletion of world reserves of fossil fuels based on current usage



Although energy can be obtained from the direct combustion of fossil fuels, electricity is a more convenient form of energy. Electricity can be carried from one place to another, switched on and off, and used as an energy source for many different devices.

abundance the relative amount of a substance

FIGURE 15.4 Sources of energy for society



Electricity is sometimes described as a **secondary fuel**, because it is not a natural source of power — it is produced from other fuels or energy sources. At present, most of Australia's electricity is generated by burning fuel at power stations.

Fossil fuels could be made to last longer, and the environmental impacts of fossil fuel combustion could be reduced through a variety of other options. These include:

- Electricity can be generated using renewable energy options such as wind, moving water, wave motion, falling water, tides and biomass reactions.
- Solar energy can be converted to electrical energy using photovoltaic (PV) cells. These are already a popular option because people can install them on the roofs of their houses. PV cells have the added benefit of feeding any excess electricity produced back into the grid.
- Electricity can be generated directly from chemical reactions, as in fuel cells, which are thought to be a possible energy source for the future. You will learn more about fuel cells in unit 3.

Table 15.2 details the energy production in Australia for 2017–18 and the ten-year average annual growth rate. While fossil fuels still comprise approximately 85 per cent of all energy produced in Australia, over the last ten years annual increases in the growth of **renewable energy** have been significant.

TABLE 15.2 Electricity production in Australia by energy source, 2017–18, and projected average annual growth rate

	2017–2018		Average annual growth	
	GWh	Share (per cent)	2017–2018 (per cent)	10 years (per cent)
Fossil fuels	**216 497**	**82.9**	**–0.5**	**–0.6**
Black coal	121 702	46.6	2.9	–0.5
Brown coal	36 008	13.8	–17.3	–0.5
Gas	53 882	20.6	6.8	4.1
Oil	4904	1.9	–7.0	4.2
Renewables	**44643**	**17.1**	**10.4**	**10.2**
Hydro	16021	6.1	–1.6	3.4
Wind	15174	5.8	20.5	16.6
Bioenergy	3518	1.3	0.5	2.6
– Bagasse	1425	0.5	–0.7	na
– Wood, woodwaste	315	0.1	–11.2	na
– Municipal, industrial waste	95	0.0	25.7	na
– Sulphite lyes, biofuels	429	0.2	–2.9	na
– Landfill biogas	1027	0.4	5.9	na
– Sludge biogas	226	0.1	1.6	na
Solar PV	9930	3.8	23.0	58.7
– Small scale	8922	3.4	20.6	56.8
– Large scale	1008	0.4	49.8	na
Total	**261140**	**100**	**1.2**	**0.6**

Source: www.energy.gov.au/sites/default/files/australian_energy_statistics_2019_energy_update_report_september.pdf

15.2.2 The natural and enhanced greenhouse effects

The natural greenhouse effect helps to keep the Earth at the appropriate temperature to support life. It begins when infrared radiation from the Sun strikes the Earth and warms its surface, which then radiates heat energy back into space. Gases in the atmosphere known as greenhouse gases (including carbon dioxide, CO_2, methane, CH_4, nitrous oxide, N_2O, ozone, O_3, and water vapour, H_2O) absorb some of this heat radiation, so the air warms up. Then, the molecules in the air radiate this energy back into space or down to Earth (see figure 15.5).

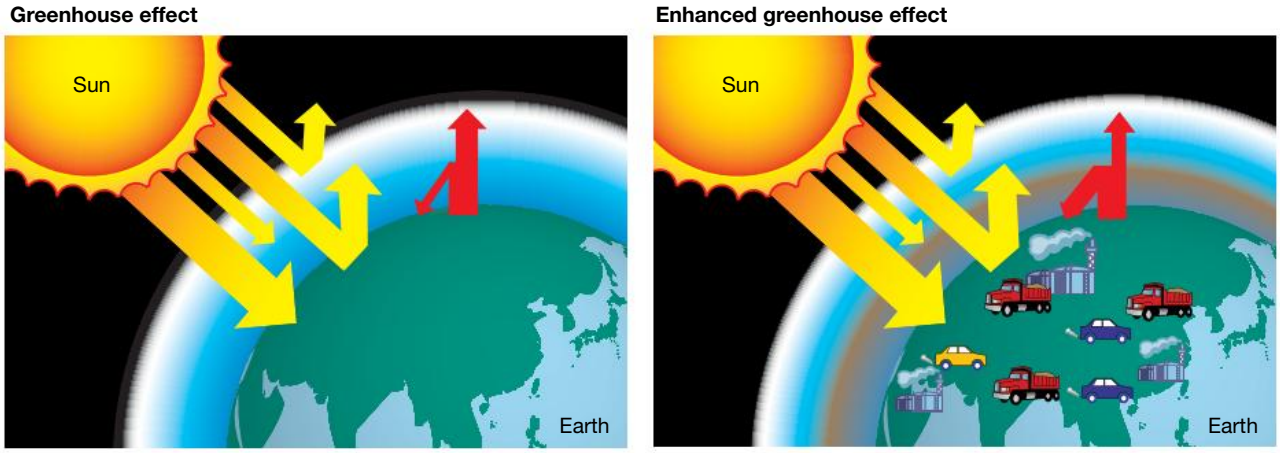
secondary fuel a fuel that is produced from another energy source

renewable energy energy sources that can be produced faster than they are used

Unfortunately, human activities have led to an increase in the amount of greenhouse gases, so that more heat is absorbed, which continues to adversely affect weather and climate. This results in an enhanced greenhouse effect, causing **global warming** and **climate change** (figure 15.6). Global warming specifically refers to the Earth's rising temperature, due mainly to the increasing concentrations of greenhouse gases in the atmosphere (figure 15.7). Climate change is a broader and more accurate term that encompasses the side effects of global warming and refers to changes in various measures of climate over a long period of time.

FIGURE 15.5 The greenhouse effect allows some heat to be trapped in the atmosphere, maintaining a constant temperature.

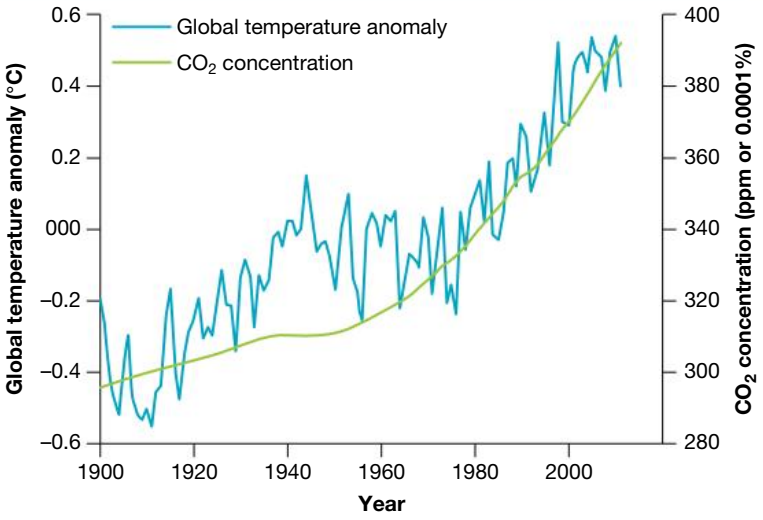
FIGURE 15.6 Excess production of greenhouse gases means the atmosphere retains more heat, increasing the average temperature of the Earth.



Greenhouse gases absorb more energy than others and contribute to global warming in the atmosphere. Carbon dioxide is the major greenhouse gas emitted by human activities and is generated during transportation, industrial processes, land use change and energy production.

Carbon dioxide is essential to life because plants absorb it, enabling them to manufacture their own sugars and produce oxygen through photosynthesis. The quantity of carbon dioxide on Earth remained almost constant for thousands of years, but now we are producing more carbon dioxide than can be converted into oxygen (figure 15.7).

FIGURE 15.7 Global temperature variation and atmospheric CO_2 concentration against time



As the quantity of carbon dioxide increases, the amount of heat retained by the atmosphere increases, thereby contributing to the greenhouse effect, creating the enhanced greenhouse effect.

global warming a gradual increase in the overall temperature of the Earth's atmosphere

climate change long-term shifts in temperatures and weather patterns

elog-1866

tlvd-4495

EXPERIMENT 15.1 online only

Modelling the greenhouse effect

Aim

To understand how the greenhouse effect warms the Earth's atmosphere

Global warming potential of gases

The ability of a gas to absorb and radiate energy is combined with its natural lifespan in the atmosphere to determine its **global warming potential (GWP)**. GWP values of gases are a comparison to carbon dioxide. One tonne of carbon dioxide has a GWP of 1 over a 100-year period. The same amount of another gas over the same time period can be compared to carbon dioxide. Methane's GWP, for example, is over 20 times greater than that of carbon dioxide, because it is better at absorbing heat, making it a more potent greenhouse gas. Nitrous oxide's GWP is over 200 times greater than that of carbon dioxide. However, carbon dioxide is the most significant greenhouse gas due to the large volumes in which it is produced. This is why GWP is a measure against carbon dioxide.

FIGURE 15.8 The impacts of greenhouse gas emissions by humans on the global climate are resulting in significant changes to ecosystems.

Water vapour is the most abundant greenhouse gas in the atmosphere but cannot be directly linked to human activities. Water vapour can absorb and radiate back infrared radiation, contributing to the greenhouse effect. However, increases in the Earth's temperature increase the amount of water vapour, creating a feedback loop, whereby human activity is indirectly increasing water vapour's contribution to the greenhouse effect.

Being able to measure gases enables countries to monitor their contribution to climate change. As you will see later in this topic, the principles of stoichiometry can be applied to quantify the effect that energy production, especially through the combustion of fossil fuels, has on the environment.

The 2021 COP26 conference resulted in parties agreeing to make the 2016 Paris Agreement operational, which involves limiting global warming to 1.5 degrees Celsius, when compared to pre-industrial levels. To do this, countries need to estimate their energy requirements, both current and future, and compare the energy sources available for the required energy production. A useful unit in such comparisons is the amount of greenhouse gas produced (usually carbon dioxide) per megajoule of energy released. Typical units are grams per megajoule, g MJ^{-1} and litres per megajoule, L MJ^{-1}. The changes to our energy sources and usage need to be significant, and countries need to act collectively to slow the effects of climate change, because these impacts are felt by all members of ecosystems.

on Resources

Weblinks Global warming potentials
Water vapour as a greenhouse gas
The legacy of our first decisions on climate change

global warming potential (GWP) the energy one tonne of gas will absorb compared to one tonne of carbon dioxide

15.2 Activities

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15.2 Quick quiz on | 15.2 Exercise | 15.2 Exam questions

15.2 Exercise

1. Describe the difference between the natural and the enhanced greenhouse effect.
2. Describe two negative effects of agriculture on greenhouse gas emissions and suggest a way these could be reduced.
3. When installing appliances into a house, why would it be beneficial to opt for those that run on electricity rather than gas?
4. Many people use the terms 'global warming' and 'climate change' interchangeably. Describe the difference between the two terms.
5. Why is using land to grow crops less impactful on greenhouse gas emissions than the same land being used to rear cattle?

15.2 Exam questions

Question 1 (1 mark)

Name the three main greenhouse gases.

Question 2 (1 mark)

MC Human activities leading to an increase of greenhouse gases have caused _____ in the Earth's energy content.

A. a net increase
B. a net decrease
C. no net imbalance
D. a doubling

Question 3 (1 mark)

MC The greenhouse effect involves the trapping of heat (infrared radiation) by what in the atmosphere?

A. Dust
B. Moisture
C. Particles
D. Gases

Question 4 (1 mark)

MC Increased levels of greenhouse gases over natural levels means _____ infrared radiation is absorbed and global temperatures _____.

A. more; rise
B. more; fall
C. less; rise
D. less; fall

Question 5 (2 marks)

How has industrialisation contributed to the enhanced greenhouse effect?

More exam questions are available in your learnON title.

15.3 Gases at standard laboratory conditions (SLC)

KEY KNOWLEDGE

- The definitions of gas pressure and standard laboratory conditions (SLC) at 25 °C and 100 kPa
- Calculations of the molar volume or molar mass of a gas produced by a chemical reaction

Source: VCE Chemistry Study Design (2023–2027) extracts © VCAA; reproduced by permission.

15.3.1 Gas behaviour

Gases have properties and exhibit behaviours that are different from liquids and solids. In any consideration of gas behaviour, the **pressure** each gas exerts, the volume that it occupies, its temperature and the number of gas molecules present in the sample must be determined.

Scientists use the **kinetic molecular theory of gases** to explain observed gas properties. The theory consists of five postulates (or points) that describe how the particles in a gas would look and behave if we could observe them directly. These are applicable to gas samples under 'moderate conditions', which is usually taken to mean pressures that are not much greater than atmospheric pressure or temperatures considerably greater than those at which the gases liquefy.

FIGURE 15.9 The air in a hot air balloon is heated. The molecules move more rapidly, hitting each other in a random chaotic motion. As they move further apart, the density of the air decreases and the balloon rises.

Kinetic molecular theory of gases

The five postulates that make up the kinetic molecular theory of gases can be summarised as follows.

1. Gases are made up of particles moving constantly and at random.
2. Gas particles are very far apart, and the volume of the particles is very small compared with the volume that the gas occupies.
3. The forces of attraction and repulsion between gas particles are practically zero.
4. Gas particles collide with each other and the walls of their container, exerting pressure. The collisions are perfectly elastic. This means that no kinetic energy is lost when they collide.
5. The higher the temperature, the faster the gas particles move, because they have increased kinetic energy.

pressure the force per unit area that one region of a gas, liquid or solid exerts on another

kinetic molecular theory of gases gas particles are in continuous, random motion; collisions between gas particles are completely elastic

15.3.2 Measuring gases

Gas pressure

Gas pressure is the force exerted by particles colliding with the walls of a container. Each tiny collision adds to the others to make up the continuous force called pressure.

The surface of an inflatable airbed, such as the one in figure 15.10, exerts a force and tries to collapse. For the airbed to stay inflated, the particles inside the airbed must be able to exert a large enough force to balance the forces exerted by the surface of the airbed and the external air pressure.

FIGURE 15.10 The particles inside the airbed create pressure.

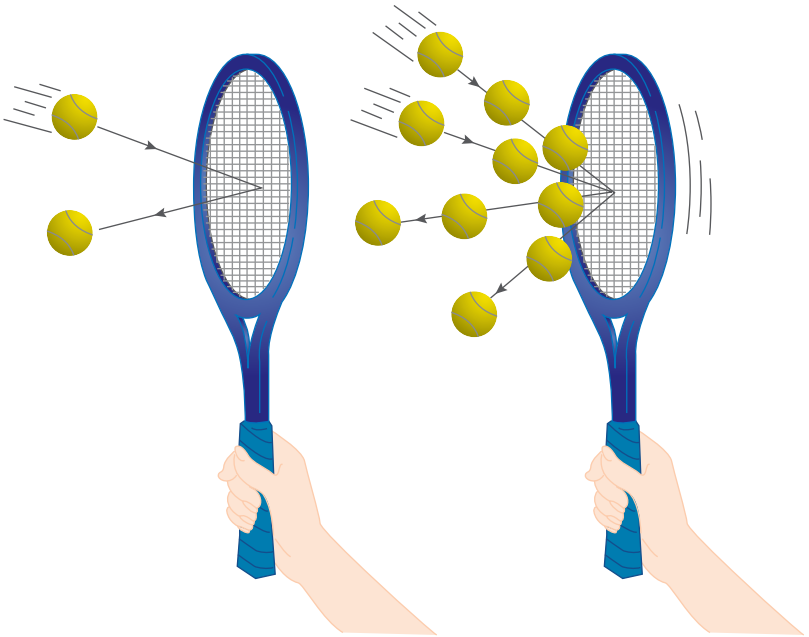
Calculating pressure

Pressure (P) is defined as *the force exerted per unit area.*

$$P = \frac{\text{force}}{\text{area}}$$

The SI unit of pressure is the **pascal (Pa)**, where one pascal is equivalent to a force of one newton exerted over an area of one square metre (Nm^{-2}).

FIGURE 15.11 If a tennis racquet is struck by a tennis ball, a small force is felt. If a stream of balls is fired at the racquet, it is felt as a continuous pressure.



Atmospheric pressure

The pressure of the atmosphere is measured by a barometer. Atmospheric pressure at sea level is 100 000 Pa. This is usually simplified to read 100 kPa (kilopascal).

In weather reports, bars and hectopascals (hPa) are used to measure gas pressure.

$$1 \text{ bar} = 100\,000 \text{ Pa} = 100 \text{ kPa}$$

$$1 \text{ millibar (mb)} = \frac{1}{1000} \text{ bar} = 100 \text{ Pa} = 1 \text{ hectopascal (hPa)}$$

While the SI unit for pressure is the pascal, this is sometimes converted to kilopascals, kPa. Occasionally, another older unit is used called atmospheres, atm.

Converting between units of pressure

$$1000 \text{ Pa} = 1 \text{ kPa}$$

$$100 \text{ k Pa} = 0.987 \text{ atm}$$

pascal (Pa) unit of pressure, measured in newtons per square metre (Nm^{-2})

Temperature

Temperature is the amount of heat present in a substance. The Celsius scale takes the freezing point of water as 0 °C and the boiling point of water as 100 °C. The space between these two fixed points is divided into 100 equal intervals, known as degrees. Temperatures below the freezing point of water are assigned negative values, such as −10 °C.

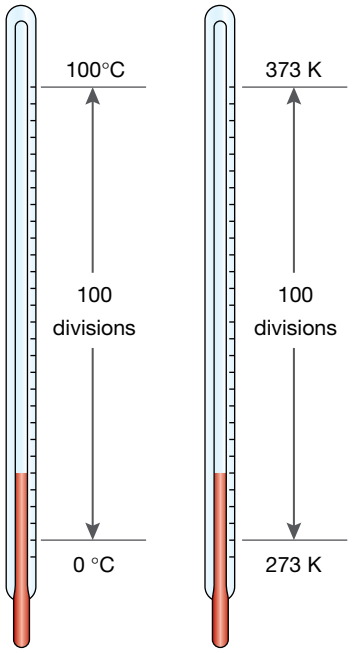
Another temperature scale is the **Kelvin**, or absolute, scale. On the Kelvin scale, the freezing point of water is 273 K and its boiling point is 373 K. A change of one degree on the Celsius scale is the same as that on the Kelvin scale. The zero point on the Kelvin scale, 0 K or **absolute zero**, is −273 °C. Thus, the relationship between the temperature on the Celsius scale and the Kelvin scale is given by the following equations.

Conversion between Celsius and Kelvin

$$K = °C + 273$$
$$°C = K - 273$$

Note that temperatures given in K do not have a ° sign.

FIGURE 15.12 These thermometers show a comparison of the Celsius and Kelvin temperature scales.



Kelvin the SI base unit of thermodynamic temperature, equal in magnitude to the degree Celsius

absolute zero the lowest temperature that is theoretically possible; 0 K

Volume

Volume is the amount of space a substance occupies. The volume of a gas is commonly measured in cubic metres (m^3), litres (L) or millilitres (mL). The conversion between these are as follows.

Conversion between units of volume

$$1\ m^3 = 1000\ L = 1\,000\,000\ mL$$
$$1 m^3 = 10^3\ L = 10^6\ mL$$

tlvd-3942

SAMPLE PROBLEM 1 Converting between units of temperature and units of volume

a. Convert 14 °C to Kelvin.
b. Convert 1.6 L to millilitres.

THINK	WRITE
a. 1. Recall that 0 °C is 273 K.	14 + 273
2. Temperatures given in K do not have a degree (°) sign.	= 287 K
b. 1. Recall that 1 L = 1000 (10^3) mL.	1.6 × 1000 = 1600
2. Because the initial value is given to 2 significant figures, the answer should be given to the same number of significant figures.	= 1.6 × 10^3 mL

PRACTICE PROBLEM 1

a. Convert 127 K to Celsius.
b. Convert 65 mL to litres.

EXTENSION: Laws to describe the behaviour of gases

The behaviour of gases has been studied for centuries. As a result, a number of laws have evolved to describe their behaviour mathematically. These laws are independent of the type of gas — it does not matter what the gas is or if the gas is a mixture such as air, these laws apply in exactly the same way to all gases. Two very useful such laws are **Boyle's law** and **Charles' law.**

Boyle's law for pressure–volume changes

Boyle's law is named after the English physicist and chemist who discovered the relationship between pressure and volume for a sample of gas. It states that, for a fixed amount of gas at constant temperature, pressure is inversely proportional to volume. Mathematically, this can be represented as:

$$P \propto \frac{1}{V}$$

where $\propto$ means 'is proportional to'.

From this relationship it can be stated that PV = a constant value.

If P_1V_1 = a constant value and P_2V_2 = a constant value, then:

$$P_1V_1 = P_2V_2$$

Boyle's law describes the relationship between pressure and volume of a gas, such that for a fixed amount of gas at constant temperature, pressure is inversely proportional to volume; $P_1V_1 = P_2V_2$

Charles' law describes the relationship between temperature and volume of a gas, such that for a given amount of gas at constant pressure, volume is directly proportional to the absolute temperature; $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Charles' law for temperature–volume changes

The relationship between temperature and volume was first identified by the French scientist Jacques Charles, after whom this law is named. While the expansion of all materials with increasing temperature is well known, Charles subjected constant amounts of various gases (at constant pressure) to changes in temperature, each time making accurate measurements of the resulting volume.

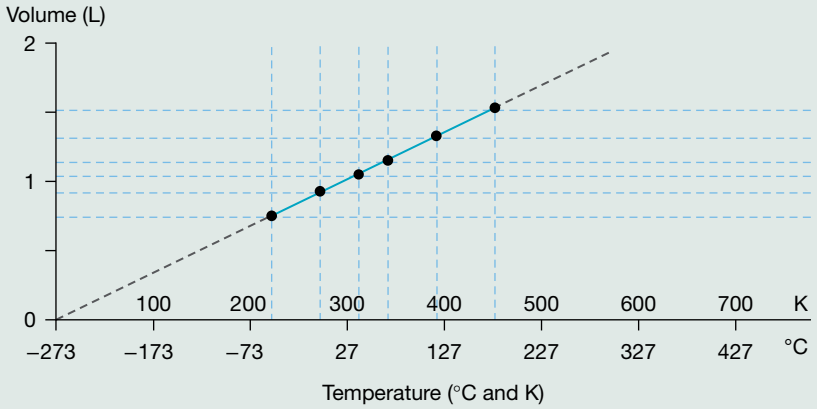
To understand this law better, consider a typical set of results as shown in table 15.3.

TABLE 15.3 Sample results showing the relationship between the temperature and volume of a gas

Temperature (°C)	Volume (L)
–50	0.75
0	0.92
40	1.05
70	1.15
120	1.32
180	1.52

If these results are graphed a pattern emerges, as shown in figure 15.13.

FIGURE 15.13 Representation of Charles' law showing the two temperature scales, Kelvin and degrees Celsius.



If the Celsius scale is used to measure temperature, a linear relationship is observed. However, this is not a directly proportional relationship because the graph does not pass through the origin. If the Kelvin scale is used, it does. We can, therefore, state Charles' law as follows.

For a given amount of gas at constant pressure, volume is directly proportional to the absolute temperature.

Mathematically, this can be represented as $V \propto T$ (where T is the absolute temperature), from which it can be stated that $\frac{V}{T}$ = a constant value.

If $\frac{V_1}{T_1}$ = a constant value and $\frac{V_2}{T_2}$ = a constant value, then:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

You will notice that figure 15.13 implies a 'lowest possible temperature' — the temperature at which it intersects the *x*-axis. This temperature is –273.15 °C or 0 K and is known as absolute zero.

elog-1867

EXPERIMENT 15.2 online only

The relationship between pressure and volume

Aim

To investigate the relationship between pressure and volume for a fixed amount of gas at constant temperature

Standard laboratory conditions (SLC)

The volume of a gas sample is sensitive to both temperature and pressure (refer to preceding extension box). This makes the comparison of gas volumes tricky. In order to make these comparisons easier, scientists have established sets of standard conditions. These are accepted worldwide and enable gas volumes to be compared meaningfully without temperature or pressure having unwanted effects.

Most experiments are carried out in the laboratory, so the standard laboratory conditions are a temperature of 25 °C (298 K) and a pressure of 100 kPa.

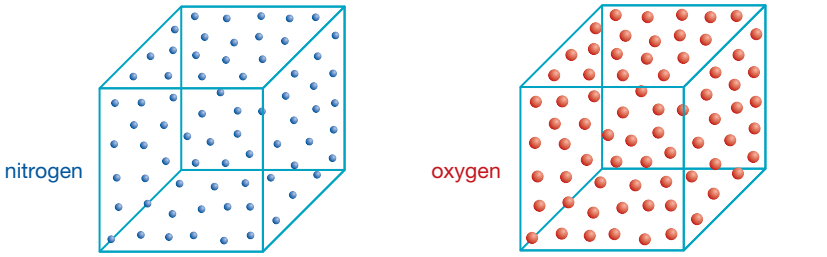
15.3.3 Molar gas volume and Avogadro's hypothesis

Amedeo Avogadro was an Italian scientist who put forward the hypothesis that 'equal volumes of all gases measured at the same temperature and pressure contain the same number of particles'. This means that, if two gases have the same temperature, pressure and volume, they must contain the same number of moles.

It has been found that 1 mole of any gas at SLC occupies a volume of 24.8 L. This volume is called the **molar gas volume** and means that one mole of any gas occupies 24.8 L at 25 °C and 100.0 kPa. This constant is included in table 4 of the VCE Chemistry Data Book.

molar gas volume the volume occupied by a mole of a substance at a given temperature and pressure; at SLC, one mole of gas occupies 24.8 litres

FIGURE 15.14 Under the same conditions of temperature and pressure, the volume of a gas depends only on the number of molecules it contains, and not on what the particles are.



V = 24.8 L
T = 298 K
P = 100.0 kPa
6.02×10^{23} molecules = 1 mol N_2

V = 24.8 L
T = 298 K
P = 100.0 kPa
6.02×10^{23} molecules = 1 mol O_2

The molar volume of a gas varies with temperature and pressure but, at any given temperature and pressure, it is the same for all gases. A direct relationship exists between the number of moles of a gas (n), its molar volume (V_m) and in the actual volume (V); where V is measured in litres.

The relationship between moles and volume of gases

At a given temperature and pressure:

$$n = \frac{V}{V_m}$$

which at SLC is:

$$n_{SLC} = \frac{V}{24.8}$$

FIGURE 15.15 A mole of hydrogen gas occupies the same volume as a mole of oxygen, but because hydrogen weighs less than oxygen, it floats upwards in the air.

tlvd-0663

SAMPLE PROBLEM 2 Calculating the mass of a gas at SLC

Find the mass of 1556.5 mL of H_2 gas that was collected at SLC.

THINK	WRITE
1. Identify the given and unknown quantities and compare the units given to those required. Convert the volume to litres.	$V = \frac{1556.5 \text{ mL}}{1000} = 1.5565 \text{ L}$ $m = ?$
2. Determine the number of moles of H_2 gas using the formula $n = \frac{V}{V_m}$.	$n = \frac{V}{24.8}$ $= \frac{1.5565}{24.8}$ $= 0.06276 \text{ mol}$
3. Determine the mass (m) of 0.06276 moles of H_2 gas using the formula $n = \frac{m}{M}$. **TIP:** The formulas $n = \frac{V}{V_m}$ and $n = \frac{m}{M}$ can be found in table 3 of the VCE Chemistry Data Book.	$n(H_2) = \frac{m}{M}$ $0.0628 = \frac{m}{2.0}$ $m = 2.0 \times 0.0627$ $m = 0.126 \text{ g}$
4. Give your answer to 2 significant figures.	The mass of H_2 gas collected was 0.13 g.

PRACTICE PROBLEM 2

a. 5345 mL of a gas was collected at SLC and weighed. Its mass was 9.50 g. Find the molar mass of the gas.

b. Given that the gas is one of the main constituents of air, identify the gas.

15.3 Activities

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15.3 Quick quiz on | 15.3 Exercise | 15.3 Exam questions

15.3 Exercise

1. Convert the following temperatures from Kelvin to Celsius or Celsius to Kelvin.
   a. 300 K b. 427 K c. 173 K
   d. 119 °C e. –200 °C
2. Use the kinetic molecular theory of gases to explain why gases exert pressure.
3. Calculate the number of moles of the following gases at SLC.
   a. 75 mL of methane, CH_4 b. 2.56 L of hydrogen, H_2 c. 250 mL of nitrogen, N_2

4. Calculate the volume of the following gases at SLC.
   a. 1.53 mol hydrogen, H_2
   b. 13.6 g of methane, CH_4
   c. 2.5×10^{30} molecules of nitrogen, N_2
5. Calculate the mass of the following gas samples at SLC.
   a. 150 mL of oxygen, O_2
   b. 4.5 L of carbon dioxide, CO_2

15.3 Exam questions

Question 1 (1 mark)

MC A and B are two identical flasks at standard laboratory conditions (SLC). Flask A contains 5.0 g of N_2 gas and flask B contains 11.4 g of an unknown gas.

Identify the gas in flask B.

A. H_2 B. SO_2 C. HBr D. C_4H_{10}

Question 2 (1 mark)

MC When hydrochloric acid is added to aluminium sulfide, the highly toxic gas hydrogen sulfide is evolved. The equation for this reaction is

$Al_2S_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2S(g)$.

If excess hydrochloric acid is added to 0.100 mol of aluminium sulfide, what will the volume of hydrogen sulfide produced at standard laboratory conditions (SLC) be?

A. 1.65 L B. 4.96 L C. 7.44 L D. 14.9 L

Question 3 (2 marks)

What is the mass (in kg) of 850 L of carbon monoxide gas measured at SLC?

Question 4 (3 marks)

A 0.953 L quantity of a gas measured at SLC has a mass of 3.20 g. What is the molar mass of the gas? What is the gas?

Question 5 (10 marks)

Consider the sample results in table 15.4, which show the relationship between pressure and volume.

TABLE 15.4 Sample results showing the relationship between the pressure and volume of a gas

Pressure (kPa)	Volume (L)
120	0.261
145	0.218
162	0.193
180	0.171
200	0.159
216	0.145
240	0.130
258	0.120

a. Convert each pressure measurement (P) into its reciprocal $\left(\frac{1}{P}\right)$. **(2 marks)**

b. Plot a graph of pressure versus volume. Put pressure on the horizontal (x) axis and volume on the vertical (y) axis. **(3 marks)**

c. Plot a second graph with the volume on the vertical axis and $\frac{1}{P}$ on the horizontal axis, and extrapolate the graph to the origin. **(3 marks)**

d. Compare and account for the two graphs. **(2 marks)**

More exam questions are available in your learnON title.

15.4 Calculations using the ideal gas equation and stoichiometry

KEY KNOWLEDGE

- Calculations using the ideal gas equation ($PV = nRT$), limited to the units kPa, Pa, atm, mL, L, °C and K (including unit conversions)
- The use of stoichiometry to solve calculations related to chemical reactions involving gases (including moles, mass and volume of gases)

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15.4.1 The ideal gas equation

The **ideal gas equation** is used to describe the behaviour of a hypothetical ideal gas to help predict the behaviour of a real gas. It is obtained by combining Boyle's law and Charles' law (refer to extension box in section 15.3.2) with Avogadro's hypothesis.

Ideal gas equation

Ideal gas equation:

$$PV = nRT$$

The **molar gas constant**, R, has a value of 8.31 J K^{-1} mol^{-1} when:

- *pressure* (P) is measured in kilopascals (kPa)
- *volume* (V) is measured in litres (L)
- *temperature* (T) is measured in degrees Kelvin (K)
- the *quantity* of gas is measured in moles (n).

ideal gas equation $PV = nRT$, where pressure is in kilopascals, volume is in litres and temperature is in Kelvin

molar gas constant R, the constant of the universal gas equation; R = 8.31 J mol^{-1} K^{-1} when pressure is measured in kPa, volume is measured in L, temperature is measured in K and the quantity of the gas is measured in moles (n)

tlvd-0664

SAMPLE PROBLEM 3 Calculating the volume of a gas using the ideal gas equation

Find the volume of 6.30 mol of carbon dioxide gas at 23.0 °C and 550 kPa pressure.

THINK	WRITE
1. Use the molar gas constant (R) taken from table 4 in the VCE Chemistry Data Book.	$R = 8.31$ J K^{-1} mol^{-1}
2. Identify the given and unknown quantities and compare the units given to those required. No unit conversion is required. **TIP** The majority of incorrect calculations when using the ideal gas equation arise from not converting other units (e.g. Pa, mL, mass or °C) in the question into the correct units of kPa, L, n and K.	$P = 550$ kPa $n = 6.30$ mol $T = 23.0 + 273 = 296$ K $R = 8.31$ J $K^{-1}mol^{-1}$ $V = ?$L
3. Rearrange the ideal gas equation to find volume, V. The ideal gas equation can be found in table 3 of the VCE Chemistry Data Book.	$PV = nRT$ $V = \frac{nRT}{P}$ $= \frac{6.30 \times 8.31 \times 296}{550}$
4. Give your answer to 3 significant figures.	$V = 28.2$ L

PRACTICE PROBLEM 3

Calculate the volume of gas, in litres, occupied by 12.8 mol of CH_4 at 9.87 atm and 60 °C.

tlvd-3493

SAMPLE PROBLEM 4 Calculating the pressure of a gas using the ideal gas equation

Calculate the pressure, in atm, of a 6.0 L container containing 13.2 mol of CO_2 at 27 °C.

THINK	WRITE
1. Identify the given and unknown quantities, and compare the units given to those required. The temperature needs to be converted to Kelvin before substituting in values.	$P = ?$ $V = 6.0\,\text{L}$ $n = 13.2\,\text{mol}$ $R = 8.31\,\text{JK}^{-1}\text{mol}^{-1}$ $T = 27°\text{C}$ $= 27 + 273$ $= 300\,\text{K}$
2. Rearrange the ideal gas equation to make P the subject and substitute all values to find the pressure, which will be in kPa.	$P = \frac{nRT}{V}$ $= \frac{13.2 \times 8.31 \times 300}{6.0}$ $= 5484.8\,\text{kPa}$
3. Convert to atm, given that 100 kPa = 0.987.	$\frac{5484.6}{100} \times 0.987 = 54\,\text{atm}$

PRACTICE PROBLEM 4

Calculate the pressure, in Pa, of 1.60 mol of carbon dioxide gas that occupies 5300 mL at 293 K.

15.4.2 The use of stoichiometry in chemical reactions involving gases

Mass–volume calculations

Many chemical reactions involve gases. For example, barbecues, furnaces and engines burn fuel and produce carbon dioxide and water vapour. In order to ensure that the reactions are efficient and as complete as possible, it is essential to calculate the volumes or masses of the gases or fuels required.

tlvd-3494

SAMPLE PROBLEM 5 Calculating the volume of products formed when mass of reactants is known

Methane, CH_4, combusts in excess oxygen, O_2, at SLC to form carbon dioxide, CO_2, and water, H_2O. If 3.2 g of methane is consumed, what volume of carbon dioxide is produced?

THINK	WRITE
1. Write a balanced symbol equation to identify the mole ratios of the reactants and products.	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

2. Calculate the mol of methane consumed using the formula $n = \frac{m}{M}$.	$n(CH_4) = \frac{3.2}{16} = 0.2 \text{ mol}$
3. From the equation determined in step 1, identify the mole ratio of $CH_4 : CO_2$. The ratio is 1 : 1.	$n(CO_2) = n(CH_4)$
4. If 0.2 mol of methane is consumed, 0.2 mol of carbon dioxide was produced.	$n(CO_2) = 0.20 \text{ mol}$
5. At SLC, calculate the volume of CO_2 formed by rearranging the formula $n = \frac{V}{V_m}$. At SLC $V_m = 24.8 \text{ L mol}^{-1}$.	$V(CO_2) = n \times V_m$ $= 0.2 \times 24.8$ $= 5.0 \text{ L}$

PRACTICE PROBLEM 5

Ethane, C_2H_6, combusts in excess oxygen, O_2, at SLC to form carbon dioxide, CO_2, and water, H_2O. If 20 g of carbon dioxide is produced, how many mL of ethane was consumed?

Volume–volume calculations

Consider the reaction between nitrogen and hydrogen gas.

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

This equation says that when one mole of N_2 reacts with three moles of H_2, two moles of ammonia will be produced. According to Avogadro's hypothesis, if the gases are at the same pressure and temperature, their *molar* ratios are equal to their *volume* ratios. Therefore, we can use volumes instead of moles and say that 10 mL of N_2 reacts with 30 mL of H_2 to form 20 mL of ammonia.

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

$$1 \text{ mol} + 3 \text{ mol} \rightarrow 2 \text{ mol}$$

$$1 \text{ vol} + 3 \text{ vol} \rightarrow 2 \text{ vol}$$

$$10 \text{ mL} + 30 \text{ mL} \rightarrow 20 \text{ mL}$$

Volume–volume calculations

$$\frac{V(\text{unknown})}{V(\text{known})} = \frac{\text{coefficient of unknown}}{\text{coefficient of known}}$$

tlvd-3495

SAMPLE PROBLEM 6 Calculating the volume of products formed when volume of reactants is known

If 100 L of ethene is burned according to

$$C_2H_4\,(g) + 3O_2\,(g) \rightarrow 2CO_2\,(g) + 2H_2O\,(g)$$

calculate the volume of the following.

a. Carbon dioxide produced

b. Oxygen consumed

(Assume all gas volumes are measured at the same temperature and pressure.)

THINK

1. Because all gas volumes are measured at the same temperature and pressure, the equation may be interpreted in terms of volume ratios.

WRITE

a. $V(CO_2) = 2\ V(C_2H_4)$
Therefore, $V(CO_2) = 2 \times 100 = 200$ L.

b. $V(O_2) = 3V(C_2H_4)$
Therefore, $V(O_2) = 3 \times 100 = 300$ L.

PRACTICE PROBLEM 6

Methane gas burns in air at room temperature and pressure, according to the following equation.

$$CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2H_2O\,(g)$$

If 25 mL of methane is burned at room temperature and pressure, find the volumes of the following reactants and products.

a. Oxygen

b. Carbon dioxide

c. Water

Stoichiometry using the ideal gas equation

The equation $n = \frac{V}{V_m}$ can only be used when reactions involving gases occur at SLC. When not at SLC, the ideal gas equation needs to be used. The principles of stoichiometry are the same — use mole ratios and the known quantity to determine the desired final quantity, be that pressure, volume or temperature.

Stoichiometry problems can be solved using the following steps.

1. Unless provided, write a balanced chemical reaction for the equation and identify the known (given) and unknown (required) quantities of substance.
2. Convert any given values into SI units. Temperature should be in K, pressure in kPa, volume in L.
3. If not provided, calculate the number of moles of the known quantity, and use the molar ratios from the chemical equation to calculate the moles of the required substance.
4. Calculate the quantity (such as volume) of the required substance.

tlvd-0665

SAMPLE PROBLEM 7 Finding the volume of a fuel when given the mass of a product

In a gas barbecue, propane is burned in oxygen to form carbon dioxide and water vapour. If 22.0 g of CO_2 is collected and weighed, find the volume of propane at 200 °C and 1.00×10^5 Pa. The balanced equation for this reaction is $C_3H_8\,(g) + 5O_2\,(g) \rightarrow 3CO_2\,(g) + 4H_2O\,(g)$.

THINK

WRITE

1. Identify the given and unknown quantities and compare the units given to those required. Convert the temperature units to Kelvin.

$$P = 1.00 \times 10^5 \text{ Pa}$$
$$m(CO_2) = 22.0 \text{ g}$$
$$T = 200 \text{ °C} + 273 = 473 \text{ K}$$
$$V = ? \text{ L}$$

2. Calculate the number of moles of the known quantity of substance using the following formula, found in table 3 of the VCE Chemistry Data Book: $n = \frac{m}{M}$.

$$n = \frac{m}{M}$$
$$M \text{ of } CO_2 = 12.0 + (2 \times 16.0)$$
$$= 44.0 \text{ g mol}^{-1}$$
$$n(CO_2) = \frac{22.0}{44.0}$$
$$= 0.500 \text{ mol}$$

3. Use the combustion equation given to find the molar ratio of unknown to known quantities, and calculate the number of moles of the required substance.

   From the given reaction, we know that three moles of CO_2 are produced from one mole of C_3H_8.

$$n(C_3H_8) = \frac{n(CO_2)}{3}$$
$$= \frac{0.500}{3}$$
$$= 0.1666$$

4. Rearrange the ideal gas equation, $PV = nRT$, to determine the volume of the propane.

$$V(C_3H_8) = \frac{nRT}{P}$$
$$= \frac{0.1666 \times 8.31 \times 473}{100} \times 100$$

5. Give your answer to 3 significant figures.

$$V = 6.55 \text{ L}$$

PRACTICE PROBLEM 7

Using the following equation, calculate the mass of hexane that is required to produce 100 L of carbon dioxide gas at SLC.

$$2C_6H_{14}\,(l) + 19O_2\,(g) \rightarrow 12CO_2\,(g) + 14H_2O\,(g)$$

15.4 Activities

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15.4 Quick quiz on | **15.4 Exercise** | **15.4 Exam questions**

15.4 Exercise

1. a. Calculate the volume occupied by 42.0 g of nitrogen gas at a pressure of 200 kPa and a temperature of 77 °C.
   b. A 5.00 L balloon contains 0.200 mol of air at 120 kPa pressure. What is the temperature of air in the balloon?
2. Calculate the volume of gas, in litres, occupied by the following.
   a. 3.5 mol of O_2 at 100 kPa and 50 °C
   b. 6.5 g of Ar at 50 kPa and 100 °C
   c. 0.56 g of CO_2 at 50 atm and 20 °C
   d. 1.3×10^{-3} g He at 60 kPa and –50 °C
   e. 1.5×10^{21} molecules of Ne at 40 kPa and 200 °C
3. At high temperatures, such as those in a car engine during operation, atmospheric nitrogen burns to produce the pollutant nitrogen dioxide, according to the following equation.

$$N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$$

   a. If 20 mL of nitrogen is oxidised, calculate the volume of oxygen needed to produce the pollutant. Assume that temperature and pressure remain constant.
   b. What is the initial volume of reactants in this combustion reaction?
   c. What is the final volume of products in the reaction?
   d. Is there an overall increase or decrease in the volume of gases on completion of the reaction?
4. Magnesium reacts with hydrochloric acid according to the following equation.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

   What mass of magnesium, when reacted with excess hydrochloric acid, would produce 5.0 L of hydrogen gas, measured at 26.0 °C and 1.2 kPa?
5. Oxygen gas can be prepared in the laboratory by the decomposition of potassium nitrate according to the following equation.

$$2KNO_3(s) \rightarrow 2KNO_2(s) + O_2(g)$$

   When 14.5 L of O_2 is formed at 1.00 atm and 25.0 °C, what mass of KNO_2 is also formed?

15.4 Exam questions

Question 1 (1 mark)

MC The equation for photosynthesis is

$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

What volume of carbon dioxide, at SLC, is required to produce 200 mL of oxygen?

A. 200 mL
B. 1200 mL
C. 33.3 mL
D. 60.0 mL

Question 2 (1 mark)

MC A helium balloon is inflated to a volume of 3.84 L and a pressure of 1.0 atm at a temperature of 23 °C.

What is the amount, in mol, of helium in the balloon?

A. 0.16
B. 6.3
C. 2.0
D. 0.002

Question 3 (2 marks)

Dilute nitric acid reacts with anhydrous sodium carbonate, Na_2CO_3, to produce carbon dioxide gas.

$$2HNO_3(aq) + Na_2CO_3(s) \rightarrow 2NaNO_3(aq) + CO_2(g) + H_2O(l)$$

In an experiment carried out at SLC, what volume of CO_2 was produced from 18 g of anhydrous Na_2CO_3?

Question 4 (4 marks)

10 g of methane, CH_4, is injected into a sealed, evacuated 6.0 L vessel at 280 kPa.

After another 5 g of methane is injected into the vessel, what is the temperature change in °C if the pressure is kept constant?

Question 5 (3 marks)

Airbags are an important safety feature of today's cars. The airbag contains a mixture of solid sodium azide, NaN_3, and potassium nitrate, KNO_3. In the event of an accident, trip sensors send an electric signal to an igniter. The heat generated causes the reactants to decompose completely according to the following equation.

$$10NaN_3(s) + 2KNO_3(s) \rightarrow 5Na_2O(s) + K_2O(s) + 16N_2(g)$$

The airbag inflates to its maximum capacity of 62.0 L at a pressure of 100 kPa. If 98.0 g of sodium azide is contained in the car's airbag, at what temperature does the airbag inflate? Give your answer in degrees celsius.

$$M(NaN_3) = 65.0 \text{ g mol}^{-1}$$

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15.5 Review

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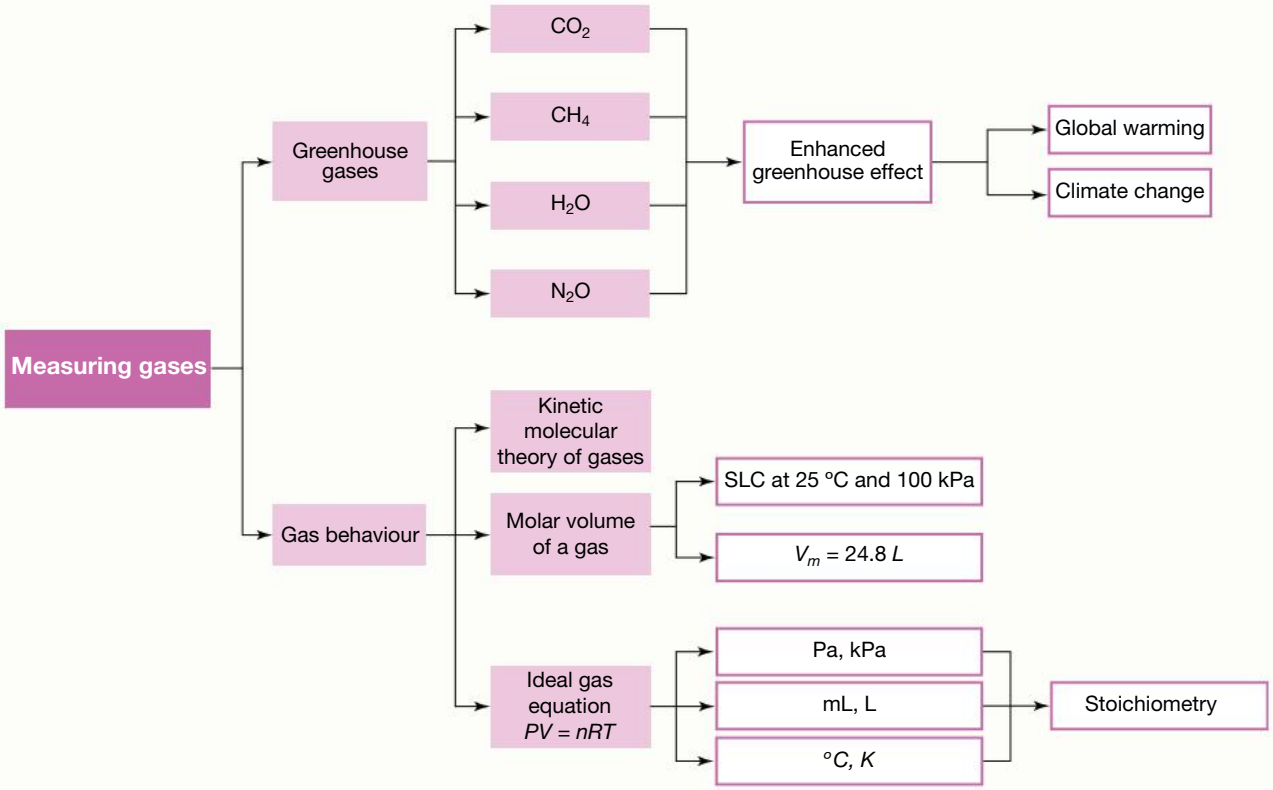
15.5.1 Topic summary



15.5.2 Key ideas summary

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15.5.3 Key terms glossary

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on Resources

Solutions	Solutions —Topic 15 (sol-0814)
Practical investigation eLogbook	Practical investigation logbook — Topic 15 (elog-1620)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 15 (doc-37095) Key ideas summary — Topic 15 (doc-37096)
Exam question booklet	Exam question booklet — Topic 15 (eqb-0096)

15.5 Activities

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15.5 Review questions

1. 'Coal, gas and oil should be made more expensive to deter people from wasting them.' Discuss this statement, offering at least two arguments for and against the proposal.
2. a. Assuming that each kilowatt hour of electricity produces about 1.44 kilograms of carbon dioxide, examine your family's electricity bills and determine how much carbon dioxide your family produces annually.
   b. Explain why the production of carbon dioxide might lead to environmental problems.
   c. Suggest three ways in which your family could decrease electricity use to contribute to the minimisation of carbon dioxide release into the atmosphere.
3. Convert the following.
   a. 392 K to Celsius
   b. 27 °C to Kelvin
   c. 56 mL to litres
4. The pressure on a gas remains constant. Its volume is 700 mL. The temperature is 27 °C. Calculate the temperature needed to change the volume to the following.
   a. 14.0 mL
   b. 420.0 mL
5. An empty 200 mL flask has a mass of 84.845 g. It is filled with a gas at 17.0 °C and 102.6 kPa pressure and then weighs 85.084 g. Calculate the molar mass of the gas. What is the gas?
6. Calculate the number of moles of gas present in each of the following gas samples.
   a. 32.3 L of argon at 102.0 kPa and 15 °C
   b. 24.3 L of nitrogen at 13.2 atm and 35 °C

15.5 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC Which of the gases listed is the most potent greenhouse gas?

A. Carbon dioxide, CO_2
B. Water vapour, H_2O
C. Methane, CH_4
D. Nitrous oxide, N_2O

Question 2

MC The absorption of heat radiation by greenhouse gases causes the air to warm up. What does this statement describe?

A. The natural greenhouse effect
B. The enhanced greenhouse effect
C. Global warming
D. Climate change

Question 3

MC A rigid container holds a fixed volume of gas at a certain temperature and pressure. In order to double the pressure of the gas inside the container, what could a person do?

A. Halve the amount of gas in the container
B. Halve the amount of gas, but double the absolute temperature
C. Double the amount of gas and double the absolute temperature
D. Double the absolute temperature

Question 4

MC The following reaction was carried out at SLC.

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

If 40 mL of hydrogen reacted completely with chlorine, what volume of hydrochloric acid was produced?

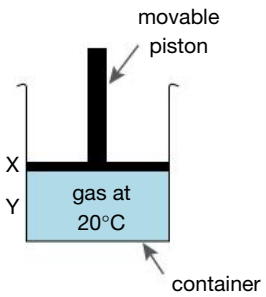
A. 40 mL **B.** 60 mL **C.** 20 mL **D.** 80 mL

Question 5

MC The diagram shows a container filled with gas and sealed by a movable piston. There is sufficient gas to support the piston at point X when the temperature is 20 °C.

The piston is pushed down to point Y and locked, so that the volume of the container is halved. What will the average number of molecules striking a unit area of the wall of the container per unit time do?

A. Double
B. Halve
C. Remain the same
D. Change in an unpredictable manner



Question 6

MC In the apparatus in question 5, assume that the piston is locked at X. If the gas is heated, what will the pressure it exerts do?

A. Increase
B. Decrease
C. Remain the same
D. Change in an unpredictable manner

Question 7

MC 10 L of hydrogen gas is collected at 110 kPa and 20 °C. It is then compressed into a 2 L container at a pressure of 1100 kPa. What will the new temperature be?

A. 40 °C
B. 200 °C
C. 273 °C
D. 313 °C

Question 8

MC Tired of working in laboratories that are either freezing or too hot, Jenny has proposed a third set of standard conditions. She has defined these as a temperature of 15 °C and a pressure of 100 kPa. Under these conditions, what would the molar volume of a gas be?

A. Less than 22.7 L
B. 22.7 L
C. Between 22.7 L and 24.8 L
D. Greater than 24.8 L

Question 9

MC Calculate the volume of gas occupied by 0.560 mol of carbon dioxide gas at 101.5 kPa and 17 °C.

A. 13.3 L
B. 24.6 L
C. 0.78 L
D. 75 mL

Question 10

MC The equation for the complete combustion of ethyne is provided.

$$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$$

What will the combustion of 0.26 g of ethyne at 23 °C and 105 kPa do?

A. Consume 0.62 L of oxygen, O_2
B. Produce 0.47 L of carbon dioxide, CO_2
C. Produce 1.76 g carbon dioxide, CO_2
D. Produce 0.18 L of water, H_2O

Section B — Short answer questions

Question 11 (2 marks)

Use the kinetic molecular theory of gases to explain the following.

a. Why gases exert pressure **(1 mark)**
b. Why, for a fixed amount of gas at constant volume, pressure increases with temperature **(1 mark)**

Question 12 (5 marks)

An aerosol can of deodorant has a volume of 120 mL. The contents exert a pressure of 9.0×10^5 Pa at 27 °C.

a. Calculate the number of moles of gas present in the can. **(2 marks)**

b. How many particles are present in the can of deodorant? **(1 mark)**

c. If the contents of the can are transferred to a 200 mL container, what would be the temperature if the pressure drops to 6.0×10^5 Pa? **(2 marks)**

Question 13 (3 marks)

Each time Nick breathes, he inhales about 400 mL of air. Oxygen makes up about 20 per cent by volume of air. How many oxygen molecules does he inhale in one breath at 25 °C and 1.0×10^4 Pa?

Question 14 (7 marks)

Butan-1-ol (density = 0.81 g mL^{-1}) burns according to the following equation.

$$CH_3CH_2CH_2CH_2OH(l) + 6O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$$

a. When 10.0 mL of butan-1-ol is burned, calculate the following.

i. The mass of water produced **(2 marks)**

ii. The volume of carbon dioxide produced at SLC **(1 mark)**

iii. The volume of carbon dioxide produced at 200 °C and 1.2 atm pressure **(1 mark)**

b. If the butan-1-ol is used to produce 100 mL of carbon dioxide at SLC, calculate the volume of butan-1-ol needed. **(3 marks)**

Question 15 (3 marks)

A number of technologies either exist or are being developed to remove carbon dioxide from flue gases. Some of these are also being investigated for the removal of carbon dioxide directly from the air. One such process uses calcium oxide and steam (at 400 °C). The overall equation for this reaction is as follows.

$$CaO(s) + CO_2(g) \xrightarrow[400\ ^\circ C]{\text{Steam}} CaCO_3(s)$$

Calculate the maximum volume of carbon dioxide at SLC that can be removed per 1.00 tonne of calcium oxide.

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16 Analysis for salts

KEY KNOWLEDGE

In this topic you will investigate:

Analysis for salts

- sources of salts found in water or soil (which may include minerals, heavy metals, organo-metallic substances) and the use of electrical conductivity to assess the salinity and quality of water or soil samples
- quantitative analysis of salts:
  - molar ratio of water of hydration for an ionic compound
  - the application of mass–mass stoichiometry to determine the mass present of an ionic compound
  - the application of colorimetry and/or UV–visible spectroscopy, including the use of a calibration curve, to determine the concentration of ions or complexes in a water or soil sample.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of VCE Chemistry. Experiments and investigations, supported by a **practical investigation eLogbook** and **teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

EXAM PREPARATION

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16.1 Overview

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16.1.1 Introduction

Water is an excellent solvent and, as such, is hardly ever found in a pure state, which increases the diversity of salts found in our soils and waterways. Water and soil analysis is an important step in determining the type and number of contaminants within each to ensure that Australia is meeting the Sustainability Development Goals of the United Nations. This enables statutory bodies such as Victoria's Environment Protection Agency (EPA) to set allowable limits and to impose penalties to protect the environment. It also allows water that might be unsuitable for one purpose to be safely used for another, thus making more efficient use of precious water resources.

FIGURE 16.1 Knowing what is in a water supply may permit its use in some situations but not in others. This allows a more efficient use of water.

In this topic, we focus on the sources of dissolved salts in soils and water — specifically, minerals, heavy metals and organometallic compounds. Analytical methods and instrumentation typically used to analyse for the dissolved contaminants are also explained, including gravimetric analysis, colorimetry and UV–visible spectroscopy.

LEARNING SEQUENCE

16.1 Overview 616
16.2 The sources of salts in soil and water 617
16.3 Quantitative analysis of salts — stoichiometry and molar ratios 624
16.4 Quantitative analysis of salts — colorimetry and UV-visible spectroscopy 642
16.5 Review 652

on Resources

Solutions	Solutions — Topic 16 (sol-0846)
Practical investigation eLogbook	Practical investigation logbook — Topic 16 (elog-1870)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 16 (doc-37913) Key idea summary — Topic 16 (doc-37914)
Exam question booklet	Exam question booklet — Topic 16 (eqb-0125)

16.2 The sources of salts in soil and water

KEY KNOWLEDGE

- Sources of salts found in water or soil (which may include minerals, heavy metals, organo-metallic substances) and the use of electrical conductivity to assess the salinity and quality of water or soil samples

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In subtopic 11.6, we investigated salts as being the products of **neutralisation** reactions. Now we will consider salts that are found in soil and water sources — either naturally or due to human activity — and how we can assess the level of salt (salinity level) to determine overall quality of the soil or water. These salts may include minerals, heavy metals and organometallic substances.

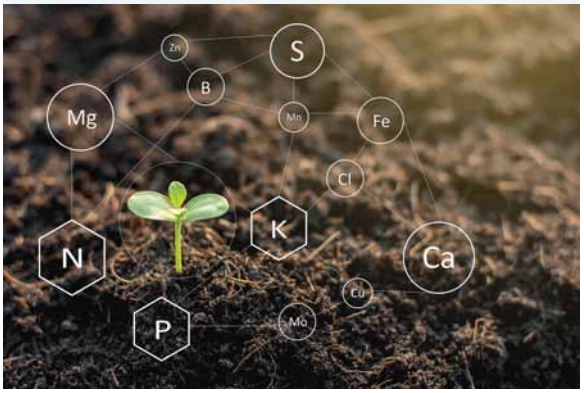
neutralisation chemical reaction in which an acid and a base react to form a salt and water

minerals naturally occurring inorganic elements or salts

16.2.1 Minerals

The salts found naturally in soil and water come from a wide variety of sources. Soluble salts found in soils include calcium, magnesium, potassium, sodium, iron (II and III), zinc, aluminium, ammonium, nitrates, chlorides, sulfates, carbonates and bicarbonates (see figure 16.2).

FIGURE 16.2 An array of minerals can be found in soil, making it suitable for plant growth when the minerals are in their desired quantities.



Many **minerals** in the ground dissolve as the water flows over the ground or filters through layers of rock underneath it. In fact, mineral water that you find on the shelves at the supermarket is obtained from underground water sources and must contain at least 250 ppm (parts per million) of total dissolved solid salts. The minerals you find on the bottle label have previously dissolved directly from the earth and are not added later (figure 16.3).

FIGURE 16.3 The label on a bottle of mineral water shows the many substances that are dissolved in it.

Sodium chloride is the most common salt found in water, and the fact that the oceans contain so much sodium chloride is evidence of this.

Remember that when salts dissolve in an aqueous solution, their ionic bonds break, forming cations (positively charged ions) and anions (negatively charged ions). However, salts do not need to dissolve in large amounts to have a significant effect on how water may be used. Consider the salt calcium carbonate, $CaCO_3$, which dissolves from rocks such as limestone and chalk as water flows over or through them. Calcium carbonate has a very low solubility (0.013 g L^{-1} at SLC). However, when it reacts with rain water, which contains dissolved carbon dioxide, calcium hydrogen carbonate (also known as calcium bicarbonate) is produced, which is much more soluble (~166 g L^{-1}). The equation for this reaction is:

$$CaCO_3(s) + CO_2(aq) + H_2O(l) \rightarrow Ca(HCO_3)_2(aq)$$

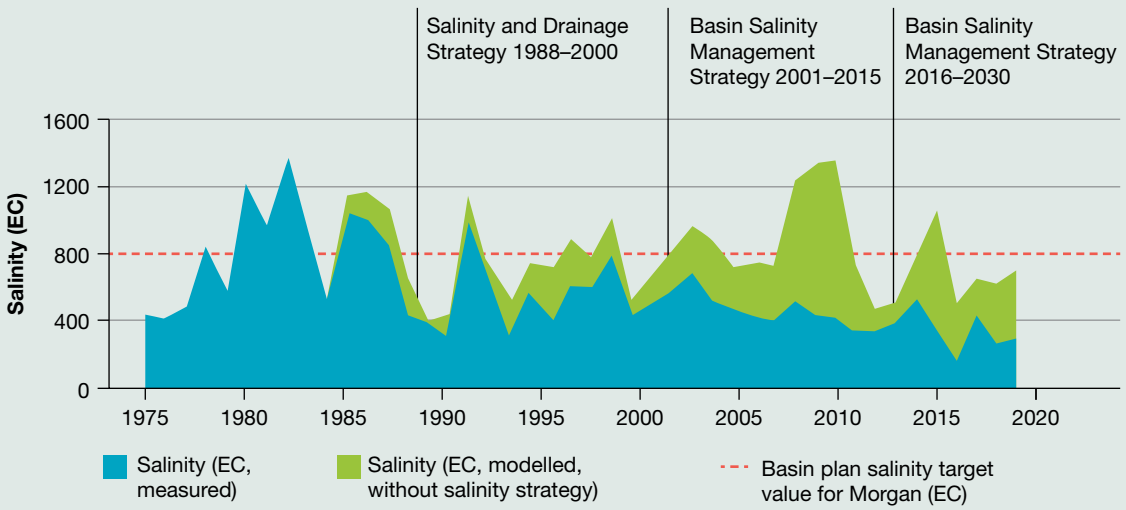
Aquatic shellfish have calcium carbonate in their shells, and climate change and ocean acidification are affecting these organisms (see subtopic 11.7 for more on this).

Have you heard of 'hard water'? This is caused from dissolved calcium ions, Ca^{2+}, and magnesium ions, Mg^{2+}. Consequently, the water is difficult to lather and also unsuitable for use in boilers and other high-temperature applications.

CASE STUDY: Murray–Darling Basin

Human activity can affect the levels of dissolved salts in water and soil. A good example of this is the increasing salinity levels in the Murray River. Until 100 million years ago, the Murray–Darling Basin was covered by a shallow sea. After this sea retreated, the salt deposits left behind were buried underground. Prior to European settlement, the salt would naturally be carried along the waterways to reach the Southern Ocean at the mouth of the Murray River in South Australia. However, since the advent of extensive irrigation, land clearing and agriculture, the extra water soaking into the soil has caused water tables to rise. This has brought large amounts of this previously undisturbed salt to, or near to, the surface — from where it can enter rivers and streams. These eventually flow into the Murray River, increasing its salinity levels. This problem has been met with responses such as improved farming and irrigation systems, and the establishment of a number of salinity management schemes — including the salt interception scheme that removes approximately half a million tonnes of salt each year, regulated by the Murray–Darling Basin Authority (MDBA) and documented within the Basin Salinity Management Strategy 2030. The strategy's focus is on maintaining appropriate salt levels 'to protect economic, environmental, cultural and social values'. Figure 16.4 shows a graph highlighting the measured salinity levels at Morgan in South Australia from nearly 50 years ago and up to 2019, and also compares these to modelled levels without a salinity strategy.

FIGURE 16.4 Measured salinity levels at Morgan in South Australia against modelled salinity levels without a salinity strategy.



Sources of dissolved salts

Dissolved salts in water can come from natural sources as well as from human activity.

16.2.2 Heavy metals

Heavy metals are metallic elements with high relative atomic masses. However, their definition can vary between sources. They can be:

- found naturally in the earth's crust in low abundance (approximately 5 per cent by weight), and then naturally making their way into soil and waterways
- essential for life in trace amounts because they are required for various metabolic processes; however, they must be acquired through diet.

If they occur locally at levels higher than their natural concentrations, heavy metals can:

- contaminate soil and water through various means — both through natural contamination (for example, volcanic eruptions) and as the result of human activity (for example, disposal of industrial waste), as they react with either the water vapour in the atmosphere or with liquid water, becoming soluble cations
- be contaminants of concern for human health because they do not degrade, are cumulative in the human body and can become toxic at high levels; some are known to induce cancer and other health problems
- be converted to organometallic compounds and bioaccumulate in organisms, working their way through the food chain because they are:
  - persistent in the environment
  - fat-soluble compounds (non-polar)
  - difficult to eliminate because they are stored in the fatty tissues of animals.

Humans are exposed to heavy metals mainly through water and food consumption; therefore, close monitoring of drinking water quality is paramount. Limits for acceptable concentrations of heavy metals in water are included in almost all standards, but specifically in the drinking water standards in Australia. A selection of the guidelines for a few metals of concern is shown in table 16.1.

heavy metal metal with high density or of high relative atomic weight, and that is a risk to health or environment

TABLE 16.1 Selection of heavy metals guideline values in drinking water

Heavy metal	Health guideline (mg/L)	Aesthetic guideline, (mg/L)	Source in drinking water
Arsenic	0.01		From natural sources and mining, industrial and agricultural wastes, including manufacture of electrical components, wood preservatives and pesticides, gold mining, processing of ceramic materials
Cadmium	0.002		From industrial or agricultural contamination; contaminant in metals used to galvanise pipes and their subsequent erosion; metallurgical industries; improper waste disposal (e.g. nickel–cadmium batteries)
Chromium	0.05		From industrial or agricultural contamination of raw water or corrosion of materials in distribution system/plumbing; metallurgical processes such as alloying and electroplating; catalysts and oxidants; paint pigments
Copper	2	1	From corrosion of copper pipes
Iron		0.3	Occurs naturally in water in small quantities; high concentrations stain laundry and fittings; iron bacteria cause blockages and corrosion, and affect taste/odour
Lead	0.01		From natural sources and domestic/industrial sources, including household plumbing containing lead (e.g. pipes, solder), paint pigments, cable sheathing, ammunition, solder, batteries

(continued)

TABLE 16.1 Selection of heavy metals guideline values in drinking water *(continued)*

Heavy metal	Health guideline (mg/L)	Aesthetic guideline, (mg/L)	Source in drinking water
Mercury	0.001		From mining and industrial wastes; fungicides; electrical equipment; batteries
Nickel	0.02		From prolonged contact of water with nickel-plated fittings
Zinc		3	From corrosion of galvanised pipes, fittings and brasses

Mercury, lead and cadmium are some of the heavy metals that pose major disposal problems and, consequently, are the main causes of exposure for heavy metal poisoning. Inorganic and organic lead compounds are extremely poisonous and are difficult to dispose of. Heavy metal contamination is commonly caused from industrial and agricultural activities due to corrosion and inadequate disposal of materials.

Cleaning up heavy metal contamination is possible through:

- precipitation reactions, where a soluble solution is added to the contaminated water source, causing the heavy metal to precipitate out as an insoluble salt
- running the mixture through a filter, which will remove the precipitate containing the heavy metal cation and clean the water sample.

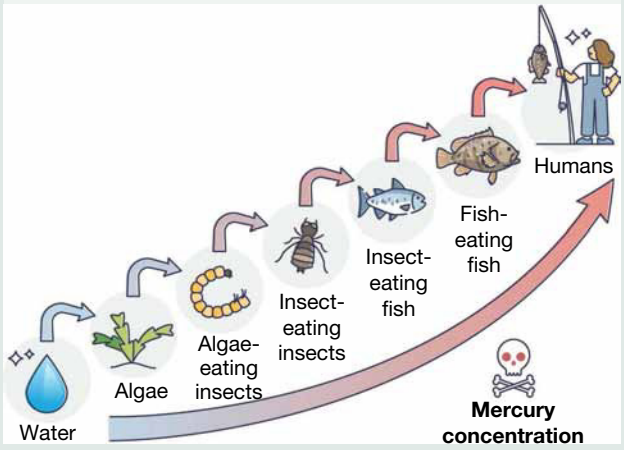
Heavy metals

Heavy metals present as soluble cations are toxic to the environment and organisms (plant and animal), including humans.

CASE STUDY: Mercury in Minamata Bay

The dangers of heavy metals were first exposed after a tragedy in Minamata Bay in Japan during the 1950s, where at least 60 people had been seriously affected by mercury poisoning prior to the cause being identified. Estimates of the cumulative death toll vary but it is generally agreed to be many thousands, and birth defects persist to this day. The source of the mercury was a chemical company that had been using mercury as a catalyst in the manufacture of acetaldehyde, a precursor to the polymer PVA (polyvinyl acetate), and was discharging large quantities of mercury waste into the sea. In the sea, bacteria converted the waste to toxic methyl mercury, which was ingested by the fish, and subsequently by the humans in a process known as **biomagnification**. This is where a substance becomes more concentrated in the tissues of organisms higher up the food chain. This results in the concentration of these substances in their bodies becoming higher than in the surrounding environment as the substances persist and accumulate.

FIGURE 16.5 The process of biomagnification



biomagnification process by which a compound (such as a pollutant or organometal) increases its concentration in the tissues of organisms as it travels up the food chain

16.2.3 Organometallic compounds

Organometallic compounds are heavy metals directly bonded to carbon atoms of organic molecules. As we saw from mercury in the preceding Minamata Bay case study, the conversion of mercury into toxic methyl mercury is an example of the formation of organometallic compounds by some aquatic microorganisms.

However, such compounds can also be useful for our society. For example, they can be:

- used as homogenous catalysts for industrial chemical reactions
- used as reagents in industry and research
- involved in the production of light-emitting diode (LED) lights
- useful in the synthesis of organic compounds.

While inorganic heavy metals are often only sparingly soluble, organometallic heavy metal compounds are much more soluble and can, therefore, enter the food chain more easily. This means they can be more environmentally hazardous due to the highly toxic nature of many of these compounds. Methyl mercury is one such example. While dimethyl mercury is virtually insoluble (and a very dangerous poison), monomethyl mercury is essentially ionic and dissolves much more readily in water. Its structure is:

$$H_3C—Hg^+X^-$$

where X is an appropriate anion. You will note from this structure that a covalent bond exists between the mercury atom and the methyl group. This is an example of a more complex type of bonding that does not fit the models that were considered in unit 1.

What makes this compound particularly dangerous is that it can also be produced when microorganisms in water combine with organic residues that may be present in the water.

organometallic compounds a compound containing a metal bonded to a carbon atom of an organic molecule

CASE STUDY: Lead the way to environmental chaos

Petrol used to be supplemented with tetraethyl lead (TEL), $(CH_3CH_2)_4Pb$, a non-polar organometallic compound that was found in the 1920s to make engines run more smoothly. This was commonly referred to 'leaded petrol' or 'super petrol', with its peak use during the 1960s and 1970s correlating with high vehicle ownership and use. Despite historical knowledge of the adverse effects of lead on human health, including neurological problems, this was ignored by Robert Kehoe, who was hired in 1924 by General Motors to advocate for TEL as an additive in petrol. Consequently, the use of the lead additive spread globally rapidly, while Kehoe continued to claim ambient lead levels in organisms (including humans) was normal and not harmful at low levels — therefore, leaded petrol did not constitute a public health hazard (which is clearly not the case)! This was finally contested in 1965; however, it still took governments globally a long time to phase out the use of leaded petrol, with Australia beginning its phase out in 1985 and stopping it completely by 2002. The environmental damage was and still is huge as a result of lead exhaust particles settling and concentrating in soils and waterways near roadways. Lead used to also be present in paint, which was also phased out and then prohibited in a similar time frame to leaded petrol. Now, mining and metal manufacturing are the largest sources of lead emission in Australia.

FIGURE 16.6 Old petrol bowser showing the pumps for unleaded petrol and super 'leaded' petrol

These environmental concerns feed into Australia's commitment towards the United Nations Sustainable Development Goals 2030 regarding:

- *Goal 6 — Clean water and sanitation*, with people having access to safe drinking water (see subtopic 10.2)
- *Goal 11 — Sustainable cities and communities*, by reducing the adverse per capita environmental impact of cities by monitoring fine particulate matter
- *Goal 12 — Responsible consumption and production*, achieving environmentally sound management of chemicals to reduce their release to air, soil and water sources, to decrease the adverse effects on human health and the environment
- *Goal 14 — Life below water*, to prevent and reduce pollution of the oceans (see subtopic 10.2).

Organometallic substances

Organometallic substances are heavy metals directly bonded to carbon atoms of organic molecules, increasing their toxicity and so causing great environmental concern.

16.2.4 Measuring salinity using electrical conductivity

Electrical conductivity provides a quick and inexpensive way of measuring the combined level of dissolved salts. Measurements are usually made in electrical conductivity units (ECs). In the Murray–Darling Basin case study in section 16.2.1, figure 16.4 outlines salinity levels at Morgan, showing salinity levels in EC units.

FIGURE 16.7 An electrical conductivity meter

Salts are ionic compounds; when they dissolve in water, they dissociate to produce mobile ions, which are free-moving charged particles capable of carrying an electrical current. This means solutions of ionic compounds can conduct electricity to an extent dependent upon the number of ions present. Electrical conductivity is measured using a handheld device (shown in figure 16.7) and is, therefore, a quick and cheap method that can be used in the field where the *total* salt content of a water sample is to be measured; it does not determine which particular salts are present.

EXTENSION: Desalination methods

Desalination is the process of removing salt and other unwanted substances from sea water and waste water. It is becoming an increasingly important method for producing water that is fit for humans use, with approximately 30 countries worldwide relying, at least in part, on desalinated water. In Australia, the drought of 1997–2009 created severe water shortages, with desalination plants for seawater being constructed as a result.

Two main methods of desalination are used around the world. These are distillation and reverse osmosis. Both of these methods use a large amount of energy. This results in the fresh water produced being higher in cost than fresh water from traditional sources.

However, with further research and attention to energy-saving measures, it is anticipated that these energy requirements will be reduced, resulting in desalination becoming an economically viable alternative. New developments use advances in membrane technology and nanotechnology. These are much more energy efficient than the traditional large-scale methods, with some of them also showing promise for small-scale, 'point-of-use' applications.

To learn more about desalination, including the processes of distillation and reverse osmosis, download the extension digital document from the Resources panel.

FIGURE 16.8 Desalination plant in the Arabian Desert emirate of Dubai

on Resources

Digital document EXTENSION: Desalination methods (doc-37921)

16.2 Activities

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16.2 Quick quiz on | **16.2 Exercise** | **16.2 Exam questions**

16.2 Exercise

1. What is salt?
2. What is the most common salt found in water?
3. Give five examples of elements that can be found dissolved in mineral water.
4. Calcium in mineral water is sourced from $CaCO_3$ in limestone. Given that calcium carbonate is almost insoluble in water, explain why calcium is found in mineral water.
5. Identify the sources of heavy metals in water.
6. What are the dangers posed by some heavy metals, such as mercury and lead, in drinking water?
7. What is an organometallic substance? Provide an example.
8. Explain why electrical conductivity provides a measurement of overall salt content rather than accurate values for particular salts.
9. What are the units for electrical conductivity?
10. Which solution will measure a higher electrical conductivity: 1 M of $NaNO_3$ or 1 M of Na_2CO_3? Explain.
11. What characteristics would a substance need to have to lead to biomagnification?

16.2 Exam questions

Question 1 (1 mark)

MC Which of the following solutions would be expected to have the highest conductivity?

A. 0.1 M $MgCl_2$
B. 0.1 M KCl
C. 1.0 M KCl
D. 1.0 M $MgCl_2$

Question 2 (1 mark)

The salinity of water is found by testing its electrical conductivity (EC). It is usually measured in μS cm^{-1} EC units where 1 μS cm^{-1} EC is equivalent to 1.6 mg L^{-1} of total dissolved salts.

An upper limit of 250 μS cm^{-1} EC was allowed for the water discharged by a particular paper mill. A sample of the discharged water was tested and found to contain 450 mg L^{-1} of total dissolved salts. Did this fall above or below the required limit?

Question 3 (1 mark)

MC Which of the following compounds would *not* increase the conductivity if added to water?

A. Glucose
B. Sodium chloride
C. Nitric acid
D. Ammonium nitrate

Question 4 (1 mark)

Hair dryers are often labelled with safety warnings telling consumers to keep electrical appliances away from water. If pure water does not tend to conduct electricity, why is this warning necessary?

Question 5 (1 mark)

Nitrates enter waterways naturally through the decomposition of living organisms. They can also be present in elevated levels due to run-off from soils treated with nitrate-rich fertilisers.

Ammonium nitrate is a fertiliser. Write the equation (including symbols of state) for the dissociation of ammonium nitrate when it dissolves in water.

More exam questions are available in your learnON title.

16.3 Quantitative analysis of salts — stoichiometry and molar ratios

KEY KNOWLEDGE

- Quantitative analysis of salts:
  - molar ratio of water of hydration for an ionic compound
  - the application of mass–mass stoichiometry to determine the mass present of an ionic compound

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16.3.1 Molar ratio of water of hydration in salts

Metabolic reactions taking place in our bodies are constantly cycling between the process of hydration and dehydration of various biologically important chemicals, including various salts and proteins. And, furthermore, the process of hydration is also utilised frequently in industry due to its importance in the production of various chemicals, including some acids, alcohols and aldehydes. **Water of hydration** refers

water of hydration water molecules present in the crystal lattice of hydrated compounds; can also be referred to as 'water of crystalllisation'

to the water molecules within the crystal lattice structure of inorganic salts where the water molecules are not structurally altered. A **hydrate** is simply the result of that compound absorbing the water from its environment.

Three different types of hydrates exist. In this subtopic, we will focus on the most common type: **inorganic salt hydrates**, in which water molecules hydrate ionic compounds to varying degrees (identified through different molar ratios). Water molecules can be easily removed (dehydrate) through processes such as heating to produce **anhydrous** salts that contain no water molecules. Common inorganic hydrated salts are listed in table 16.2.

hydrate any compound that absorbs water in the form of H_2O molecules into their chemical structure

inorganic salt hydrates inorganic salts where water molecules have been incorporated into their crystal structure

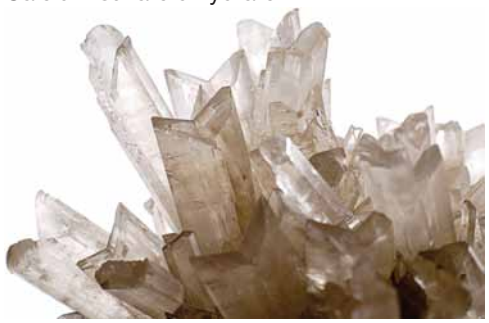
anhydrous contains no water molecules in its chemical structure

TABLE 16.2 Common inorganic salt hydrates

Inorganic salt hydrate	Hydrated chemical formula	Uses
Sodium carbonate decahydrate	$Na_2CO_3 \cdot 10H_2O$	Washing soda; sometimes still used as a detergent
Cobalt(II) chloride hexahydrate	$CoCl_2 \cdot 6H_2O$	Anhydrous form is used to detect humidity in the air due to its vivid pink colour change when hydrated
Copper(II) sulfate pentahydrate	$CuSO_4 \cdot 5H_2O$	Can be used as a colouring agent in paintings; used as a fungicide and herbicide
Magnesium sulfate heptahydrate	$MgSO_4 \cdot 7H_2O$	Epsom salts; used for soothing sore muscles (in bath salts) and as a soil additive for plant growth

(continued)

TABLE 16.2 Common inorganic salt hydrates *(continued)*

Inorganic salt hydrate	Hydrated chemical formula	Uses
Sodium tetraborate decahydrate	$Na_2B_4O_7 \cdot 10H_2O$	Used in slime, cosmetics, detergents and fire retardants
Calcium sulfate dihydrate 	$CaSO_4 \cdot 2H_2O$	Gypsum salts; used as a fertiliser for plants

Nomenclature of inorganic salt hydrates

Have you started to notice the pattern of how these hydrated salts are named? Just like you learned when naming organic molecules in topic 7, certain rules need to be followed to name these compounds correctly. The water molecules are only incorporated within the structure and have not changed the chemical structure; however, they do change the molar mass and properties of the compound and, therefore, need to be present in the empirical formula and name of the hydrated salt. Here are the steps to determine the degree of hydration and appropriately name the compound:

1. Name the ionic salt compound as described in subtopic 4.3.
2. Stoichiometrically determine how many water molecules are incorporated within the crystal lattice using the molar ratio by mass.
3. Identify the prefix associated with the number of water molecules (see table 16.3).
4. Add the word 'hydrate' after the prefix.

TABLE 16.3 Water of hydration prefixes for naming inorganic hydrates

Number of water molecules (degree of hydration)	Prefix
1	mono
2	di
3	tri
4	tetra
5	penta
6	hexa
7	hepta
8	octa
9	nona
10	deca

Naming hydrated salts

1. Name the anhydrous ionic salt compound.
2. Determine the degree of hydration and the required prefix.
3. Add the word 'hydrate' after the prefix.

tlvd-0520

SAMPLE PROBLEM 1 Examining the formulas and names of hydrated ionic compounds

a. Name the compound ZnCl2·4H2O.
b. Give the formula for the compound calcium sulfate dihydrate.

	THINK	WRITE
a.	Name the binary ionic compound in the usual way, and because four water molecules are in the compound, use the prefix *tetra-* followed by the word 'hydrate'. Because zinc only has one possible charge (2+), it is often seen without the (II).	Zinc(II) chloride tetrahydrate
b. 1.	Identify the cation and anion and write their symbols.	Cation: Ca^{2+} Anion: SO_4^{2-}
2.	Balance the charges so the net charge is zero, keeping the polyatomic sulfate ion intact. Since the ions are equally charged, the formula for calcium sulfate is $CaSO_4$.	$CaSO_4$
3.	determine the number of water molecules using the prefix: *di-* means 'two',	$CaSO_4 \cdot 2H_2O$

PRACTICE PROBLEM 1

a. Name the compound $CoBr_2 \cdot 6H_2O$.
b. Give the formula for the compound tin(II) chloride dihydrate.

Determining the degree of hydration and empirical formula of hydrated inorganic salts

As mentioned, some ionic compounds crystallise from an aqueous solution to form a hydrated ionic compound that holds a specific number of water molecules per salt compound. Hydrated copper(II) sulfate, for example, appears as a blue crystalline solid and has the formula $CuSO_4 \cdot 5H_2O$ (see figure 16.9). This means that, for each Cu^{2+} ion and SO_4^{2-} ion in the crystal lattice structure, five water molecules are also included. Heating the copper(II) sulfate crystals evaporates and removes the water molecules to leave a white powder known as anhydrous copper(II) sulfate. The formula of anhydrous copper(II) sulfate is $CuSO_4$ (figure 16.9). The mass of the water of hydration, and its percentage contribution to the mass of the ionic salt compound, can be calculated when the masses of both the hydrated and anhydrous compound are known or if the degree of hydration (table 16.3) is shown in the formula.

FIGURE 16.9 (left) Hydrated copper(II) sulfate, $CuSO_4 \cdot 5H_2O$, is also called copper(II) sulfate pentahydrate (right). Dehydrated (anhydrous) copper(II) sulfate, $CuSO_4$, is white in colour. When water is added, it turns the characteristic blue colour of hydrated copper sulfate, $CuSO_4 \cdot 5H_2O$.

Steps to finding an empirical formula of a hydrated salt

1. Write down the *chemical formulas* of the anhydrous salt and water.
2. Write down the *mass* of the anhydrous salt and water.
3. Convert masses to *moles* using $n = \frac{m}{M}$.
4. Find the simplest whole number *molar ratio* by dividing both moles by the number of moles of anhydrous salt.

tlvd-3496

SAMPLE PROBLEM 2 Calculating molar ratios of water of hydration of inorganic salts

Cobalt(II) chloride is often used to detect the presence of water vapour. Anhydrous cobalt(II) chloride crystals are blue but turn pink upon hydration.
To determine the degree of hydration, a sample of hydrated cobalt(II) chloride was analysed. A sample of mass 7.253 g was heated to drive off all the water content. At the end of this process, the remaining anhydrous cobalt(II) chloride had a mass of 3.953 g.

a. Calculate the empirical formula of hydrated cobalt(II) chloride.

b. Name the hydrated salt.

THINK

a. 1. Write headings for cobalt(II) chloride and water.

2. Write down the anhydrous mass and determine the mass of water removed.

3. Calculate the number of moles of the anhydrous salt and the moles of water, using the formula $n = \frac{m}{M}$.

4. Divide by the number of moles of anhydrous salt to find the whole number ratio of moles of water per mole of salt.
(*Note:* in this version of the calculation you divide through by the moles of anhydrous salt to give a value of 1. This then allows for easy comparison of the moles of water.)

5. Write empirical formula.

b. i. Name the salt.

ii. Identify the prefix of the water of hydration.

iii. Add 'hydrate'.

iv. Put it all together.

WRITE

Component	$CoCl_2$	H_2O
Masses	3.953 g	$m(H_2O) = 7.253 - 3.953$ $= 3.300$ g
Moles	$n(CoCl_2) = \frac{m}{M}$ $= \frac{3.953}{(58.9 + 2 \times 35.5)}$ $= \frac{3.953}{129.9}$ $= 0.0304$	$n(H_2O) = \frac{m}{M}$ $= \frac{3.300}{(2 \times 1.0 + 16.0)}$ $= \frac{3.300}{18.0}$ $= 0.183$
Ratio (divide by smallest)	$= \frac{0.0304}{0.0304}$ $= 1$	$= \frac{0.183}{0.0304}$ $= 6.02 \approx 6$

$CoCl_2 \cdot 6H_2O$

Cobalt(II) chloride

hexa

hydrate

Cobalt(II) chloride hexahydrate

PRACTICE PROBLEM 2

Copper(II) sulfate is a common chemical found in school laboratories. You have been given a sample and told it is anhydrous copper(II) sulfate, which you know should be a white powder. However, the powder is blue, which indicates water has been absorbed by the salt and it has become hydrated.

To determine the degree of hydration, your sample of blue copper(II) sulfate was analysed. A sample of mass 4.148 g was heated to drive off all the water content. At the end of this process, the remaining anhydrous copper(II) sulfate had a mass of 2.644 g.

a. Calculate the empirical formula of hydrated copper(II) sulfate.

b. Name the hydrated salt.

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tlvd-0622

EXPERIMENT 16.1 online only

The empirical formula of a hydrated salt

Aim

To determine the empirical formula of a hydrated salt

16.3.2 Mass–mass stoichiometry

Mass–mass **stoichiometry** involves solving a problem in which the mass of at least one reactant or product is given. You are then asked to calculate the mass of another reactant consumed or mass of product produced. Mass–mass stoichiometry requires the conversion of masses of substances to moles, or moles of substances to masses, through the application of the formula for determining the number of moles of a substance.

Mass–mass stoichiometry formulas

$$\text{moles}\,(n) = \frac{\text{mass}\,(m)}{\text{molar mass}\,(M)}$$

$$\text{mass}\,(m) = \text{moles}\,(n) \times \text{molar mass}\,(M)$$

TIP: The formula $n = \frac{m}{M}$ can be found in table 3 of the VCE Chemistry Data Book.

Solving mass–mass stoichiometric problems in four steps

1. Write a balanced chemical equation for the reaction, identifying the known (given) and unknown (required) quantities of substance.
2. Calculate the number of moles of the known quantity of substance present.
3. From the balanced equation, find the molar ratio that states the proportion of unknown to known quantities in the reaction. Convert this to a fraction of coefficient of unknown divided by coefficient of known. To calculate the number of moles of unknown required substance, multiply the number of moles of the known substance by this fraction.
4. Calculate the quantity (mass) of the required substance.

stoichiometry the amount and ratio of reactants and products in a chemical reaction

tlvd-0597

SAMPLE PROBLEM 3 Determining the mass of a product from complete combustion of a reactant

Sulfur is present in coal in the form of pyrite, FeS_2. This substance is also known as 'fool's gold'. When pyrite burns, it pollutes the air with one of its combustion products, sulfur dioxide, and produces solid iron(III) oxide. What mass of iron(III) oxide is formed from the complete combustion of 183.5 g of pyrite?

THINK

1. Write the balanced equation and identify known (given) and unknown (required) moles.

WRITE

$4FeS_2(s) + 11O_2(g) \rightarrow 2Fe_2O_3(s) + 8SO_2(g)$

Pyrite is the known quantity and iron(III) oxide is the unknown quantity.

2. Calculate the number of moles of the known quantity of substance, FeS_2, using the formula $n = \frac{m}{M}$.

$$n(FeS_2) = \frac{183.5}{55.8 + (2 \times 32.1)}$$
$$= \frac{183.5}{120.0}$$
$$= 1.529 \text{ mol}$$

3. Find the molar ratio, $Fe_2O_3 : FeS_2$ from the equation, and use it to calculate the number of moles of the required substance, Fe_2O_3.

The ratio $Fe_2O_3 : FeS_2$ is 2 : 4, which becomes

$$n(Fe_2O_3) = \frac{2}{4} \times n(FeS_2) = \frac{n(FeS_2)}{2}$$
$$= \frac{1.529}{2}$$
$$= 0.7645 \text{ mol}$$

4. Calculate the required quantity (mass) of Fe_2O_3 formed by the reaction.

$$m(Fe_2O_3) = n \times M$$
$$= 0.7645 \times (2 \times 55.8 + 3 \times 16.0)$$
$$= 122.0 \text{ g}$$

PRACTICE PROBLEM 3

When 2.864 g of potassium iodide reacts completely with a solution of lead nitrate, a yellow precipitate of lead(II) iodide forms according to the following equation.

$$\mathbf{2KI(s) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)}$$

Calculate the mass of lead(II) iodide that precipitates.

Calculations with limiting or excess reactants

In real life, you will rarely have reactions taking place where the reactants are present in their exact molar ratios. Most of the time you will have either:

- not enough of one reactant. This is the **limiting reactant** and it will be completely used up in the reaction.
- too much of the other reactant. This is the **excess reactant**, which will be left over after the reaction is complete because its consumption is limited by the small amount of the first reactant. The amount of one of the reactants is not available in the required **mole ratio** (as per reaction). This means that the second reactant will not be used up completely in the reaction.

Amount of product formed due to limiting reactant

The amount of product formed by the reaction is limited and determined by the amount of the limiting reactant.

Consider the reaction between methane and oxygen:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

Recall that a balanced equation indicates the mole ratios in which the reactants are used up and the products are formed. According to this equation, one mole of methane gas reacts with two moles of oxygen gas to produce one mole carbon dioxide and two moles of water.

limiting reactant the reactant that is completely used up in a reaction

excess reactant reactant left over when the limiting reagent has been used up

mole ratio represents the relative number of moles of each substance involved in the reaction

If one mole of methane and three moles of oxygen are mixed and allowed to react, all the methane and two moles of oxygen will be used up. This means that one mole of oxygen is left over. In this case:

- The limiting reactant that has been completely used up is methane.
- The excess reactant, of which some remains at the end of the reactions, is oxygen.
- The amount of product formed by the reaction is limited by the amount of limiting reactant. In this case, the addition of more oxygen to the mixture would have no effect on the amount of product formed, because no more methane is available to react with it.

Note: the same stoichiometry steps are used in questions asking to identify the excess reactant. However, any further calculations will only be using the limiting reactant.

Calculations with limiting reactants

In a limiting reactant calculation, first determine which substance is completely used up (the limiting reactant) and which is left over (the excess reactant). Calculations are then based on the moles of the limiting reactant.

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SAMPLE PROBLEM 4 Determining limiting reactant and reactant(s) in excess

Magnesium and oxygen react to form magnesium oxide according to the equation:

$$2Mg\,(s) + O_2\,(g) \rightarrow 2MgO\,(s)$$

If 24.31 g of magnesium and 20.00 g of oxygen are available for reaction, determine the limiting reactant, the reactant in excess, and the mass of magnesium oxide formed.

THINK	WRITE
1. Convert the mass to moles.	$n(\text{Mg}) = \frac{m}{M}$ $= \frac{24.31}{24.3}$ $= 1.00 \text{ mol}$ $n(O_2) = \frac{m}{M}$ $= \frac{20.00}{32.0}$ $= 0.625 \text{ mol}$
2. Compare this mole ratio with the mole ratio in the equation to identify which reactant is limiting.	According to the equation, the ratio of Mg : O is 2 : 1. Therefore, if 1.00 mol of Mg were to react, we would need 0.500 mol of O_2. We have 0.625 mol O_2, which is more than enough. Therefore, Mg is the limiting reactant and O_2 is the excess reactant.
3. Find the molar ratio of the required product to the limiting reactant, MgO : Mg, and use it to calculate the number of moles of the required substance, MgO.	The ratio MgO : Mg is 2 : 2 which means $n(\text{MgO}) = \frac{2}{2} \times n(\text{Mg}) = n(\text{Mg})$ $= 1.00 \text{ mol}$
4. Calculate the mass of MgO produced.	$m(\text{MgO}) = n \times M$ $= 1.00 \times 40.3$ $= 40.3 \text{ g}$

PRACTICE PROBLEM 4

Aluminium burns in oxygen to produce an intense burst of light and is often used in sparklers and fireworks. Aluminium oxide is produced in the process. The equation for the reaction is:

$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$

Calculate the mass of aluminium oxide produced from 5.00 g of aluminium and 10.0 g of oxygen.

16.3.3 Mass–volume stoichiometry

Many chemical reactions involve interactions between solids and solutions. These reactions may be referred to as mass–volume stoichiometry and could be:

- two aqueous solutions mixed to form a precipitate
- solids dissolved in aqueous solutions to form new products.

Stoichiometric calculations that involve solids and solutions require the use of two formulas to calculate moles.

Mass–volume stoichiometry

When working with solids, we use the formula

$$n = \frac{m}{M}$$

When working with solutions, we use the formula

$$n = c \times V$$

Solving mass–volume stoichiometric problems in four steps

1. Write a balanced chemical equation for the reaction, identifying the known (given) and unknown (required) quantities of substance.
2. Calculate the number of moles of the known quantity of substance present using the formulas shown in the preceding box.
3. From the balanced equation, find the molar ratio that states the proportion of known to unknown quantities in the reaction. Convert this to a fraction of coefficient of unknown divided by coefficient of known. To calculate the number of moles of unknown required substance, multiply the number of moles of the known substance by this fraction.
4. Calculate the quantity of the required substance.

Note: mass–volume stoichiometry involves the same steps as in mass–mass stoichiometry, except that the formula used is $n = c \times V$.

TIP: Determining mole ratios

When determining mole ratios, always set up the mole ratio as your *unknown* coefficient divided by the *known* coefficient; determine the number of moles of the *unknown* by multiplying this ratio by the *known* moles.

$$n\,(\text{UNKNOWN}) = \left[\frac{\text{coefficient of UNKNOWN}}{\text{coefficient of KNOWN}}\right] \times n\,(\text{KNOWN})$$

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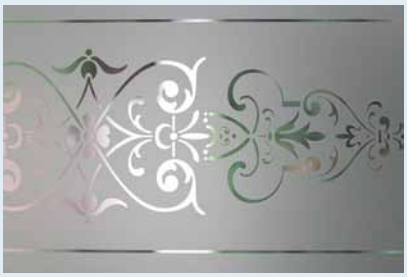
SAMPLE PROBLEM 5 Determining the mass of a reactant using the concentration and volume of the other reactant

Art may be created on glass using hydroflouric acid, in a process called etching. The reaction that occurs is:

$$SiO_2(s) + 4HF(aq) \rightarrow SiF_4(g) + 2H_2O(l)$$

A quartz sculptor has a 500 mL container of 22.50 M hydrofluoric acid. What mass of quartz could be etched by the acid?



THINK	WRITE
1. Write the equation, identifying the known and unknown quantities.	$SiO_2(s) + 4HF(aq) \rightarrow SiF_4 + 2H_2O(1)$ *unknown* *known*
2. Calculate the number of moles of the known quantity of substance, HF.	$n\,(HF) = cV$ $= 22.50 \times 0.500$ $= 11.25$ mol
3. Find the molar ratio from the equation and use it to calculate the number of moles of SiO_2 required.	$n(SiO_2) : n(HF)$ is $1 : 4$ $n(SiO_2) = \frac{1}{4} n(HF)$ $= \frac{1}{4} \times 11.25$ $= 2.81$ mol
4. Calculate the required quantity of SiO_2, using $m = n \times M$.	$m = n \times M$ $= 2.81 \times (28.1 + 2 \times 16.0)$ $= 169$ g

PRACTICE PROBLEM 5

A piece of aluminium is placed in a beaker containing 500 mL of H_2SO_4 solution and hydrogen gas is evolved. Given that the initial mass of Al was 15.14 g and its final mass was 9.74 g, calculate the concentration of the acid.

16.3.4 Gravimetric analysis

Gravimetric analysis is a form of quantitative analysis in which the amount of a substance is established through the measurement of a mass. It is a well-established technique that has been carried out in analytical laboratories for centuries.

At its simplest, it may involve drying a product to determine its moisture content. Consider water being added to frozen chickens in order to increase the apparent weight at which they are sold. Gravimetric analysis can be used to check whether this practice has occurred. In agriculture, a soil scientist often needs to know the moisture content of a soil sample in order to determine the effectiveness of various soil conditioning procedures.

gravimetric analysis analysis in which the amount of a substance is established through the measurement of a mass

Gravimetric analysis can also be a much more sophisticated procedure in which knowledge of chemical reactions, solubilities and stoichiometry is used to determine the amount of a substance in a sample. Using chemical knowledge in this way makes gravimetric analysis a powerful tool that can be used by analytical chemists in a wide range of situations.

on Resources

Video eLesson Gravimetric analysis (eles-2491)
Interactivity Gravimetric analysis (int-6357)

tlvd-3497

SAMPLE PROBLEM 6 Using gravimetric analysis to determine the amount of water in a sample

A 23.1 g sample of soil is heated to 105 °C for 60 minutes. It is then weighed and the process repeated a number of times. A constant mass of 21.0 g is eventually obtained, which does not change upon further heating.
Calculate the percentage of water in this sample of soil.

THINK	WRITE
1. Identify the change in water mass.	The mass of water lost from the sample is $23.1 - 21.0 = 2.1$ g.
2. Calculate the percentage of water in the sample.	$\frac{2.1}{23.1} \times \frac{100}{1} = 9.1\%$

PRACTICE PROBLEM 6

The soil sample from sample problem 5 was kept in inappropriate conditions and it re-absorbed moisture from air. A week later, its new mass was found to be 22.1 g. Calculate the percentage of water re-absorbed by the dried soil sample.

Extending the gravimetric method

Gravimetric analysis can be used to quantitatively determine a component in a solid or aqueous mixture.

This procedure is especially suitable for:

- determining the quantity of metal cations in a mixture
- determining the quantity of anions in a mixture such as sulfate, which form some easily prepared insoluble salts (such as barium sulfate).

When designing a gravimetric procedure, knowledge of precipitation reactions (refer to topic 4) and, therefore, the solubilities of ionic reaction products is important when choosing what solution to add and in predicting the identity of the precipitate formed. To refresh your memory, the solubility in water of some common compounds is shown in table 16.4.

TABLE 16.4 The solubility in water of compounds of common anions

Anion	Cations forming soluble compounds	Cations forming insoluble compounds
Nitrate, NO_3^-	All	—
Chloride, Cl^- Bromide, Br^- Iodide, I^-	Most	Ag^+, Pb^{2+} ($PbCl_2$ is moderately soluble in hot water.)
Sulfate, SO_4^{2-}	Most	Ba^{2+}, Pb^{2+} (Ag_2SO_4 and $CaSO_4$ are slightly soluble.)
Carbonate, CO_3^{2-}	Na^+, K^+, NH_4^+	Most
Phosphate, PO_4^{3-}	Na^+, K^+, NH_4^+	Most
Sulfide, S^{2-}	Na^+, K^+, NH_4^+	Most (MgS, CaS, BaS, Al_2S_3 and Fe_2S_3 decompose in water.)
Hydroxide, OH^-, and oxide, O^{2-}	Na^+, K^+, Ba^{2+}	Most ($Ca(OH)_2$ is slightly soluble.)

Gravimetric analysis (practical component)

This technique often involves the following steps (see also figure 16.10):

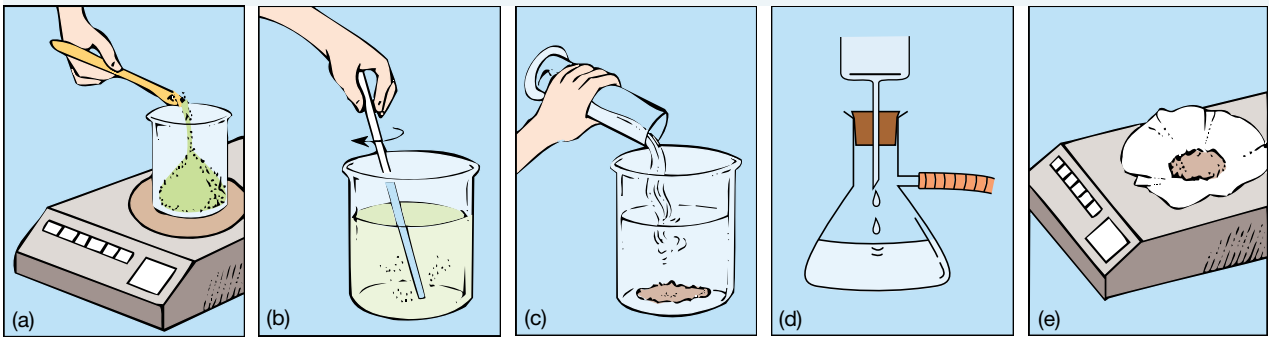
a. weighing the sample to be analysed (for example, a soil sample)
b. dissolving the sample in water (if there are insoluble components, such as in soil, you will need to filter these components out now)
c. adding excess of a suitable chemical to form a precipitate (it is this precipitate that removes the ions required from the solution so the amount of ions may later be determined)
d. weighing the filter paper and then filtering to collect the precipitate
e. repeated drying and weighing until a constant mass of precipitate is obtained (from this mass and the original mass of material dissolved, the component may be determined quantitatively).

Gravimetric analysis (stoichiometric calculations component)

- Simple stoichiometry (following the steps outlined previously) is used to deduce the amount of the component in the original mixture.
  Note: sometimes the substance being analysed is already dissolved, as with water samples. In such cases, the method starts with adding a suitable chemical to a known volume to form a precipitate (see figure 16.10, step c).
- An additional step is to determine the percentage by mass, or the percentage purity of the target ion per original sample.

Note: If dilutions have been made, you must include the dilution factors in your calculations.

FIGURE 16.10 The steps involved in gravimetric analysis



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SAMPLE PROBLEM 7 Calculating the percentage purity of a substance

A pharmaceutical company wishes to test the purity of some commercial barium chloride. A sample of the commercial barium chloride was weighed and found to have a mass of 10.0 g. After the sample was dissolved in water, excess sulfuric acid was added to form a precipitate of barium sulfate. This precipitate was then filtered, dried and weighed. It was found to have a mass of 10.55 g. Calculate the percentage purity of the barium chloride.

THINK	WRITE
1. Write the balanced equation.	$BaCl_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + 2HCl(aq)$
2. Calculate the moles of the known quantity of product.	$m(BaSO_4) = 10.55\text{ g}$ $M(BaSO_4) = 233.4\text{ g mol}^{-1}$ $\therefore n(BaSO_4) = \frac{m}{M}$ $= \frac{10.55}{233.4}$ $= 0.045\,20\text{ mol}$
3. Use the mole ratio from the equation to calculate the moles of $BaCl_2$ used in reaction.	$n(BaCl_2) : n(BaSO_4) = 1 : 1$ $\therefore n(BaCl_2) = 0.045\,20\text{ mol}$
4. Calculate mass of $BaCl_2$.	$m = n \times M$ $M(BaCl_2) = 208.3\text{ g mol}^{-1}$ $n(BaCl_2) = 0.045\,20\text{ mol}$ $\therefore m(BaCl_2) = 0.045\,20 \times 208.3\text{ g}$ $= 9.415\text{ g}$
5. Calculate the percentage purity of the barium chloride.	$\therefore$ % barium chloride (original sample) $= \frac{9.415}{10.0} \times 100 = 94.2\%$ (3 significant figures)

PRACTICE PROBLEM 7

To determine the percentage of arsenic in a particular brand of pesticide, a 2.15 g sample is treated to precipitate all the arsenic as its sulfide, As_2S_3. If 0.353 g of precipitate is obtained, calculate the percentage of arsenic in the pesticide.

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SAMPLE PROBLEM 8 Calculating the level of ions in a sample

A sample of groundwater from an abandoned lead battery manufacturing site was analysed gravimetrically for its lead content.
A 2000 mL sample was treated with sodium sulfate solution until no further precipitate was observed to form. After filtering and drying to constant weight, 1.21 g of lead(II) sulfate was obtained.
Calculate the level of lead(II) ions in the groundwater in mg L^{-1}.

THINK	WRITE
1. As we are interested in the Pb^{2+} ions only, an ionic equation is more practical than a full equation. We can, therefore, write the equation and list the known information.	$Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$ $m(PbSO_4) = 1.21\ g$ $M(PbSO_4) = 207.2 + 32.1 + 4 \times 16.0$ $= 303.3\ g\ mol^{-1}$
2. Calculate moles of $PbSO_4$.	$n(PbSO_4) = \frac{1.21}{303.3}$ $= 0.003\ 99\ mol$
3. Use the mole ratio from the equation to calculate moles of Pb^{2+}.	$n(Pb^{2+}) : n(PbSO_4) = 1 : 1$ Therefore, $n(Pb^{2+}) = 0.003\ 99\ mol$
4. Calculate mass of Pb^{2+}. TIP: Remember to use the non-rounded value for the number of moles.	$m\left(Pb^{2+}\right) = 0.003\ 99 \times 207.2$ $= 0.827\ g$
5. Calculate Pb^{2+} in mg L^{-1}.	Since the amount analysed was 2000 mL (2 L), the concentration of Pb^{2+} equals: $\frac{0.827}{2} = 0.413\ g\ L^{-1}$ $= 413\ mg\ L^{-1}$

Alternative solution:
The workings shown rely on the assumption that all the lead in the original sample precipitates.

When the molar mass of $PbSO_4$ is evaluated, it becomes clear that, in the 303.3 g that represents one mole, 207.2 g is due to lead.

The fraction of lead in $PbSO_4$ is therefore $\frac{207.2}{303.3}$.

It follows that in 1.21 g of $PbSO_4$ are $\frac{207.2}{303.3} \times 1.21 = 0.827$ g of lead.

The rest of the calculation is the same as shown.

PRACTICE PROBLEM 8

The zinc content in a water sample was analysed as follows.

A 1000 mL sample was treated with a solution of ammonium hydrogen phosphate, $(NH_4)_2HPO_4$, to precipitate all the zinc ions present. A precipitate of NH_4ZnPO_4 was obtained, which was then decomposed by heating to produce 2.918 g of $Zn_2P_2O_7$.

Calculate the concentration of zinc in the sample in g L^{-1}.

Sources of error in gravimetric analysis

Although most gravimetric procedures have a clearly recognisable series of steps, the properties of the chemicals involved may necessitate some modifications to the method in particular circumstances.

A thorough appreciation of the method is also important in evaluating the limits of your results. If certain things happen during the analysis, it is important to predict how these may affect the final result. Table 16.5 shows some possible sources of error.

TABLE 16.5 Sources of error in gravimetric analysis

Action	Effect on analysis result	Reason
Insoluble materials not filtered out before forming the precipitate	Overestimated	The apparent mass of the precipitate will increase.
Not enough of the precipitate-forming chemical added	Underestimated	Not enough of the precipitate will form because some of the ions that are being analysed will remain in solution.
Forming a precipitate that is too soluble	Underestimated	Not all of the ions being analysed will be in the precipitate.
Forming extra precipitate due to the presence of other competing ions	Overestimated	Too much precipitate will form.
Weighing the precipitate before it is dry	Overestimated	The water present will increase the apparent mass of the precipitate.
Not rinsing the precipitate before drying it	Overestimated	As the precipitate dries, other soluble chemicals will begin to crystallise out of the small amount of solution still trapped in the precipitate. These will add to the mass.*
Adding too much of the precipitate-causing chemical	No effect	This is a necessary part of the method to make sure that all of the required ions are in the precipitate. The chemical must be in excess.
Using too much water for the initial dissolving	No effect	This is a practical consideration — the more water you have, the longer the filtering step.

**Note*: To avoid this situation, the precipitate should always be washed with a small amount of pure solvent before it is dried and weighed. Too much water or solvent could cause some of the precipitate to redissolve. In this case, the final result will be underestimated.

16.3 Activities

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16.3 Quick quiz on | 16.3 Exercise | 16.3 Exam questions

16.3 Exercise

1. a. What is a hydrated salt?
   b. What is an anhydrous salt?
   c. How can you determine the empirical formula of a hydrated salt?
2. MC In an experiment, some hydrated calcium chloride was heated. The process produced 0.0263 mol of the anhydrous salt and 0.0553 mol of water was removed.
   What is the formula of the hydrated calcium chloride?
   A. $CaCl_2 \cdot H_2O$
   B. $CaCl_2 \cdot 2H_2O$
   C. $CaCl_2 \cdot 5H_2O$
   D. $CaCl_2 \cdot 6H_2O$
3. Upon analysis, 18.759 g of hydrated strontium chloride is found to contain 3.470 g of water. Calculate the empirical formula of hydrated strontium chloride.
4. A 1.124 g sample of $CaSO_4$ crystals was heated to drive off the water of crystallisation. When completely dry, a residue of 0.889 g was obtained. Determine the empirical formula of the hydrated compound.
5. A 0.942 g sample of $MgSO_4$ crystals was heated to drive off the water of crystallisation. When completely dry, a residue of 0.461 g was obtained. Determine the empirical formula of the hydrated compound.
6. A solution is made by dissolving an unknown amount of barium chloride in water. It is then added to a solution of potassium carbonate such that all the barium chloride reacts. The equation for this reaction is

$$K_2CO_3(aq) + BaCl_2(aq) \rightarrow 2KCl(aq) + BaCO_3(s)$$

   a. What information in the equation tells you that barium carbonate is a precipitate?
   b. If 4.582 g of barium carbonate is formed, calculate the mass of barium chloride that was originally dissolved.
7. Consider the following balanced equation:

$$Cd(NO_3)_2(aq) + Na_2S(aq) \rightarrow CdS(s) + 2NaNO_3(aq)$$

   a. Calculate the mass of CdS produced from 235 mL of a 0.178 M solution of Na_2S with excess $Cd(NO_3)_2$ present.
   b. What is the significance of $Cd(NO_3)_2$ being present in excess?
8. In a laboratory experiment, a strip of magnesium weighing 2.56 g was placed into 200 mL of a hydrochloric acid solution. Bubbles of gas were observed and identified as hydrogen. The next day, no bubbles were observed in the beaker so the magnesium strip was removed from the beaker, dried and reweighed. Its mass was recorded as 0.350 g.
   a. Why were no bubbles observed in the beaker on the second day of the experiment?
   b. Calculate the concentration of the acid.
   c. Identify the sources of error in this experiment.
   d. Outline the safety precautions that should be taken in this experiment.

9 The calcium content of a sample can be determined gravimetrically by first precipitating the calcium ions present as calcium oxalate. When this precipitate is heated, it decomposes completely to produce calcium oxide.
In testing the purity of a calcium chloride sample, a chemist used gravimetric analysis and obtained 8.81 g of calcium oxide from an original 18.0 g sample of calcium chloride.
Calculate the percentage purity of the calcium chloride.

10. MC In gravimetric analysis, the precipitate is dried in an oven and cooled before weighing. What is the main reason the precipitate must be cooled before weighing?
A. So that the balance does not melt or catch on fire
B. To minimise convection currents that can cause the balance to weigh inaccurately
C. To ensure that the crucible containing the precipitate can be held in the hand
D. Because substances weigh less at higher temperatures

16.3 Exam questions

Question 1 (1 mark)

The degree of hydration refers to the number of water molecules (water of crystallisation) associated with certain ionic compounds — for example, $CuSO_4 \cdot 5H_2O$ has a degree of hydration of 5.

The degree of hydration can be found by finding the simplest whole number ratio of moles of water to that of the ionic compound.

A 1.634 g sample of hydrated calcium chloride, $CaCl_2 \cdot xH_2O$, was heated until all the water was removed, leaving a solid of mass 1.234 g.

Calculate the ratio of calcium chloride : water to find the degree of hydration, x.

Question 2 (1 mark)

In an experiment in their chemistry class, Ming was left with a white powder in a desiccator. It was suspected to be either anhydrous sodium carbonate or anhydrous sodium hydrogen carbonate.

The next day, another group member carefully weighed out 7.359 g of the powder and reacted it with acid. Subsequently, it was determined that 0.0875 mol was present in the sample.

What was the white powder?

Question 3 (1 mark)

4.6 g of anhydrous zinc sulfate with an M of 161.5 g mol^{-1} was obtained by evaporating the water from 8.2 g of the crystalline hydrated salt with empirical formula $ZnSO_4 \cdot xH_2O$. Calculate the value of x.

Question 4 (1 mark)

MC Some students performed a gravimetric analysis of salt in water and found that their mass percentage of salt was significantly higher than that of others in their class. Which of the following would *not* explain this?
A. They had allowed some of the precipitate to pass through the filter paper.
B. They had not dried their precipitate sufficiently to ensure no further mass loss occurred.
C. They had not rinsed their precipitate with portions of water before filtering.
D. They had forgotten to subtract the mass of the filter paper used.

Question 5 (4 marks)

An experiment was carried out to find the percentage of sodium chloride by mass in a lake situated in a district with salinity problems. A water sample of mass 100.0 g was collected at the site and used to prepare three trials where 25.0 g of lake water was reacted with an excess of silver nitrate solution to form a silver chloride precipitate.

The results for the three trials are shown in the table provided.

Trial	1	2	3
Mass of AgCl (g)	2.42	2.39	2.43
Mass of lake water (g)	25.0	25.0	25.0

Assuming all chloride present in the water is due to sodium chloride, use the average mass of precipitate to calculate %(NaCl) in the water.

More exam questions are available in your learnON title.

16.4 Quantitative analysis of salts — colorimetry and UV–visible spectroscopy

KEY KNOWLEDGE

- The application of colorimetry and/or UV–visible spectroscopy, including the use of a calibration curve, to determine the concentration of ions or complexes in a water or soil sample

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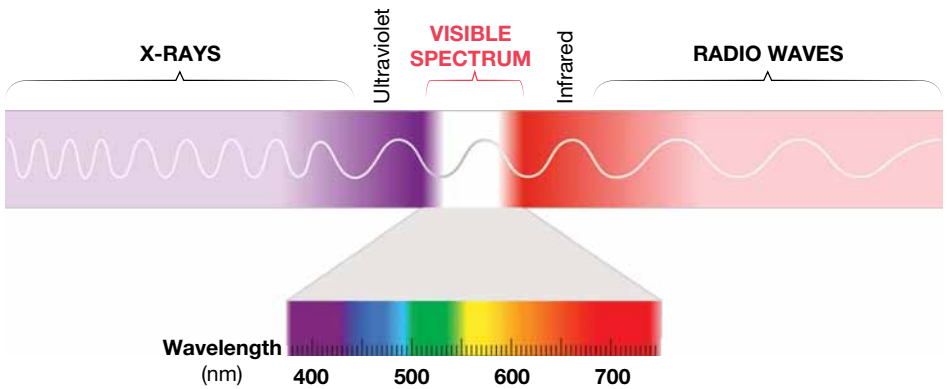
16.4.1 The electromagnetic spectrum

Both colorimetry and UV–visible spectroscopy rely on the electromagnetic spectrum — specifically, the region between ultraviolet light (UV light) and the visible light spectrum (see figure 16.11). Humans can detect colours only within the visible spectrum; anything outside this region is invisible or colourless.

FIGURE 16.11 The electromagnetic spectrum, highlighting the range of visible light between 380 nm and 780 nm. Outside of this range, humans are unable to detect any colour.



16.4.2 Colorimetry

Colorimetry is a technique in which colour levels are compared with a set of standards, where the intensity of the colour depends on the identity and concentration of the component being analysed.

If a solution is naturally coloured, or can be coloured by adding certain chemicals, colorimetric analysis may be performed by eye. Two examples of this are:

- the use of universal indicator to measure pH, where the colour produced is compared by eye with a standard chart
- estimating the level of chlorine in swimming pool water, where the sample is coloured by adding chemicals and the results compared with a chart.

Since the colour a substance appears is related to how it absorbs various wavelengths of light, more accurate results can be quantitatively obtained by using an instrument called an **instrumental colorimeter**. This compares how the light is absorbed by the colour in the *test sample* with how the light is absorbed by the colours in the *standards* (samples of known concentration that have been treated identically to the sample).

From the standard absorbance readings, a calibration curve (a graph of absorbance versus concentration, see figure 16.12) is produced, allowing the concentration of the sample to be interpolated directly from this graph.

The basic design of a colorimeter is shown in figure 16.13.

FIGURE 16.12 Example calibration curve for the absorbance of known concentrations of a coloured solution

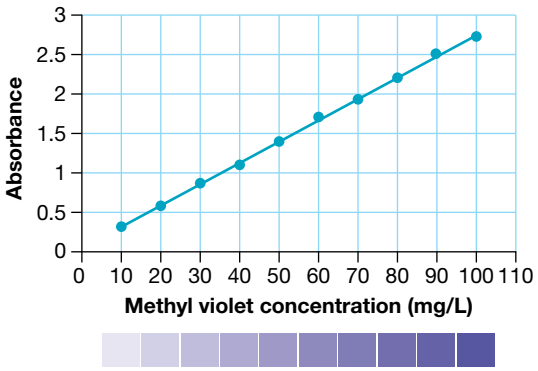
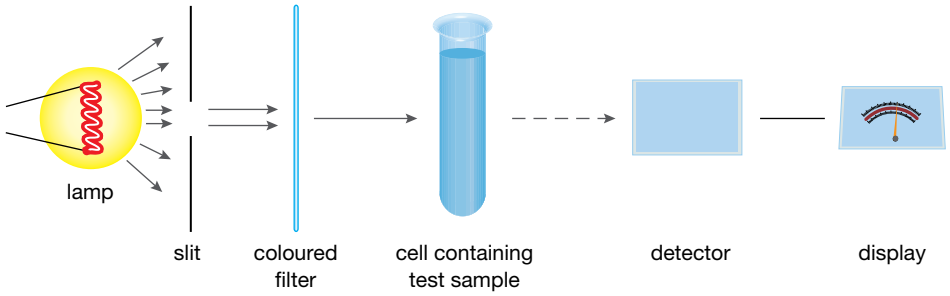


FIGURE 16.13 The essential features of the colorimeter



In a colorimeter, the detector measures the amount of light that passes through the test cell and converts this into an absorbance reading. While normal white light can be used, more accurate results are obtained using a complementary colour (by changing the colour of the filter) to the test sample. For example, if copper sulfate (blue–green) levels are being measured, the best colour to shine through such samples is red light, as shown in figure 16.14. Examples of complementary colours are summarised in table 16.6 and figure 16.15.

instrumental colorimeter instrument used to determine the concentration of coloured compounds in solution through comparison of a test sample against samples of known concentration through production of a calibrations curve (absorbance versus concentration)

FIGURE 16.14 A blue–green substance strongly absorbs red light.

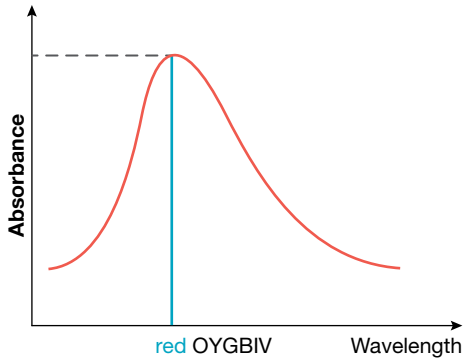
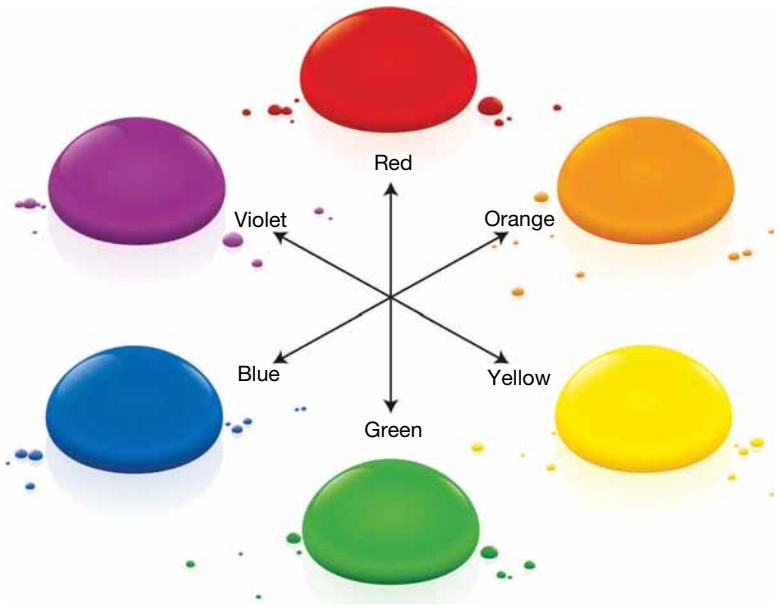


TABLE 16.6 Observed colours and their complementary colours

Colour observed	Colour absorbed
Green	Deep red
Blue–green	Red
Blue	Orange
Blue–violet	Yellow
Violet	Yellow–green
Deep red	Green
Red	Blue–green
Orange	Blue
Yellow	Blue–violet
Yellow–green	Violet

FIGURE 16.15 Complementary colour wheel



TIP: Complementary colours

A quick way to remember which colours are complementary to each other is to use the rainbow acronym ROYGBIV (**R**ed, **O**range, **Y**ellow, **G**reen, **B**lue, **I**ndigo, **V**iolet). Eliminating indigo, split the list of colours in half and place them like a fraction:

$$\frac{\text{R O Y}}{\text{G B V}}$$

This tells us that the pairs of complementary colours are:

- red and green
- orange and blue
- yellow and violet.

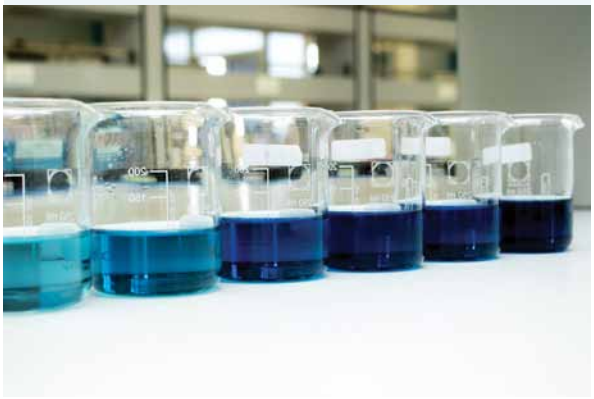
To summarise, colorimetry:

- is relatively cheap
- is useful for measuring coloured species
- can measure a colourless species if a coloured derivative can be prepared
- needs care to be taken to ensure that no other coloured species are present in the sample that could also absorb the particular colour of light being used, causing errors with results
- can produce results to within 1 to 2 per cent accuracy.

Example of colorimetry use for determining phosphate levels

Phosphates are important nutrients for plants but can become an environmental problem when they enter water systems in large amounts. As we have already seen, this can lead to eutrophication if certain other conditions are also present. Sources of phosphate include fertilisers such as ammonium phosphates (typically $(NH_4)H_2PO_4$ and/or $(NH_4)_2HPO_4$) and superphosphate (a mixture of calcium dihydrogen phosphate, $Ca(H_2PO_4)_2$, and calcium sulfate, $CaSO_4$), which are applied in large amounts as part of current agricultural practices. Until recently, laundry detergents were another source of this nutrient; they were added to remove hardness and buffer the wash water to maintain a slightly alkaline pH. Environmental awareness by consumers, however, has now led to phosphate removal from virtually all such detergents.

FIGURE 16.16 Testing for a blue phosphate complex



Testing for phosphate levels is a relatively easy application of colorimetry. The process involves the addition of ammonium molybdate, a chemical that forms a blue compound if phosphate is present — and the more phosphate, the more intense the blue colour of the solution. If a set of standards containing known phosphate levels is produced in the same way as the unknown, the level in the unknown sample may then be determined. This can be done roughly by eye using colour matching, or by determining a calibration curve produced from an instrumental colorimeter.

16.4.3 Calibration curves

Creating a calibration curve is an important step in analytical chemistry to establish a relationship between the absorbance and the concentration of the substance being analysed.

Preparing and using a calibration curve

1. Prepare a series of coloured standard solutions of known concentrations. (The standard *must* be prepared under the same conditions as the test sample.)
2. Using a complementary coloured filter, read the absorbance.
3. Plot the calibration curve of absorbance (y-axis) versus concentration (x-axis) and include a line of best fit.
4. Interpolate your test sample data to determine its concentration using its absorbance reading.

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SAMPLE PROBLEM 9 Determining the absorbance of a solution using a calibration curve

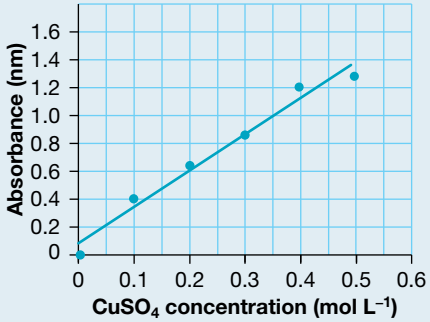
The absorbance of several solutions of $CuSO_4$ was determined through colorimetry, and the results are shown in the table provided. Using a calibration curve, determine the absorbance of a solution of 0.23 mol L^{-1} $CuSO_4$.

Concentration $CuSO_4$ (M)	Absorbance (nm)
0	0
0.1	0.406
0.2	0.638
0.3	0.854
0.4	1.202
0.5	1.276

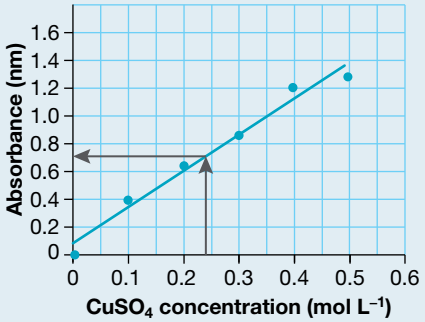
THINK

1. Plot the graph of absorbance versus concentration and the line of best fit. The line of best fit represents the calibration curve for this experiment.

WRITE



2. Draw a vertical line from the *x*-axis to the calibration line and then from the line to the *y*-axis to determine the corresponding absorbance for a solution of 0.23 mol L^{-1} $CuSO_4$.



The absorbance of the 0.23 mol L^{-1} solution is 0.7 nm.

PRACTICE PROBLEM 9

With the data from sample problem 9, determine the concentration for a sample of $CuSO_4$ with an absorbance of 0.2.

16.4.4 UV–visible spectroscopy

Compared to colorimetry, **UV–visible spectroscopy** measures the absorption of light at a particular wavelength, and in turn increases the accuracy. It can also be used to determine the concentration of many colourless compounds, such as caffeine, which do not absorb in the visible part of the electromagnetic spectrum. In many ways,

UV–visible spectroscopy technique used to study compounds that absorb light in the UV–visible region

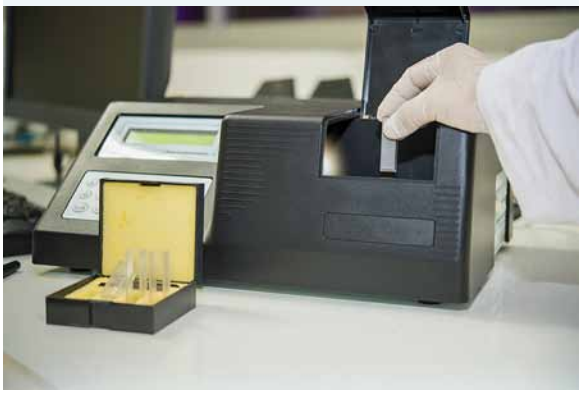
UV–visible spectroscopy is a more sophisticated development of instrumental colorimetry. The basic idea is the same — the amount of light absorbed is proportional to the concentration of the substance being tested. However, this technique is far more selective, and, therefore, less likely to suffer interference from similarly coloured compounds, due to using light at a specific wavelength. For example, although two compounds in a test sample appear blue, one of these might absorb more strongly at a specific wavelength compared to the other, enabling you to distinguish between them.

UV–visible spectroscopy is suitable for:

- coloured solutions (same as colorimetry)
- solutions that can be made coloured (same as colorimetry)
- solutions that absorb in the ultraviolet region (unique to UV–visible spectroscopy); you just need to consider the optimal wavelength to ensure you will get the maximum absorbance for your sample.

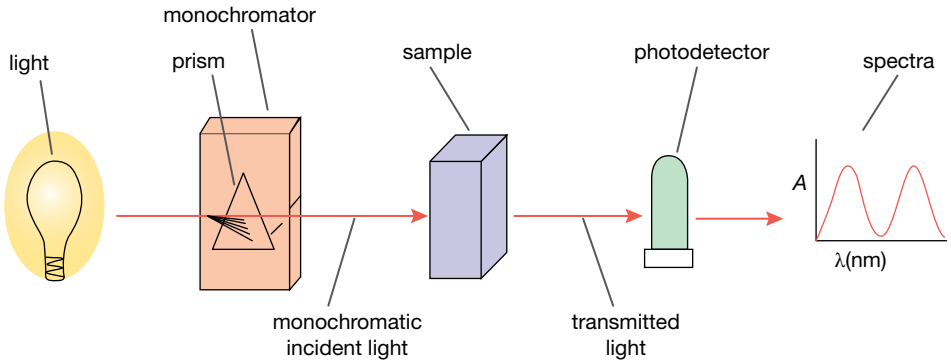
Just as many substances absorb light from the visible section of the electromagnetic spectrum (and, therefore, appear coloured), so do substances that absorb radiation from the ultraviolet region of the spectrum. Our eyes are not able to detect this UV radiation, which means substances absorbing in this region may not appear coloured. UV–visible spectroscopy is, therefore, suitable for many colourless substances as well as coloured ones.

FIGURE 16.17 UV–visible spectrophotometer



A UV–visible spectrophotometer is schematically illustrated in figure 16.18. Light is passed through a monochromator with a wavelength selection device, such as a prism. Light of this selected wavelength then passes through the sample and is received by a photodetector and recorded by computer. The resulting spectrum is given as a graph of absorbance (A) versus wavelength (λ).

FIGURE 16.18 UV–visible spectrophotometer schematic



The UV–visible spectrophotometer can be used both qualitatively and quantitatively.

Qualitative analysis using a UV–visible spectrophotometer

For **qualitative analysis**:

- The sample to be analysed is dissolved using a suitable solvent.
- A spectrum is obtained by measuring the absorbance against a range of wavelengths.
- This spectrum can be compared to known spectra for the suspected substances in the sample.
- In practice, UV–visible spectroscopy is used only as supporting evidence in an identification, not as a qualitative analysis method in its own right.

qualitative analysis use of experimental procedures to determine which elements are in a substance

Quantitative analysis using UV–visible spectroscopy

In **quantitative analysis**:

- The spectrum is measured of a pure sample of the substance to be measured.
- From the spectrum, the wavelength of highest absorbance (largest clean peak) is determined.
- A set of standards of known concentration is produced, and the absorbance at the determined wavelength is measured.
- A calibration curve is created from the standard absorbances.
- The absorbance of the test sample is measured using the same wavelength. (This allows a direct comparison to the standards to allow you to find the unknown test samples concentration.)
- The test sample absorbance is interpolated to find its concentration.

For example, to measure the level of glucose in a sample of urine, the spectrum of pure glucose would be obtained and a suitable wavelength chosen. Glucose samples of known concentration would then be tested at this wavelength and their absorbances noted. Finally, the sample of urine would be tested and the absorbance compared to that of the standards to obtain the urine glucose concentration. Note that in choosing a suitable wavelength for such an analysis, care must be taken to choose a wavelength at which the glucose in the urine is the only substance that absorbs (that is, a single clean peak).

quantitative analysis use of experimental procedures to determine the percentage composition of a compound or the percentage of a component in a mixture

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SAMPLE PROBLEM 10 Deriving the concentration of an unknown solution using absorbance

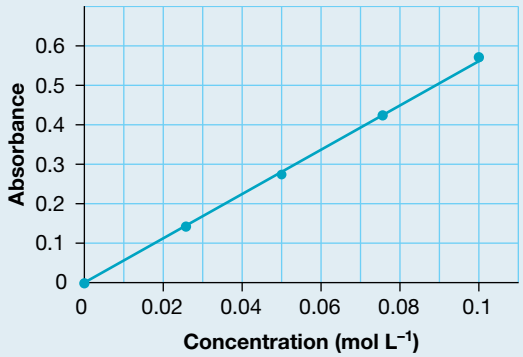
Standard solutions of $Cu(NO_3)_2$ were analysed with UV–visibility and their absorbance was determined, as shown in the provided table. Derive the concentration of the unknown solution.

Concentration (mol L^{-1})	Absorbance
0.025	0.135
0.050	0.270
0.075	0.420
0.100	0.570
unknown	0.340

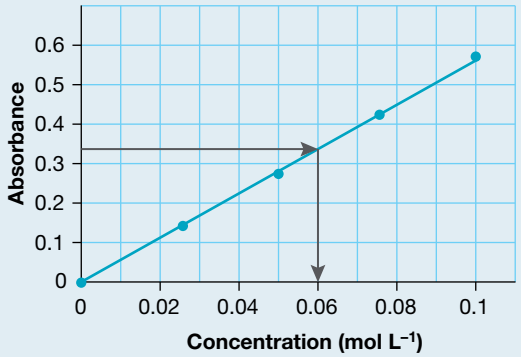
THINK

1. Plot the graph of absorbance versus concentration and the line of best fit. The line of best fit represents the calibration curve for this experiment.

WRITE



2. Determine the unknown concentration by drawing a horizontal line from the *y*-axis at the 0.340 nm absorbance value given to the calibration line, and from this to the *x*-axis to determine the corresponding concentration.



The unknown concentration is 0.060 mol L^{-1}.

PRACTICE PROBLEM 10

With data from sample problem 10, determine the concentration of a sample of $Cu(NO_3)_2$ with an absorbance of 0.23.

16.4 Activities

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16.4 Quick quiz on | 16.4 Exercise | 16.4 Exam questions

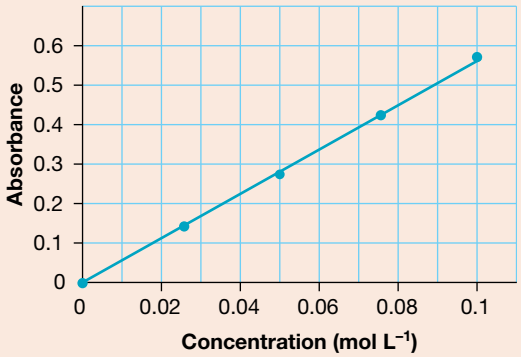
16.4 Exercise

1. What is a calibration curve and what is its purpose?
2. You are working in a biotechnology laboratory and have extracted some plasmid DNA from your bacterial cultures. The instructions state to test the concentration of DNA at 260 nm and 280 nm. Which analytical machine do you need to use? Explain.
3. Ca^{2+} ions in body fluid are commonly tested by forming a violet-coloured complex with the reagent ortho-cresolphthalein complexone. Standards of known calcium concentration produced the absorbance readings shown in the following table.

Ca^{2+} concentration (mM)	Absorbance
0.0	0.00
0.8	0.07
1.6	0.15
2.4	0.24
3.2	0.33
4.0	0.39

a. Create a calibration curve from the data.
b. You are testing a patient's blood serum, which returned an absorbance of 0.20 nm. What is the concentration of calcium in their blood?
c. The normal range for calcium in the blood is 2.2–2.6 mM. The patient was suspected of having hypocalcaemia (low blood serum calcium levels), which equates to a level less than 2.2 mM. Were the doctor's suspicions correct?
d. When performing colorimetry, what can you do to increase the accuracy and precision of your result?

4. Compare and contrast the use of an instrumental colorimeter and a UV–visible spectrometer.
5. MC Use the calibration curve provided of copper(II) nitrate standards to answer the following question.



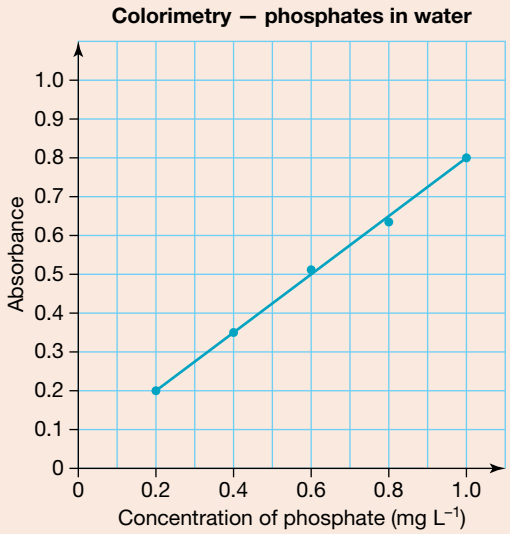
A test sample returned an absorbance of 0.1 nm. What is the concentration of copper(II) nitrate in the sample?
A. 0.58 mol L^{-1} **B.** 0.017 mol L^{-1} **C.** 0.1 M **D.** 0.58 nm

16.4 Exam questions

Question 1 (1 mark)

MC The phosphate concentration of water in a dam on a rural property was tested after excessive algal growth had been observed. A 500 g sample was collected from the dam. As initial measurements proved to be high, the water was diluted by a factor of 5 before testing by colorimetry.

If the absorbance reading at 840 nm of the diluted water sample was found to be 0.32, what was the phosphate concentration in the *undiluted* sample?



A. 0.07 mg L^{-1}
B. 0.32 mg L^{-1}
C. 0.35 mg L^{-1}
D. 1.75 mg L^{-1}

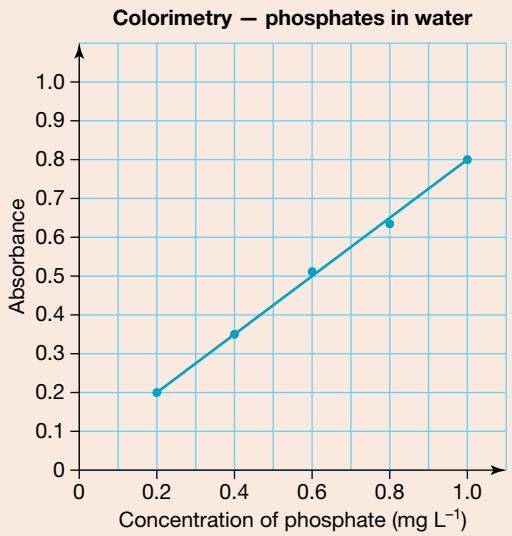
Question 2 (1 mark)

MC Two students were measuring the phosphate concentration in water by colorimetry. If they obtained a concentration that was higher than the rest of the class, which of the following is *not* consistent with their results?

A. They inadvertently used a container that had not been thoroughly washed since it was last used for water collection.
B. They did not seal their water sample and some water evaporated between collection and measuring absorbance.
C. They used a larger sample of water in their reaction to produce a coloured compound.
D. The cuvette (cell) they used for recording absorbance had some water in it.

Question 3 (1 mark)

A student performed a colorimetry experiment to determine the concentration of phosphorus present as phosphate in pond water. Five standard solutions were prepared and their absorbances were measured and plotted on the calibration curve shown.



If the absorbance of the (undiluted) water sample was 0.67, what was the concentration of phosphorus (include units)?

Question 4 (1 mark)

A dam on a rural property had a particularly high concentration of phosphate contaminant due to a fertiliser spill. As such, a sample of water had to be diluted as follows before testing in the colorimeter.

10.0 mL of the contaminated water was pipetted into a 250 mL volumetric flask and made up to the mark with distilled water.

If the phosphate concentration of the diluted water was found to be 0.45 mg of phosphorus per litre, what was the phosphate concentration in the undiluted sample?

Question 5 (1 mark)

MC How is colorimetry used to measure concentrations?

A. By monitoring the light absorbed by a coloured compound and comparing it to the absorbance of a set of standards
B. By measuring the temperature change produced when a substance is combusted
C. By recording the length of time a compound takes to pass through a gel-packed column
D. By measuring the light emitted by a metallic compound when it is heated in a flame

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16.5 Review

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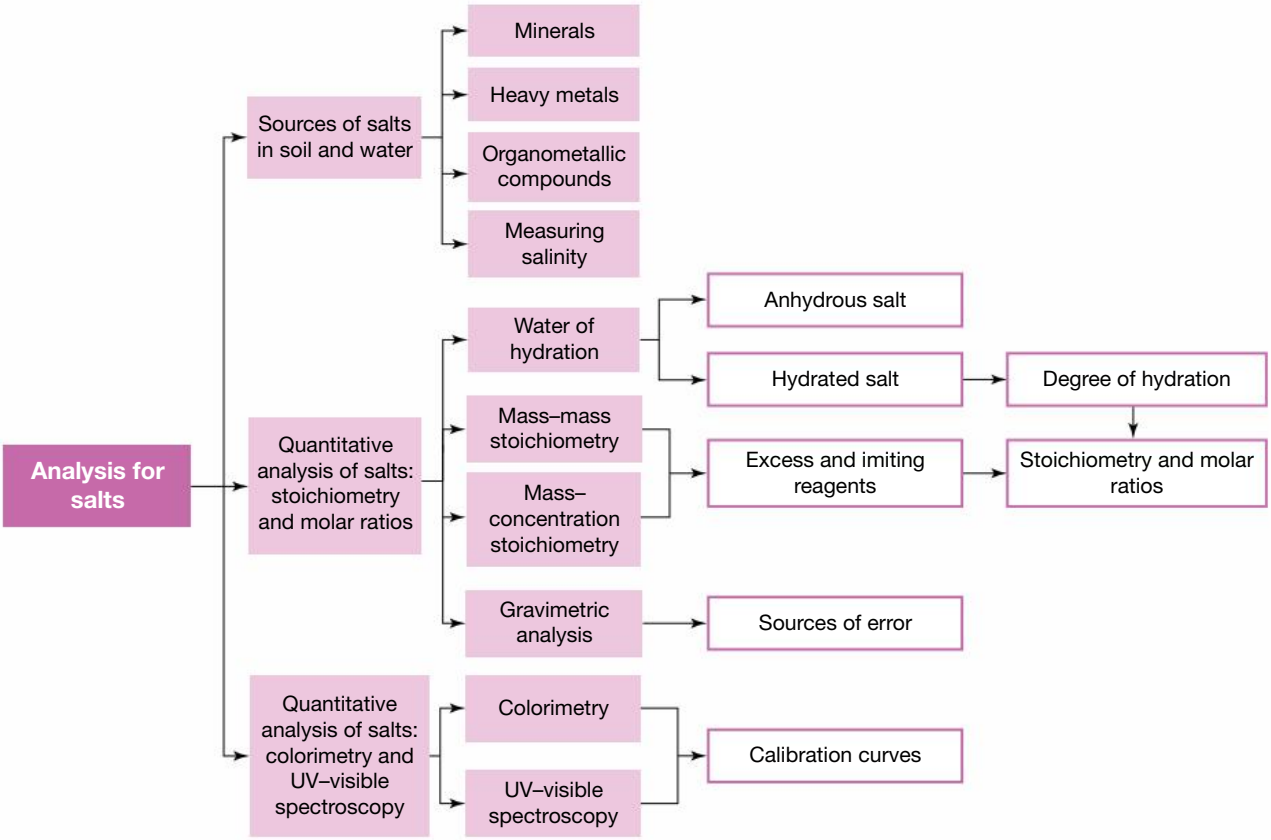
16.5.1 Topic summary



16.5.2 Key ideas summary

online only

16.5.3 Key terms glossary

online only

Resources

Solutions	Solutions — Topic 16 (sol-0846)
Practical investigation eLogbook	Practical investigation logbook — Topic 16 (elog-1870)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 16 (doc-37913) Key idea summary — Topic 16 (doc-37914)
Exam question booklet	Exam question booklet — Topic 16 (eqb-0125)

16.5 Activities

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16.5 Review questions

1. A solution containing 5.10 g of barium nitrate reacts completely with a solution of sodium sulfate. The unbalanced equation for this reaction is:

$$_____\,Ba(NO_3)_2(aq) + _____\,Na_2SO_4(aq) \rightarrow _____\,NaNO_3(aq) + _____\,BaSO_4(s)$$

   a. Balance the equation by inserting the necessary coefficients and identify the precipitate formed.
   b. Calculate the mass of precipitate formed.
2. Barbecues burning charcoal briquettes are unsafe for indoor use because of the colourless, odourless, poisonous gas produced.
   a. Find the $n(O_2)$ gas that reacts with 3.5 g of charcoal briquettes (assume pure C) to produce carbon monoxide.
   b. If there is a plentiful supply of air, a safe colourless, odourless gas is produced. Find the mass of this gas produced if the same amount of charcoal is burned. You will need to write another equation.
3. In respiration, the equation for the reaction that produces energy in our bodies is

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

   If 8.90 g of glucose is used, determine the following.
   a. The mass of oxygen needed
   b. The mass of carbon dioxide produced
4. The silver used in jewellery and tableware becomes tarnished when exposed to air containing small amounts of hydrogen sulfide. The tarnish is a layer of silver sulfide.

$$4Ag(s) + 2H_2S(g) + O_2(g) \rightarrow 2Ag_2S(s) + 2H_2O(g)$$

   Calculate the mass of the tarnish when 0.025 g of silver is reacted.
5. The metal tungsten, used to make the filaments for incandescent light bulbs, can be obtained from its oxide by reduction with hydrogen.

$$WO_3(s) + 3H_2(g) \rightarrow W(s) + 3H_2O(g)$$

   a. What mass of tungsten can be obtained from 200 g of its oxide?
   b. What mass of hydrogen is required?
6. What mass of copper is required to react completely with 250 mL of 0.100 M $AgNO_3$?

$$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$

7. As an emergency procedure, the *Apollo 13* astronauts used lithium hydroxide to remove carbon dioxide from the interior of their crippled spacecraft as it returned from the moon. Like all hydroxides, lithium hydroxide forms a metal carbonate when it reacts with carbon dioxide.
The equation for this reaction is:

$$2LiOH(s) + CO_2(g) \rightarrow Li_2CO_3(aq) + H_2O(l)$$

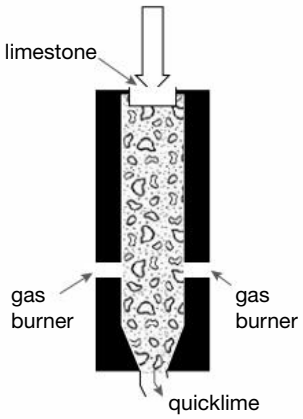
A possible alternative chemical for this process might have been the more readily available sodium hydroxide.

a. Calculate the mass of carbon dioxide that could be removed per kilogram of lithium hydroxide.
b. Write the equation for the reaction between sodium hydroxide and carbon dioxide.
c. From part **b**, calculate the mass of carbon dioxide that can be removed per kilogram of sodium hydroxide.
d. Use your answers to parts a and c to suggest a reason for the choice of lithium hydroxide rather than sodium hydroxide in a spacecraft.
e. Derive the ionic equations for both the reactions mentioned in this question.

8. Limestone is an important raw material for industry. When heated in a kiln, it decomposes to form quicklime, CaO, and carbon dioxide.
The equation for this reaction is

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

A typical kiln is shown in the figure. The production of lime from such kilns was one of Australia's earliest industries.

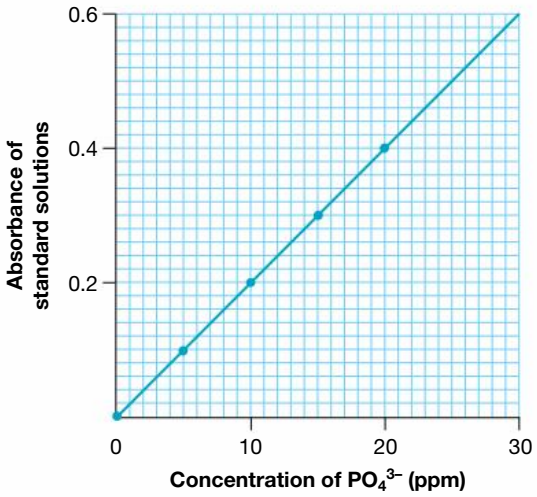


a. If limestone containing 83.5 per cent calcium carbonate is used, calculate the mass of quicklime that would be produced from 100 tonnes of this limestone.
b. This process is sometimes called 'lime burning'. Why is this not a correct term to use?

9. Examine figure 16.4, showing measured salinity levels at Morgan in South Australia.

a. Explain why electrical conductivity can be used to measure the level of salt in the river.
b. Describe the trends in salt content before the early 1980s and after the early 1980s.
c. Why is it incorrect to interpret these results as evidence of sodium chloride content?
d. Why is it more correct to interpret the results in terms of total salt content?

10. The level of phosphorus (as phosphate) in a detergent may be determined colorimetrically by reacting it with a molybdate solution to develop an intense blue compound. In the analysis of one particular brand of detergent, a 1.000 g sample was dissolved in 1000 mL. A carefully measured 20.0 mL sample of this was then treated with molybdate solution and heated to develop the blue colour. When analysed in an instrumental colorimeter, a reading of 0.260 was obtained.
Four standard solutions containing phosphorus levels of 5, 10, 15 and 20 ppm were similarly treated and their absorbances measured to produce the graph provided.



a. What is the concentration of phosphorus in the diluted solution?
b. What is the percentage by mass of phosphorus in the detergent?
c. What colour was the light used for this analysis?

11. In each of the following situations, we want to analyse gravimetrically the component written in bold type. From the list in brackets, choose the ions that should be added to achieve this.
    a. **Sulfate** (magnesium, aluminium, lead, ammonium)
    b. **Chloride** (sodium, barium, calcium, silver)
    c. **Carbonate** (potassium, lithium, barium, ammonium)
    d. **Calcium** (nitrate, ethanoate, sulfate, hydroxide)
12. 'Every compound is soluble to a greater or lesser extent.' Discuss the implications of this statement for gravimetric procedures. Would this lead to a systematic or a random error? Explain.
13. The following steps were considered by a student when designing a gravimetric procedure to estimate the level of magnesium carbonate in a sample of Epsom salts. The steps have been written down in random order. Rearrange these steps into an order that would be acceptable, remembering that some steps might be repeated more than once and that some steps might not be required.
    A. Filter and collect the precipitate produced.
    B. Add a solution of potassium chloride.
    C. Accurately weigh out a sample of Epsom salts.
    D. Accurately weigh the precipitate.
    E. Dissolve the Epsom salts in a volume of water.
    F. Dry the precipitate.
    G. Measure out an accurate volume of water.
    H. Add a solution of barium chloride.

16.5 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC Which of the following would be most likely to be a contaminant in a sample of surface water?

A. Cl^-
B. SO_4^{2-}
C. Ca^{2+}
D. Hg^+

Question 2

MC A field technician collected a creek water sample (CW) and added two preserving agents (A and B) to her sample. In her field kit was also a container of deionised water (DW).

The protocol also called for the preparation of a blank.

Which substances should the blank contain?

A. A, B and CW
B. A, B and DW
C. A, B
D. DW only

Question 3

MC A calibration curve is a graph of instrumental reading against what?

A. Mass
B. Time
C. Volume
D. Concentration

Question 4

MC Before digital printing, sodium thiosulfate, $Na_2S_2O_3$, known as 'hypo' by photographers, was used to remove excess silver bromide, AgBr, in the liquid-based film-developing process according to the following equation:

$$2Na_2S_2O_3(aq) + AgBr(s) \rightarrow Na_3Ag(S_2O_3)_2(aq) + NaBr(aq)$$

In the *balanced* equation for the reaction between sodium thiosulfate and silver bromide

A. 1 mole of sodium thiosulfate solution produces 1 mole of sodium bromide solution.
B. 3 moles of sodium thiosulfate solution produce 2 moles of $Na_3Ag\ (S_2O_3)_2$.
C. 1 mole of sodium thiosulfate solution produces 2 moles of sodium bromide solution.
D. 1 mole of silver bromide produces 1 mole of $Na_3Ag(S_2O_3)_2$.

Question 5

MC Phosphorus may be prepared from calcium phosphate according to the following equation:

$$2Ca_3(PO_4)_2(s) + 6SiO_2(s) + 10C(s) \rightarrow P_4(s) + 10CO(g) + 6CaSiO_3(s)$$

How much phosphorus can be produced if 1000 kg of calcium phosphate is used completely?

A. 50 kg
B. 100 kg
C. 200 kg
D. 500 kg

Question 6

MC A gelatinous precipitate of iron(III) hydroxide may be prepared according to the following equation:

$$FeCl_3(aq) + 3NaOH(aq) \rightarrow 3NaCl(aq) + Fe(OH)_3(s)$$

A solution containing 16.23 g of $FeCl_3$ is mixed with 300 mL of a 1.00 M NaOH solution. Which of the following statements is correct?

A. $FeCl_3$ is in excess.
B. NaOH is in excess.
C. $FeCl_3$ is the limiting reagent.
D. Neither reagent is in excess.

Question 7

MC We want to analyse a sample of salty water for its chloride ion content. This is to be done using gravimetric analysis. Which of the following chemicals could be added for the formation of a precipitate?

A. $CaCO_3$
B. $Ba(NO_3)_2$
C. K_2SO_4
D. $AgNO_3$

Question 8

MC A sample of lawn food is analysed gravimetrically. As part of the procedure, a precipitate of barium sulfate is formed. However, when this is later weighed, it is still slightly damp. As a result of this, what would the calculated percentage of sulfate be?

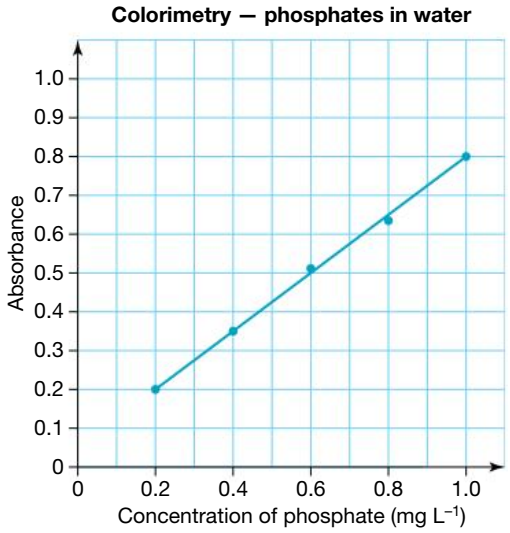
A. Increased
B. Decreased
C. Unchanged
D. Varied in a random manner

Question 9

MC The phosphate levels in a dam were tested using colorimetry. The provided calibration curve was produced.

The sample gave an absorbance reading of 0.30. What was the concentration of phosphorus (in the form of phosphate) in the water?

A. 0.27 mg L^{-1}
B. 0.30 M
C. 0.33 M
D. 0.33 mg L^{-1}



Question 10

MC The sodium chloride concentration of bore water was determined by gravimetric analysis. Adding an excess of silver nitrate to 100.0 mL of bore water produced a precipitate of silver chloride. Assuming all chloride ions present are due to NaCl, what was the concentration of salt in g L^{-1} if 1.556 g of precipitate was formed?

A. 0.6348
B. 0.3852
C. 3.852
D. 6.348

Section B — Short answer questions

Question 11 (6 marks)

A sample of limestone is analysed for its calcium carbonate, $CaCO_3$, content as follows. A 25.00 g sample is crushed and heated to a high temperature. It is then cooled, weighed and reheated until a constant mass is obtained. The mass remaining at the end of this process is 11.64 g.

a. What gas is evolved during this process? **(1 mark)**
b. Write the equation for the decomposition of calcium carbonate in the limestone at high temperature. **(1 mark)**
c. Use the results obtained to calculate the mass of calcium carbonate in the limestone, and hence its percentage purity. **(3 marks)**
d. State one important assumption in this method. **(1 mark)**

Question 12 (4 marks)

Phosphorus is an important plant nutrient and is, therefore, present in many fertilisers. It can be determined gravimetrically by carrying out a number of steps, eventually leading to its precipitation as insoluble $Mg_2P_2O_7$. In one such analysis, a 14.298 g sample of fertiliser yielded 4.107 g of $Mg_2P_2O_7$ precipitate.

a. Calculate the mass of phosphorus in the precipitate. **(2 marks)**

b. Hence, calculate the percentage of phosphorus in the fertiliser. **(1 mark)**

c. State one important assumption that is made in this method. **(1 mark)**

Question 13 (7 marks)

As part of a research project into tidal effects in a river estuary, a student takes water samples at various locations and determines the concentration of salt (as sodium chloride) in each. Her method involves collecting 10.0 L of water from each location and boiling it to reduce the volume. Excess silver nitrate solution is then added to precipitate all the chloride ions present as silver chloride. This precipitate is then collected, dried and weighed to constant mass.

Following this procedure, one particular sample produced 4.463 g of precipitate.

a. Write a balanced equation for the reaction that leads to the formation of the precipitate. **(1 mark)**

b. Calculate the mass of sodium chloride in this particular sample. **(3 marks)**

c. Hence, calculate the concentration of sodium chloride in this sample, in g L^{-1}. **(1 mark)**

d. In what way does boiling the sample to reduce its volume affect your calculation? **(1 mark)**

e. Some important assumptions are made by the student in this procedure. Name one. **(1 mark)**

Question 14 (2 marks)

4.6 g of anhydrous zinc sulfate with an M of 161.5 g mol^{-1} was obtained by evaporating the water from 8.2 g of the crystalline hydrated salt with empirical formula $ZnSO_4 \cdot xH_2O$. Calculate the value of x.

Question 15 (3 marks)

The degree of hydration refers to the number of water molecules (water of crystalisation) associated with certain ionic compounds; for example, $CuS_4 \cdot 5H_2O$ has a degree of hydration of 5.

The degree of hydration can be found by finding the simplest whole number ratio of water to the ionic compound.

A 1.634 g sample or hydrated calcium chloride $(CaCl_2 \cdot xH_2O)$ was heated until all the water was removed, leaving a solid of mass 1.234 g.

Calculate the ratio of calcium chloride : water to find the degree of hydration, x.

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AREA OF STUDY 2 How are chemicals measured and analysed?

OUTCOME 2

Calculate solution concentrations and predict solubilities, use volumetric analysis and instrumental techniques to analyse for acids, bases and salts, and apply stoichiometry to calculate chemical quantities.

PRACTICE EXAMINATION

STRUCTURE OF PRACTICE EXAMINATION		
Section	**Number of questions**	**Number of marks**
A	20	20
B	6	30
	Total	**50**

Duration: 50 minutes

Information:

- This examination consists of two parts; you must answer all question sections.
- Pens, pencils, highlighters, erasers, rulers and a scientific calculator are permitted.
- You may use the VCE Chemistry Data Book for this task.

on Resources

Weblink VCE Chemistry Data Book

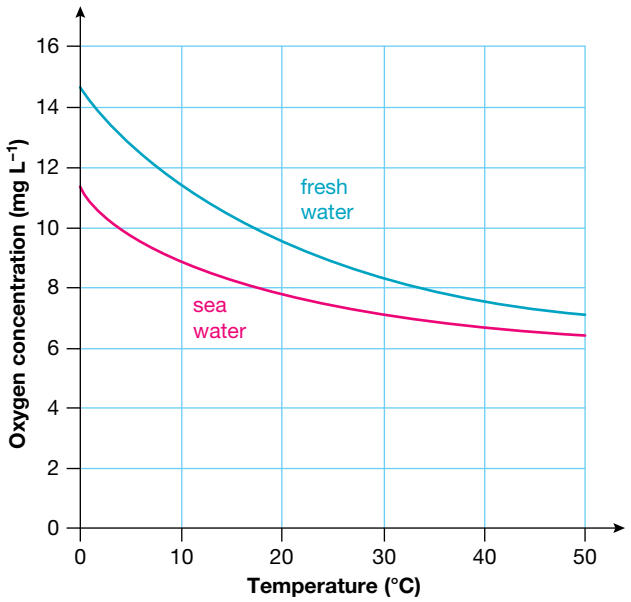
SECTION A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

1. According to the label on a bottle of mineral water, the concentration of the hydrogen carbonate (or bicarbonate) ion is 248 mg L^{-1}. If a 350 mL bottle of this water was consumed, what would the mass of HCO_3^- ingested in grams be?

A. 0.0868
B. 0.141
C. 1.41
D. 86.8

2. The following graph shows the concentration of dissolved O_2 in both fresh and sea water.



According to the graph, at 15 °C, what would the mass of dissolved oxygen in 250 mL of fresh water be closest to?

A. 0.0021 g
B. 0.0026 g
C. 2.1 g
D. 2.6 g

3. A popular alcoholic beverage lists its alcohol content as 35.0%(v/v). The mass of ethanol present in one 'standard drink' is 10.0 g, which corresponds to a volume of 8.90 mL of pure ethanol. What volume of the alcoholic beverage would a person need to consume to reach the alcohol content of a standard drink?

A. 8.9 mL
B. 25 mL
C. 29 mL
D. 35 mL

4. The following table shows the allowable level of heavy metal contaminants in drinking water sources.

Metal	Allowable level (mg L^{-1})
Arsenic	0.010
Chromium	0.050
Cadmium	0.002
Lead	10

A 100 mL sample of drinking water was analysed by atomic absorption spectroscopy and found to contain 0.0010 mg of cadmium and 0.0002 mg of arsenic. What do these results mean?

A. The sample is above the allowable limit for both arsenic and cadmium.
B. The sample is above the allowable limit for arsenic but below the limit for cadmium.
C. The sample is above the allowable limit for cadmium but below the limit for arsenic.
D. The sample is below the allowable limit for both arsenic and cadmium.

5. Which of the following sucrose solutions has the highest amount of solute?

A. 10 mL of a 0.40 mol L^{-1} solution
B. 120 mL of a 0.010 mol L^{-1} solution
C. 40 mL of a 0.12 mol L^{-1} solution
D. 25 L of a 1.00×10^{-4} mol L^{-1} solution

6. A solution of nitrate ions was prepared by mixing 100 mL of 1.0 mol L^{-1} sodium nitrate solution with 100 mL of 1.0 mol L^{-1} aluminium nitrate solution. What was the resulting nitrate ion concentration?

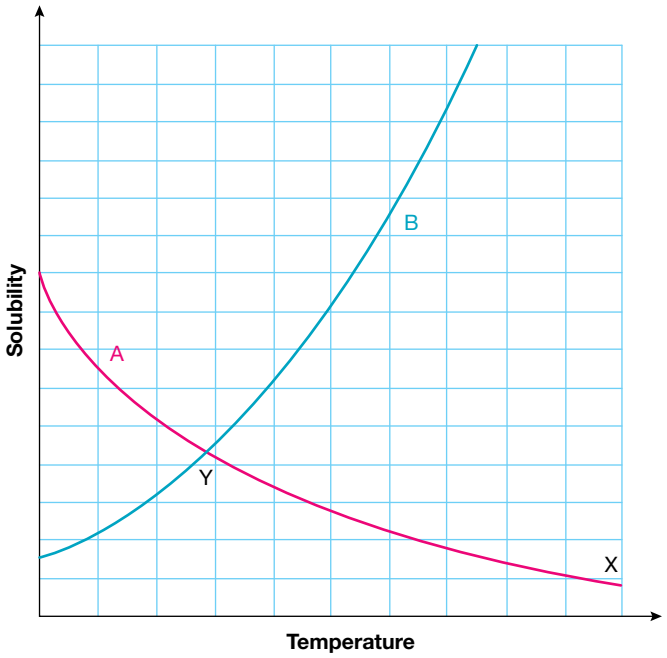
A. 4.0 mol L^{-1}

B. 3.0 mol L^{-1}

C. 2.0 mol L^{-1}

D. 1.0 mol L^{-1}

7. The solubility curves for two different substances are shown in the provided graph.



From the information provided, which of the following statements is likely to be true?

A. A is a gas while B is a solid.

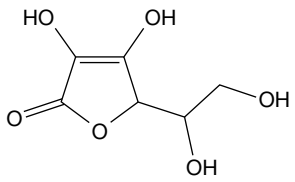
B. A is more soluble than B at high temperatures.

C. Point X represents an unsaturated solution for substance A.

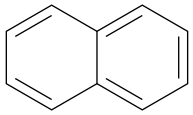
D. Point Y represents a supersaturated solution for both A and B.

8. Which of these organic compounds would be expected to dissolve well in water?

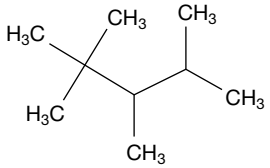
A.



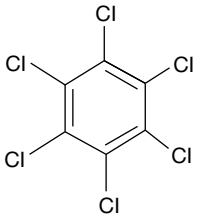
B.



C.



D.



A. Compound A
B. Compound B
C. Compound C
D. Compound D

9. Which of the following impurities would *not* be removed from water supplies by precipitation?
A. Calcium ions
B. Suspended solids
C. Phosphates
D. Salinity

10. A student weighed out 2.396 g of anhydrous sodium carbonate and made this up to the mark in a 250.0 mL volumetric flask. What was the molarity of the sodium carbonate solution?
A. 0.000090 42 mol L^{-1}
B. 0.09042 mol L^{-1}
C. 0.01155 mol L^{-1}
D. 0.5990 mol L^{-1}

11. A beaker contains 50 mL of 0.15 mol L^{-1} NaCl solution. What volume of water must be added to this solution to change its concentration to 0.050 mol L^{-1}?
A. 20 mL
B. 50 mL
C. 100 mL
D. 150 mL

12. A student needs to prepare a standard solution of anhydrous sodium carbonate for a volumetric analysis. Which set of equipment should be used for this purpose?
A. Electronic balance, measuring cylinder, beaker and spatula
B. Volumetric flask, burette, spatula and pipette
C. Electronic balance, pipette, funnel and spatula
D. Volumetric flask, funnel, spatula and electronic balance

13. Which of the following statements about greenhouse gases is true?
A. Water vapour in the atmosphere makes no contribution to the natural greenhouse effect.
B. Carbon dioxide contributes only to the enhanced greenhouse effect.
C. Methane gas is an effective absorber of infrared radiation.
D. Carbon dioxide absorbs ultraviolet radiation but does not absorb infrared radiation.

14. In an experiment involving the decomposition of hydrogen peroxide, Amy collected 3.20 g of evolved oxygen at 20 °C and 101 kPa pressure.
What would the volume of oxygen collected be?
A. Slightly greater than 2.48 L
B. Slightly less than 2.48 L
C. Equal to 2.48 L
D. Not enough information is given to decide.

15. A student reacted 0.200 mol of hydrochloric acid with excess magnesium and collected the hydrogen evolved. The temperature and pressure in the laboratory at the time of the experiment were 22 °C and 100 kPa.
The equation for the reaction is

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g).$$

What was the volume of hydrogen collected, in litres?
A. 2.45
B. 2.48
C. 4.90
D. 4.96

16. A 0.165g sample of gas collected in an experiment was found to have a volume of 120 mL at SLC. What is the molar mass of this gas closest to?

A. 29 $gmol^{-1}$

B. 34 $gmol^{-1}$

C. 65 $gmol^{-1}$

D. 80 $gmol^{-1}$

17. One of the ways of measuring salinity makes use of conductivity measurements. These can detect the overall concentration of ions but not the particular ions present. Which of the following solutions would be expected to have the highest conductivity?

A. 0.005 mol L^{-1} $CaSO_4(aq)$

B. 0.005 mol L^{-1} $AlCl_3(aq)$

C. 0.008 mol L^{-1} $NaCl(aq)$

D. 0.006 mol L^{-1} $MgCl_2(aq)$

18. Which of the following substances is *not* an organometallic contaminant?

A. Methyl iodide

B. Tetraethyl lead

C. Tributyl tin

D. Methyl mercury

19. Upon heating, a hydrated salt produced 1.05 mol of water and 0.15 mol of anhydrous salt. What is the degree of hydration of this salt?

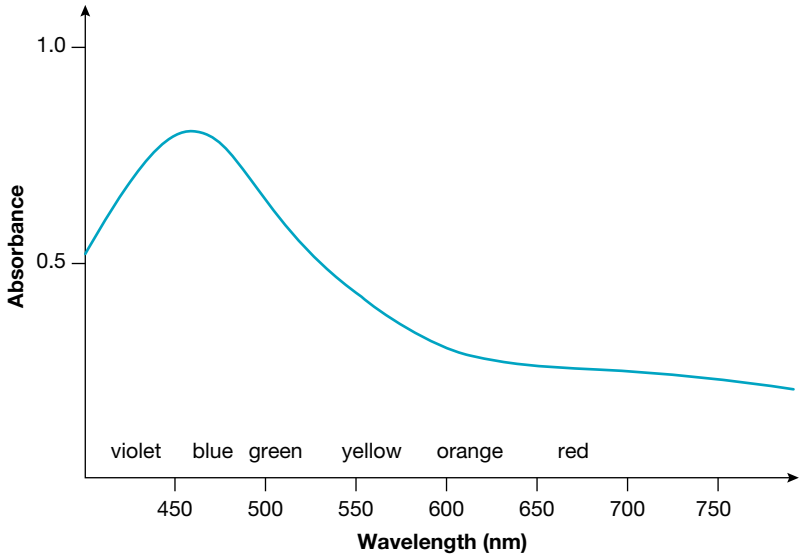
A. 5

B. 6

C. 7

D. 10

20. The UV–visible absorption spectrum of a transition metal solution is shown in the provided figure. The colours corresponding to the wavelengths are also shown.



Which of the following statements is correct?

A. The solution is blue and blue light should be used to measure absorbance.

B. The solution is orange and orange light should be used to measure absorbance.

C. The solution is blue and orange light should be used to measure absorbance.

D. The solution is orange and blue light should be used to measure absorbance.

SECTION B — Short answer questions

Question 21 (4 marks)

Express the following solutions in the concentration unit given.

a. A 375 mL can of soft drink contains 4.05 g of sucrose. What is the sucrose concentration in g L^{-1}? **(1 mark)**

b. The National Health and Medical Research Council (NHMRC) of Australia specifies a maximum level of 1.5 milligrams per litre for fluoride in drinking water.

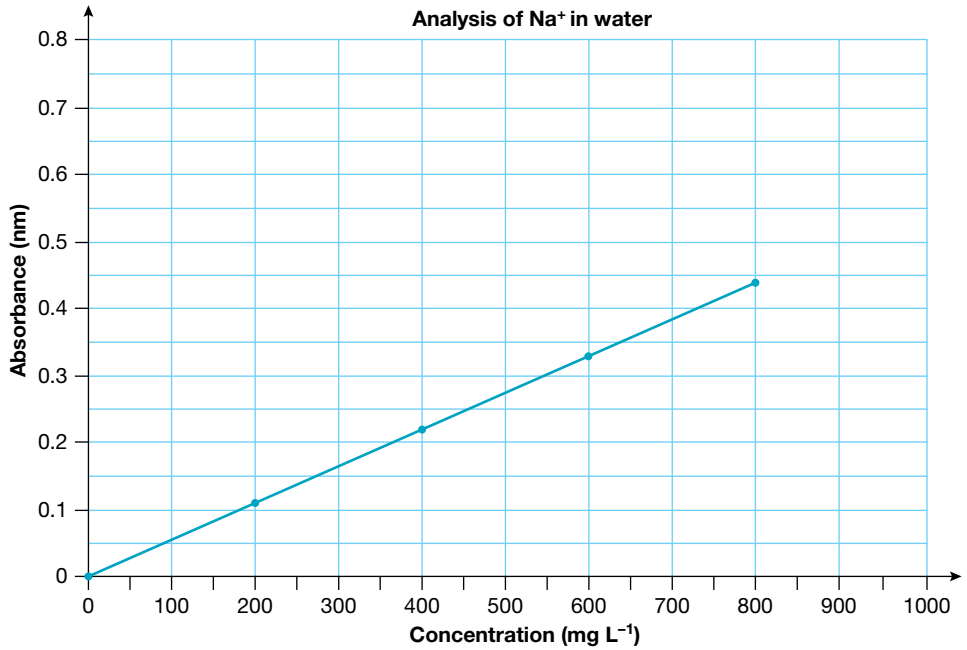
i. What mass of fluoride, in grams, would be present in a 250 mL glass of this water? **(1 mark)**

ii. What would be the molar concentration of fluoride at this level of concentration? **(1 mark)**

iii. What would this concentration be in ppm? **(1 mark)**

Question 22 (4 marks)

An analysis was made of sodium ions present in the stream water running through a region known to be affected by salinity. A suitable instrumental method was used for the analysis. An initial test of the water in the stream gave a reading that was off scale and so the samples to be tested were diluted to one-tenth of their original concentration before being reanalysed. The calibration curve used for the analysis is shown in the provided figure.



a. What is the highest concentration (in mg L^{-1}) that can be determined using this calibration graph? **(1 mark)**

b. The diluted water samples gave an absorbance reading of 0.30.

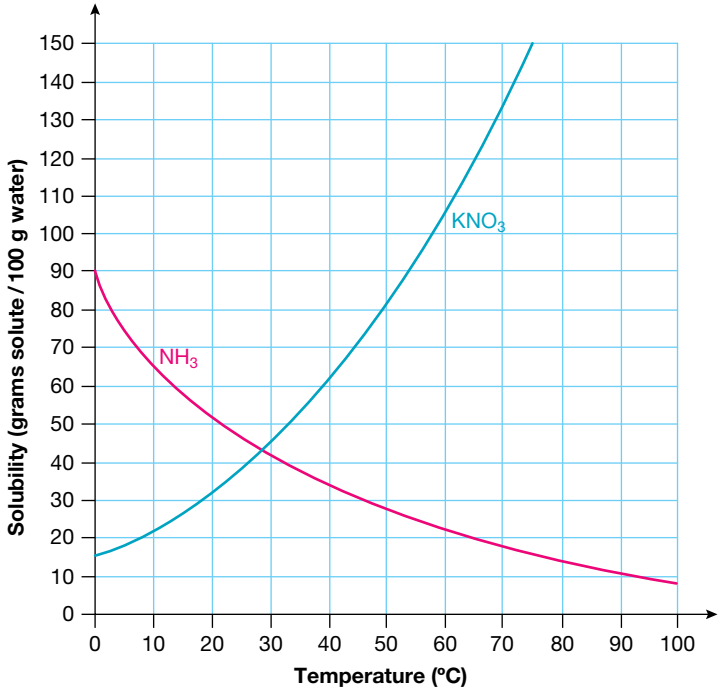
i. What is the concentration of sodium ions in the stream water measured in mg L^{-1}? **(1 mark)**

ii. What mass of sodium in kg is present in 3.5 megalitres (ML) of the stream water? **(1 mark)**

c. What volume of distilled water would have been added to 50.0 mL of the stream water in order to dilute it to one-tenth of its original concentration? **(1 mark)**

Question 23 (3 marks)

The solubility curves for ammonia and potassium nitrate are shown in the provided figure.



According to these curves:

a. What would be the maximum amount of ammonia that would dissolve in 100 g H_2O at 65 °C? **(1 mark)**

b. At what minimum temperature would 46 g of KNO_3 dissolve in 50 g of water? **(1 mark)**

c. A saturated solution of potassium nitrate was prepared in 200 g water at 40 °C. The solution was then cooled to 0 °C. What mass of solute would crystallise? **(1 mark)**

Question 24 (8 marks)

The concentration of a barium hydroxide solution was determined by titrating it with a previously standardised solution of hydrochloric acid. The concentration of the acid was 0.0500 M. In the titration, the hydrochloric acid was placed in a pre-rinsed burette and 20.00 mL aliquots of the barium hydroxide solution were delivered to a conical flask using a pre-rinsed pipette. The titration was repeated four times and the following titres were obtained: 18.64 mL, 18.15 mL, 18.18 mL and 18.10 mL.

a. With what substance should the conical flask have been rinsed before use? **(1 mark)**

b. Calculate the average of the concordant titres. **(1 mark)**

c. Write the equation for the reaction between hydrochloric acid and barium hydroxide. **(1 mark)**

d. Calculate the concentration of the barium hydroxide solution based on the given data. **(3 marks)**

e. The titration was carried out using bromothymol blue as an indicator. Bromothymol blue is yellow in acidic solutions and blue in basic solutions. What were the colours of the indicator, in order, seen throughout the titration? **(1 mark)**

f. If the burette had been rinsed with water, how would this have affected the value of the concentration of barium hydroxide calculated in this analysis? **(1 mark)**

Question 25 (4 marks)

In an experiment to determine the molar mass of a gas, David and Kannika collected 210 mL of carbon dioxide. This was found to have a mass of 0.386 g. After conversion to SLC, the volume of gas was found to be 216 mL.

a. Calculate the number of mol of carbon dioxide collected. **(1 mark)**

b. Calculate the molar volume (at SLC) that was determined in this experiment. **(2 marks)**

c. A group working at the next bench obtained a molar volume result of 24.9 L using 250 mL of oxygen gas instead of carbon dioxide. Explain why the two results are so similar. **(1 mark)**

Question 26 (7 marks)

Compounds of zinc can cause water to appear turbid (cloudy) in appearance and they also give the water an unpleasant taste. The concentration of Zn^{2+} ions in some turbid water was analysed gravimetrically by precipitating the zinc ions as zinc phosphate using sodium phosphate to form the precipitate.

When excess sodium phosphate solution was added to 250.0 mL of the zinc-contaminated water, 1.844 g of precipitate was formed.

a. Write the *ionic* equation for the formation of the zinc phosphate precipitate from the zinc ions and phosphate ions. Include symbols of state. **(1 mark)**

b. **i.** Calculate the amount in mol of zinc phosphate precipitated. **(1 mark)**

ii. Calculate the molar concentration of Zn^{2+} ions present in the water sample. **(2 marks)**

c. What experimental technique should be used to ensure that the precipitate is fully dried before weighing? **(1 mark)**

d. **i.** What is one assumption made about the precipitate formed? **(1 mark)**

ii. What would have been the effect on the calculated Zn^{2+} concentration if 10 mL more of the sodium phosphate solution had been added? Explain. **(1 mark)**

PRACTICE SCHOOL-ASSESSED COURSEWORK

ASSESSMENT TASK — A report of an application of chemical concepts to a real-life context

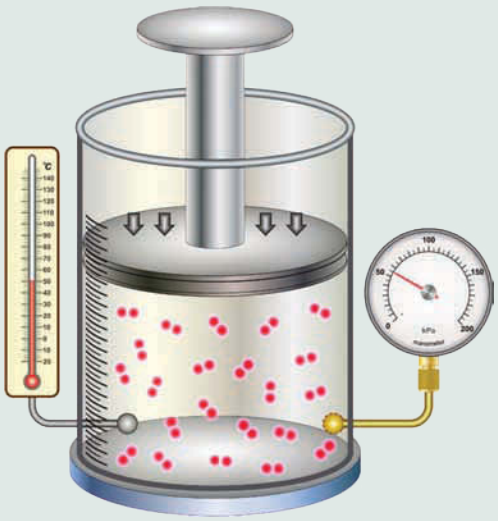
In this task, you will research and write a report on the application of the ideal gas equation to a real-life context.

- Computers with internet access and other research sources are permitted.
- Pens, pencils, highlighters, erasers, rulers and a scientific calculator are permitted.
- Use of the VCE Chemistry Data Book is permitted.

Total time: 1 hour (30 minutes researching, 30 minutes writing)

Total marks: 24 marks

Word count: 300–400 words



THE IDEAL GAS EQUATION

Task

The ideal gas equation is studied in VCE Chemistry, in which the following equations are utilised for various calculations:

$$PV = nRT \qquad n = \frac{m}{M} \qquad n = \frac{V}{V_m}$$

However, the true impact and significant role of the ideal gas equation in our everyday lives is often not realised. This assessment requires you to research and write a report on an application of the ideal gas equation to a real-life context of your choice. You will concentrate on one application, and will be given 30 minutes to research a context and 30 minutes to communicate the application and significance of the ideal gas equation in your chosen scenario.

The assessment of your report will be based on the following:

- discussion of relevant chemical information, ideas, concepts, theories or models and the connections between them
- evaluation of the significance and implications of the ideal gas equation within the real-life context
- use of the ideal gas equation (***PV*** = ***nRT***)
- use of stoichiometry to solve calculations related to chemical reactions
- use of clear, coherent and concise expression to communicate scientific ideas
- acknowledgment of all sources of information, via standard scientific referencing conventions.

on Resources

Digital document U2AOS2 School-assessed coursework (doc-38066)

17 Scientific investigations

KEY KNOWLEDGE

online only

In this area of study, you will adapt or design and then conduct a scientific investigation related to chemical equations and/or analysis, which must include the generation of primary data. You will organise and interpret the data and reach a conclusion in response to your research question.

Investigation design

- chemical science concepts specific to the selected scientific investigation and their significance, including the definition of key terms
- scientific methodology relevant to the selected scientific investigation, selected from the following: classification and identification; controlled experiment; fieldwork; modelling; product, process or system development; or simulation
- techniques of primary qualitative and quantitative data generation relevant to the investigation
- accuracy, precision, repeatability, reproducibility, resolution and validity of measurements in relation to the investigation
- health, safety and ethical guidelines relevant to the selected scientific investigation

Scientific evidence

- the distinction between an aim, a hypothesis, a model, a theory and a law
- observations and investigations that are consistent with, or challenge, current scientific models or theories
- the characteristics of primary data
- ways of organising, analysing and evaluating generated primary data collected to identify patterns and relationships, and to identify sources of error
- the use of a logbook to authenticate generated primary data
- the limitations of investigation methodologies and methods, and of data generation and/or analysis

Science communication

- the conventions of scientific report writing, including scientific terminology and representations, standard abbreviations and units of measurement
- ways of presenting key findings and implications of the selected scientific investigation.

KEY SCIENCE SKILLS

- Develop aims and questions, formulate hypotheses and make predictions
- Plan and conduct investigations
- Comply with safety and ethical guidelines
- Generate, collate and record data
- Analyse and evaluate data and investigation methods
- Construct evidence-based arguments and draw conclusions
- Analyse, evaluate and communicate scientific ideas

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This topic is available online at **www.jacplus.com.au**.

17.1 Overview

Hey students! Bring these pages to life online

Watch videos

Engage with interactivities

Answer questions and check results

Find all this and MORE in jacPLUS

17.1.1 Introduction

How do chemists investigate? The ability to conduct scientific investigations to answer questions, address hypotheses, generate primary data and communicate findings is fundamental to all aspects of science.

In this topic, you will find guidelines for carrying out practical tasks safely and ethically, and for preparing thorough reports and scientific posters using appropriate scientific conventions. You will learn how to develop a question and hypothesis, plan and conduct a methodology, and collate, analyse and evaluate primary data and scientific ideas using appropriate chemical terminology.

FIGURE 17.1 Students conducting a scientific investigation in the classroom

You will develop all of these skills during your Unit 2 AOS 3 practical investigation, where you will design and conduct a practical investigation related to the production of gases, acid–base or redox reactions, or the analysis of substances in water. As it is a student-designed investigation, you will have the opportunity to solve and explore chemical questions of your choice and then communicate your findings and the conclusions drawn, making use of both key knowledge and key science skills.

LEARNING SEQUENCE

online only

17.1 Overview
17.2 Key science skills and concepts in chemistry
17.3 Characteristics of scientific methodology and primary data generation
17.4 Health, safety and ethical guidelines
17.5 Quality of data and measurements
17.6 Ways of organising, analysing and evaluating primary data
17.7 Challenging scientific models and theories
17.8 The limitations of investigation methodology and conclusions
17.9 Options related to analysing substances in water, acid–base or redox reactions, and production of gases
17.10 Presenting findings using scientific conventions
17.11 Review

on Resources

Solutions	Solutions — Topic 17 (sol-0847)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 17 (doc-37915) Key idea summary — Topic 17 (doc-37916)
Exam question booklet	Exam question booklet — Topic 17 (eqb-0126)

17.2 Key science skills and concepts in chemistry

KEY KNOWLEDGE

- Chemical science concepts specific to the selected scientific investigation and their significance, including the definition of key terms
- The distinction between an aim, a hypothesis, a model, a theory and a law
- The use of a logbook to authenticate generated primary data
- The conventions of scientific report writing, including scientific terminology and representations, standard abbreviations and units of measurement

KEY SCIENCE SKILLS

- Develop aims and questions, formulate hypotheses and make predictions
- Plan and conduct investigations
- Comply with safety and ethical guidelines
- Generate, collate and record data

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17.2.1 Why do we conduct investigations?

Conducting investigations and scientific inquiry in chemistry is fundamental to our understanding of the world around us. For the field of chemistry to progress, whether through building on existing theories or testing new theories, research is fundamental. Through its findings, our knowledge, understanding and practical applications are increased. For example, only recently a new plastic, polyhydroxybutyrate (PHB), was discovered. This is a natural, bacterially produced biodegradable natural polymer — meaning, unlike traditional plastics, it is not produced from crude oil. Everything you touch and see has been affected at some point by research investigations. Pesticides and insecticides have been developed to help grow the vegetables and fruit you eat, desalination plants have allowed us to drink safe water, and polymers have been developed for specific purposes — including Teflon non-stick pans, kevlar bulletproof jackets and clingfilm wrap for your sandwiches, to mention but a few.

Research is a continual and fluid process. At each stage of new evidence, we push the boundaries of our knowledge and understanding by refining our existing theories or realising that our existing theories do not hold, based on that new evidence. For example, the theory of the electron was refined with the introduction of quantum mechanics; prior to this, electrons were considered to be minute balls.

Why is it important to be able to conduct investigations? How are changes in our understanding of aspects of chemistry achieved? What might cause the rejection of an earlier theory? How might refinements of an existing theory come about?

FIGURE 17.2 Chemists conducting experiments, making observations and recording data

Carefully planned investigations is one process by which chemical knowledge is advanced. These investigations may be either:

- experimental studies that are carried out
  - in a laboratory
  - in the field
- observational studies.

Both experimental and observational studies generate data that can be analysed, from which conclusions can be drawn.

Experimental studies

In an experiment, an investigator deliberately exposes a substance or system of interest to a chosen factor, and observes the effect of that change. For example, chemists might investigate the effect of temperature on the yield for a reaction producing biofuel. To do this, the reaction would proceed at a set temperature and this would berepeated for at least five temperatures, with the amount of biofuel produced measured and recorded along with observations. The data would then be analysed to address the proposed theories.

FIGURE 17.3 Observational studies differ from experimental studies.

Observational studies

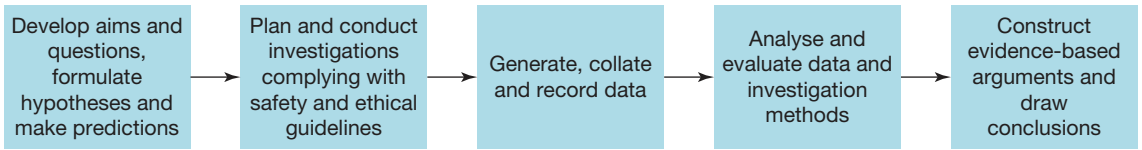
In an observational study, an investigator collects data about the object of study but does not change the existing conditions. For example, a scientist might gather samples of water from several locations downriver from an industrial site. These samples would be taken to the laboratory for analysis, to determine the concentrations of heavy metals in the river water over some distance.

17.2.2 The scientific method

In Unit 2, Outcome 3 you can use your developing experimental skills to investigate a question related to the production of gases, acid–base or redox reactions, or the analysis of substances in water. You will be conducting a practical investigation that uses laboratory or fieldwork to respond to a question. A practical investigation involves considerable planning and expertise in working scientifically to design a method, record results, analyse findings and communicate conclusions.

The diagram shown in figure 17.4 summarises this process of practical investigations and scientific method.

FIGURE 17.4 The scientific method



These key science skills are examinable. While questions about your specific investigation will not be asked, general questions about scientific investigations and inquiry are likely to be asked. You should ensure you are prepared for this, not just for Units 1 and 2, but also for Units 3 and 4.

This investigation will draw upon a number of key skills that you have been developing in Unit 2, Areas of Study 1 and 2, and the specific skills in Unit 2, Area of Study 3. As part of this investigation, you will be required to produce a report. This is usually in the form of a scientific poster, but your teacher may ask for another format such as a practical report, oral presentation or digital presentation. You will also be required to produce a logbook.

In this investigation, you will have the opportunity to show your skill and imagination in experimental design, commitment to a task and your communication ability to explain your **results**. You will need to develop a question, plan actions to answer this question, undertake an investigation and interpret the data to form a conclusion.

This Unit 2 Outcome 3 task requires a significant amount of class time, including time to both conduct investigation and communicate findings. Your teacher will likely set aside multiple weeks for the activity, so some planning and organisation on your part will be needed to achieve a personally satisfying outcome. Table 17.1 will assist with your planning. The timeline may be different depending on your school and when they conduct the investigation. Your teacher may also have set checkpoints regarding when you are required to submit work and what specific components need to be included.

results section in a scientific report in which all data obtained is recorded, usually in the form of tables and graphs

TABLE 17.1 Sample schedule for your investigation

Task	Due date
• Introduction of task • Time to formulate your topic and question for the investigation	Lesson 1, Week 1
• List of possible topics and questions submitted to teacher for approval	Week 1

• Complete a full methodology for your investigation (in your logbook) including:
  • a list of required materials and equipment
  • a clear, repeatable method
  • the completion of any risk assessments and understanding of ethical guidelines.
Your teacher may decide to make this a formal task, done under test conditions in class and assessed, but with feedback provided afterwards on aspects that might need to be addressed before you begin.
	End of Week 1

• Your requested equipment is assembled by the teacher and lab technician. You may be required to make modifications based on resource availability.	By the end of the week before your experiment begins

• Your experiment begins:
  • First period: Set up your equipment, take some preliminary data, fine-tune your procedure, and troubleshoot any difficulties with the equipment and the taking of measurements.
  • Second and third period: Begin the cycle of measurements and data analysis. Progressively graph your results, evaluate trends and adjust your procedure.
	Week 2

• Communication of findings:
  • Analysis and evaluation of data
  • Construction of evidence-based arguments and discussion of your results
  • Critical evaluation of your methodology and data
  • Conclusions drawn from your investigation
  • Submission of your logbook and report in format designated by your teacher
	Week 3

Note: This schedule is an example only and may be altered to suit your class and your school's timetable.

17.2.3 Using a logbook

As part of your scientific investigation (as well as all practical experiments throughout the year), you are required to keep a **logbook**. This, alongside your report, is assessed for Unit 2, Area of Study 3. Your teacher will check this regularly for authentication and assessment purposes.

- The use of a logbook is standard scientific practice.
- A logbook is used to record background information, plan the design of the investigation (including management of risks), record the data and the preliminary analysis of results.

Usually this logbook is a bound exercise book; however, your teacher may request a digital logbook instead. It is vital to show all aspects of your practical investigation within your logbook using the scientific approach.

This logbook will be assessed by your teacher. You must date all work in your logbook to show when it was completed and assist in validating your work. It should be filled out as you progress through your investigation, not after all observations are made.

Your logbook should be written in non-erasable pen, with any mistakes crossed out (do not use whiteout). It should be in a bound book (or appropriate digital format) with numbered pages and dates.

FIGURE 17.5 All observations should be recorded in a logbook.

The main components of your logbook are listed in table 17.2. Further information on these components are found in later subtopics.

logbook a record containing all the details of progress through the steps of a scientific investigation

TABLE 17.2 Components of a logbook

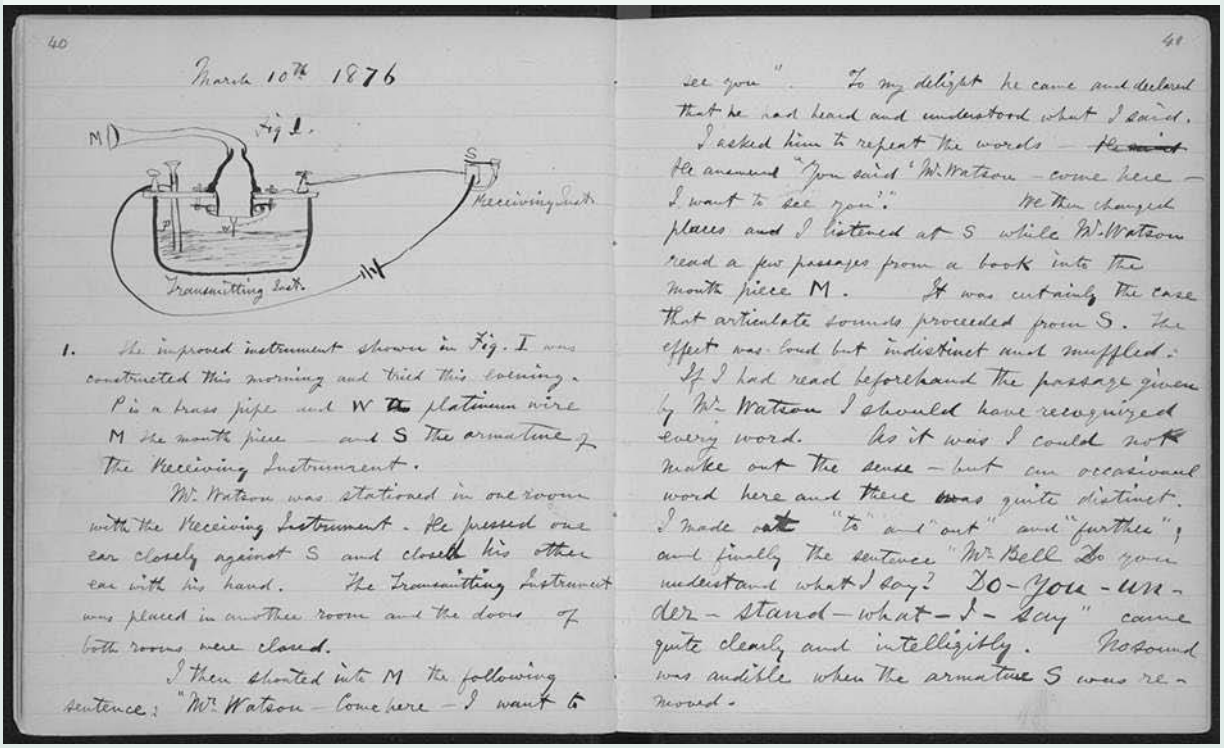
Component	Features
Chosen question (as the title)	Information about your topic, how you chose it and the question you have selected.
Introductory material	This will include background data on your topic, diagrams, notes and tables, information about key terms and similar past experiments. Be sure to record not only the data items themselves, but also their sources, so that you can easily locate and revisit them, and appropriately reference them.
Hypothesis and aim	A clear hypothesis and aim should be recorded, and any variables should be identified.
Methodology	Show all equipment you plan to use and a clear method you plan to follow, with detailed steps that could be reproducible by someone else. This should include any health, safety and ethical guidelines.
Results	Observe and record results in an appropriate form. Tables are particularly useful. You may also include diagrams and photos. You should ensure every result has an associated date, so when it was collected is clear.
Discussing and analysing results	Refer to your results and carefully evaluate them, referring back to your hypothesis and questions. You may have set discussion questions to answer to help scaffold your thoughts and ideas. This will form a basis for your final communication.

CASE STUDY: The importance of logbooks

Figure 17.6 shows pages from the logbook of Alexander Graham Bell, the inventor of the telephone.

Clearly, many features are included that are vital for a logbook, such as dates of findings (in this case, these pages were from 10 March 1876), a clear outline of methodology, design of the equipment used and a summary of findings. Despite this logbook being over 100 years old, it clearly shows the scientific method being used.

FIGURE 17.6 Extract of logbook from Alexander Graham Bell, written while he was inventing the telephone

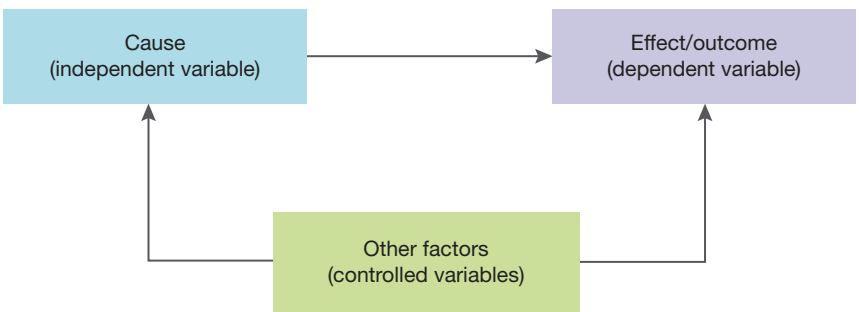


17.2.4 Variables

Independent, dependent and controlled variables

In an experiment, a variable is any factor that the researcher can control or change or measure. Three kinds of variables are commonly recognised (see figure 17.7). For some variables, you will set the value at the start of each experiment; others will be determined by your experiment; and sometimes there may be variables that you calculate using your measurements.

FIGURE 17.7 The relationships between variables in an experiment



Types of variables

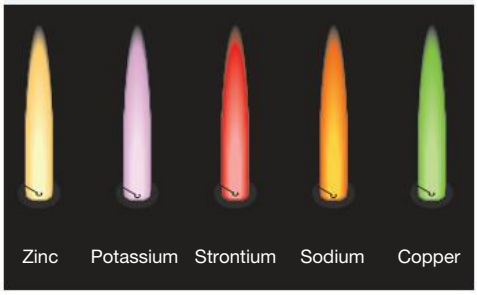
- An **independent variable** is a factor that is *deliberately manipulated by the investigator* and affects the dependent variable. For example, you may be examining the temperature of six different water samples. The independent variable is the different water samples. This may include specific differences in the water samples, such as the source or type of the sample (for example, comparing carbonated water to still water). When graphing results, the independent variable is always placed on the horizontal axis.
- A **dependent variable** is the factor that *the investigator measures*. The dependent variable is affected by the independent variable. In the investigation mentioned in the preceding point, the dependent variable would be the temperature. The dependent variable is always placed in the vertical axis of a graph.
- **Controlled variables** are all the other factors that the investigator must *maintain at constant values through the course of an experiment*. If these factors are not kept constant, they can confound the experimental results because they can cause changes in the dependent variable. Continuing the same investigation example, controlled variables would include the volume of water tested and the instrument used to record data. Often, controlled variables also include environmental factors such as humidity and air temperature, but these are harder to control.

Variables example

Say Allira and Hunter are investigating the colour a flame turns when exposed to different metal ions. In this investigation, the variables are as follows:

- *Independent variable:* The factor that is being manipulated is the *type of metal ion*.
- *Dependent variable:* The factor that is being measured is the *colour of the flame*.
- *Controlled variables:* The factors to control are ensuring the same amount of metal ion, the same Bunsen burner is used, and the temperature of the flame is the same (using the roaring flame for all tests).

FIGURE 17.8 Different metal ions being investigated



Variables can also be considered as numerical (quantitative) or categorical (qualitative). Quantitative includes any value that is numerical; for example, temperature, pH or mass. Qualitative involves variables that are not numerical; for example, colour, the source of the water or the type of salt being examined. See section 17.5.2 for further detail on this.

on Resources

Video eLesson SkillBuilder — Controlled, dependent and independent variables (eles-4156)

Interactivities SkillBuilder — Controlled, dependent and independent variables (int-8090)
Variables (int-7731)

independent variable variable that is changed or manipulated by an investigator

dependent variable variable that is influenced by the independent variable; the variable that is measured

controlled variable variable that is kept constant across different experimental groups

investigation question the focus of a scientific investigation in which experiments act to provide an answer

aim a statement outlining the purpose of an investigation, linking the dependent and independent variables

17.2.5 Developing questions and aims

In Unit 1, you were permitted to research using questions provided; however, in the Unit 2 research investigation you need to come up with a topic, create an **investigation question** that is the focus of your scientific inquiry, and develop an experimental **aim**.

Both your aim and question should show a clear link between the independent and dependent variables being examined.

Coming up with a topic

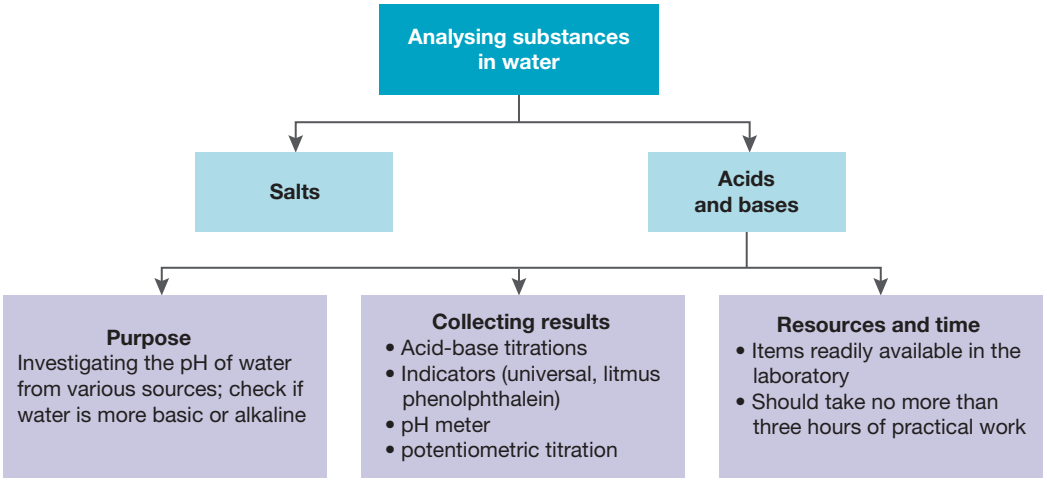
Choosing a topic is not an immediate process — it takes time and careful consideration. It is important you don't just pick a topic that sounds interesting, but instead pick one that is reasonable to complete in the provided time frame and using the available resources.

The topic of your investigation needs to be related to the production of gases, acid–base or redox reactions, or the analysis of substances in water, so review the work you completed in Unit 2 Areas of Study 1 and 2 to help brainstorm ideas.

FIGURE 17.9 Water can vary in pH

You may wish to create a mind map or a diagram outlining the different aspects of water quality and the different ways that you measure each of these. Research the time each measurement would take and what other research you could conduct. This will help you get your head around the different topics and likely requirements.

FIGURE 17.10 Example of the beginnings of a flow chart that may help with selecting a topic



Some examples of topics may be:

- Salts found in water, including minerals, heavy metals and organometallic substances
- Sources of acids and bases found in water, such as dissolved CO_2

Creating an investigation question

Turning the topic into a question focuses your mind on what you want to find out.

Requirements for an investigation question

- Can be investigated through scientific method
- Practicable, given your knowledge, time and school resources
- Asked in a way that indicates what you will do.

Once you have determined a topic, you need to create an investigation question that allows you to answer and solve a specific inquiry question. This needs to link a specific independent variable with a specific dependent variable.

As an example of formulating a question from a topic, say your chosen topic is 'Exploring the amount of sodium in water'. A question you could form from this could be 'Do samples of water from different taps contain varying levels of sodium?' Or, 'How can electrical conductivity be used to determine how sodium concentration differs between various sources of water?' The first question revolves around examining the variable of sodium concentration in different taps, but doesn't focus specifically on how it is measured. The second question focuses more on how the sodium concentration is measured, but is less specific with the water source.

You can formulate a question from a topic in many different ways. Just make sure it is something that can be measured, explored and answered in the scope of your practical investigation.

Developing an aim

Often, developing an aim of an investigation is done at the same time as formulating a question from your topic. The aim outlines the purpose or the key objective of the investigation. It outlines what you are trying to discover from your investigation.

It is important that your aim:

- links your independent and dependent variables
- is succinct (no more than two lines)
- links clearly to your investigation question.

Two different ways to format your aim

1. To [determine/investigate/compare] how the **dependent variable** is affected by the **independent variable**
2. To [determine/investigate/compare] how the **independent variable** affects the **dependent variable**

Examples of aims include:

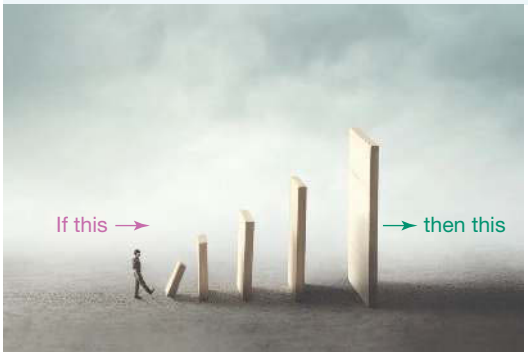
- To determine the **solubility** of **copper chloride** in 100 g of water
- To investigate the **pH** of five **different brands of soft drink** using universal indicator
- To compare the **copper concentration** between **water from indoor taps and the water from outdoor taps**
- To observe how **changing the temperature** of water affects the **solubility of copper sulfate**
- To determine if **blowing bubbles** into a bottle of water alters its **pH**

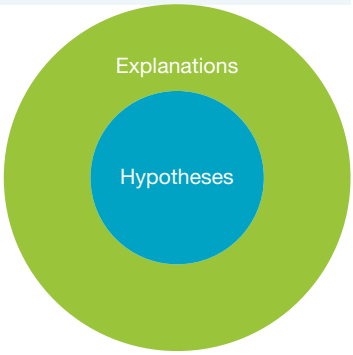
In all the provided example aims, a clear link exists between the independent variable (shown in plum) and the dependent variable (shown in green).

17.2.6 Formulating hypotheses and making predictions

Formulating a **hypothesis** is an important step in the scientific method.

FIGURE 17.11 A hypothesis is a testable explanation for a concept.





hypothesis a tentative, testable and falsifiable statement for an observed phenomenon that acts as a prediction for the investigation

A hypothesis is a tentative, testable and falsifiable statement for an observed phenomenon, which predicts the relationship between two variables or predicts the outcome of an investigation.

A hypothesis usually predicts the relationship between independent and dependent variables, providing a tentative, testable and falsifiable prediction of what will be the findings of the investigation outlined in the aim and hypothesis.

- **Tentative** means that a hypothesis is not certain, but an attempt to explain a phenomenon based on theory.
- **Testable** means that a hypothesis can easily be tested by observations and/or investigations.
- **Falsifiable** means that a way exists to invalidate a hypothesis; that is, to prove a hypothesis wrong.

A hypothesis can be written in many acceptable ways.

tentative not fixed or certain; may be changed with new information
testable able to be supported or proven false through the use of observations and investigation
falsifiable able to be proven false using evidence

To write a hypothesis, a good tip is to use the following format:

IF [statement involving the **change in independent variable**], THEN [**prediction involving the dependent variable**] DUE TO [**tentative explanation** for the predicted effect on the dependant variable].

Your prediction often includes reference to whether the variable will increase or decrease.

Table 17.3 shows some examples of hypotheses using the if, then, due to format.

TABLE 17.3 Examples of good working hypotheses

IF	THEN	DUE TO
IF a straw is used to **blow bubbles of carbon dioxide** into a glass of water	THEN the **pH will increase**	DUE TO the **production of carbonic acid.**
IF the **water quality from an old household tap** is examined	THEN it will have **higher levels of copper**	DUE TO **contamination,** compared to water from newer pipes.
IF the **concentration of HCl is increased** in a reaction with magnesium	THEN **the volume of hydrogen gas will be produced at a faster rate**	DUE to a **greater availability of hydronium ions.**

FIGURE 17.12 What happens if bubbles are blown into a glass of water?

FIGURE 17.13 Will water from older pipes contain more copper?

FIGURE 17.14 Will the concentration of the acid affect how gas is produced?

Now consider the following statements and decide if each is an example of a well-formatted hypothesis.

- Statement 1: 'Boiling water will increase its pH.'
  No. This is simply a testable prediction. It does not include a tentative explanation.
- Statement 2: 'IF **water is boiled**, THEN the recorded **pH will increase**.'
  No. This does not identify a tentative explanation. The statement is a method followed by a predicted outcome.
- Statement 3: 'IF **water with a pH below 7 is boiled** and **dissolved gases are removed**, THEN the **pH will increase and become closer to neutral**.'
  Yes. This identifies a tentative hypothesis (explanation) and a predicted outcome by which the hypothesis can be tested.

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SAMPLE PROBLEM 1 Writing an aim, hypothesis and research question

Polly is putting the kettle on. Polly is very curious about science and wants to see how she can change the speed the water boils. She has heard rumours that salt causes water to boil faster. She has four different types of salts in her house: table salt, sea salt, Himalayan pink salt and chicken salt. Write an appropriate research question, aim and hypothesis for this scenario.

THINK	WRITE
1. Determine the variables to help write an aim, hypothesis and research question. The factor that Polly is manipulating (the independent variable) is the type of salt. The factor that Polly is measuring (the dependent variable) is the time it takes for the water to boil.	The independent variable is the salt type. The dependent variable is the time it takes for the water to boil.
2. Create a research question based on Polly's problem. Make sure that the question is one that is testable and clearly outlines what is occurring in the investigation.	Does the type of salt added to water affect the time it takes for water to boil?
3. Write an aim that clearly outlines the purpose of the investigation. Be sure to link the independent variable and the dependent variable.	To determine if different types of salts affect the time it takes for water to boil.
4. Write a hypothesis in the 'IF . . . THEN' format. Remember that a hypothesis needs to link the independent variable and dependent variable. Your hypothesis may not be correct, but it must be testable. In this experiment, you may also specify which salt you think will affect the dependent variable the most.	If table salt, sea salt, Himalayan salt or chicken salt is added to water, then the time taken for the water to boil will decrease, with pure table salt causing the largest decrease in time.

PRACTICE PROBLEM 1

Jack and Jill want to know if changing the material on an incline will affect the speed at which a ball rolls down the incline.

Write an appropriate research question, aim and hypothesis for this scenario.

on Resources

Video eLesson SkillBuilder — Writing an aim and forming a hypothesis (eles-4155)

Interactivities SkillBuilder — Writing an aim and forming a hypothesis (int-8089)
Formatting a hypothesis (int-7732)

17.2.7 Concepts specific to investigations in chemistry

As part of an investigation, including key chemistry concepts that are relevant and clearly explaining their significance is vital. This may include:

- key background knowledge
- key terms
- techniques used in an investigation
- chemical representations
- scientific notation.

This shows a clear link to your understanding of an investigation, and allows others to see the connection between theory and practical applications.

Concepts should be researched prior to commencing your investigation, recorded in your logbook (and referenced). This background information also will form part of your introduction in your report.

Key background knowledge

Concepts that are relevant to your investigation include:

- explanations of key formulae
- detail about the theories being examined
- information about other practical investigations exploring similar concepts.

An example of this may be in calculating the number of molecules in a set amount of water. In your background information, it would be important to describe previous findings and experiments conducted, identify use and application of the investigation, and describe key formulae relevant to it (for example, $N = n \times N_A$), including the identification of the symbols used.

You should also have clear concepts links to theory in your **discussion** section of the investigation.

discussion detailed area of a scientific report in which results are discussed, analysed and evaluated, relationships to concepts are made, errors, limitations and uncertainties are assessed and suggestions for future improvements are outlined

Key terms

In practical investigations, defining any key terminology is vital.

This can be done in two ways:

- within a report itself
- as part of an appendix or glossary at the end of the report.

CASE STUDY: Key terms within a report

The following excerpt shows an introduction from a scientific report written by a student. This investigation was conducted to explore different types of polymers.

> *Polymers are large molecules made by joining smaller molecules (monomers) together. They form a wide range of substances, both natural and synthetic. Polymers are formed by the addition polymerisation of alkenes, which are hydrocarbons that contain one double bond between carbon atoms.*

This student has clearly defined key terms as part of their introduction within their report itself. What terms have they defined?

CASE STUDY: Key terms as part of an appendix or glossary

Read the following excerpt of an introduction from a scientific report written by a different student, investigating the same practical.

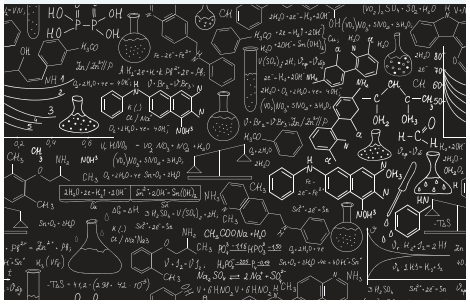
> *In this investigation, different types of **polymers** are being investigated. Polymers that are formed through the addition polymerisation of **alkenes** are specifically being explored in this investigation.*
>
> *Glossary of key terms:*
>
> ***alkenes:** hydrocarbons that contain one double bond between carbon atoms*
> ***polymers:** large molecules made by joining smaller molecules (monomers) together*

In this situation, the student has not defined the terms in their introduction itself, but has bolded key words that later appeared in their glossary.

Chemical representations

A variety of representations are used in chemistry. This includes the use of models, sketches, graphs, equations, formulae, symbols and diagrams. As well as this, many vital conventions exist regarding the use of numerical data, including significant figures and scientific notation. Perhaps the most common chemical representation is the use of chemical formulas. Care should be taken with capital letters and subscripts and superscripts when representing atoms and ions. For example, CO is carbon monoxide, while Co is the metal cobalt.

FIGURE 17.15 Representations form a vital part of chemistry reporting.



Formulas that demonstrate structure

Structures are often drawn in a skeletal form in chemistry, particularly in organic chemistry. One compound can be represented in various ways, as seen in figure 17.16. Some common structural formulas are shown in figure 17.17.

FIGURE 17.16 Different representations of methane: (a) ball-and-stick model, (b) diagram showing bond angles, (c) structural formula, (d) shape diagram

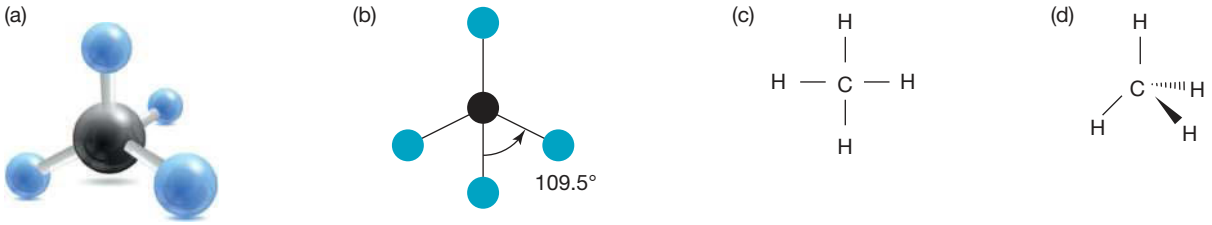
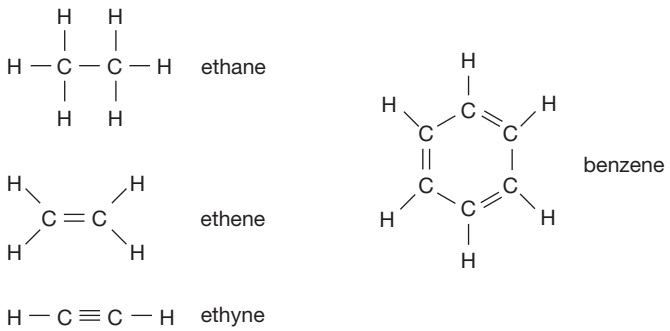
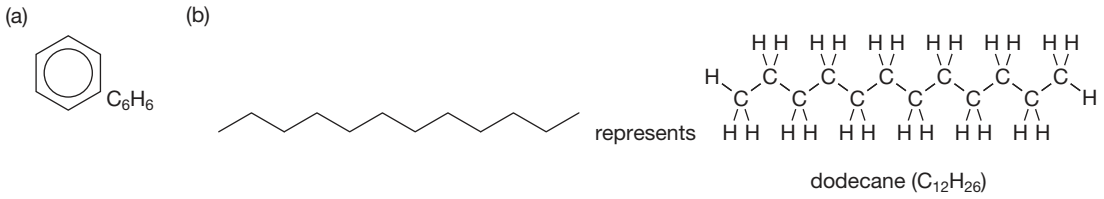


FIGURE 17.17 Some common structural formulas representing various molecules



Lines between atoms represent the number of bonds present. Sometimes we can shorten these representations even further, as shown in the skeletal formulas for the benzene ring and dodecane in figure 17.18. Skeletal formulas do not show specific carbon and hydrogen atoms or the bonds connecting them.

FIGURE 17.18 Skeletal formulas of different molecules: (a) benzene and (b) dodecane



Scientific notation

Very large and very small quantities can be more conveniently expressed in scientific notation. In scientific notation, a quantity is expressed as a number between 1 and 10 multiplied by a power of 10.

To write in scientific notation, follow the form $N \times 10^a$, where N is a number between 1 and 10 and a is an integer (positive or negative).

Steps to convert a number into scientific notation

1. Determine where the decimal point needs to go so that N is between 1 and 10.
2. Count the number of places the decimal point is moved to determine a (the power of 10 or the exponent). If the decimal point was moved to the left, a will be positive; if it was moved to the right, a will be negative.
3. Write the number in scientific notation.

For example, the average distance between the Earth and the moon is 380 000 000 m. This is more conveniently expressed as 3.8×10^8 m, in which the decimal point was moved eight places to the left. The radius of a lead atom in metres is 0.000 000 000 175. This is more conveniently expressed as 1.75×10^{-10} m, in which the decimal point was moved 10 places to the right.

As you can see, very large numbers will have a positive exponent (a), whereas very small numbers will have a negative exponent. For example, 5×10^{-3} can be written out as 0.005, whereas 5×10^3 is written out as 5000.

In chemistry, scientific notation is generally used for numbers less than 0.01 and greater than 1000.

Quantities in scientific notation can be entered into your calculator using the EXP button or ^ button.

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SAMPLE PROBLEM 2 Using scientific notation

a. The average distance between Earth and the sun is 149 600 000 kilometres. Write this in scientific notation.

b. The mass of a proton is 0.000 000 000 000 000 000 000 001 67 g. Write this in scientific notation.

THINK	WRITE
a. 1. Determine the position of the decimal point for the number to be between 1 and 10 and remove any zeros that are not between non-zero digits. The decimal point would need to go between 1 and 4 to form 1.496.	1.496

2. Determine the exponent (the power of 10) by counting the number of places the decimal point was moved.
If the decimal point was moved to the left, the exponent will be positive; if the decimal point was moved to the right, the exponent will be negative.
The decimal point was moved eight places to the left so the exponent is 8. 10^8

3. Write the number in scientific notation, remembering to include the units. 1.496×10^8 km

b. 1. Determine the position of the decimal point in order for the number to be between 1 and 10 and remove any zeros that are not between non-zero digits.
The decimal point would need to go between 1 and 6 to form 1.67. 1.67

2. Determine the exponent by counting the number of places the decimal point was moved.
If the decimal point was moved to the left, the exponent will be positive; if the decimal point was moved to the right, the exponent will be negative.
The decimal point was moved 24 spots to the right so the exponent is –24. 10^{-24}

3. Write the number in scientific notation, remembering to include the units. 1.67×10^{-24} g

PRACTICE PROBLEM 2

Express the following quantities in scientific notation.

a. The diameter of Saturn's rings, 282 000 km

b. The number of metres that sound travels in one hour, 1 235 000

c. The uncertainty of a highly precise clock, 0.000 000 000 000 000 003 seconds

17.2 Activities

learnon

Students, these questions are even better in jacPLUS

Receive immediate feedback and access sample responses

Access additional questions

Track your results and progress

Find all this and MORE in jacPLUS

17.2 Quick quiz on | 17.2 Exercise | 17.2 Exam questions

17.2 Exercise

1. Using the following topics, create a testable question that could be used for a practical investigation.
   a. Examining how the pH of water differs at varying temperatures
   b. Exploring if the amount of salt in water affects its boiling point
   c. Calculating the solubility of different compounds in water

2. What is the purpose of a logbook in practical investigations?
3. Describe the difference between a dependent and an independent variable.
4. Why is it important to control variables in an investigation?
5. A student conducted an experiment to measure the effect of changing the amount (measured in moles) on the volume of a sample of gas. This was done in three stages using different amounts of the gas at various temperatures.
   Stage 1: 0.01 mol of gas at 10 °C
   Stage 2: 0.02 mol of gas at 20 °C
   Stage 3: 0.03 mol of gas at 30 °C
   The results from each trial were then analysed to produce an overall conclusion.
   a. State the independent and the dependent variables in this experiment.
   b. What were the controlled variables in this experiment?
   c. In regards to controlling variables, why would the results of this test be difficult to interpret?
6. The following table outlines an investigation topic with some variables identified. Complete the table for the three other topics listed.

Topic	Independent variables		Dependent variables
	Numerical	**Categorical**	
Bouncing basketball	• Drop height • Pressure of the ball	Surface ball lands on, ball type	Rebound height, impact time, energy loss, change in momentum, average force of impact
Conductivity of metal			
Concentration of salt in water			
Boiling point of different soft drinks			

7. After some preliminary reading, a student has become intrigued by the possibility that hydrocarbons with double bonds (alkenes) have lower boiling points than those with single bonds (alkanes). Therefore, she proposes the question: Do alkanes and alkenes have different boiling points?
   Write a reasonable hypothesis that she could test experimentally based on this question.
8. Which of the following is a characteristic of a good hypothesis?
   A. It must be proven true.
   B. It must be testable by observation or experiment.
   C. It must be based upon experiments done by other scientists.
9. Explain two ways in which key terms can be defined in your report.
10. Express the following quantities in scientific notation.
    a. A red blood cell, about 0.000 008 m across
    b. A flea, about 0.0013 m long
    c. The moon, 384 400 000 metres from the Earth

17.2 Exam questions

Question 1 (1 mark)

MC How do scientific processes begin?

A. By performing experiments
B. By making observations
C. By constructing hypotheses
D. By designing experiments

Question 2 (1 mark)

MC When you test a hypothesis and the data that you obtain does not support the hypothesis, what should you do?

A. Change the data to suit your hypothesis
B. Repeat the experiment until you get data to support the hypothesis
C. Decide you have made a mistake when carrying out the experiment
D. Repeat the experiment and if the same data is obtained decide that the hypothesis is not supported

Question 3 (1 mark)

MC Consider the following data obtained in an experiment. In the experiment, the amount of starch present in two test tubes was measured at different times and temperatures. Both test tubes started with the same amount of starch and each contained the same volume and concentration of the enzyme amylase.

Amylase catalyses the breakdown of starch to maltose.

Time (minutes)	Temperature 25 °C	Temperature 37 °C
0	100%	100%
10	95%	70%
15	90%	50%
20	85%	30%

What is the hypothesis being tested in this experiment?

A. At a temperature of 25 °C, starch is not broken down to maltose.
B. At a temperature of 37 °C, amylase is not needed to break down starch to maltose.
C. Amylase will break down starch to maltose.
D. The temperature of the environment will affect the activity of the enzyme amylase.

Question 4 (6 marks)

A student decided to investigate the effect of temperature on the rate of the following reaction.

$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

Part of the student's experimental report is provided.

Effect of temperature on the rate of production of carbon dioxide gas

Aim

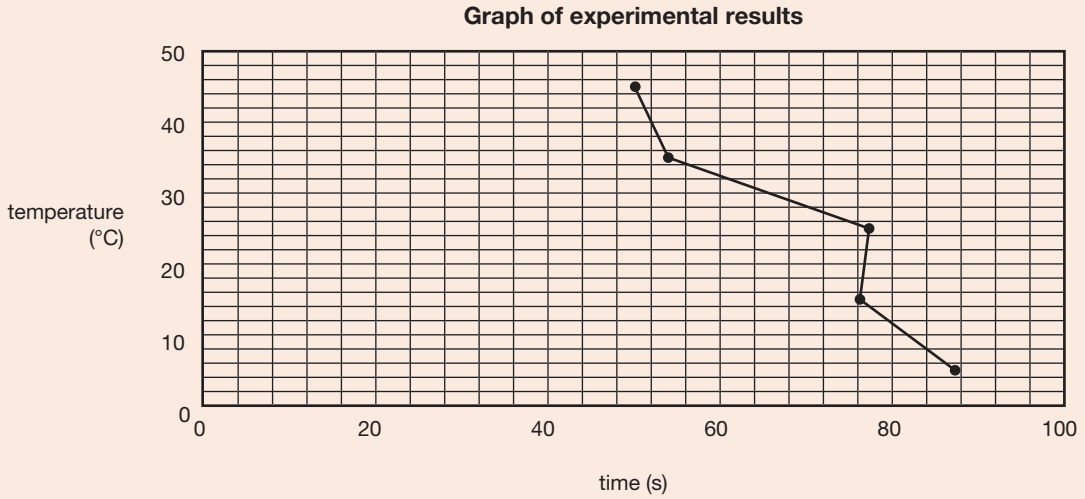
To find out how temperature affects the rate of production of carbon dioxide gas, CO_2, when a solution of hydrochloric acid, HCl, is added to chips of calcium carbonate, $CaCO_3$

Method

1. Put 0.6 g of $CaCO_3$ chips into a conical flask.
2. Put a reagent bottle containing 2 M HCl into a water bath at 5 °C.
3. When the temperature of the HCl solution has stabilised at 5 °C, use a pipette to put 10.0 mL of the HCl solution into the conical flask containing the $CaCO_3$ chips.
4. Put a balloon over the conical flask and begin timing.
5. When the top of the balloon has inflated so that it is 10 cm over the conical flask, stop timing and record the time.
6. Repeat steps 1 to 5 using temperatures of 15 °C, 25 °C, 35 °C and 45 °C.

Results

The following graph gives the experimental results.



a. What is the independent variable? **(1 mark)**

b. What is the dependent variable and how is it measured? **(2 marks)**

c. Predict the relationship between the independent variable and the dependent variable. Explain your prediction. **(3 marks)**

More exam questions are available in your learnON title.

17.3 Characteristics of scientific methodology and primary data generation

KEY KNOWLEDGE

- Scientific methodology relevant to the selected scientific investigations, selected from the following: classification and identification; controlled experiment; correlational study; fieldwork; modelling; product, process or system development; or simulation
- Techniques of primary qualitative and quantitative data generation relevant to the investigation
- The characteristics of primary data

KEY SCIENCE SKILLS

- Plan and conduct investigations
- Generate, collate and record data

Source: VCE Chemistry Study Design (2023–2027) extracts © VCAA; reproduced by permission.

Carefully following scientific method when conducting practical investigations is important. This helps you ensure that your results are **precise**, **accurate**, **repeatable**, **reproducible** and **valid**. This includes minimising errors and uncertainties in data in order to draw **conclusions** in relation to your question.

Each type of **scientific methodology** has its specific purposes, procedures, advantages and **limitations**. The researcher's choice depends on which method is most appropriate for the specific topic of research interest and hypothesis being tested.

precise how close multiple measurements of the same investigation are to each other; a measure of repeatability or reproducibility

accurate describes an experimental measurement that is close to a true value

repeatable how close the results of successive measurements are to each other in the exact same conditions

reproducible how close the results of measuring the same quantity are to each other in changed conditions (different observer, method, instrument, time etc.)

valid describes an experiment that investigates what it sets out to investigate (via appropriately controlling variables, eliminating bias etc.)

conclusion section at the end of a scientific report that relates back to the question, sums up key findings and states whether the hypothesis was supported or rejected

scientific methodology the type of investigation being conducted to answer a question and resolve a hypothesis

limitations factors that affect the interpretation and/or collection of findings in a practical investigation

Components of scientific inquiry

Most scientific inquiries involve most, if not all, of following:

- formulating a question and hypothesis to be tested
- controlling variables
- using control groups and experimental groups
- completing a logbook outlining the introduction, methodology, results, discussion and conclusion of an investigation
- ensuring that methods are being used that allow for validity, accuracy, precision and reliability
- ensuring that methods are being used that reduce uncertainties and errors
- collecting data accurately in an appropriate form that best suits the question being investigated.

17.3.1 Types of scientific investigation methodologies

Different types of scientific inquiry and research methods that can be used are shown in table 17.4. These are also outlined in the VCE Chemistry Study Design.

TABLE 17.4 Types of scientific investigation methodologies

Methodology	Use
Case study	To investigate a particular event or problem that contains a real or hypothetical situation including the complexities that would be encountered in the real world
Classification and identification	To arrange phenomena, objects or events into manageable sets, and to recognise phenomena or objects as belonging to particular sets or possibly being part of a new or unique set
Experiment	To demonstrate a known fact, test a hypothesis or make a discovery. It may include investigating the relationship between an independent variable and a dependent variable, controlling all other variables, as part of the scientific method
Fieldwork	To generate site-specific data, recorded in the student's logbook, to solve a problem or to investigate an issue at a specific location
Literature review	To collate and analyse secondary data related to other people's scientific findings and/or viewpoints in order to answer a question or provide background information

Modelling	To construct a physical or conceptual model, representing a system involving concepts that help people know, understand or simulate the system; or a mathematical model, describing a system using mathematical equations involving relationships between variables, which can be used to make predictions
Product, process or system development	To design an artefact, process or system to meet a human need
Simulation	To study the behaviour of a real or theoretical system using a model when the variables cannot be easily controlled (e.g., system too complex, too dangerous or too inaccessible)

Many of these methods have very specific uses in various aspects of scientific investigation.

You will find that for Unit 2, Outcome 3 you will probably be conducting a controlled experiment within a classroom setting.

17.3.2 Characteristics of the scientific method

Following a set scientific method in your investigation is very important. What is the difference between 'scientific investigation methodologies' and the 'scientific method'?

Difference between 'scientific investigation methodologies' and the 'scientific method'

Scientific investigation methodology is a technique used to make predictions and produce answers. The scientific method is a particular scientific methodology that shows the steps and the process involved for answering questions.

So, the scientific methodology is the overarching *what* you are going to do — the type of investigation you are going to carry out, such as a case study, a controlled experiment or modelling. The scientific method is *how* you are going to do it — the steps you will follow to conduct your investigation.

The scientific method is a set process that involves many distinct steps that allow you to easily conduct an experiment and communicate your findings in response to the scientific methodology you explored (see figure 17.19).

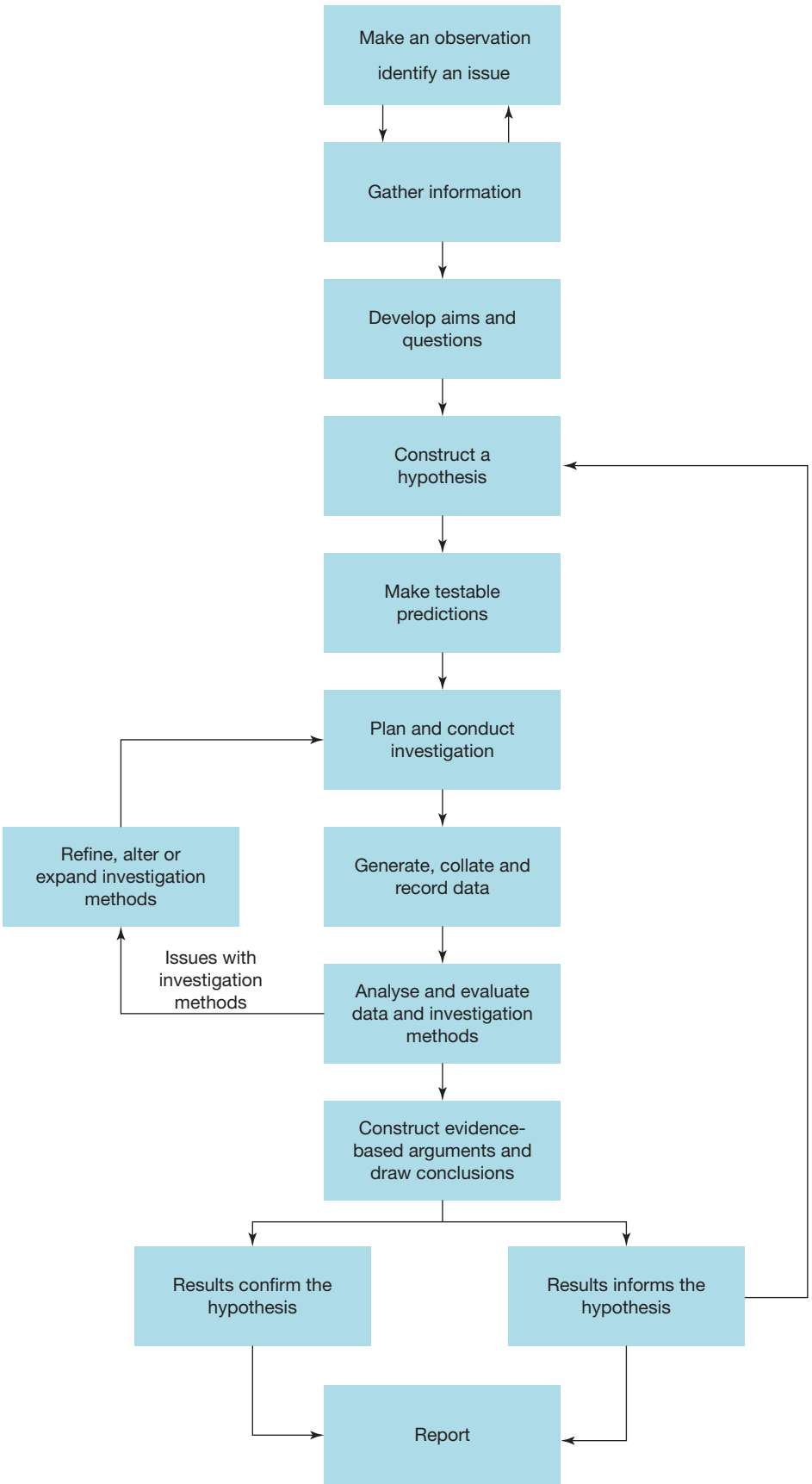
17.3.3 Designing an investigation

Designing an investigation means constructing a detailed experimental plan to test a hypothesis in advance of doing the experiment. In VCE Chemistry, your investigation may be either student-adapted or student-designed.

- *Student-adapted:* Altering an already designed or conducted investigation to suit your investigation question
- *Student-designed:* Creating a novel task from scratch to explore your investigation question

scientific investigation methodology the principles of research based on the scientific method

FIGURE 17.19 The scientific method includes many components.



A typical plan might include details under the following headings, which you would record in your logbook.

Designing an investigation — a typical plan

Title of experiment: Usually your scientific question.

Planning: A section to brainstorm ideas and show your planning.

Aim: Information about the purpose of the experiment.

Hypothesis: Your hypothesis about what you expect will happen in your experiment.

Background information: Some information about KEY KNOWLEDGE being explored in your investigation, and details around these. This may include other investigations or practicals you have researched.

Materials: This includes the following items that should be checked with your teacher:
- the laboratory equipment and the consumables needed for your experiment
- any personal safety equipment required and their availability.

Health, safety and ethical considerations:
- List any safety issues associated with the conduct of your experiment; for instance, handling potentially hazardous substances and/or using potentially hazardous equipment such as some electrical equipment.
- For each safety issue identified, list in your logbook the safety controls and precautions to be taken. This can be done as a risk assessment.
- List any ethical issues associated with the conduct of your experiment; for example, the need to apply appropriate protocols to access information about Aboriginal and Torres Strait Islander peoples' knowledge or techniques.

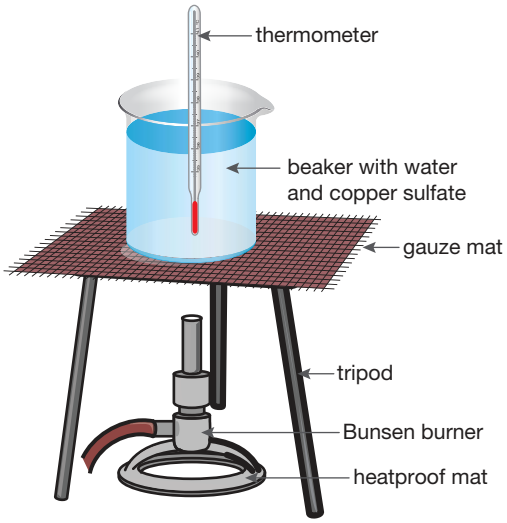
Method:
- Identify your independent variable and how you will change it during the experiment (you should aim to have five variations of your independent variable).
- Identify the dependent variable expected to respond to these changes and identify how you will measure the changes.
- Identify all the controlled variables to be kept constant throughout the experiment.
- Outline a clear step-by-step method.
- Include a diagram of your experimental set up where appropriate.

FIGURE 17.20 Planning a practical investigation

Practical investigation planning	
Name	Jill
Title	Exploring the solubility of copper sulfate in water at various temperatures
Planning	Three starting questions I want to answer: 1. Is there an optimal temperature for copper sulfate solubility in water? 2. Is copper sulfate soluble at all temperatures? 3. What is the solubility of copper sulfate at room temperature?
Aim	To determine if the solubility of copper sulfate in water changes at varying temperatures
Hypothesis	If the solubility of copper sulfate is affected by temperature, then the solubility of copper sulfate will increase when the temperature is increased.

(continued)

FIGURE 17.20 Planning a practical investigation *(continued)*

Practical investigation planning	
Background information	Solubility is the extent to which a solute dissolves in a solvent. All sulfates are soluble in water except those formed with silver, lead, calcium, strontium and barium. In water, copper sulfate dissociates as follows: $CuSO_4(s) \rightarrow Cu^{2+}(aq) + SO_4^{2+}(aq)$ The solubility of copper sulfate is known to be 20.5 g per 100 g of water at 20 °C.
Materials	• Copper sulfate solid • Watch glass • Spatula and stirring rod • Scales • Bunsen burner, heat-proof mat, gauze mat and tripod • Thermometer • 200 mL beaker • Deionised water • Container to dispose of copper sulfate
Health, safety and ethical considerations	Copper sulfate **Hazards** • May be harmful if swallowed • Causes mild eye and skin irritation • Toxic to aquatic life **Safety precautions** • Wear gloves, safety glasses and a lab coat • Do not dispose chemical down sink — dispose of in container • Avoid breathing in large amounts of fumes **First aid** • Swallowed: Seek medical attention; induce vomiting per medical advice if more than 15 minutes • Eye: Wash with running water • Skin: Wash with water and soap • Inhaled: Fresh air; if breathing becomes difficult, give oxygen and seek medical attention
Method	Independent variable: Temperature of the water Dependent variable: Solubility of copper sulfate, mass of copper sulfate able to dissolve Control variables: • Type of solvent (water) • volume of solvent (100mL) 1. Fill the beaker with 100 mL of water. 2. Place copper sulfate on a watch glass and weigh it. 3. Record the temperature of the room. 4. Slowly add copper sulfate into the water, mixing with a stirring rod. 5. Determine when no more copper sulfate can be dissolved and record the mass. 6. Fill another beaker with 100 mL of water and heat to 40 °C. Repeat Steps 4 and 5. 7. Repeat Step 6 at 60 °C, 70 °C, 80 °C and 100 °C. 

Other factors that are important to consider during the design of your investigation include:

- the use of control and experimental groups
- sample size
- how you are going to ensure accuracy, precision and validity
- how you will repeat the investigation to obtain more data and show reproducibility
- how you will control variables.

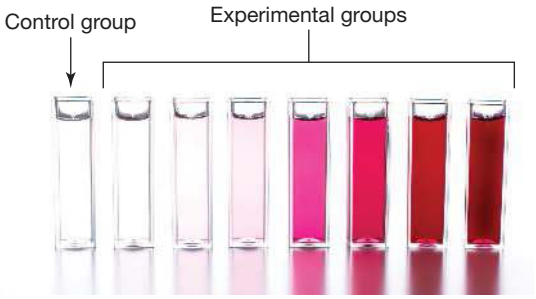
Control and experimental groups

The design of many experiments includes a **control group** as well as one or more **experimental groups**.

- The experimental groups are exposed to the changing conditions determined by the independent variable.
- The control group is not affected by the independent variable.

For example, a scientist may be trying to measure the absorbance of light by various concentrations of cobalt chloride, $CoCl_2$, solutions. The independent variable in such an investigation would be the concentration of cobalt chloride, and the dependent variable would be the absorbance reading obtained. However, a number of other variables may affect the result. These include the nature of the solvent, the type of glass that the containers holding the solutions are made from, the distance the light has to travel (especially through the solution) before it is measured and the temperature of the solution.

FIGURE 17.21 Comparing control and experimental groups



A convenient way to control all these variables, and maybe even some that you aren't aware of, is to use a control. Everything about the control, from the way it is prepared to how it is manipulated and measured in the experiment, is the same as for the test solutions containing cobalt chloride. The only difference is that no cobalt chloride (the independent variable) is in the control. This, therefore, allows the scientist to isolate the amount of absorption in each reading that is due to the cobalt chloride alone because it is the only variable left responsible for any differences in absorbance readings.

Another example of control and experimental groups being used may occur in a practical investigation exploring the conductivity of different concentrations of sodium chloride, NaCl, solution. The experimental groups would be the different concentrations of sodium chloride being tested. The control group would be testing the circuit with pure water (0% sodium chloride), to examine the conductivity when no NaCl is present.

The control group serves several purposes:

- it shows that the experiment is working; that is, the change in the dependent variable is due to the independent variable
- it provides a baseline result against which the results of the experimental group can be compared.

You need to decide on the number and size of the experimental groups, and whether a control group is required. For instance, in an experiment concerned with the effect of temperature on the change of state of a substance, temperature control groups are often forgone because the effect of temperature cannot be removed. It can only be manipulated.

control group a group that is not affected by the independent variable and is used as a baseline for comparison

experimental group test group that is exposed to the independent variable

Sample size

The size of the control and experimental groups is an important factor in experimental design. This is known as the **sample size**. The size of each group must be sufficiently large that:

- replicate results can be obtained, suitable averages can be calculated, and trends can be observed to allow comparisons between the outcomes in the control and the experimental groups
- the results from the experimental group can be seen as applicable to the larger population.

The upper limit on sample size is determined by cost and space considerations and other practicalities.

Complicated formulae exist for calculating minimum sample sizes. However, one simple general rule is that ten observations are required for each experimental variable. This may not be possible in the scope of your investigation but is an important point of discussion.

Selecting appropriate equipment and techniques

When selecting appropriate techniques, it is important to ensure the following:

- the technique can be performed in an appropriate time frame
- the technique is appropriate to your investigation and serves a purpose to answering the question and supporting or rejected your hypothesis
- the data is easily recorded, measured and interpreted, with a particular emphasis on quantitative data
- the technique can be safely performed; this is particularly important in a school environment, where health and safety restrictions are closely regulated
- the equipment used in the technique is available and cost effective; if not available in a school, it can be used with permission at other locations
- the technique allows for the control of other variables; if too many factors cannot be controlled and will affect results, the technique isn't a great choice for an investigation.

Equipment

You should also consider the most appropriate equipment to use for a particular purpose.

For example, if a liquid volume of 25 mL is required, what would be the most appropriate piece of equipment to measure it? If a high-precision measurement is required (that is, providing a lower **uncertainty**), a 25.00 mL pipette would probably be used. If, however, this level of precision is not required, a measuring cylinder or even a 100 mL beaker may be more appropriate.

Your school will have a range of measuring instruments, and these will vary in precision and ease of use.

You won't always need to use the most accurate instrument. A simple instrument that allows for quick measurements will be enough more often than not. Sometimes a simple stopwatch is just as good as an electronic timer, for example, or a voltmeter may compare well to a more accurate multimeter.

FIGURE 17.22 Different types of measuring equipment vary in precision

Some instruments that you might consider are as follows, listed based on what they measure.

Equipment to measure mass

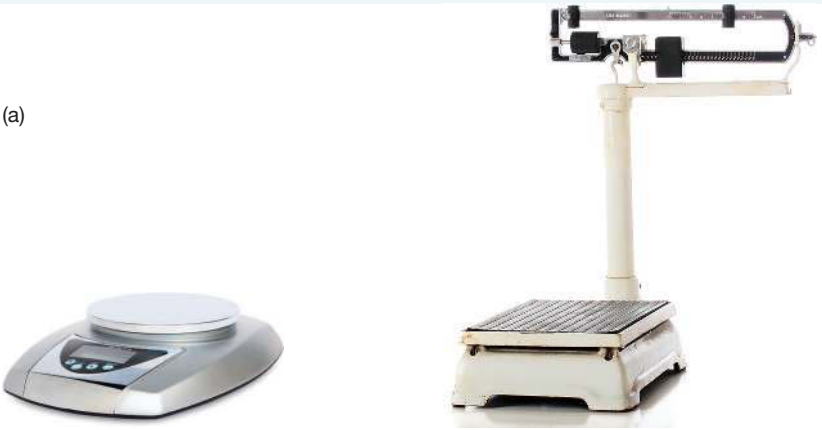
- *Top loading balance:* Very accurate; very good for small masses; simple to use. With equipment set up above the balance, it can be used to measure small variations in attractive and repulsive forces such as magnetic force, electric force and surface tension. If the balance sits on a laboratory jack, force against distance can be easily measured.
- *Beam balance:* Accurate with a large range of values; can be time-consuming to measure several masses.

sample size the number of trials in an investigation

uncertainty a limit to the precision of data obtained; a range within which a measurement lies

FIGURE 17.23 (a) Top loading balance and (b) beam balance

(a)



Equipment to measure volume

- *Micropipette* (figure 17.24a): Much more precise than other equipment, but can only measure set volumes. Easy swapping of tips allows for a reduced chance of contamination. Micropipettes are easy to use, but take a bit more practice then pipettes, beakers and measuring cylinders.
- *Pipette* (figure 17.24b): Can only measure a few particular volumes. A pipette is more precise than beakers and measuring cylinders, but making a mistake can be easy during use, such as through the introduction of air bubbles.
- *Measuring cylinder* (figure 17.24c): Depending on the increments, these are less precise than pipettes and micropipettes, but more precise than beakers. They are very easy to use.
- *Beaker* (figure 17.24d): Depending on the increments on the beaker, these are relatively imprecise and can lead to huge uncertainty in measurements, but they are easy to use.

FIGURE 17.24 Equipment used to measure volumes: (a) micropipette, (b) pipette, (c) measuring cylinder and (d) beaker

(a) (b) (c) (d)

Equipment to measure pH

- *Titrations:* Used to determine the concentration of a solution, particularly through acid–base titrations. Detailed descriptions of titration equipment are provided in topic 14.
- *pH meter* (figure 17.26a): Can be more time-consuming and expensive than other methods, but provides much more accurate data. Can break if not maintained or stored correctly.
- *Indicators* (figure 17.26a, b and c): quick to use, but rely on colour interpretation. Some indicators, such as universal indicator, give more information about the specific pH, whereas others (such as litmus), give a wide range of pH. Samples of solutions should normally be used when using indicators and the colour change may make it difficult to observe other results.
- *Litmus paper:* Allows for a quick visual to determine if a solution is acidic or alkaline; low precision.

FIGURE 17.25 Equipment used in titrations. (a) A volumetric flask is used to prepare a standard solution. (b) A conical flask holds the solution of unknown concentration. (c) A pipette is used to add the unknown solution to the conical flask. (d) The burette holds the standard solution, which is added to the conical flask.

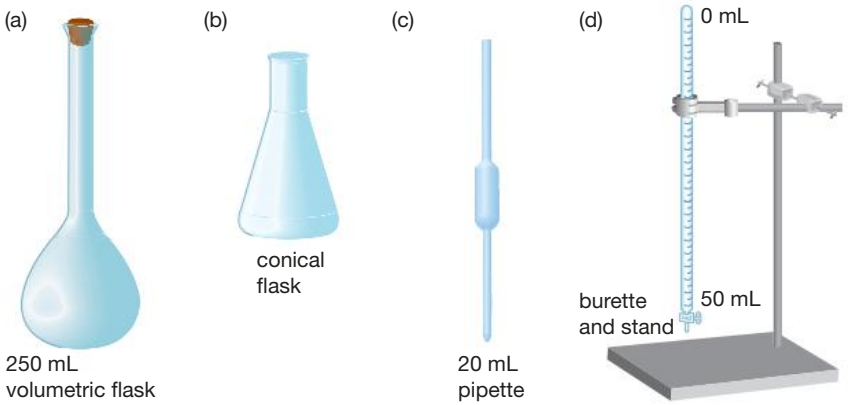
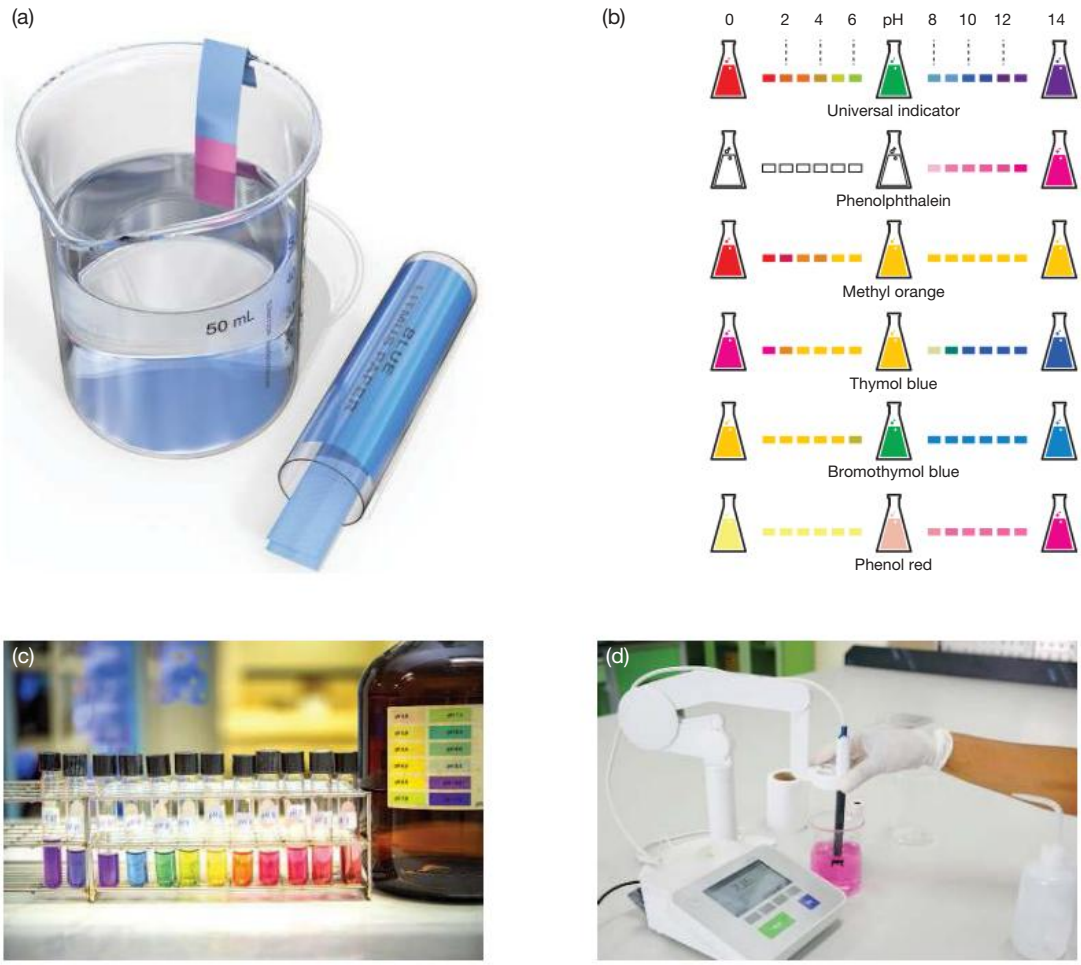


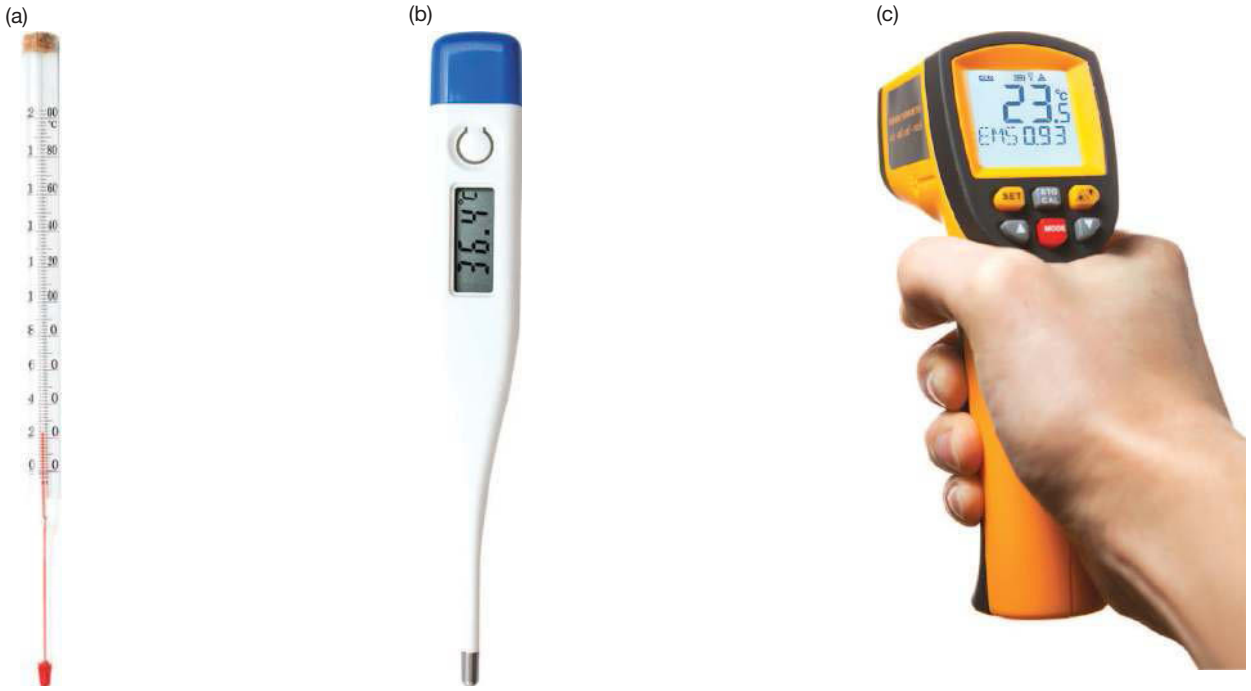
FIGURE 17.26 (a) Litmus paper, changes red in acid and blue in bases (b) Different indicators that can be used to determine pH (c) Universal indicator (d) pH meter



Equipment to measure temperature

- *Thermometer* (figure 17.27a): Easy to use but often can only be read to 1 °C.
- *Digital thermometer* or *temperature probe* (figure 17.27b): Easy to use and more precise, but more expensive.
- *Infrared thermometer* (figure 17.27c): Easy to use, but can be affected by external temperature more; however, it is better when measuring the temperature of gases and solids.

FIGURE 17.27 Equipment used to measure temperature: (a) alcohol thermometer, (b) digital thermometer and (c) infrared thermometer



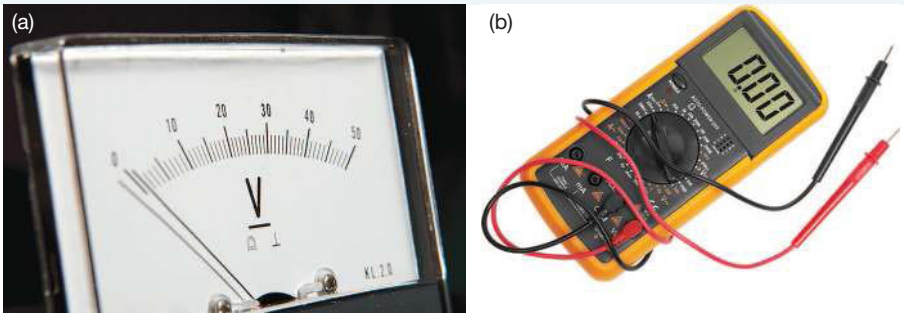
Equipment to measure time

- *Stopwatch:* Simple to use; accurate down to your response time; not reliable for short time intervals.
- *Electronic timer:* Requires some instruction; very accurate; best suited for short time intervals; can be used with electrical contacts and photogates.

Equipment to measure electrical current

- *Meters:* Includes voltmeters, ammeters, galvanometers; easy to set up, but care is needed to ensure the meter is wired into the circuit correctly, or the meter can be damaged; large range of values; usually analogue displays.
- *Multimeters:* Easy to set up; more tolerant of incorrect use, but can be damaged if incorrectly connected to a high current; large range of values; usually digital displays.

FIGURE 17.28 Equipment used to measure electrical current: (a) voltmeter and (b) multimeter



Specialised equipment to measure concentration

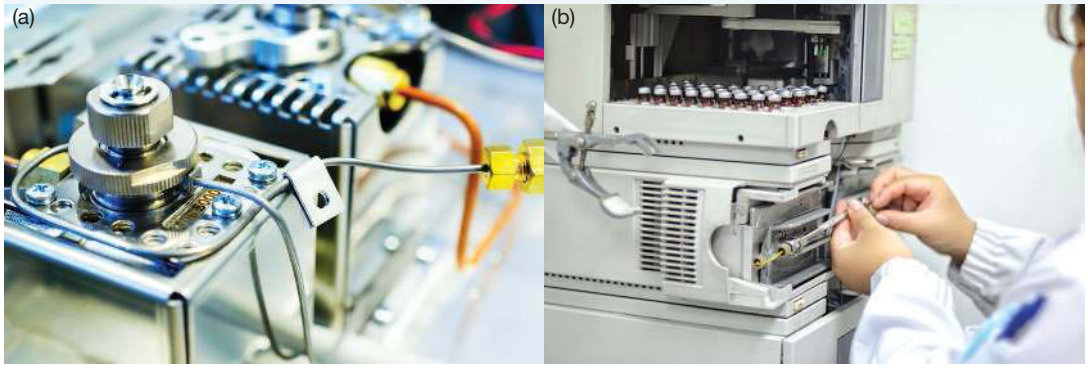
You may also have access to specialised equipment you can use in your practical investigation. Although these are unlikely to be available in a general school laboratory, it is important to note that the following equipment is all highly accurate and precise. Discuss with your teacher if any of these are available for use at your school or nearby. Alternatively, you may wish to explore these as a point of discussion when outlining uncertainties and possible errors in the data and results you obtained. Specific uses of this specialised equipment is discussed in topics 15 and 16.

- *Mass spectrometer:* An analytical instrument that determines the relative isotopic masses of the different isotopes of an element and abundance.
- *Instrumental colorimeter:* An instrument that compares the colour in the test sample with the colours produced in samples of known concentration, allowing for the concentration in the sample to be determined based on absorption.
- *Gas chromatograph:* Measures the content of various components in an injected sample.
- *HPLC (high-pressure liquid chromatography):* Used to measure the concentration of organic substances.

FIGURE 17.29 Mass spectrometer

FIGURE 17.30 Colorimeter

FIGURE 17.31 Specialised equipment used to measure concentration: (a) gas chromatograph and (b) high-pressure liquid chromatography



17.3.4 Conducting investigations

When conducting investigations, it is vital to:

- follow all health and safety protocols
- ensure you know how to use any chosen equipment correctly to minimise errors — ask if you are not sure!
- carefully follow your method and, if any changes are required, note these down in your logbook
- make sure you are controlling variables outside your independent variable to keep your results valid, accurate and precise
- clearly record any results obtained, along with the date; this includes any results that did not go according to plan and any results for both control and experimental groups
- make sure that you carefully pack up equipment after use; if equipment is required to be set up for a few days, ensure it will be in a location where it cannot be affected by other individuals or environmental factors
- repeat your experiment (if time allows) to improve accuracy and minimise errors to improve repeatability.

FIGURE 17.32 Recording your method and results obtained is vital in investigations.

17.3.5 Techniques of primary qualitative and quantitative data collection

Types of data

Data is a set of facts that are collected, observed or generated. Typically, data that you collect is raw data that must later be analysed and interpreted to produce useful information.

Primary data are data collected firsthand. This data provides direct or firsthand evidence about some phenomenon — such as, for example, a research investigation. Your completed logbook will be a **primary source** of data about pursuing an investigation into your research question.

Secondary data are summaries and commentary on the primary data of another individual. **Secondary sources** include review articles in newspapers and popular science magazines written by one person who summarises and comments on the research of others.

Qualitative and quantitative data

Qualitative data (or categorical data) are expressed in words. This type of data is descriptive and not numerical and can be easily observed but not measured. Bar graphs or pie graphs are often used to display the frequencies of categorical variables.

The two types of qualitative data are:

- **Ordinal data**, which can be ordered or ranked; for example, ionisation energies (first, second, third) or opinion polls (strongly agree, agree, disagree, strongly disagree).
- **Nominal data**, which cannot be organised in a logical sequence; for example, types of sub-atomic particles (proton, neutron or electron) or the colour of a solution after a metal salt has been dissolved (blue, clear, white, yellow).

primary data direct or firsthand evidence about some phenomenon, obtained from investigations or observations

primary source a document which is a record of direct or firsthand evidence about some phenomenon

secondary data researchers' comments on or summaries and interpretations of primary data

secondary source a document that comments on, summarises or interprets primary data

qualitative data categorical data that examines the quality of something (e.g. colour or gender) rather than numerical values

ordinal data qualitative data that can be ordered or ranked

nominal data qualitative data that has no logical sequence

Quantitative data (or numerical data) can be precisely measured and have values that are expressed in numbers. Line graphs or scatter plots are often used to display the frequencies of numerical variables.

The two types of numerical data are:

- **Continuous data**, which can take any numerical value, such as the temperature of a solution or the mass of a substance. Any data that is measured is continuous data.
- **Discrete data**, which can only take on set values (integers) that can be counted, such as the number of protons in an atom or the number of electron shells.

Table 17.5 shows examples of how some attributes can be expressed both qualitatively and quantitatively.

TABLE 17.5 Examples of data types

Attribute	Qualitative	Quantitative
Colour	Green	520 nm
Sound	Loud	85 decibels
Speed	Fast	120 kph
Temperature	Warm	40 °C

When collecting data, it is vital to consider what is most appropriate for your investigation. Normally, the best evidence is primary quantitative data, and for a majority of your investigations this is what should be collected and recorded. However, sometimes an investigation may allow for the collection of only qualitative data. Qualitative data can be subjective — for example, identifying a colour may vary between individuals. Thus, it is better to obtain quantitative data, which also allows for trends and patterns to be more easily observed.

It is important that results are carefully checked to make sure that recorded data is correct. Many people interpret measurements slightly differently, or use the wrong units, so make sure you are double-checking data. All your collected data should be recorded in your logbook. Ensure you note down all observations (usually in a table), regardless of whether or not you think they are important.

17.3.6 Generating and collecting primary data

While conducting your investigation, you will generate and collect primary data. Information about how to best present this data in covered in subtopic 17.6.

When generating and collecting primary data, it is important to:

- clearly record all observations, regardless of whether they support your hypothesis
- ensure you provide the date each piece of primary data is collected
- outline the conditions when you collected your primary data (for example, the room temperature, humidity or other conditions that may have affected your results)
- note down both qualitative and quantitative data
- use tables to help organise your data
- ensure you are using the most appropriate equipment to help gather the primary data
- note down other factors that may affect results; for example, if you are working in a group, did the same group member take each measurement, or was it a different group member?
- ensure that all trials are recorded, if completing multiple trials.

quantitative data numerical data that examines the quantity of something (e.g. length, time); also known as numerical data

continuous data quantitative data that can take any continuous value

discrete data quantitative data that can only take on set values

How can you generate and collect this data?

The collection of data is often made through observation and measurements. Observations can be made directly through your senses; for instance, by using your eyes to record the depth of colour in a solution or bubbles forming from a reaction, or using your olfactory sense to note the smell of a product (only if it is safe to do so, and remember to waft the air above the substance toward you; never directly sniff it).

Scientific observations are more often made using instruments that extend the power of the unaided senses, permit accurate measurements, or enable collection of data that may otherwise be undetectable. Examples include thermometers to measure temperature, micrometers to record thicknesses, scales to measure mass, and various probes that measure pH, temperature, conductivity, oxidation-reduction potential (ORP) or CO_2 gas.

Not all observations are made using instruments in a laboratory setting. Familiar tools that can be used to gather data include digital audio recorders, video recorders and digital cameras. Can you think of other tools for making observations and gathering data?

17.3.7 Adapting and extending processes in investigations

During your investigation, you may need to adapt or extend the processes as you conduct your experimentation. You might find that the technique you were planning to use does not work in the way it should have, or that you were unable to control variables.

You might need to make changes during your investigation to enable primary data to be collected. You might also find you only gain a limited amount of data that does not allow you to illustrate trends and patterns, so you may choose to extend your investigation and gain more data points.

It is important that you clearly describe any adaptions or extensions in your logbook, explaining both why and how the modifications were made. It is important when recording data that you also clearly show which technique or piece of equipment you were using if you made any changes.

17.3 Activities

learnon

Students, these questions are even better in jacPLUS

Receive immediate feedback and access sample responses

Access additional questions

Track your results and progress

Find all this and MORE in jacPLUS

17.3 Quick quiz on | 17.3 Exercise | 17.3 Exam questions

17.3 Exercise

1. Explain the difference between the following.
   a. Scientific methodology and scientific method
   b. Primary and secondary data
   c. Continuous data and discrete data
2. Give an example of a strength and a weakness of quantitative and qualitative data.

3. Testing a scientific question by experiment involves a number of stages. These are shown by the following statements. Use the letters to put these stages into their correct order.
   A. Formulate the hypothesis. **B.** Decide on the question. **C.** Analyse the results.
   D. Communicate the results. **E.** Plan the experiment. **F.** Carry out the experiment.
4. Why is it important to control variables in an investigation instead of testing multiple independent variables at one time?
5. A student is conducting an investigation to explore how the concentration of salt affects conductivity. As part of this, she wants to use 100 mL of water, with each container having increasing salt amounts in around 5 g increments. Identify what piece of glassware would be most useful in this investigation and justify your response.

17.3 Exam questions

Question 1 (1 mark)

Explain the purpose of including a control group when designing an investigation.

Question 2 (3 marks)

Describe three factors you need to consider when selecting equipment for your investigation.

Question 3 (4 marks)

Consider an apple.

a. Identify pieces of three quantitative data that you could observe. **(1 mark)**
b. What instruments, if any, would you need to make these observations? **(1 mark)**
c. Identify three pieces of qualitative data that you could observe. **(1 mark)**
d. What instruments, if any, would you need to make these observations? **(1 mark)**

Question 4 (7 marks)

You have decided to conduct an investigation exploring the relationship between the temperature and volume of a balloon by heating and cooling it and measuring the circumference.

a. What is the independent and dependent variable in this investigation? **(1 mark)**
b. Identify each variable as quantitative or qualitative. **(1 mark)**
c. Write an aim and hypothesis for this experiment. **(2 marks)**
d. List two variables that need to be controlled in this investigation. **(2 marks)**
e. MC Of the different scientific investigation methodologies listed, which is most appropriate for this investigation? **(1 mark)**
   A. Case study **B.** Literature review **C.** Simulation **D.** Experiment

More exam questions are available in your learnON title.

17.4 Health, safety and ethical guidelines

KEY KNOWLEDGE

- Health, safety and ethical guidelines relevant to the selected scientific investigation

KEY SCIENCE SKILL

- Comply with safety and ethical guidelines

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17.4.1 Health and safety guidelines

Part of the enjoyment of a practical investigation is that the topic may be unconventional or use an innovative method. Such situations, however, can present some risk, so special care needs to be taken to ensure you and others are safe.

General safety rules

Some general safety precautions help to ensure that you and others are not injured in the laboratory. These precautions include the following:

FIGURE 17.33 Safety equipment, such as safety glasses, gloves and lab coats, is vital in various experiments.

- Wear protective clothing. This might include laboratory coat, safety glasses and gloves.
- Be aware of the position of safety equipment such as the fire blanket, fire extinguisher, safety shower and eye wash.
- Ask if you are unsure how to operate equipment or how to use apparatus.
- Read labels carefully to confirm contents and concentration of chemicals.
- Clean and return all equipment to the correct places, ensuring lids are placed back on containers when not in use.
- Check for the correct disposal of equipment and chemicals, including damaged equipment (such as broken glassware).
- Read instructions carefully before commencing an experiment.
- Prepare a **risk assessment** for required chemicals and equipment.
- Do the investigation as outlined in your approved plan. Don't vary your plan without approval from your teacher.
- Don't do experimental work unsupervised unless you have prior approval from your teacher.
- When first setting up electrical experiments, ask your teacher to check the circuit.
- Don't interfere with the equipment set-up of others.

Creating a risk assessment

It is important to address health and safety concerns through the use of a risk assessment.

A risk assessment is a procedure for identifying hazardous chemicals, what the risks are and how to work safely with them.

The risk assessment also assesses potential hazards with equipment being used and outlines standard handling procedures to ensure the health and safety of individuals and the environment.

Risk assessments should also take into considerations correct disposal of equipment and chemicals to adhere to safety and bioethical guidelines. Many chemicals are harmful to the environment, so correct disposal is paramount. Table 17.6 lists the usual requirements for a written risk assessment.

risk assessment a document that examines the different hazards in an investigation and suggested safety precautions

TABLE 17.6 Requirements for a written risk assessment

Requirement	Information included
Outline of investigation	Title, date and location of task
Summary of method	Brief list of steps indicating how the chemicals and equipment will be used
Equipment/chemicals used	List of materials used in the experiment
Equipment/chemical risk and hazards	List of hazards associated with each of the materials being used
Risk control measures	Precautions taken to limit risks, including identification of safety equipment used

An example of a risk assessment is shown in figure 17.34.

FIGURE 17.34 An example risk assessment

ACTIVITY	Cross-linking an addition polymer to make slime

SUMMARY OF EXPERIMENT	
AIM	To investigate how the properties of a linear polymer may be altered by the introduction of weak cross-linking between its chains.
METHOD	1. Pour the polyvinyl alcohol into the beaker and add a few drops of the food dye (optional). 2. Add the borax solution and stir with the paddle pop stick. It will take a few minutes for the slime to appear. 3. Perform tests on the product that will enable you to describe its properties and how these are different to the original polymer.

PROTECTIVE MEASURES				
GLASSES	GLOVES	DUST MASK	LAB COAT	FUME HOOD
x	x		x	

SAFETY INFORMATION

REACTANT

Polyvinyl alcohol (solution, 6%)

Hazards
- Flammable
- Irritating to the eyes

Safety precautions
- Wear gloves, safety glasses and a lab coat
- Keep away from sources of ignition
- Use in well-ventilated areas

FIRST AID	
SWALLOWED	Rinse mouth out with water immediately and repeat until all traces are removed. Seek medical attention.
EYE	Flush out with water. Seek medical attention if pain or irritation persists.
SKIN	Wash with soap and water.
INHALED	Move into fresh air, give oxygen if required. Seek medical attention if breathing if difficult.

REACTANT

Borax (sodium tetraborate, solution, 4%)

Hazards
- Not classified as hazardous substance at the concentration used
- However, at concentrations above 4.5%, may damage fertility and unborn child

Safety precautions
- Wear gloves, safety glasses and a lab coat
- Should not be handled by pregnant women. Those of reproductive age should also avoid the chemical.
- Wash hands after use, even if gloves were worn

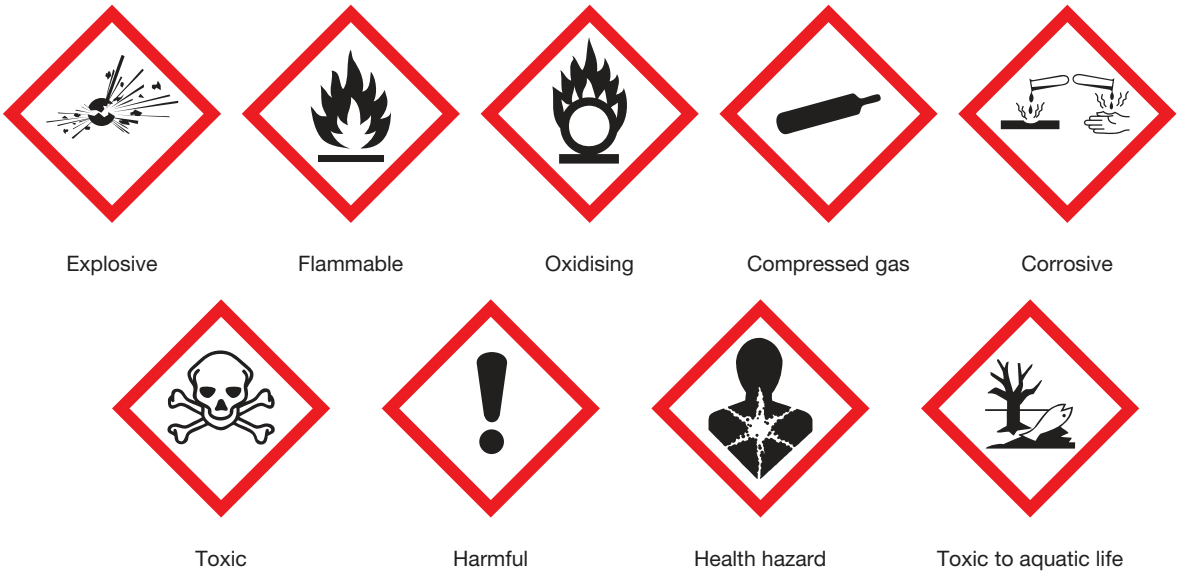
FIRST AID	
SWALLOWED	Rinse mouth out with water immediately and repeat until all traces are removed. Seek medical attention.
EYE	Flush out with water. Seek medical attention if pain or irritation persists.
SKIN	Wash with soap and water. Seek medical attention if pain or irritation persists.
INHALED	Move into fresh air, give oxygen if required. Seek medical attention if breathing if difficult.

CONCLUSION

- Wear gloves, glasses and a lab coat for the duration of this experiment
- Make sure that the area is well-ventilated

Signed: _____ **Date:** _____

FIGURE 17.35 GHS pictograms can be found throughout risk assessments to visually show any possible hazards.



17.4.2 Ethics

Ethics are the principles of acceptable and moral conduct. They apply not only to scientific investigations but also to many aspects of life, determining what is 'right' and what is 'wrong'.

Science interacts with ethics in several ways, including:

- the way an experiment is conducted
- confidentiality and morality around research
- conflicts with religious and personal beliefs.

Ethical standards and considerations also apply to any type of research or data collection method involving people (or animals).

Ethics are particularly obvious in drug trials, both with animal testing and human trials. It is important that individuals give permission and are made aware of all possible side effects and risks associated with treatments.

FIGURE 17.36 Ethics relate to the idea of moral choice.

While drug trials are not relevant for your practical investigation, ethical considerations still need to be made. For example, you may be conducting surveys on issues relating to opinions on water quality. The confidentiality of individuals who provide responses needs to be considered and permissions gained from respondents.

Being mindful of individuals in regards to personal beliefs is also important. While drug trials have minimum ethical standards for the use of animals in trials, for some individuals, differing personal beliefs may affect experimentation and interpretation of data. This is an ethical consideration that needs to be evaluated and understood when researching and reporting on these topics. If your investigation is focused on Aboriginal and Torres Strait Islander peoples' traditional knowledge, techniques or artefacts, you will need to apply appropriate cultural protocols to access information. See topic 9 for more information.

ethics acceptable and moral conduct determining what is 'right' and 'wrong'

17.4 Activities

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17.4 Quick quiz on | 17.4 Exercise | 17.4 Exam questions

17.4 Exercise

1. MC A scientist makes a mistake when calculating the conductivity of a new nanomaterial. The mistake is carried over into a published paper in a scientific journal. What should the scientist do?
   - A. They should ignore their mistake and hope no-one notices.
   - B. They should pretend the journal introduced a typo.
   - C. They should report their mistake to the editor of the scientific journal and ask that a correction be made in the next edition of the journal.
   - D. They should hang their head in shame and never do science again.
2. Research and identify possible hazards and suggest safety precautions for the following equipment and chemicals.
   - a. 2.0 mol L^{-1} hydrochloric acid
   - b. Burette
   - c. Boiling water
   - d. Thermometer
3. Provide two examples of when ethics may be important in a chemistry investigation.
4. List three purposes of a risk assessment.
5. Look around your laboratory and note its safety features and equipment. Then answer the following.
   - a. Does it have any stored pressure fire extinguishers? How are these identified? On what types of fire can these be used and on what types of fire should they not be used?
   - b. Does it have any dry chemical extinguishers? How are these identified? On what types of fire can these be used and on what types of fire should they not be used?
   - c. Where is/are the fire blanket(s) located? Describe a scenario in which a fire blanket would be used and how you would use it.
   - d. Where are the master (emergency) shut-offs for gas and electricity located?
6. The SDS for a chemical to be used in an experiment contains the following risk phrases:
   - Irritating to eyes
   - Skin/flammable/vapours may cause dizziness

   Suggest appropriate methods to reduce these identified risks.

17.4 Exam questions

Question 1 (1 mark)

The Material Safety Data Sheet (MSDS) for a concentrated solution of glycolic acid states that it is corrosive to the eyes, skin and respiratory system, and that it is harmful if a concentrated solution of it is ingested or inhaled.

Outline one safety precaution that should be taken when handling this compound.

Question 2 (2 marks)

A student decided to investigate the effect of temperature on the rate of the following reaction.

$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

Part of the student's experimental report is provided here.

Effect of temperature on the rate of production of carbon dioxide gas

Aim

To find out how temperature affects the rate of production of carbon dioxide gas, CO_2, when a solution of hydrochloric acid, HCl, is added to chips of calcium carbonate, $CaCO_3$.

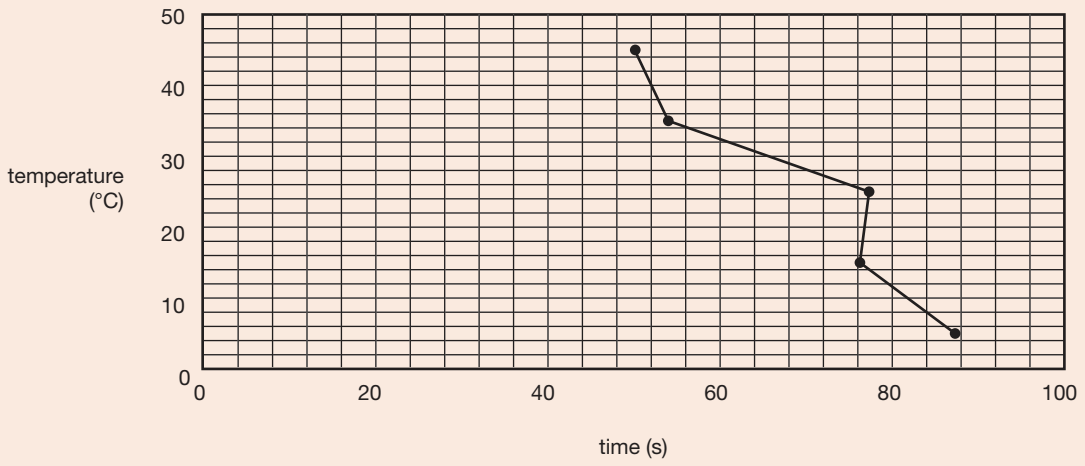
Method

1. Put 0.6 g of $CaCO_3$ chips into a conical flask.
2. Put a reagent bottle containing 2 M HCl into a water bath at 5 °C.
3. When the temperature of the HCl solution has stabilised at 5 °C, use a pipette to put 10.0 mL of the HCl solution into the conical flask containing the $CaCO_3$ chips.
4. Put a balloon over the conical flask and begin timing.
5. When the top of the balloon has inflated so that it is 10 cm over the conical flask, stop timing and record the time.
6. Repeat steps 1 to 5 using temperatures of 15 °C, 25 °C. 35 °C and 45 °C.

Results

The following graph gives the experimental results.

Graph of experimental results



What does the student need to do to ensure that they comply with all applicable safety guidelines during the investigation?

Question 3 (2 marks)

A student designed an experiment to investigate current efficiency during the electrolysis of a sodium chloride, NaCl, solution.

This reaction is modelled by the equation:

$2Na^+(aq) + 2Cl^-(aq) + 2H_2O(l) + \rightarrow H_2(g) + Cl_2(g) + 2NaOH(aq)$

a. Identify a safety risk associated with the chemicals produced during the experiment. **(1 mark)**

b. What are the safety measures required to reduce the safety risk identified? **(1 mark)**

Question 4 (2 marks)

A group of students designed and carried out an experiment to investigate if tartaric acid, $C_4H_6O_6$, that was bought commercially is 99 per cent pure, as claimed by the manufacturer. The experiment involved titrating $C_4H_6O_6$ with sodium hydroxide, NaOH, solution, calculating the percentage purity of $C_4H_6O_6$ and comparing the experimental value to the manufacturer's stated value.

Part of the report submitted by one of the students is shown here.

Method

Part A — Preparation of tartaric acid solution

1. Purchase tartaric acid, $C_4H_6O_6$, powder.
2. Prepare a solution of $C_4H_6O_6$ by accurately measuring 30.0 g of the powder, placing it in a 500.00 mL volumetric flask and then making it up to 500.00 mL with de-ionised water.

Part B — Titration

1. Collect stock solution of 0.5 M sodium hydroxide, NaOH, and use this to fill a burette.
2. Deliver a 10.00 mL aliquot of $C_4H_6O_6$ solution into a conical flask. Add four drops of phenolphthalein indicator.
3. Carefully titrate 0.5 M NaOH into the $C_4H_6O_6$ solution until a permanent pink colour remains.
4. Record the volume of the titre.
5. Repeat the titration until concordant titres are obtained.

The material safety data sheet (MSDS) for $C_4H_6O_6$ powder includes the following statement: 'Warning! This product causes eye, skin and respiratory tract irritation.'

Apart from a laboratory coat, what personal protective equipment (PPE) should be used by the students in each of the following situations?

a. Preparing the $C_4H_6O_6$ solution **(1 mark)**

b. Conducting the titration **(1 mark)**

More exam questions are available in your learnON title.

17.5 Quality of data and measurements

KEY KNOWLEDGE

- Accuracy, precision, repeatability, reproducibility, resolution, and validity of measurements in relation to the investigation

KEY SCIENCE SKILLS

- Plan and conduct investigations
- Analyse and evaluate data and investigation methods

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For your investigation, ensuring that your measurements are accurate, precise, repeatable, reproducible and valid is vital. In order to improve these in your measurements, you need to ensure that your investigation is both reproducible and repeatable. You should be able to identify all of these characteristics in experiments to evaluate the quality of experimental data.

17.5.1 Accuracy and precision

Accuracy

Accuracy refers to how close an experimental measurement is to a known value. If an archer is accurate, for example, their arrows hit close to the target. Consider an experimental calculation of the boiling point of water, which is known to be 100 °C. A student who obtained an experimental value of 99 °C is more accurate than a student who obtains a value of 105 °C.

Table 17.7 shows two investigations by different students. Student 1 has more accurate results, because their results are much closer to the actual boiling point of water. Their measurements are no more than 2 °C from the known value.

TABLE 17.7 Two investigations by different students to determine the boiling point of water

Investigation by Student 1		Investigation by Student 2	
	Temperature (°C)		Temperature (°C)
Trial 1	100	Trial 1	95
Trial 2	102	Trial 2	94.8

Precision

Precision refers to how close multiple measurements of the same investigation are to each other. Note that results that are precise may not be accurate. It is often difficult to have completely precise results due to **random errors**.

Table 17.8 again shows two investigations by different students. Student 1 has more precise results, because the range of their measurements (1.2 °C) is much smaller when compared to Student 2 (7 °C).

TABLE 17.8 Two investigations by different students measuring the point in which water boils

Investigation by Student 1		Investigation by Student 2	
	Temperature (°C)		Temperature (°C)
Trial 1	98.5	Trial 1	100
Trial 2	98.6	Trial 2	102
Trial 3	99	Trial 3	95
Trial 4	98	Trial 4	99
Trial 5	99.2	Trial 5	106

FIGURE 17.37 Results vary in precision when recording temperature.

Comparing accuracy and precision

Sometimes an individual with the most accurate data does not have the most precise data. In order to obtain the best experimental data, we want results that are both accurate and precise. This can be improved by minimising errors — the accuracy of results is affected by systematic errors and the precision of results is affected by random errors.

accuracy how close an experimental measurement is to a known value

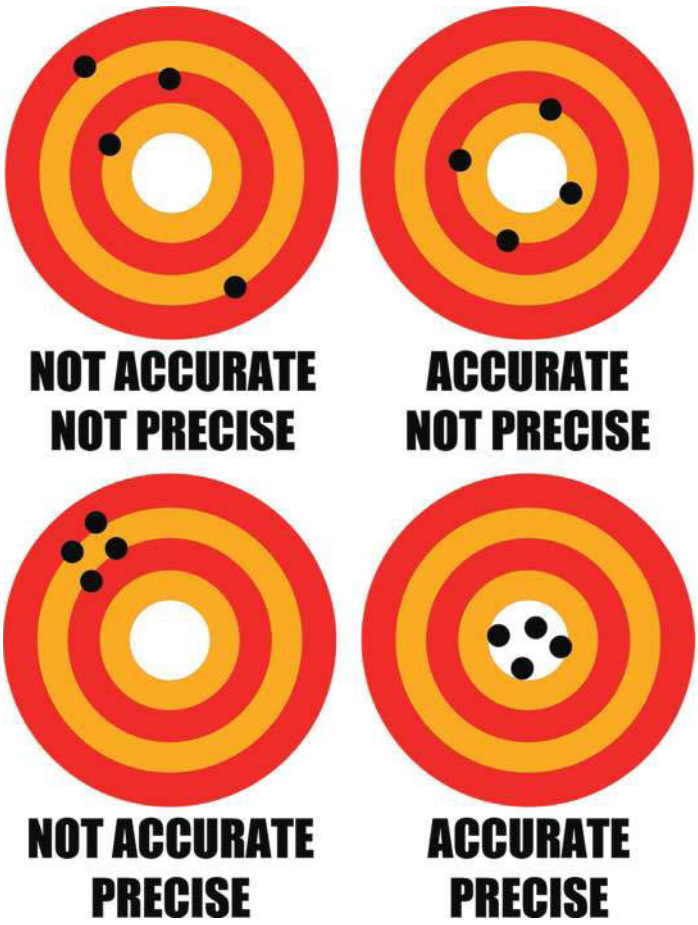
random errors chance variations in measurements; result in a spread of readings

Accuracy versus precision

A good way to remember the difference between *accurate* and *precise* is to use word association:

- **A**ccurate data is close to the **a**ctual value.
- **P**recise data is when all the different **p**oints of data are close together.

FIGURE 17.38 Comparing precision and accuracy



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SAMPLE PROBLEM 3 Evaluating data for precision and accuracy

Students conducted an experiment to determine the temperature of a substance as it changed from a solid to a liquid. They repeated the experiment four times and achieved the following results:

Student 1: 56.5 °C, 58.0 °C, 60.0 °C, 55.0 °C
Student 2: 60.5 °C, 61.0 °C, 60.5 °C, 62.0 °C
Student 3: 56.5 °C, 58.5 °C, 57.0 °C, 56.0 °C

Students were then provided with the exact value of the melting temperature of the substance, which was found to be 56.48 °C.

a. Which student had the least accurate data?
b. Which student had the least precise data?
c. Was the student with the most precise data also the student with the most accurate data? Explain your answer.

THINK	WRITE
a. 1. Review what accuracy means.	Accuracy refers to how close a measurement is to a known value.
2. Explore the data of the three students.	Student 1 had data 1.48 °C lower and 3.52 °C higher than the actual data. Student 2 had data that was up to 5.52 °C higher. Student 3 had data that was 0.48 °C lower and data that was 2.02 °C higher.
3. Determine which student had the least accurate data.	Student 2 had the least accurate data, because their values were the furthest away from the actual value.
b. 1. Review what precision means.	Precision refers how close multiple measurements of the same investigation are to each other.
2. Explore the data of the three students.	Student 1 had a data range of 5.0 °C. Student 2 had a data range of 1.5 °C. Student 3 had a data range of 2.5 °C.
3. Determine which student had the least precise data	Student 1 had the least precise data.
c. 1. Identify the students with the most accurate and most precise data.	Using the results from parts **a** and **b** it can be seen that Student 3 had the most accurate data and Student 2 had the most precise data.
2. Respond to the question and explain your answer	The student who had the most precise data was not the same student who had the most accurate data. Students may have measurements very close together (precise), but these may not be accurate. This may be to errors in their measuring device or their interpretation of the melting point (when the solid is a liquid). Data may also be accurate without being precise: you can be close to the target, but the readings are inconsistent. For reliable and valid results, data should be both accurate and precise.

PRACTICE PROBLEM 3

Students conducted an experiment to determine the temperature of a substance as it changed from a liquid to a gas. They repeated the experiment four times and achieved the following results:

Student 1: 85.4 °C, 92.0 °C, 82.0 °C, 75.5 °C

Student 2: 83.5 °C, 85.0 °C, 85.5 °C, 86.5 °C

Student 3: 85.5 °C, 90.0 °C, 89.5 °C, 81.0 °C

Students were then provided with the exact value of the boiling temperature of the substance, which was found to be 85.4 °C.

a. **Which student had the least accurate data?**
b. **Which student had the least precise data?**
c. **Was the student with the most precise data also the student with the most accurate data? Explain your answer.**

17.5.2 Repeatability, reproducibility and resolution

Repeatability

Repeatability refers to how close the results of successive measurements are to each other in *exactly the same conditions*.

Conditions that should be the same include:

- observer
- way of measuring results
- measuring instrument
- location
- laboratory conditions
- time.

Consider the following investigation to explore how temperature affects the rate of diffusion using 16 beakers.

- 4 beakers of 50 mL water at 20 °C
- 4 beakers of 50 mL water at 40 °C
- 4 beakers of 50 mL water at 60 °C
- 4 beakers of 50 mL water at 80 °C

A student places exactly 3 mL of food dye in each beaker. They then record the time taken for the food dye to completely spread through the water. If they get similar results in their data within each of the test temperatures, the method is said to be repeatable. The more times an experiment is repeated, the closeness of the data produced directly reflects the degree of precision.

EXTENSION: Statistical analysis of results to calculate repeatability

Calculating the standard deviation of the variable you are investigating gives you a measure of the repeatability of your experiment.

If you repeat the same experiment n times, measuring the value of the same variable x, the formula for the standard deviation of a sample is:

$$\sigma_{\text{sample}} = \sqrt{\frac{\sum |x - \bar{x}|^2}{n - 1}}$$

where $\bar{x}$ is the mean.

In the example of a student measuring the rate of diffusion, if the values measured for 10 trials at 60 °C are, in seconds, 12.5, 11.4, 12.3, 12.6, 12.5, 11.6, 12.2, 12.0, 11.7 and 11.9, then the standard deviation of this sample is 0.42 (to two decimal places).

The closer to zero the standard deviation of your sample is, the more repeatable your results are.

Reproducibility

Reproducibility is how close results are when the same variable is being measured, but *under different conditions*.

Consider the investigation just described. The next day, the other students in the class (that is, different observers) use the procedure set out by the first student, following the same method except that they measure the 50 mL in a measuring cylinder (different equipment) before adding it to the beaker. In each beaker, they place exactly 3 mL of red food dye (since the colour wasn't specified) and record the time taken for the food dye to completely spread through the water. If these student gets similar results to the first student, the method is said to be reproducible.

Reproducibility is a key part of the scientific process. It is important to check the reproducibility of your data or experiments because, if they are not reproducible, it might be due to systematic errors affecting the accuracy of your measurements.

Resolution

Resolution refers to the smallest change of measurement that a particular equipment can detect.

For example, if a burette is marked every 0.1 mL, sometimes the meniscus will fall exactly on a marked line (for example, 9.20 mL); sometimes it will fall between two marked lines (such as between 9.10 and 9.20 mL). In this case, the measurement can only be stated as 9.15 mL — a measurement such as 9.13 mL, for example, cannot be claimed. The resolution of a burette with increments of 0.10 mL is half of that, or 0.05mL.

17.5.3 Validity

The **validity** of data refers to whether the experiment investigates what it claims to investigate. Validity factors in both experimental design and implementation.

If your experimental method clearly relates to the purpose of the investigation, you control any potentially confounding variables, take care to be precise in your measurements and thorough in your analysis, your results should be valid and meaningful. Valid experiments also have minimised factors such as **experimental bias**.

Validity can be:

- *Internal:* the degree to which the experimental procedures measure what they are supposed to measure. Testing internal validity asks questions such as: Can the results be trusted? Could another unknown variable have influenced the results? Could any cause and effect relationships identified be explained by other factors? Internal validity is favoured when an experiment is carefully designed and a scientific approach is used.
- *External:* the extent to which research findings can be generalised to the greater population. Testing external validity asks questions such as: Is the sample of the population that was used in the research study representative of the greater population? Can we be reasonably sure that the results of the research are applicable to the greater population? External validity is favoured by an experimental design that includes the use of a control group, has control and experimental groups of sufficient sizes, and randomly assigns subjects to the control and test groups.

resolution the smallest change of measurement that a particular equipment can detect

validity how accurately an experiment investigates the claim it is intended to investigate

experimental bias a type of influence on results in which an investigator either intentionally or unintentionally manipulates results to get a desired outcome

bias the intentional or unintentional influence on a research investigation

Minimisation of experimental bias

Bias is an intentional or unintentional influence on a research investigation as a result of systematic errors introduced by the researcher into the sampling or the testing procedures of an experiment. These biases will prejudice the research findings and raise questions about their validity and reliability.

Numerous types of bias are possible in experiments — some of which apply more to chemistry than others. Types of bias include:

- ***Measurement bias*:** This bias occurs when experimenters manipulate results in order to get a desirable outcome. Sometimes this can be unintentional (for example, if an experimenter consistently records the boiling point earlier than they should, leading to a lower recorded temperature). Often, however, it is through the deliberate actions of an individual. (For example, when measuring the rate of a reaction, an experimenter might deliberately stir one reaction to make the rate appear higher to better suit their hypothesis.)
- ***Selection bias*:** This type of bias can arise when test subjects are not randomly assigned to the experimental and control groups.
  An example of selection bias is in clinical trials of a new synthetic drug. A doctor may choose family members to receive a drug being tested and have individuals he doesn't know receive a placebo. Selection bias can be minimised by randomly and equally allocating subjects to each group.
- ***Sampling bias*:** This type of bias can arise if the subjects chosen for the study are not representative of the target population. If this occurs, the research results cannot be generalised to that population.
  For example, the average height of students at a school is calculated, but due to time constraints only 50 out of the 600 students are measured. If only year 7 students are measured, this will not be representative of the target population and is an example of sampling bias.
  Sampling bias can be minimised by ensuring that the participants in the study are a reasonable representation of the target population.
- ***Response bias*:** This type of bias arises when only certain members of the target population respond to an invitation to participate in a scientific trial, resulting in an unrepresentative sample of the larger population.
  Similar to sampling bias, response bias can be minimised by ensuring that the subjects responding to be included in the study are a reasonable representation of the target population.

measurement bias a type of influence on results in which an experiment manipulates results to get a desired outcome; may be unintentional (i.e. through the placebo effect) or intentional

selection bias a type of influence on results in which test subjects are not equally and randomly assigned to experimental and control groups

sampling bias a type of influence on results in which participants chosen for a study are not representative of the target population

response bias a type of influence on results in which only certain members of the target population respond to an invitation to participate in the clinical trial, resulting in an unrepresentative sample of the larger population

FIGURE 17.39 Selection bias

FIGURE 17.40 Sampling bias

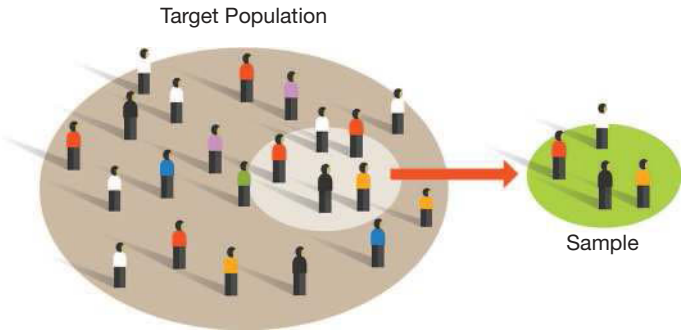

FIGURE 17.41 Response bias

17.5 Activities

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17.5 Quick quiz on | 17.5 Exercise | 17.5 Exam questions

17.5 Exercise

1. Distinguish between validity and reproducibility.
2. Give an example of when results would not be considered repeatable.
3. Under what circumstances can it be said that the conclusions or findings of research are 'valid'?
4. List two procedures that could have an adverse impact on the internal validity of an experiment.
5. Explain, with reference to an example, how an experiment can be reproducible and repeatable but not valid, whereas an experiment that is valid must also be reproducible.
6. Explain, with examples, the difference between precision and accuracy.

17.5 Exam questions

Question 1 (1 mark)

MC What does using large sample sizes in an experiment increase?

A. Repeatability
B. Precision
C. Validity
D. Uncertainty

Question 2 (1 mark)

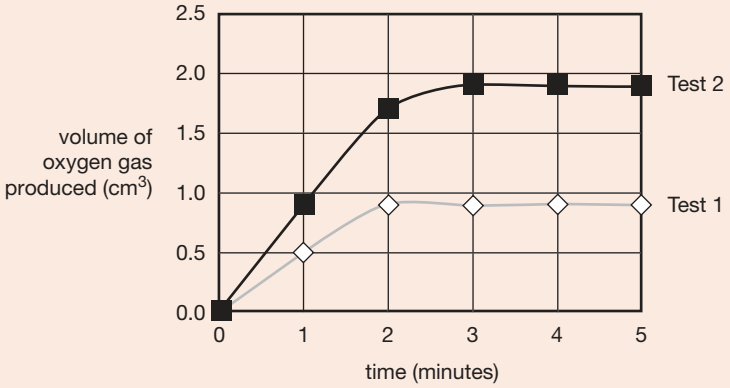
MC When can experimental bias occur?

A. When all variables but one is kept constant in an experiment
B. When a large representative sample size is used
C. When the results from one experimental group are ignored
D. When an experiment is repeated many times

Question 3 (1 mark)

MC Hydrogen peroxide is a toxic by-product of many biochemical reactions. Cells break down hydrogen peroxide into water and oxygen gas with the help of the intracellular enzyme catalase. The optimum pH of catalase is 7.

A Chemistry student measured the activity of catalase by recording the volume of oxygen gas produced from the decomposition of hydrogen peroxide when a catalase suspension was added to it. The catalase suspension was made from ground, raw potato mixed with distilled water. The student performed two tests and graphed the results



Test 1 used 5 mL of 3% hydrogen peroxide solution and 0.5 mL of catalase suspension, and was conducted at 20 °C in a buffer solution of pH 7. Test 2 was carried out under identical conditions to Test 1, except for one factor that the student changed.

During the experiment, the student measured the varying pH levels using a digital pH meter. The student calibrated the meter using a pH 7 buffer solution.

Why did the student calibrate the pH meter?

A. To ensure a random error would not influence the results
B. To eliminate the effect of all uncontrolled variables
C. To enable the use of the instrument with precision
D. To allow the pH to be measured accurately

Question 4 (8 marks)

Four students conducted an investigation to determine the time taken to produce a set volume of hydrogen from an acid metal reaction at room temperature. The expected time was around 65 seconds.

The results of each student are shown in the provided table.

	Measurement 1	Measurement 2	Measurement 3
Aurelia	57 s	61 s	62 s
Sienna	71 s	73 s	74 s
Levi	55 s	71 s	52 s
Cruz	65 s	88 s	88 s

a. Which student had the most accurate data? How do you know? **(2 marks)**
b. Which student had the least accurate data? How do you know? **(2 marks)**
c. Which student had the most precise data? How do you know? **(2 marks)**
d. Which student had the least precise data? How do you know? **(2 marks)**

More exam questions are available in your learnON title.

17.6 Ways of organising, analysing and evaluating primary data

KEY KNOWLEDGE

- Ways of organising, analysing and evaluating generated primary data collected to identify patterns and relationships, and to identify sources of error
- Ways of presenting key findings and implications of the selected scientific investigation

KEY SCIENCE SKILLS

- Generate, collate and record data
- Analyse and evaluate data and investigation methods
- Construct evidence-based arguments and draw conclusions

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17.6.1 Organising primary data

Scientists gather raw data or plain facts from their observations. For your investigations across Units 1–4, you need to present your secondary and primary data — as, for example, text entries, sketches, tables and diagrams — in logbooks or in field notebooks. These may be supplemented by audio and video recordings.

Some of the most common ways to organise and present primary data include:

- tables
- schematic diagrams
- flow charts
- graphs.

FIGURE 17.42 Using logbooks is an easy way to organise data

As part of your investigation and representations and analysis of primary data, you may need to do some simple calculations — which may include calculations of percentages, mean, percentage change and ratios — to be able to provide a richer comparison.

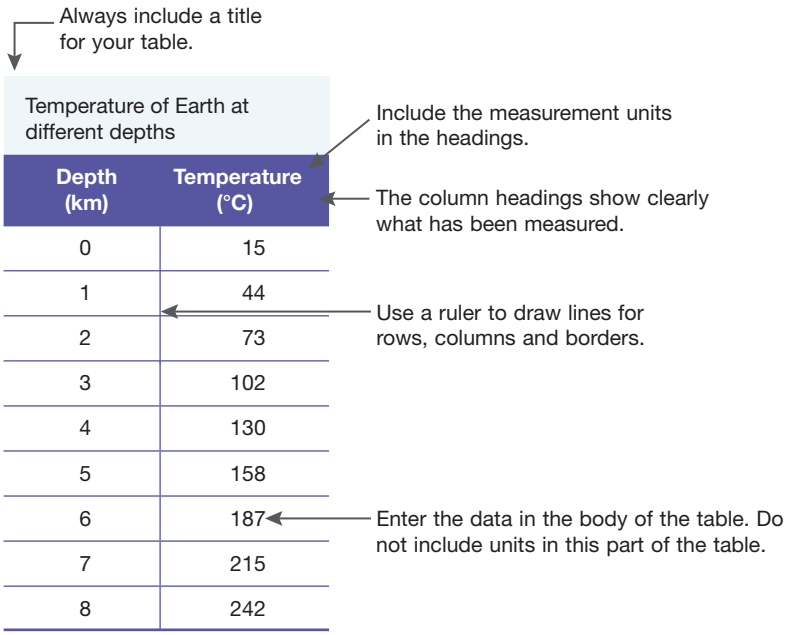
Using a table

Tables should be used when you initially record data; they help separate and organise your information. This is usually the most appropriate technique to gather your data for your logbook.

All tables should:

- have a heading
- display the data clearly, with the independent variable in the first column and the dependent variable in later columns
- include units in the column headings and not with every data point
- be designed to be easy to read; if a table becomes too complicated, it is better to break it down into a number of smaller tables
- use appropriate significant figures (or decimal places) that are consistent across data sets (for example, if one data point is 2.5, the other data point cannot be just 1, it should be 1.0).

FIGURE 17.43 Format of a scientific table



Using a graph

Presenting results as a graph makes it easier to see patterns and trends in your data, allowing more accurate result analysis. While you will usually use a table to record results in your logbook, processing your data into graphs is recommended to identify and illustrate trends, and is preferred on your scientific poster.

When drawing graphs:

- decide on the type of graph to be used. Different types of information are better suited to different types of graphs.
  - If both the independent and dependent variables are quantitative, a line graph or scatter plot is preferred.
  - Bar graphs are used when one piece of data is qualitative and the other is quantitative.
  - Histograms are used when intervals and frequency are being explored.
- include a title; this should link the dependent and independent variables that are shown in the graph
- assign axes correctly; the independent variable should be on the horizontal (x) axis, and the dependent variable is on the vertical (y) axis
- rule axes and label each clearly; those displaying numerical variables should have a clearly marked scale and units
- make sure your scale is suitable and the numbers are evenly distributed
- draw a line (or curve) of best fit as required (for example, for continuous variables); this is a smooth curve or line that passes as close as possible to all the plotted points
- include the origin, the zero value for the variables, on both axes.

The most common graphs you will use in chemistry include:

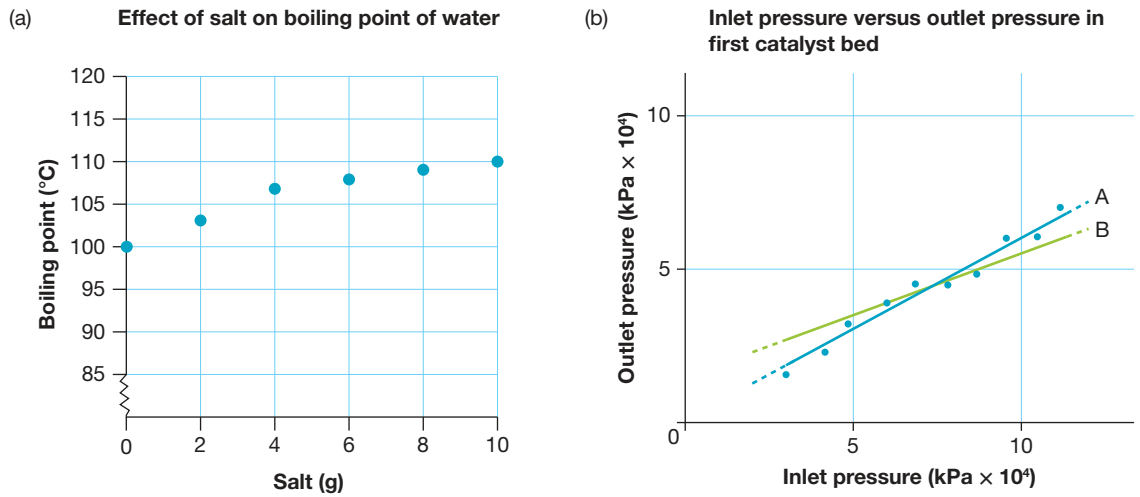
- scatterplots
- line graphs
- bar/column graphs
- histograms.

Scatterplots

Scatterplots require both sets of data to be numerical. Each dot represents one observation, recorded in regards to the independent and dependent variable. A scatterplot can easily show trends between data sets, and correlations can be identified.

scatterplots graph in which two quantitative variables are plotted as a series of dots

FIGURE 17.44 Examples of (a) a scatterplot graph and (b) a scatterplot graph with a line of best fit

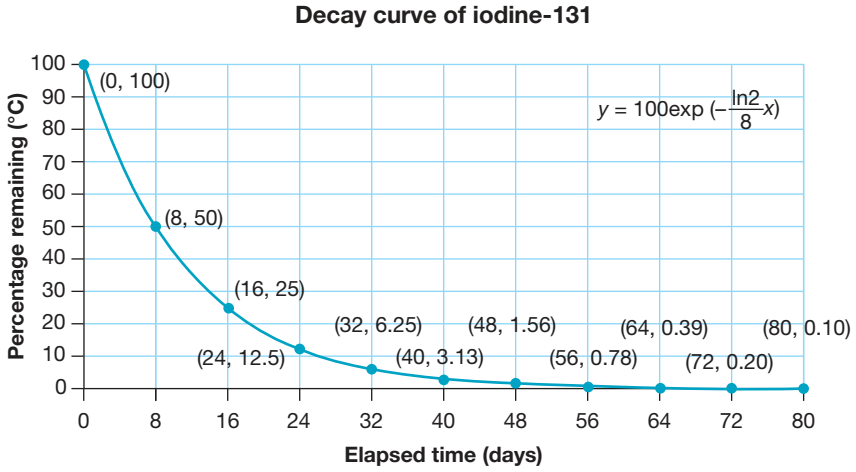


BACKGROUND INFORMATION: Drawing a line of best fit

A **line of best fit** can be used to show the general trend of data in a scatterplot graph, and provides a quick summary.

The line of best fit doesn't need to pass through each data point. Although you should try to draw the line through each data point if possible, you may not be able to go through all of them. As a general rule, try to have as many data points above your line as you have below. Don't assume your line must pass through the origin. For example, graph (b) in figure 17.44 shows two possible lines of best fit. Although line A does not pass through any points, it is a better fit than line B. For some data, a curve of best fit may be more appropriate.

FIGURE 17.45 Example of scatterplot with an exponential line of best fit



line of best fit trend line added to a scatterplot to best express the data shown; these are straight lines, and are not required to pass through all points

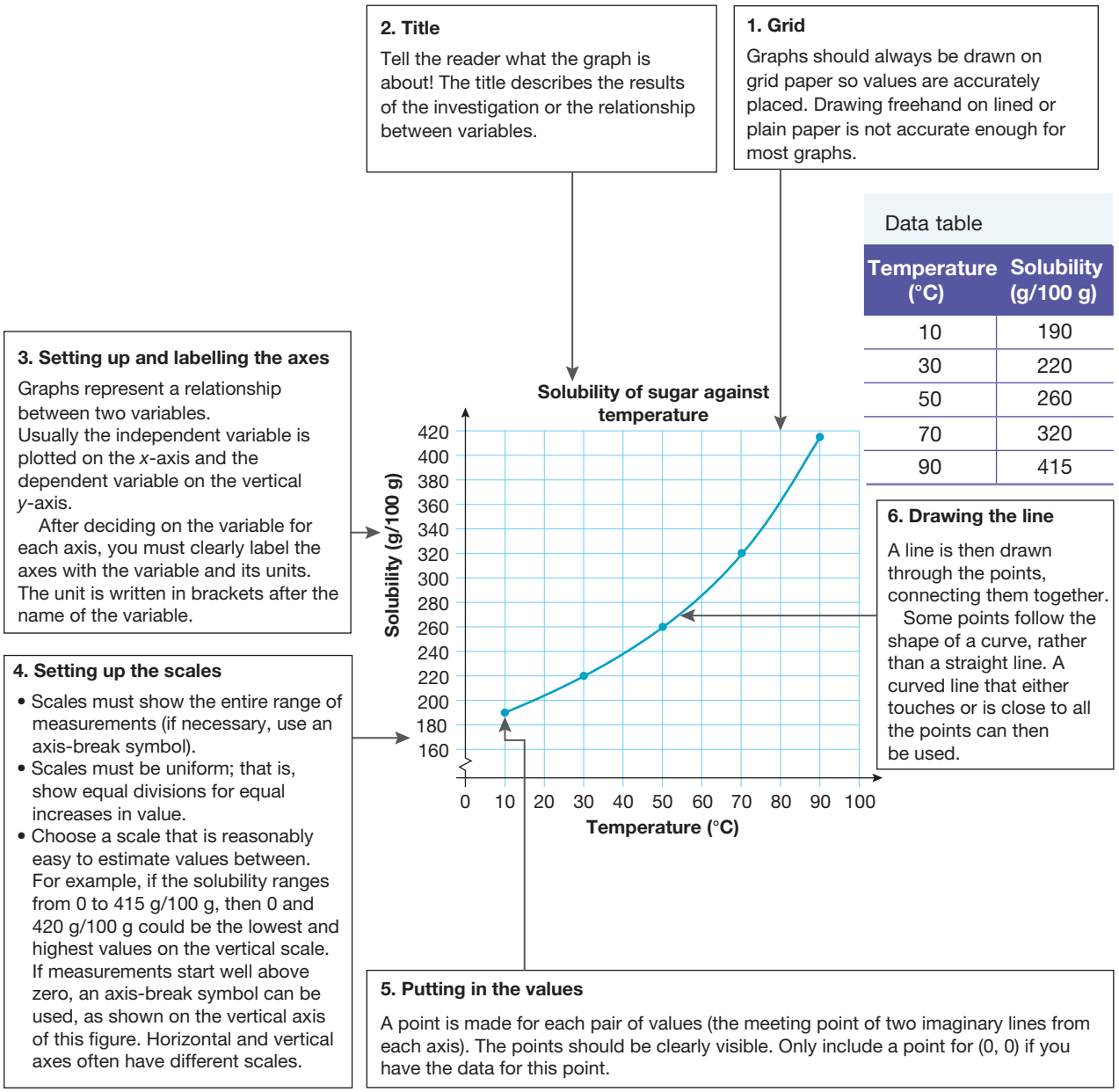
line graph graph in which points of data are joined by a connecting line; used when both pieces of data are quantitative (numerical)

Line graphs

In a **line graph**, a series of dots represents the values of a variable, and the dots are joined using a straight line (this is different from a line of best fit, in which the line is straight and does not have to go through each point).

Line graphs are often used to show changes over a continuous period of time, or over space. In particular, line graphs can identify patterns, trends and turning points in a dataset. Line graphs are sometimes curved rather than being straight point to point.

FIGURE 17.46 Setting up a line graph



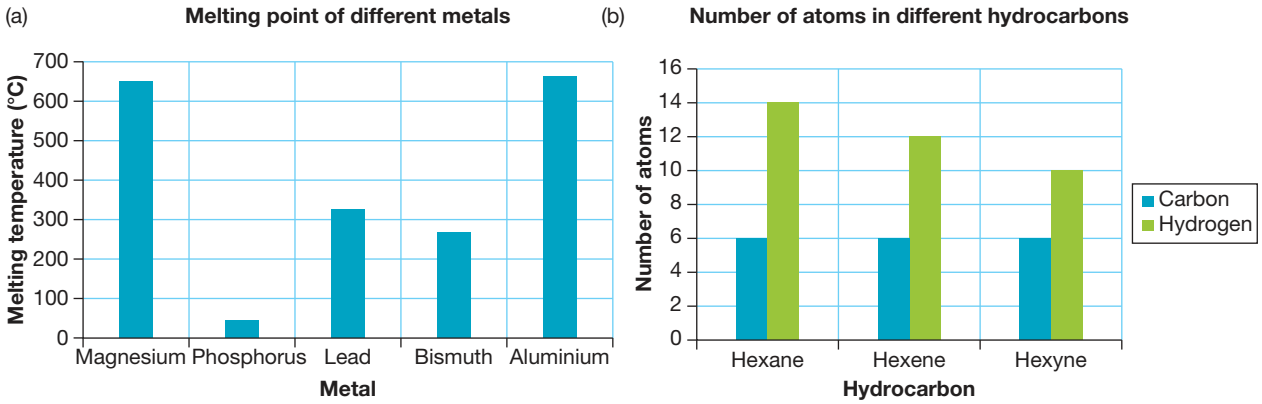
Bar graphs

Bar graphs are often used when one piece of data is qualitative and the other is quantitative. The bars are separated from each other. The horizontal or *x*-axis has no scale because it simply shows categories. The vertical or *y*-axis has a scale showing the units of measurements.

bar graph graph in which data is represented by a series of bars; usually used when one variable is quantitative and the other is qualitative

Bar graphs can also be used to compare two sets of data by using side-by-side bars, as shown in graph (b) in figure 17.47.

FIGURE 17.47 Examples of (a) a bar graph showing differences in melting points of different metals, and (b) a side-by-side bar chart comparing number of hydrogen and carbon atoms

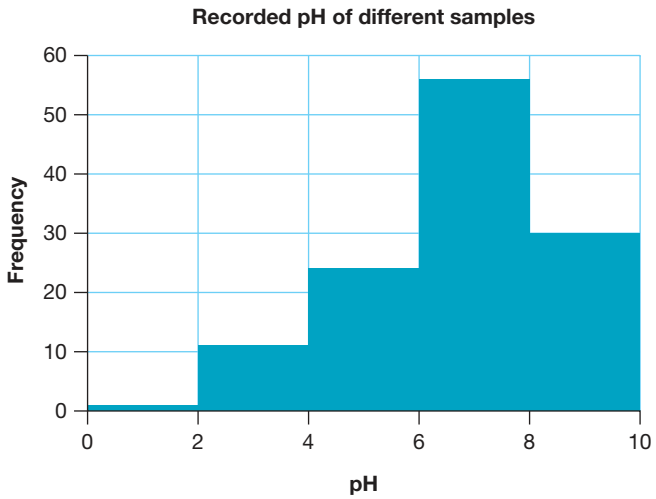


Histograms

A **histogram** is a special kind of bar graph, showing continuous categories. The bars are not separated, unlike in some bar graphs. Histograms are often used when examining frequency.

In figure 17.48, the exact values cannot be determined, because data is displayed in intervals. For example, it can be seen that 30 samples had a pH between 8 and 10. However, we do not know what specific values these are.

FIGURE 17.48 Histogram showing frequency of pH of different samples



Using Microsoft Excel

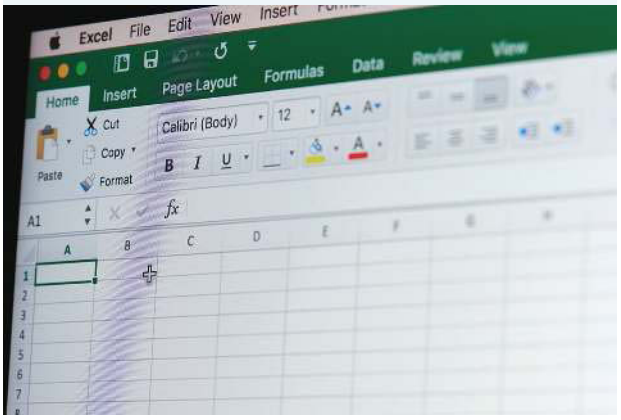
While you may very carefully hand draw your graph, being able to create digital graphs is important, especially for neat presentation on your poster or report. Microsoft Excel is extremely helpful for this. It can:

- store your measurements. Make sure you save your data every few minutes when you are working on it, and back up your computer.
- calculate any derived physical quantities, such as speed and acceleration of a parachute or the percentage of energy lost by a bouncing ball. The 'Fill down' command is a time saver.

histogram graph in which data is sorted in intervals and frequency is examined; used when both pieces are data are quantitative

- be a powerful graphing tool, but must be controlled by the user. You will have to select the graph and choose what aspects of your graph you want to show. For example, what scale on the axis do you want? What do you want to label your axes with? Do you want the data displayed on the graphed points?
- generate a line of best fit. If you right-click on any data point, a window pops up with the option 'Add Trendline'. This is the Excel command to create a line of best fit. Once selected, you have several choices. If your graph looks like a straight line, choose 'Linear'.
- create error bars. However, in Excel all error bars are usually the same for each data point, rather than calculated separately.

FIGURE 17.49 Excel is a useful data tool



on Resources

Interactivity Selecting a graph (int-7733)

17.6.2 Analysing primary data

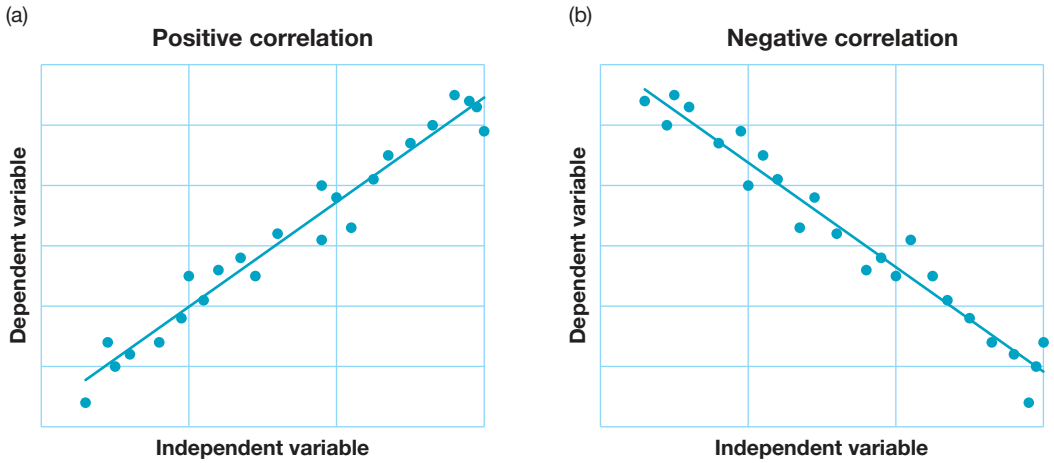
When analysing primary data, it is important to explore trends and patterns that can be seen.

This may include asking questions such as:

- Is there a clear positive or negative correlation in the data?
  - Positive correlation (figure 17.50a) is when one variable increases in response to another increasing variable (for example, increasing the temperature increases the rate of diffusion).
  - Negative correlation (figure 17.50b) is when one variable decreases in response to the other variable increasing (for example, increasing the amount of insulin in blood decreases the blood glucose level).
- Are there any **outlier** (unusual data)?
- What results would you expect for specific data that you didn't observe experimentally?
- Can you calculate the average for your data?

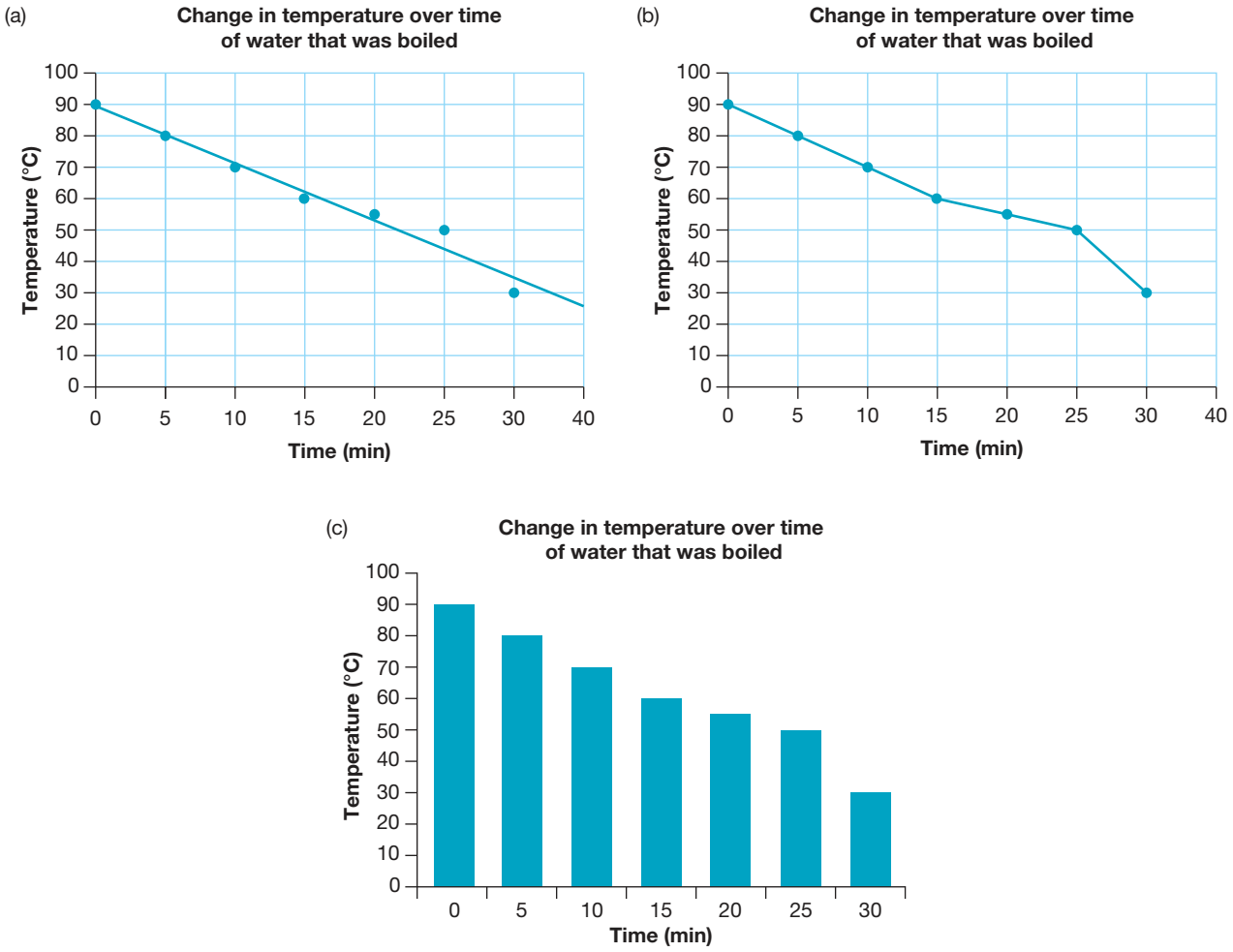
outlier result that is a long way from other results and seen as unusual

FIGURE 17.50 (a) Positive correlation between variables (b) Negative correlation between variables



Analysis of your data often depends on the type of graph selected, because it alters the way that trends and patterns can be seen. For example, the graphs in figure 17.51 show the same data presented in three ways.

FIGURE 17.51 Various graphs showing temperature change over time: (a) scatterplot using a line of best fit, (b) line graph and (c) bar graph



If you were analysing the data from each of the three graphs, the information shown in table 17.9 may be revealed.

TABLE 17.9 Trends observed from three graphs

Graph	Trends observed
Graph (a)	There is a clear downwards trend in data, as temperature decreases over time. However, the data at 25 °C is slightly higher than the expected pattern based on trends, and the data at 30 °C is slightly lower than expected.
Graph (b)	There is a clear downwards trend in data, as temperature decreases over time. The rate of temperature drop slows after 15 minutes, before the rate increases again between 25 and 30 minutes.
Graph (c)	There is a clear downwards trend in data, as temperature decreases over time. The temperature is lowest at 30 minutes, where it is half the temperature seen at 15 minutes.

Graphs (a) and (b) are more powerful representations of the given data compared to graph (c) (A bar graph is not recommended for continuous data.) Regardless of the graph type used, it is important that it shows any clear trends and patterns seen in the data, and that any outliers (unusual data) can be seen.

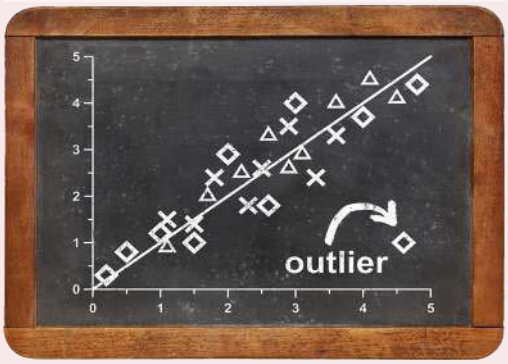
Outliers

Outliers are results that are far removed from other results and seen as unusual.

They should be accounted for and analysed but are not often included when averages are calculated. However, it is important to consider why outliers have occurred as part of your discussion and evaluation of data.

If you had an outlier, what did you do about it? Rather than ignoring it, you should try to account for it. Most commonly, it will be a systematic error, a random error or a personal error in measurement or in calculation, and so can be dropped. Occasionally, an outlier can be a legitimate observation that warrants further investigation.

FIGURE 17.52 Outliers in data are unusual results.



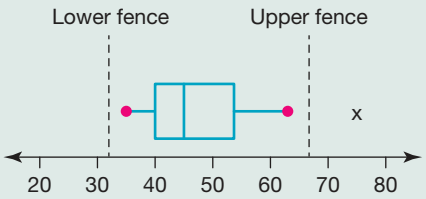
EXTENSION: Calculating outliers

If sufficient data is available, comparing the value to reasonable upper and lower fences may help you classify it as an outlier. To calculate the lower fence and upper fence of the data set, we first need to calculate the interquartile range (IQR). Once this has been calculated, the lower and upper fences are given by the following rules:

$$\text{Lower fence} = Q_1 - 1.5 \times \text{IQR}$$

$$\text{Upper fence} = Q_3 + 1.5 \times \text{IQR}$$

If a data value lies outside the lower or upper fence, it can be considered an outlier.



Interpolation and extrapolation

Graph analysis can also be used to predict and make assumptions about data that was not gathered experimentally. This can be through **interpolation** (predicting data points within the data set that were not measured) or **extrapolation** (predicting data points outside data set based on predicted relationship).

Based on the graphs in figure 17.51 you might estimate:

- the temperature at 45 minutes (extrapolation)
- the temperature at 13 minutes (interpolation)
- the time that the temperature was 45 °C.

The prediction you get may vary greatly between all the graphs, as shown in table 17.10, so it is important to carefully consider which you use.

interpolation estimation of a value within the range of data points tested

extrapolation estimation of a value outside the range of data points tested

TABLE 17.10 Analysing data from different graph types

	Graph (a)	Graph (b)	Graph (c)
Temperature at 45 minutes	20 °C	13 °C	10 °C
Temperature at 13 minutes	65 °C	62 °C	66 °C
Time the temperature was 45 °C	24 minutes	27 minutes	26 minutes

17.6.3 Evaluating primary data

When you evaluate data, it is important to link back to your question of investigation. Evaluating builds on the analysis of data. While analysis is mostly about interpreting the data obtained, evaluating is about determining the significance of data in relation to the investigation question.

How to evaluate primary data

Some of the questions you should explore when evaluating data include:

- Does the data provide an answer to the question of your investigation?
- Does the data support or refute your hypothesis?
- If any outliers, errors or uncertainty were present in your data, why may these have occurred?
- Can your data be linked to different models and theories that are presently supported?
- Could you make further adjustments to improve your data in future investigations that may reduce errors or limitations?

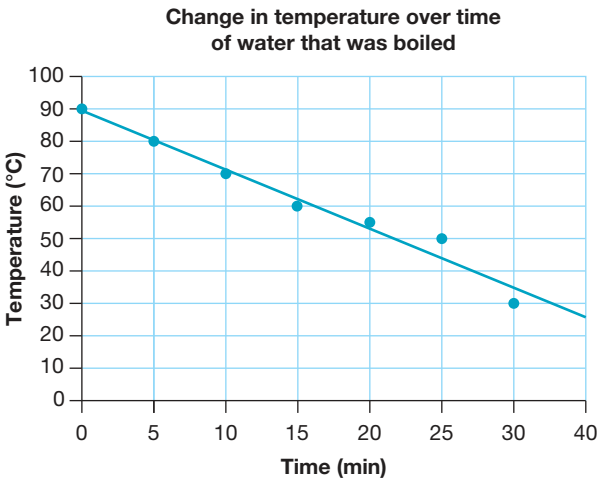
It is important when evaluating data that you can explain and justify this in relation to your question.

Sample evaluation of data

Recall the experiment measuring the change in temperature over time of water that was boiled, discussed previously. Two examples of evaluating the data from this experiment are provided here.

Student 1: It can be seen from the results that temperature decreased over time.

Student 2: From graph (b), as shown in figure 17.51, it is clear that temperature decreases over time, dropping from an initial temperature of 90 ºC to a final recorded temperature of 30 ºC, which supports the theory of heat loss through convection. While there is a clear trend in temperature decreasing over time, the rate of decrease was inconsistent, particularly between 25 and 30 minutes. This may be due to a decrease in external air temperature, causing the rate of heat loss through convection to change.



It is clear that student 2 had a better grasp on evaluating the data, being able to describe trends and a link to theory. What else should they add?

Suggesting improvements

When you are evaluating your primary data, it is important to understand how it is affected by error and uncertainty, and how these are affected by factors such as precision, accuracy and validity.

At this point, you should be able to suggest improvements to increase accuracy and precision and to reduce the likelihood of random and systematic errors. This may involve improvements to:

- the use of equipment
- the number of samples
- the techniques used
- the collection of data.

The improvements you suggest should have an effect on the data you obtain, and should help ensure that another person repeating your investigation can follow these suggestions to improve the accuracy and precision of their result. An improvement should be able to have a quantifiable impact that is focused on minimising *systematic* and *random errors*.

Improvements should not be targeted towards personal errors (for example, 'take more care; ask the teacher for more time'), because these are not errors with the experimental design but mistakes made by an experimenter.

17.6.4 Sources of error and uncertainty

In nearly all investigations, error and uncertainty are very difficult to avoid and can have a significant impact on results. It is important to minimise errors and uncertainty in your investigations.

Error

Errors are differences between a measurement taken and the true value that is expected. These lead to a reduction in the accuracy of the investigation.

Several sources of error can be identified in an investigation:

- **Mistakes** (personal errors) often result from carelessness and should not be included in your final report or analysis and evaluation of data. Rather, the experiment should be repeated correctly.
  An example might be gross misreading of an instrument and writing the wrong result in your logbook, such as 40 instead of 4.0.
- Random errors are chance variations in measurements that affect the precision of measurements and are always present in measurements of continuous data. They can be reduced through repeating measurements and calculating an average. For example:
  - a slight variation in eye height when measuring 1 mL in a pipette leading to a change in apparent position (this is known as parallax error)
  - temperature, wind and position changes on sensitive instruments, such as a top-loading balance
  - judgements when reading the smallest division on the scale of a measuring instrument, such as a ruler.
- **Systematic errors** are errors that affect the accuracy of a measurement that cannot be improved by repeating an experiment. Any error that causes data to differ from the true value in the same way each time (consistently too high or too low) is a systematic error. They are usually due to equipment limitations, incorrect calibration or inappropriate methodology. For example:
  - An instrument, such as a weighing balance, is uncalibrated and incorrectly set to zero, causing all measurements to be slightly too high.
  - The scale on a ruler is slightly off, and every 1 cm on the ruler is actually 1.1 cm. Therefore, all measurements will be slightly lower than they actually are.
- Parallax errors are both systematic and random errors. They are considered systematic because they are due to a limitation of the equipment and so are unavoidable. They are also considered random errors because of their effect on the data — they will cause data to be sometimes higher than the true value and sometimes lower.

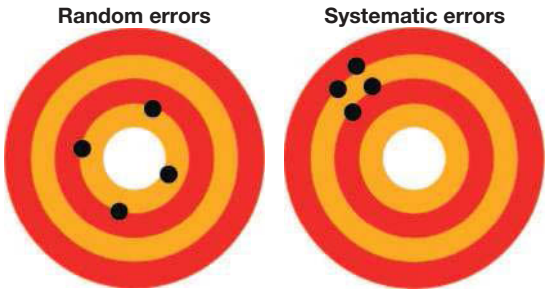
FIGURE 17.53 Measurements of liquids should be read at eye level from the bottom of the meniscus to minimise errors.

mistakes human errors or personal errors that can impact results, but should not be included in analysis; instead, the experiment should be repeated correctly

systematic errors errors that affect the accuracy of a measurement that cannot be improved by repeating an experiment; usually due to equipment or system errors

Both systematic and random errors are known as measurement errors. These need to be discussed in your evaluation of results and accounted for. As shown in figure 17.54, systematic errors mostly affect the accuracy of the data obtained, whereas random errors affect the precision of data.

FIGURE 17.54 Comparing the effect systematic and random errors have on results



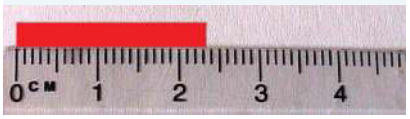
Uncertainty

A degree of uncertainty exists in any physical measurement. The uncertainty is a range within which the true measurement lies. The uncertainty can be due to human error or to the limitations of the measuring instrument. In most physical measurements, the last significant figure shows a small degree of uncertainty. For example, the length of an Olympic competition swimming pool is correctly expressed as 50.00 m. The last zero has a small degree of uncertainty. An error bar is a way of representing that uncertainty graphically.

Reading rulers

A metre ruler has lines to mark each millimetre, but there is space between these lines. You could measure a length to the nearest millimetre, but because of the space between the lines, if you look carefully, you can measure to a higher precision. You can measure to the nearest 0.5 mm.

FIGURE 17.55 Errors are also possible when using rulers



The best estimate for the length of the red line shown in figure 17.55 is 2.35 cm. The actual length is closer to 2.35 cm than it is to either 2.30 cm or 2.40 cm. The measurement of 2.35 cm says the actual length is somewhere between 2.325 cm and 2.375 cm.

The way to write this is:

$$\text{The length of the red line} = 2.35 \pm 0.025 \text{ cm}$$

The 0.025 represents the resolution or uncertainty in the measurement (half of the precision that can be recorded).

Reading top-loading balances

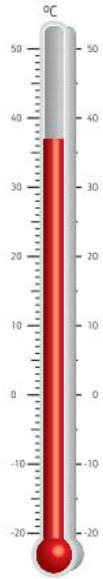
Top-loading balances measure to a certain number of significant figures. Say the reading on a digital scale is 8.94 grams. This means the mass is not 8.93 g or 8.95 g. The actual mass is somewhere between 8.935 and 8.945 grams. In this example, the smallest unit of measurement is 0.01 g. Therefore, the tolerance is half of this (0.005 g), because the measurement can be 0.005 g below or above the recorded measurement. The way to write this is:

$$\text{Mass} = 8.94 \pm 0.005 \text{ g}.$$

FIGURE 17.56 Top-loading balance

FIGURE 17.57 What is the uncertainty in these measuring instruments?

(a) (b)



tlvd-0104

SAMPLE PROBLEM 4 Recording readings on measurement instruments with uncertainty

a. Record the reading on the scales in figure 17.57(a), including the uncertainty due to the resolution (tolerance) of the instrument.

b. Record the reading on the thermometer in figure 17.57(b), including the uncertainty due to the resolution (tolerance) of the instrument.

THINK	WRITE
a. 1. Determine the reading on the scale.	128.93
2. Considering the rounding of the true measurement, determine the realistic range of the measurement.	The measurement can be between 128.925 and 128.935
3. Considering the instrument resolution, determine the uncertainty by halving the smallest unit that can be measured.	The smallest measurement possible is 0.01 g. $\frac{0.01\text{ g}}{2} = 0.005\text{ g}$
4. Record the reading, including the uncertainty. (Notice how this matches the realistic range determined above.)	128.93 ± 0.005 g
b. 1. Determine the reading on the scale.	47 °C
2. Determine the range of the measurement. In this case, you can see a difference of just 0.5°C due to the gaps in the thermometer, so our range above and below the read measurement is less than 0.5.	The measurement can be between 46.75 and 47.25.
3. Considering the instrument resolution, determine the uncertainty by halving of the smallest unit that can be measured.	The smallest unit that can be measured is 0.5°C. $\frac{0.5\text{ °C}}{2} = 0.25\text{ °C}$
4. Record the reading, including the uncertainty.	47 ± 0.25 °C

PRACTICE PROBLEM 4

a. Record the reading on scales, including the uncertainty, that show a reading of 0.12 grams.
b. Record the reading on scales, including the uncertainty, that show a reading of 0.195 grams.

17.6 Activities

learnon

Students, these questions are even better in jacPLUS

- Receive immediate feedback and access sample responses
- Access additional questions
- Track your results and progress

Find all this and MORE in jacPLUS

17.6 Quick quiz on | 17.6 Exercise | 17.6 Exam questions

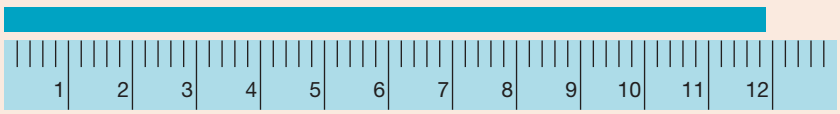
17.6 Exercise

1. Describe the difference between a random and systematic error and provide two examples of each.
2. List two ways that you can minimise uncertainty in an investigation.
3. Determine the length of each line in the following diagram, showing the uncertainty in each case.

(a)

(b)

(c)



4. Identify which graph type would be most appropriate for the following investigations, and justify your choice.
   a. Comparing the pH of different household liquids
   b. Showing how pH changes with temperature
   c. Measuring the temperature inside a car every five minutes for an hour
   d. Showing the frequencies of different test mark intervals for 400 students

5. Using the data provided in the following table, construct an appropriate graph.

Time (mins)	Volume of ice cube (cm^3)
0	30
5	25
10	21
15	18
20	14
25	10

Once you have constructed the graph, analyse and evaluate the data shown.

6. Describe how an outlier should be treated in analysing and evaluating data.
7. A student is designing an experiment that involves measuring liquid volumes at various stages. A number of glassware items are available for this purpose, as shown in the following table.

Item	Uncertainty (mL)
Volumetric flask (250 mL)	±0.1
Beaker (100 mL)	±10
Measuring cylinder (100 mL)	±0.1
Burette (50 mL)	±0.02
Conical flask (250 mL)	±25

Which piece(s) would be most appropriate for the following?

a. Rinsing a burette with 10 mL of water
b. Producing 250 mL of a solution of accurately known concentration
c. Producing 250 mL of a solution of approximately known concentration
d. Adding water to a conical flask to dissolve a previously weighed tablet, prior to titration

17.6 Exam questions

Question 1 (1 mark)

MC Which one of the following statements about conducting an experiment is the most correct?

A. Precise results may be biased.
B. Accuracy is assured if sensitive instruments are used.
C. A method is valid if it identifies all controlled variables.
D. Repeating a procedure will remove the uncertainty of the results.

Question 2 (1 mark)

A student designed an experiment to investigate current efficiency during the electrolysis of a sodium chloride, NaCl, solution.

This reaction is modelled by the equation:

$2Na^+(aq) + 2Cl^-(aq) + 2H_2O(l) + \rightarrow H_2(g) + Cl_2(g) + 2NaOH(aq)$

The results for part 1 of the procedure are provided here.

Part 1 — Distilled water

Trial	Initial volume (mL)		Final volume (mL)		Current (A)
	Negative electrode	Positive electrode	Negative electrode	Positive electrode	
1	170.0	170.0	100.2	135.3	2.0
2	170.0	170.0	100.3	135.3	2.0
3	170.0	170.0	99.9	135.0	2.0
4	170.0	170.0	99.8	134.8	2.0
5	170.0	170.0	100.1	135.1	2.0
Average	**170.0**	**170.0**	**100.1**	**135.1**	**2.0**

Are the results in part 1 precise? Justify your answer.

Question 3 (2 marks)

Consider the following data.

Electrical conductivity of some metals

Material	Electrical conductivity (S m^{-1})
Silver	62.1
Copper	58.7
Gold	44.2
Aluminium	37
Platinum	90

a. Identify which graph type would be most appropriate to represent this data. **(1 mark)**
b. Can you recommend any changes or improvements for this table? **(1 mark)**

Question 4 (8 marks)

A student conducted an experiment examining the relationship between temperature and the pH of water. The results from this investigation are shown in the table.

Temperature (°C)	pH
0	7.7
10	7.5
20	7.3
30	7.0
40	6.7
50	6.4

a. Complete the following:
- Plot the graph showing the data provided, ensuring that pH is shown on the vertical axis.
- Label the graph and axes as appropriate.
- Draw a line showing the relationship between the points. **(4 marks)**

b. Describe the trends seen in your data. **(2 marks)**
c. Based on the graph, explain if you could accurately determine the pH of the solution at 90 °C. **(2 marks)**

More exam questions are available in your learnON title.

17.7 Challenging scientific models and theories

KEY KNOWLEDGE

- Observations and investigations that are consistent with, or challenge, current scientific models and theories

KEY SCIENCE SKILLS

- Construct evidence-based arguments and draw conclusions
- Analyse, evaluate and communicate scientific ideas

Source: VCE Chemistry Study Design (2023–2027) extracts © VCAA; reproduced by permission.

17.7.1 Using strong evidence

In order to support or refute a hypothesis, model or theory, it is important to use strong evidence. Strong evidence usually comes from investigations and the collection of data.

Investigations that have strong evidence include the following features:

- a basis in facts derived from studies with high validity and minimal bias
- statistical evidence to support conclusions
- a clear distinction between **correlation** and **causation** — two variables may often have some correlation (they both increase, for example), but have no causation (one variable does not cause the change in value in the other)
- data from investigations that have a repeatable and reproducible method — which include those that are **randomised** and have a control group
- peer-reviewed research formed from scientific ideas.

17.7.2 Supporting and refuting hypotheses

As part of your analysis and evaluation of data and constructing your conclusion, you should be addressing your hypotheses.

If the prediction from your hypothesis was validated by your experimental results, you should say that your hypothesis is 'supported'. If your hypothesis was not supported by results, you say that it is 'refuted' (or rejected).

It is important to remember that we do not say a hypothesis is proven. This is because as new technologies and information become available, evidence can change and be interpreted in different ways. This may then disprove a previously supported hypothesis. We support a hypothesis based on the information we have available at that time.

A good example of this is to consider the evolution of the atomic model. J.J. Thompson proposed the plum pudding model followed by Ernest Rutherford's model of electrons circling around a positive centre. The evidence available at that time supported these models. Later, Bohr applied quantum physics to theorise the existence of shells in which the electrons move. This then became the widely accepted view, until Schrodinger employed quantum mechanics to propose the existence of electron clouds. At each discovery, as techniques and technology advanced, the model of the atom changed.

correlation measure of a relationship between two or more variables

causation when one factor or variable directly influences the results of another factor or variable

randomised assignment of individuals to an experiment or control group at random, not influenced by external factors

So while nothing was wrong with supporting the hypothesis at the time because evidence supported it, we cannot prove a hypothesis as 100 per cent correct or valid.

FIGURE 17.58 Strong evidence involves many facets.

TABLE 17.11 Simple examples of results that can support or disprove a hypothesis

Observation	Hypothesis	Test of hypothesis	Result	Conclusion
The phone is not charging.	The charger is faulty.	Use a different charger.	The phone still does not charge.	Hypothesis rejected
The car won't start.	The battery is flat.	Replace the battery.	The car still won't start.	Hypothesis rejected
The measurement on the scales is far higher than expected.	The scales were not set to zero before use.	Reset the scales back to zero.	The reading is closer to the expected value.	Hypothesis supported
The water did not freeze at 0 °C.	There was a contaminant in the water.	Freeze water that is pure (distilled).	The water froze at 0 °C.	Hypothesis supported

17.7.3 Supporting and refuting models and theories

As part of your investigation, you may be exploring various models and theories, and finding evidence that may refute or support them.

Science involves both being able to challenge current scientific theories and models, or gain results that are consistent with and support current models and theories. Scientific investigation is always being developed in this way.

Recall that:

- models are representations of ideas, phenomena or scientific processes
- theories are well-supported explanation of phenomena; they are based on facts that have been obtained through investigations, research and observations.

How do we support or refute models and theories?

Similar to supporting or refuting hypotheses, we need to use strong evidence to support or refute models and theories. The strongest evidence is your first-hand evidence that was obtained through scientific investigation using accurate, precise and valid methods.

In your investigation, you should research different models and theories relating to your topic. Through collecting primary data, as part of your discussion, you should state if your evidence challenges or is consistent with models and theories, providing a reason why this may have occurred. This is exactly the way that models and theories have been supported or disputed in the past.

17.7 Activities

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17.7 Quick quiz on | 17.7 Exercise | 17.7 Exam questions

17.7 Exercise

1. Why do we say that we 'support' a hypothesis rather than 'prove' it?
2. Provide three examples of strong evidence.
3. Why is it important to show the relationship between your results and concepts such as the kinetic theory of matter or the ideal gas law?
4. Why is it important in science that models and theories are constantly being challenged and revised?

17.7 Exam questions

Question 1 (1 mark)

MC Which of the following does not provide strong evidence that can be used to support models and theories?

A. Minimal bias and high validity
B. Supporting research from journals that are not peer reviewed
C. Based on scientific evidence
D. Reproductible and reliable methods

Question 2 (1 mark)

MC A group of students conducted an experiment to investigate the effect temperature has on the pressure of an airtight container. To which of the following theories or laws must the students compare their results?

A. Brønsted-Lowry theory
B. Ideal gas law
C. VESPR theory
D. Atomic theory

Question 3 (2 marks)

Explain how a model differs from a theory.

Question 4 (3 marks)

A student is exploring the theory that 'substances with a smaller surface area to volume allow for a greater rate of reaction'.

a. Outline two pieces of evidence that would be consistent with this theory. **(1 mark)**
b. Outline two pieces of evidence that would challenge this theory. **(1 mark)**
c. If you had results that challenged the theory, does that mean the theory could be wrong? Justify your response. **(1 mark)**

More exam questions are available in your learnON title.

17.8 The limitations of investigation methodology and conclusions

KEY KNOWLEDGE

- The limitations of investigation methodologies and methods, and of data generation and/or analysis

KEY SCIENCE SKILLS

- Analyse, evaluate and communicate scientific ideas
- Construct evidence-based arguments and draw conclusions

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When we conduct investigations, it is important to consider that we often need to make many **assumptions** that allow us to work around limitations in:

- scientific methodology
- method (both in investigation and analysis)
- data and data generation
- models and theories
- conclusions.

Limitations are factors that affect the interpretation and/or collection of findings in a practical investigation.

You need to consider limitations when analysing data, methodology, methods, model, theories and drawing conclusions. Limitations should be factors that are out of your control, but should be discussed in regards to how they might affect your results.

17.8.1 Limitations and assumptions in scientific methodology and methods

In the scientific methodology you select, a variety of limitations may exist. This may affect your designed method.

assumptions ideas that are accepted as true without evidence in order to overcome limitations in experiments

FIGURE 17.59 Limitations of your method depend on the equipment available. For example, if you have access to a micropipette, your measurements will be more accurate. However, if you only have access to a 1 mL pipette, limitations exist on how accurate your measurements will be.

As you are conducting your investigation, it is important to be aware of limitations that may affect your methodology. Examples include:

- not having access to equipment that limits issues with errors, uncertainty, accuracy and precision (for example, not having access to micropipettes to collect accurate quantities)
- not being able to properly control variables to a high level (such as room temperature, contaminants or humidity)
- not having adequate time to observe results (for example, it may take five weeks to obtain the best results, or you may need to check results every 20 minutes over 10 hours, which may not be possible in a school environment)
- not having the opportunity to repeat investigations and show both repeatability and reproducibility
- difficulty creating a hypothesis that is both testable and falsifiable, which will limit how a method can be designed
- limitations to access of data (for example, if conducting a case study or correlational study)
- when analysing data, not including factors such as error bars or statistical hypotheses (for example, chi-squared tests and standard deviation exploration, which are well beyond the scope of what is expected in your task).

Therefore, assumptions are often made about the methods used and the validity of results. The following is usually assumed.

- All variables (other than the independent variable) are completely controlled.
- The equipment used provides precise and accurate data.
- Repetitions of the investigation will result in similar or identical results, except where results involved obvious human error or clear outliers.

Making these assumptions is fine in the scope of your task. However, it is important that you comment on these limitations in your logbook and scientific report or poster and suggest possible improvements to the method and methodology used that could lead to an investigation with improved validity, precision and accuracy.

17.8.2 Limitations and assumptions in data

The data that is gathered from the experimental results will have limitations. Limitations arise from several sources that can affect the quality of the data, such as the following:

- Experiments create artificial situations that do not necessarily represent real-life situations.
- Although every effort may be made to identify controlled variables and keep them constant throughout the course of an experiment, it is not always possible to identify and control every variable.
- The degree to which results obtained in the laboratory can be generalised to other situations and applied in the real world is limited.

We need to make assumptions that the data was obtained in a way that best reflects real-life scenarios and that variables were all controlled.

17.8.3 Limitations in models and theories

As well as data limitations, limitations exist in the models and theories that we use. It is important to be able to use models and theories to allow us to understand a variety of observed phenomena, but it is also important to understand the limitations of our models and to consider that theories may change as more observation and research is conducted.

When linking your primary data to established models and theories, you need to take these considerations into account and outline why these limitations may have caused differences between your findings and the expected results of theories and models.

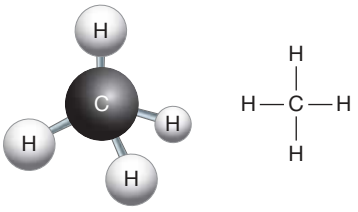
Models

Models are representations of ideas, phenomena or scientific processes. Although they are very useful, models have their limitations.

Some of the limitations of models are as follows:

- *Missing details:* because of the complexity of the processes, models cannot include all the details of the processes or the things that they represent. For instance, we often rely on the periodic table as an ideal model of elements and how they behave; however, the group 1 & 2 metals differ significantly in melting points and the metalloids do not follow the predicted properties in their periodic table position.
- *Approximation:* models are not necessarily approximations of the real world; for example, when we draw the subatomic particles in the atom, we represent them as the same size but, in reality, the electrons are much, much, much smaller than protons and neutrons. Also, many models in chemistry are based on having SLC (25°C and 1 atm) when, in reality, this would rarely be the case.

FIGURE 17.60 Different models of methane. How do you think these differ from a 'real' methane molecule?



CH_4 methane

- *Simplicity:* models often have limits in their accuracy and are often simplified and stylised; for example, a ball-and-stick model of methane (see figure 17.60) is useful but is a very highly simplified and stylised representation that reduces bonds to sticks and atoms to solid balls.
- *Subject to change:* models are based on current observations and knowledge at the time. This means that they aren't definite and can change as observations allow for different ideas to come to light. Climate models are an example of models that have changed drastically in the past decades.

Theories

A theory is a well-supported explanation of a phenomena. It is based on facts that have been obtained through investigations, research and observations.

Theories can have limitations. These may include the following:

- *Over-reliance on theories:* during practical investigations, some individuals rely on theories rather than observations, or might be tempted to ignore observations that contradict existing theories.
- *Imperfection:* a theory is the best explanation to date of an observed aspect of the natural world. As we have seen in subtopic 17.7, theories can be disproven and can be improved as observations and evidence come to light.
- *Boundary conditions:* theories often rely on a very specific set of conditions to be met. For example, in chemistry, Standard Laboratory Conditions (SLC) may be required, or in organic synthesis, a ketone is produced from a secondary alcohol in the presence of acidified manganate, MnO_4/H^+.
- *Choice of theory:* often, a phenomenon needs to be described using multiple theories. Sometimes, aspects of theories may be complementary, or they might contradict each other, such as the valence bond theory and molecular orbital theory to explain aspects of bonding.

Many individuals rely on theories rather than observations. Often, theories are treated as perfect when, in fact, they are able to be disproved as observations and evidence come to light. It is important to consider the limitations outlined here and understand that theories are not absolute.

17.8.4 Limitations in conclusions

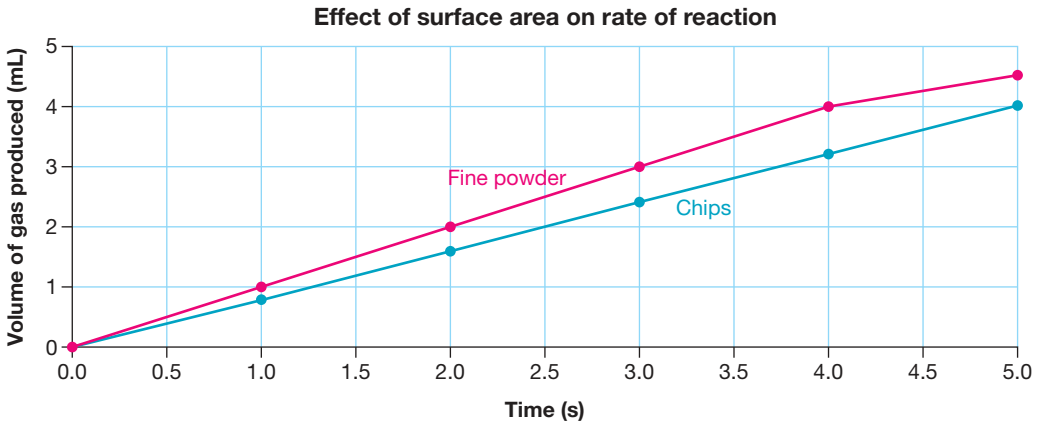
When drawing conclusions, it is important to consider their limitations. Limitations that affect the methodology, method, data, models and theories of an investigation will also affect the conclusions drawn.

As part of your investigation and critical evaluation, you need to be able to identify, describe and explain the limitations of your conclusions, including identification of further evidence required. This should be discussed clearly within the discussion section of your report or poster.

Examples of limitations in conclusions

You may be exploring the effect surface area has on the rate of reaction when calcium carbonate chips versus fine powder react with hydrochloric acid by observing the volume of carbon dioxide produced over time.

FIGURE 17.61 The effect of surface area on rate of reaction



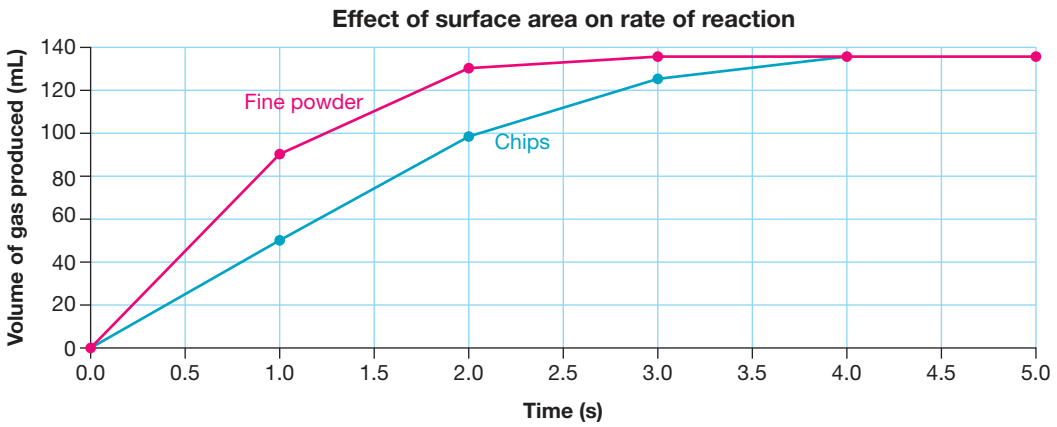
As shown in figure 17.61, the rate of reaction for the larger and smaller surface area looks almost the same for this duration of time (5 seconds). Looking at this data, you may draw the conclusion that surface area does not significantly affect the rate of the reaction. You may also conclude that a reaction processes at a consistent (linear) rate over time. This may seem like a fair assumption and conclusion based on the primary data collected, but it is important to consider limitations of this. How accurate is this trend line?

Extending the range of data collection (to five minutes) might paint a completely different picture, as shown in figure 17.62.

Here, it can be seen that the difference in surface area has caused the reaction with fine powder to proceed much faster, with the reaction completed in close to half of the time of the reaction with the chips. This wider range of data also shows that the reactions slowed down before they finished, and formed a curved shape rather than linear one, as it appeared in the previous figure.

This example shows limitations in the conclusions drawn, as you assume that the data obtained tells the entire story. It is important that in your reporting (or poster), you identify and describe further evidence that is required and how you would obtain this further evidence. In the preceding investigation, you may describe further evidence as exploring a greater range of data (such as 0–5 minutes instead of 0–5 seconds).

FIGURE 17.62 The effect of surface area on rate of reaction over 5 minutes



17.8 Activities

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17.8 Quick quiz on | **17.8 Exercise** | **17.8 Exam questions**

17.8 Exercise

1. Describe three limitations of models.
2. When analysing and evaluating data, you should describe further evidence that would be required. Why is this important, and how does this address the limitations of the investigation?
3. Why might it be difficult to control variables such as temperature when conducting investigations?
4. Outline limitations that can occur in data collection and generation. Provide three examples.

17.8 Exam questions

Question 1 (1 mark)

MC Which of the following is not a limitation of models?

A. They are oversimplifications of concepts.
B. They may not reflect real-world conditions.
C. They allow abstract concepts to be more tangible.
D. They are usually specific to certain conditions.

Question 2 (2 marks)

Scientists make conclusions based on the evidence that they have collected in an experiment. Explain whether a conclusion is ever changed by a scientist.

Question 3 (6 marks)

A student conducted an investigation to explore the different boiling points of organic molecules containing three carbons.

The student's report is provided.

Introduction: *In this experiment, various three-carbon organic compounds were investigated. The substances being explored were propanoic acid, propan-1-ol, propan-2-ol, propanone and propanal. This investigation was conducted in the laboratory and all substances were slowly heated to a maximum temperature of 100 °C and the boiling point was recorded. Propane was not included in this investigation because it was already a gas at room temperature.*

Aim: *To examine if different carbon-containing compounds have different boiling points.*

Hypothesis: If various three-carbon containing organic molecules are heated, then propanoic acid will have the highest boiling point.

Method:

1. *Set up five beakers. Label each with the organic compound.*
2. *Pour the organic compound into the appropriate beaker.*
3. *Place a thermometer in the liquid and record the initial temperature.*
4. *Place the first beaker above the Bunsen burner, using a tripod and gauze mat.*
5. *Heat the liquid to 100 °C, recording the temperature of boiling.*
6. *Repeat with the next liquid.*

	Initial temperature	Boiling point
Propanoic acid	20.0 °C	>100 °C
Propan-1-ol	25.0 °C	95.0 °C
Propan-2-ol	24.0 °C	83.5 °C
Propanone	23.5 °C	59.0 °C
Propanal	23.5 °C	49.5 °C

a. In the experiment, the substances were only heated to a maximum of 100 °C. Explain why this may lead to errors in the data obtained. **(2 marks)**

b. Students were only able to use a thermometer in which temperature could only be measured to the nearest 0.5 °C. Identify the tolerance of this device and describe the uncertainty expected in the data. **(2 marks)**

c. Outline two limitations in the experimental method or data collection process that would affect the conclusions drawn. **(2 marks)**

Question 4 (3 marks)

A student investigated the effect of different catalysts on the energy released from the decomposition reaction of hydrogen peroxide.

Report — Effect of different catalysts on the enthalpy of a reaction

Background

Different catalysts, such as manganese dioxide, MnO_2, and iron(III) nitrate solution, $Fe(NO_3)_3$, will increase the rate of decomposition of hydrogen peroxide.

$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$

Purpose

This experiment investigated the effect of using different catalysts on the molar enthalpy of the decomposition of hydrogen peroxide.

Procedure

The temperature change was measured when MnO_2 catalyst was added to a volume of hydrogen peroxide in a beaker. The procedure was repeated using $Fe(NO_3)_3$ solution as a catalyst.

Results

	Trial 1	Trial 2
Volume H_2O_2	100 mL	200 mL
Concentration H_2O_2	2.0 M	4.0 M
Catalyst	0.5 g MnO_2	50 mL 0.1 M $Fe(NO_3)_3$
Temperature change °C	3.0	10.1

Conclusion

The change in temperature using the $Fe(NO_3)_3$ catalyst was greater than the change in temperature using the MnO_2 catalyst. This demonstrates that the molar enthalpy for the decomposition reaction depends on the catalyst used.

The student's conclusion is not valid because the experimental design is flawed.

Critically review the student's experimental design. In your response, you should identify and explain three improvements or modifications that you would make to the experimental design.

More exam questions are available in your learnON title.

17.9 Options related to analysing substances in water, acid–base or redox reactions and production of gases

KEY KNOWLEDGE

- Options related to the production of gases, acid–base or redox reactions, or the analysis of substances in water

KEY SCIENCE SKILLS

- Plan and undertake investigations
- Conduct investigations to collect and record data
- Analyse and evaluate data, methods and scientific models
- Communicate and explain scientific ideas

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Options, strategies and solutions regarding production of gases, acid–base or redox reactions, and analysis of substances in water were examined in topics 10 through 16, and these should be explored and reviewed to assist with your investigation. For your investigation, it is vital to be aware of limitations with the technique or techniques that you select, and take appropriate measures to reduce uncertainties and errors that may occur and which will affect the accuracy and precision of your results.

17.9.1 Key factors when analysing substances in water

As examined in Unit 2, water is the most widely used solvent on Earth, vital for survival and life as we know it. As part of Unit 2, Area of Study 3, one of the choices for your investigation involves exploring substances in water.

The water that we drink and use is not pure, and contains many contaminants. Some of these can adversely affect our health. Therefore, regularly examining water quality to make sure that it is safe for consumption is vital.

Water may be contaminated by substances such as:

- heavy metals (including lead, chromium and arsenic)
- organic pollutants (including fertilisers, pesticides and pharmaceuticals)
- microorganisms (including bacteria, algae and parasites).

It is important to monitor water for factors such as:

- salts (refer to topic 16)
- acids and bases (refer to topic 11).

Some key factors to remember when analysing substances in water for your practical investigation are as follows:

- Avoid sampling bias when collecting water samples, through making sure that the sample is representative of all water in the area. For example, if you collect water only from the surface, the substances may be quite different to those further down. Try to collect multiple samples or make sure that the sample is well mixed (below 100 mm from the surface).
- Make sure that your water can be stored over multiple days of the experiment. If you have to continue to collect new samples, the quality of the water may have changed.
- Use repetition to improve the accuracy and precision of your investigation.
- Make sure you choose the most appropriate technique to sample and analyse your water. Your choice will often depend on the properties of the material under investigation. Consider time and cost effectiveness, safety and the availability of resources in your school. Some examples of possible analysis techniques are shown in table 17.12.

TABLE 17.12 Water analysis techniques

	Techniques	Notes on technique
Analysing salts	Gravimetric analysis	Useful for metal ions and anions
	Instrumental colorimetry	Useful to determine concentration by detecting light
	UV–visible spectrometry	Suitable for coloured solutions or those that absorb in the ultraviolet region
	High-performance liquid chromatography (HPLC)	Highly sensitive technique*

*Your school is unlikely to have access to HPLC. Access to specialised equipment should be determined before choosingthis technique.

You must be aware of limitations with the technique or techniques that you select, and take appropriate measures to reduce uncertainties and errors that may occur and which will affect the accuracy and precision of your results.

Finally, as part of your investigation, considering strategies and implications of your research on the analysis of substances in water is vital. What are some ways that we can improve water quality? How might we be able to ensure more individuals have access to clean drinking water free of contaminants? These aspects are an important part of Area of Study 3, and looking at them will allow you to receive the best results possible.

17.9.2 Key factors when analysing acids–base or redox reactions

Some key factors to remember when conducting your practical investigation in acid–base or redox reactions are as follows.

- When analysing acid–base reactions, careful selection of the indicator over the desired pH range is vital. Additionally, using indicator paper or universal indicator to measure pH are less precise strategies compared to using a pH meter.
- When investigating redox reactions, ensure that your choice of chemicals (for example, for the reactivity series of metals, metal displacement reactions or a simple cell) will provide you with enough data. When studying redox reactions you may compare three or four reactions for your investigation to generate a sufficient data set to work with.
- Use repetition to improve the accuracy and precision of your investigation.
- Make sure you choose the most appropriate technique for your acid–base or redox reaction. Your choice will often depend on the properties of the material under investigation. Consider time and cost effectiveness, safety and the availability of resources in your school. Some examples of possible analysis techniques are shown in table 17.13.

TABLE 17.13 Acid–base and redox analysis techniques

	Techniques	Notes on technique
Analysing acids and bases	Acid–base titrations	Measures concentration based on neutralisation reactions
	Indicators	Quick and easy technique to measure pH, but offers reduced precision
	Potentiometric titration	Uses two electrodes and measures the changes in voltage
	pH meters	Easy for quick measurements
Analysing redox reactions	Metal displacement reactions	Easy and quick but need to do several for sound data set
	Simple cell	Repeatability of data can be difficult

17.9.3 Key factors with production of gases

Some key factors to remember when conducting your practical investigation in production of gases are as follows:

- When investigating volume of a gas, ensure that all possible leaks have been sealed; otherwise, calculations will be significantly affected.
- When analysing the production of gases — for example, in the combustion of a fuel — try to ensure that heat loss to the surroundings in the experimentation is minimised. In this case, this can be done by using a heat shield and by covering the equipment with foil.
- Use repetition to improve the accuracy and precision of your investigation.
- Your choice of experiment will often depend on the properties of the material under investigation. Consider time and cost effectiveness, safety and the availability of resources in your school. Some examples of possible analysis techniques are shown in table 17.14.

TABLE 17.14 Analysis techniques for production of gases

	Techniques	Notes on technique
Analysing production of gases	Combustion of a fuel	Set-up takes time; need to minimise heat loss for accuracy of results
	Direct gas volume measurement (e.g. metal + acid reaction)	Set-up easy but if not sealed correctly can lose gas through leakage

17.9 Activities

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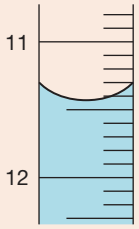
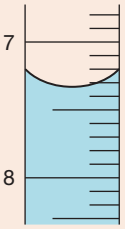
17.9 Exercise

1. Why is it important to test water?
2. Provide some examples of the following.
   a. Heavy metals that may contaminate water
   b. Microorganisms that may contaminate water
3. Would it matter what technique you used to detect salts or would all techniques lead to the same results?
4. Record the reading on the burette in the following diagram (remember that the measurements on a burette go downwards, unlike in a measuring cylinder).



5. Explain why a student reading this burette as 8.67 mL is incorrect.



17.9 Exam questions

Question 1 (1 mark)

The gravimetric analysis procedure requires the final heating, cooling and weighing step to be repeated until 'constant mass is achieved'. How might the final result be affected if this is not done?

Question 2 (9 marks)

A student conducted an experiment to determine the concentration of barium chloride, $BaCl_2$, dissolved in a sample of water via gravimetric analysis.

Excess potassium carbonate, K_2CO_3, solution was added to 500 mL of the water sample to form a precipitate. This precipitate was then collected, dried and weighed to constant mass.

Following this procedure, one particular sample produced 0.463 g of precipitate.

a. Write a balanced equation for the reaction that leads to the formation of the precipitate. **(1 mark)**
b. Calculate the mass of barium chloride in this particular sample. **(3 marks)**
c. Hence, calculate the concentration of barium chloride in this sample, in g L^{-1}. **(1 mark)**
d. Some important assumptions are made by the student in this procedure. Name one. **(1 mark)**
e. Name two sources of error in gravimetric analysis. **(2 marks)**
f. Suggest an improvement for this analysis. **(1 mark)**

Question 3 (2 marks)

A solution of citric acid, $C_3H_5O(COOH)_3$, was analysed by titration.

25.0 mL aliquots of the $C_3H_5O(COOH)_3$ solution were titrated against a standardised solution of 0.0250 M sodium hydroxide, NaOH. Phenolphthalein indicator was used and the average titre was found to be 24.0 mL.

a. MC Based on the titration, what was the concentration of $C_3H_5O(COOH)_3$ in the solution?
A. 8.0×10^{-3} M B. 8.7×10^{-3} M C. 2.6×10^{-2} M D. 7.2×10^{-2} M

b. MC Which one of the following would have resulted in a concentration that is higher than the actual concentration?
A. The pipette was rinsed with NaOH solution.
B. The pipette was rinsed with $C_3H_5O(COOH)_3$ solution.
C. The conical flask was rinsed with NaOH solution.
D. The conical flask was rinsed with $C_3H_5O(COOH)_3$ solution.

Question 4 (3 marks)

Methane gas burns according to the following equation.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

a. If 4 L of methane is burned, calculate the volume of the following.
i. Carbon dioxide gas produced **(1 mark)**
ii. Gaseous water produced **(1 mark)**
b. Why might a student's final volume be less than this when they measure it experimentally? **(1 mark)**

More exam questions are available in your learnON title.

17.10 Presenting findings using scientific conventions

KEY KNOWLEDGE

- The conventions of scientific report writing, including scientific terminology and representations, standard abbreviations and units of measurement
- Ways of presenting key findings and implications of the selected scientific investigation

KEY SCIENCE SKILL

- Analyse, evaluate and communicate scientific ideas

17.10.1 Why is communication so important in science?

The ability to clearly communicate findings, conclusions and evaluations is vital. It ensures that individuals can be properly informed and have access to strong evidence that is data driven. Effective science communication helps improves public health, advances scientific knowledge throughout society, drives advances in STEM and can also inspire others. This may involve encouraging others through meaningful discussions or providing people with the foundation to develop further research.

Science communication should be predictable and easy to understand. No misunderstanding or misconceptions should occur when effective science communication is used. Individuals should be able to read both a logbook and a poster, and gain understanding about an answer to a scientific question, easily observing if a hypothesis was supported or refuted. This subtopic details some conventions used in chemistry as well as the components of scientific reporting and scientific posters.

Components of scientific report writing

It is vital in scientific report writing to follow a set structure that is clear and predictable. This will also be the same structure of headings that you use when presenting the results of your scientific investigation. These are outlined in table 17.15.

TABLE 17.15 Aspects of a written report

Section	Description
Title	A precise and complete description of what you investigated.
Abstract	An optional section used to outline the key findings and information.
Introduction	A paragraph explaining the relevant chemical and background concepts and relationships, and how they apply to this investigation. It should explore any prior investigations conducted on this topic, and should also include a clear aim and hypothesis.
Methodology	A detailed section that describes your selection of equipment and measuring instruments, and your step-by-step method. This may include diagrams and photos. You should refer to how you controlled variables, achieved the desired accuracy, and overcame, avoided or anticipated difficulties. The methodology explained should be clear enough to enable someone else (at your level) to repeat your experiment. Do not forget to highlight how the relevant ethical and safety concerns have been addressed. This section may be more detailed in your logbook and summarised in your final report.
Results	A clear representation of your results, including your data and graphs. If you have too much data, refer to your logbook for the full set. Make sure you present your results in an organised and clear manner, following accepted conventions (such as numbering sequentially). Make sure you include the appropriate units and use the correct number of significant figures. Try to organise your results in such a way that any patterns or relationships start to become obvious, thus making it easier to analyse in the next section. However, the purpose of this section is to present results only; no interpretation of data should be included in this section. Show sample calculations if required.
Discussion	A detailed analysis and evaluation of your results. How does your data support your initial intentions and link to relevant chemistry concepts? What trends and relationships are apparent as a result of your investigation? How much is your analysis limited by uncertainties? Were any uncertainties occur in your investigation and how were they treated? What limitations and sources of error were present in your investigations and how would you improve this in future repetitions? What would your next steps in the investigation be if you had more time?
Conclusion	This should relate to the aim and must be based entirely on the evidence obtained in the experiment. It should state whether the hypothesis is supported, summarising the meaning of your results in response to your question. No new information should be included. Limitations and future work can be touched on here.
References and acknowledgements	You should acknowledge the sources of any content that you include that is not your own original work. Unaltered tables, diagrams and graphs are examples that fit this description, as are direct quotes. In your introduction, you may have mentioned previous work that your investigation is based on. This also needs to be acknowledged, along with any sources that inform your discussion of concepts and theory in a more general sense. This section is not counted in your final word count.

These components should be included regardless of the format of your report, whether it be:

- a scientific poster
- an article for a scientific publication
- a practical report
- an oral presentation
- a multimedia presentation
- a visual representation.

Ensure you present your task in the format designated or approved by your teacher.

Remember, your teacher will also mark both your scientific report (or poster) and the work in your logbook, so make sure that all required work is there. You may have some of the outlined information that appears only in your logbook, some that appears in both and some that appears only in your report. Your logbook and report are all part of practical investigation skills.

Important aspects of scientific writing

- Try to avoid subjective language; where possible, use third-person language.
- Don't just record the data you believe supports your hypothesis — you should also include any errors, uncertainties and outliers.
- If you used any calculations, show your workings.
- Use subheadings throughout your report as shown in table 17.15. Make the report clear to read.
- Provide headings for all graphs, tables and figures and label them sequentially (graph 1, graph 2 and so on).

It is important throughout your report to use clear and concise terminology relevant to the related chemistry concept. The use of key terms was covered earlier in this topic (see section 17.2.7).

Furthermore, chemical representations are required to be accurate and use common conventions. These representations were also discussed in section 17.2.7. Other representations used may be appropriate to your specific investigation. These should be used consistently in both your logbook and your poster.

17.10.2 The key findings of investigations

In your investigation, you will need to write a question and a hypothesis, create a clear reproducible methodology, conduct an experiment and collect and analyse data.

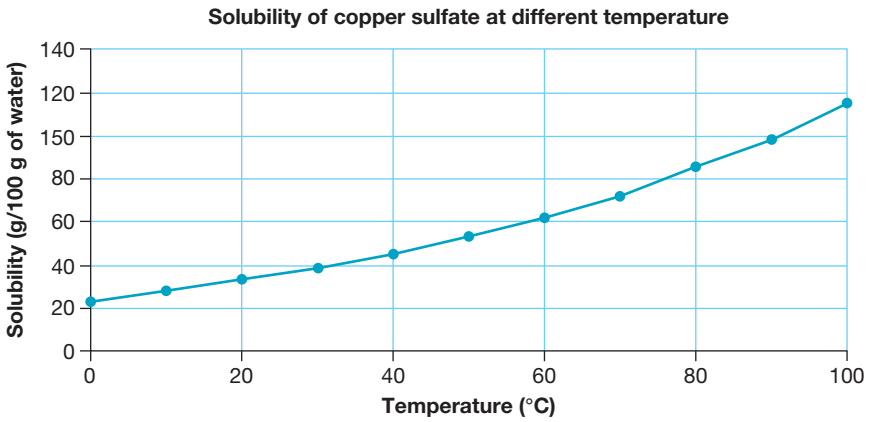
Once these steps are complete, it is then important to determine the key findings of your investigation.

Your key findings should include:

- information about the data obtained in the practical investigation and any patterns and trends
- relationship of your findings to chemical concepts, including solubility, concentration, acid–base and/ or redox concepts
- an answer to the question of your investigation.

For example, your question may have been, 'How does temperature affect the solubility of copper sulfate?' Figure 17.63 shows possible data from your experiment.

FIGURE 17.63 Solubility of copper sulfate



You could use the graph shown in figure 17.63 to link this data to the concept of solubility and determine your key findings. For example:

> *Copper sulfate, $CuSO_4$, is soluble in water, allowing for the formation of an aqueous solution. The expected solubility at 20 °C is 20.5 g/100 g water.*
>
> *In the conducted experiment, it can be seen that at 20 °C, the solubility was recorded at 33 g/100 g water, which is significantly higher than expected. The solution produced was likely to be supersaturated and not saturated as expected.*
>
> *It is also clear that as temperature was increasing, the solubility of copper sulfate also increased. This relationship can be linked to the particle model, as the increase of kinetic energy from higher temperatures causes particles to spread out and move faster. As well as that, higher temperatures lead to an increase in intermolecular bonds breaking, causing the copper sulfate to more easily be dissolved in 100 grams of water.*
>
> *From the investigation, it can be observed that the temperature has a significant impact on solubility as observed in the graph shown; however, this relationship is not linear. This supports the hypothesis of this investigation.*

17.10.3 Equations and formulas

When using equations and formulas in your scientific investigation, it is important to:

- define all variables
- provide any figures for constants (for example, Avogadro's number).

Some key formulas in chemistry are included here, and many can be found in your VCE Chemistry Data Book. You should become very familiar with this data book prior to your exam because it contains a huge amount of useful information you do not need to memorise.

Quantifying chemistry

- Relative atomic mass: $A_r = \dfrac{\text{(relative isotopic mass of first isotope} \times \text{abundance)} + \text{(relative isotopic mass of second isotope} \times \text{abundance)} + \ldots}{100}$
- n (number of moles) $= \dfrac{\text{mass}}{\text{molar mass}} = \dfrac{m}{M}$
- n (number of moles) $= \dfrac{\text{number of particles}}{6.02 \times 10^{23}} = \dfrac{N}{N_A}$
- Molecular formula = $n \times$ empirical formula

pH

- $K_w = [H_3O^+][OH^-] = 10^{-14}$
- $pH = -\log_{10}[H_3O^+]$

Measuring solubility and concentration

- Concentration (c) in grams per litre (g L^{-1}) = $\dfrac{\text{mass of solute in grams}}{\text{volume } (V) \text{ of solution in litres}} = \dfrac{m}{V}$
- Parts per million (ppm) = $\dfrac{\text{mass of solute in micrograms}}{\text{mass of solution in grams}}$
- Concentration %(m/m) = $\dfrac{\text{mass of solute in grams}}{\text{mass of solution in grams}} \times 100\%$
- Concentration %(m/v) = $\dfrac{\text{mass of solute in grams}}{\text{volume of solution in mL}} \times 100\%$
- Concentration %(v/v) = $\dfrac{\text{volume of solute in mL}}{\text{volume of solution in mL}} \times 100\%$
- Molar concentration (c) = $\dfrac{\text{quantity of solute (moles)}}{\text{volume of solution (litres)}} = \dfrac{n}{V}$

17.10.4 Units of measurements

In chemistry, it is vital to use the correct unit of measurement for accurate and clear scientific communication.

Prefixes

Table 17.16 shows the prefixes for units of measurement. These values are from a base unit, which does not have a prefix. Examples of base units include metre, seconds or grams. These base units have the value of 10^0 (or 1). All other prefixes are compared to this base unit using 10^n, in which n varies, as shown in table 17.16.

TABLE 17.16 Prefixes for units of measurement, where the base unit of measurement is the metre

Pico (p)	Nano (n)	Micro (μ)	Milli (m)	Centi (c)	Deci (d)	Kilo (k)	Mega (M)	Giga (G)
10^{-12}	10^{-9}	10^{-6}	10^{-3}	10^{-2}	10^{-1}	10^3	10^6	10^9

Understanding the different prefixes allows the correct units to be used in practical investigations, and also allows for easy conversion between different units.

The following formula is used when converting between units:

$$\frac{\text{initial unit } (10^a)}{\text{new unit } (10^b)} \times \text{value}$$

tlvd-0105

SAMPLE PROBLEM 5 Converting between units

Convert the following.

a. 12.412 millilitres to microlitres

b. 26153 milligram to decigram

c. 8.7 metres to nanometres

THINK	WRITE
a. 1. Determine the conversion between the units. $\frac{\text{millilitres}}{\text{microlitres}}$	$\frac{10^{-3}}{10^{-6}} = 10^{3}$
2. Multiply this by the value to be converted.	$10^{3} \times 12.412 \text{ mL} = 12412$
3. Add the new unit.	$12412\ \mu\text{L} = 1.2412 \times 10^{4} \mu\text{L}$
b. 1. Determine the conversion between the units. $\frac{\text{milligram}}{\text{decigram}}$	$\frac{10^{-3}}{10^{-1}} = 10^{-2}$
2. Multiply this by the value to be converted.	$10^{-2} \times 26153 \text{ mg} = 261.53$
3. Add the new unit	$261.53 \text{ dg} = 2.6153 \times 10^{2} \text{ dg}$
c. 1. Determine the conversion between the units. $\frac{\text{metre}}{\text{nanometre}}$ *Note:* 10^{0} is used for metres because it is our standard unit, so is equal to 1.	$\frac{10^{0}}{10^{-9}} = 10^{9}$
2. Multiply this by the value to be converted.	$10^{9} \times 8.7 \text{ m} = 8\,700\,000\,000$
3. Add the new unit (round if required).	$8\,700\,000\,000 \text{ nm} = 8.7 \times 10^{9} \text{ mn}$

PRACTICE PROBLEM 5

Convert the following.

a. 7823 decigrams into kilograms

b. 213 microlitres into picolitres

SI units of measurement

SI units or Système Internationale is our metric system of measurements. This system is internationally standardised. There are seven base units (see table 17.17) from which all other units are derived.

TABLE 17.17 The seven SI base units

Quantity	Unit	Symbol
Length	Metre	m
Mass	Kilogram	kg
Time	Second	s
Temperature	Kelvin	K
Electric current	Ampere	A
Amount of substance	Mole	mol
Luminous intensity	Candela	cd

Derived units

Derived units are units of measurements derived from the SI units. Table 17.18 shows some commonly used derived SI units.

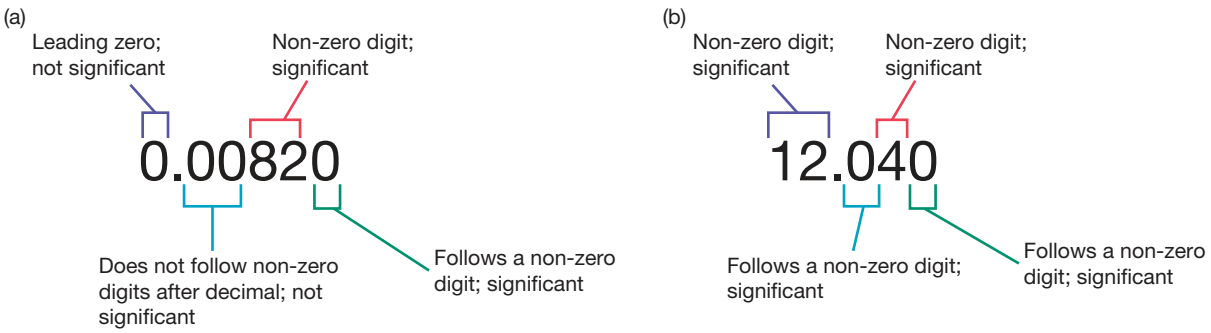
Speed is an example of a quantity that is measured in derived SI units. The SI unit of speed is the metre per second, written as m/s or m s^{-1}.

TABLE 17.18 Common derived SI units used in chemistry

Quantity	Unit	Symbol	Unit in terms of other units
Temperature	Degrees Celsius	°C	+273 K
Capacity	Litre	L	$m^3/10^3$
Force	Newton	N	$(m \times kg)/s^2$
Pressure	Pascal	Pa	N/m^2
Energy and work	Joule	J	N/m
Density		*d*	kg/m^3

Significant figures

FIGURE 17.64 Examples of significant figures: (a) three significant figures and (b) five significant figures



Working with significant figures

When multiplying or dividing, the answer is written to the least number of significant digits.

For example, if you knew a substance had a mass of 7.6 g and a molar mass of 18.5 g mol^{-1}, the amount in mol would be given by:

$$\begin{aligned} \text{mol} &= \frac{\text{mass}}{\text{molar mass}} \\ &= \frac{7.6}{18.5} \\ &= 0.4108108 \text{ mol} \\ &= 0.41 \text{ mol} \end{aligned}$$

We can be confident of the result up to two significant figures because this is the least number of significant figures on which the calculation was based. Hence, the result should be rounded to two significant figures.

When quantities are added or subtracted, the result should be expressed to the minimum number of decimal places used in the data. For example, if you measured three consecutive volumes of 23.4 mL, 24.63 mL and 20.123 mL, the total volume would be given by:

$$\begin{aligned} 23.4 + 24.63 + 20.123 &= 68.153 \\ &= 68.2 \text{ mL} \end{aligned}$$

The result should be rounded off to one decimal place because the minimum number of decimal places used in the data is one, in the volume of 23.4 mL.

on Resources

Video eLesson Determining significant figures (eles-2559)

tlvd-0106

SAMPLE PROBLEM 6 Calculations using the correct number of significant figures

In determining the density of a particular liquid, a student measured the volume of a sample as 8.3 mL. She then weighed the same sample and obtained a mass of 7.2136 g. Calculate the density to the correct level of significant figures.

THINK	WRITE
1. Determine the number of significant figures provided.	7.2136 = five significant figures 8.3 = two significant figures
2. Determine the least number of significant figures — this is what your answer will be in.	Two significant figures.
3. Calculate the density by recalling the formula: density $= \frac{\text{mass}}{\text{volume}}$	Density $= \frac{7.2136}{8.3} = 0.8691$
4. Round down to the appropriate number of significant figures. In this case, 0.8691 needs to be rounded down to two significant figures.	0.87 g mol^{-1}

PRACTICE PROBLEM 6

In determining the density of a particular liquid, a student measured a sample's volume as 21.1 mL. She then weighed the same sample and obtained a mass of 9.762 g. Calculate the density to the correct level of precision.

17.10.5 Referencing and acknowledgement of sources

In-depth scientific reporting requires a depth of research for concepts relating to an investigation. This may include:

- using other sources for definitions and background material
- finding examples of similar investigations
- research on the obtained results to link to scientific understanding.

Acknowledgements are usually used to thank individuals or organisations that provided assistance, whether it be the provision of specific materials, experimental assistance or intellectual assistance.

References are used when you are sourcing information and intellectual property that is not your own. You will find that your references will be longer than your acknowledgements.

If you use any material that is the work of another person, you must reference its source. Do not claim it as your own work.

Acknowledgements and references come in two formats:

- a short version when it occurs in the body of your poster (known as in-text referencing)
- a longer version when it occurs in the *Reference and acknowledgements* section at the end of your poster or scientific investigation.

You should include both forms of referencing.

You can make such acknowledgements in different ways, and various institutions and publications use different systems. Details of these systems can be found online and can be quite complicated. You should check with your teacher as to how this is expected to be done. Section 9.5.3 in topic 9 provides examples of some of the common systems. Many online generators can also assist you with referencing in the correct style.

17.10.6 Conventions of scientific poster presentation

Your logbook will form a key part of your assessment. You will also be required to show a scientific report. This may be in the form of a poster, an article for a scientific publication, a practical report, an oral presentation, a multimedia presentation or a visual representation.

Some important considerations for your poster are as follows:

- The poster should not exceed 600 words (features such as tables, graphs, image captions, references and acknowledgements are not included in the word count).
- It may be produced electronically or in a hard-copy form as determined by your teacher.
- It should address the key aspects without going into too much detail:
  - title
  - introduction
  - methodology and methods
  - results
  - discussion
  - conclusion
  - references and acknowledgements.
- It does not need to contain every single graph and table showing your data, but rather a subset of these that best suits your investigation question.
- It should be easy to read, with clear and succinct communication used throughout.
- It should be able to be understood in conjunction with your logbook — it does not need to be a carbon copy of your logbook.

A sample format of the poster is shown in figure 17.65 and similar sections will be found in other reports.

FIGURE 17.65 Example format of a scientific poster

Title — The question under investigation, with a clear link to the independent and dependent variables

Student name

Introduction
- Purpose of the investigation
- Aim
- Hypothesis
- Background information

Methodology and methods
- Outline of the materials and method used that can be authenticated by logbook entries
- Information about the choice of equipment and how variables were controlled
- Identification and management of relevant risks, including health, safety and ethics

Results
- Presentation of collected data in an appropriate format that allows for trends, patterns and relationships to be easily seen

COMMUNICATION STATEMENT
A one-sentence summary reporting the key finding of your investigation
(This should take up 20–25% of the total poster space)

Discussion
- Analysis and evaluation of primary data — does it support your initial intentions and relevant theory?
- Identification of outliers and their treatment
- Sources of error and uncertainty
- Identification of limitations, with suggested improvements
- Link to relevant chemical concepts
- Linking of results to investigation question, aim and hypothesis
- Implications and information about further investigations — what would the next steps be?

Conclusion
- Conclusion that provides a response to the question referring to the aim and hypothesis

Reference and acknowledgements
- A clear list of references and acknowledgements to source all information

17.10.7 Reviewing your investigation

As part of your outcome, you will need to submit both your logbook and your final report. Before you do, use the following checklist to make sure you have completed all requirements in your report and/or your logbook.

Practical investigation checklist

- ☐ Your name, the title and the aim/hypothesis are listed.
- ☐ An introduction describes the purpose and outlines the investigation in a logical and concise manner. Key terms are defined and variables are stated clearly. Relevant theory is addressed.
- ☐ The method is outlined clearly in step form, including a consideration of ethics, health and safety. This is clearly summarised in your final report or poster.
- ☐ A risk assessment is provided.
- ☐ Your logbook contains dates, headings and complete records.
- ☐ Any abbreviations are explained.
- ☐ Results are presented in an organised way, and in a table if possible. All relevant measurements are recorded with appropriate accuracy and units.
- ☐ Observations are clear and concise, as are all diagrams, graphs and tables.
- ☐ Any calculations are shown.
- ☐ A concise summary and interpretation of key findings is included, outlining trends and any unexpected results with connection to theory.

- ☐ The experimental design is evaluated and possible improvements are included.
- ☐ Suggestions are included for future investigations.
- ☐ The conclusion concisely provides a clear response to the investigation question and states whether it supports or refutes your original hypothesis.
- ☐ All sources are acknowledged and references correctly cited.
- ☐ The use of key terms, symbols and equations is appropriate.

A poster or other form of scientific report should address the sections in your logbook without going into too much detail. For example, you would display only a subset of the data to convey your findings and accuracy. Similarly, not all your graphs need appear.

17.10 Activities

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17.10 Quick quiz on | **17.10 Exercise** | **17.10 Exam questions**

17.10 Exercise

1. Describe the following aspects of a scientific report.
   - a. Introduction
   - b. Discussion
   - c. Conclusion
2. Convert the following units to the SI base unit shown.
   - a. 1984 mL to L
   - b. 1.231 km to m
   - c. 53153 µg to g
3. Convert the following to the unit shown.
   - a. 0.123 kg to mg
   - b. 167283 µL to mL
4. Calculate the following and express in the appropriate number of significant figures.
   - a. 4.25 + 9
   - b. 0.04 + 3.7
   - c. 5.640 + 70.435
   - d. 0.80 – 0.3
   - e. 840 – 627.03
   - f. 12.01 + 6.7
5. Solve the following problems using the correct number of significant figures.
   - a. Density = 27.8 grams/1.2 mL =
   - b. Mass = 23.45 grams + 5.332 grams =
   - c. Moles = 1.221M × 2.6 L =
   - d. m(Cu) = 0.45 mol × 63.5 g mol^{-1} =
   - e. The total time of a reaction to occur from the following three split times: 14.23 seconds, 10.1 seconds, 11.29 seconds.
6. Write a reference in Harvard style for this textbook. Show both the in-text referencing and the full reference for the reference list.

17.10 Exam questions

Question 1 (1 mark)

MC Which of the following features is not required in a scientific poster?

A. A list of references and acknowledgements
B. Identification of sources of error
C. A summary of relevant background concepts
D. A copy of all results, tables and graphs

Question 2 (1 mark)

MC What should the conclusion section of a report do?

A. Include new information that has not been covered in other sections of the report
B. Include graphs that best outline the data
C. Link the findings of the investigation back to the hypothesis and question of the investigation
D. Describe any outliers and errors that occurred in the data

Question 3 (4 marks)

a. Explain why concise and clear communication is vital in a scientific poster. **(1 mark)**
b. Identify three factors that should be included in your discussion. **(3 marks)**

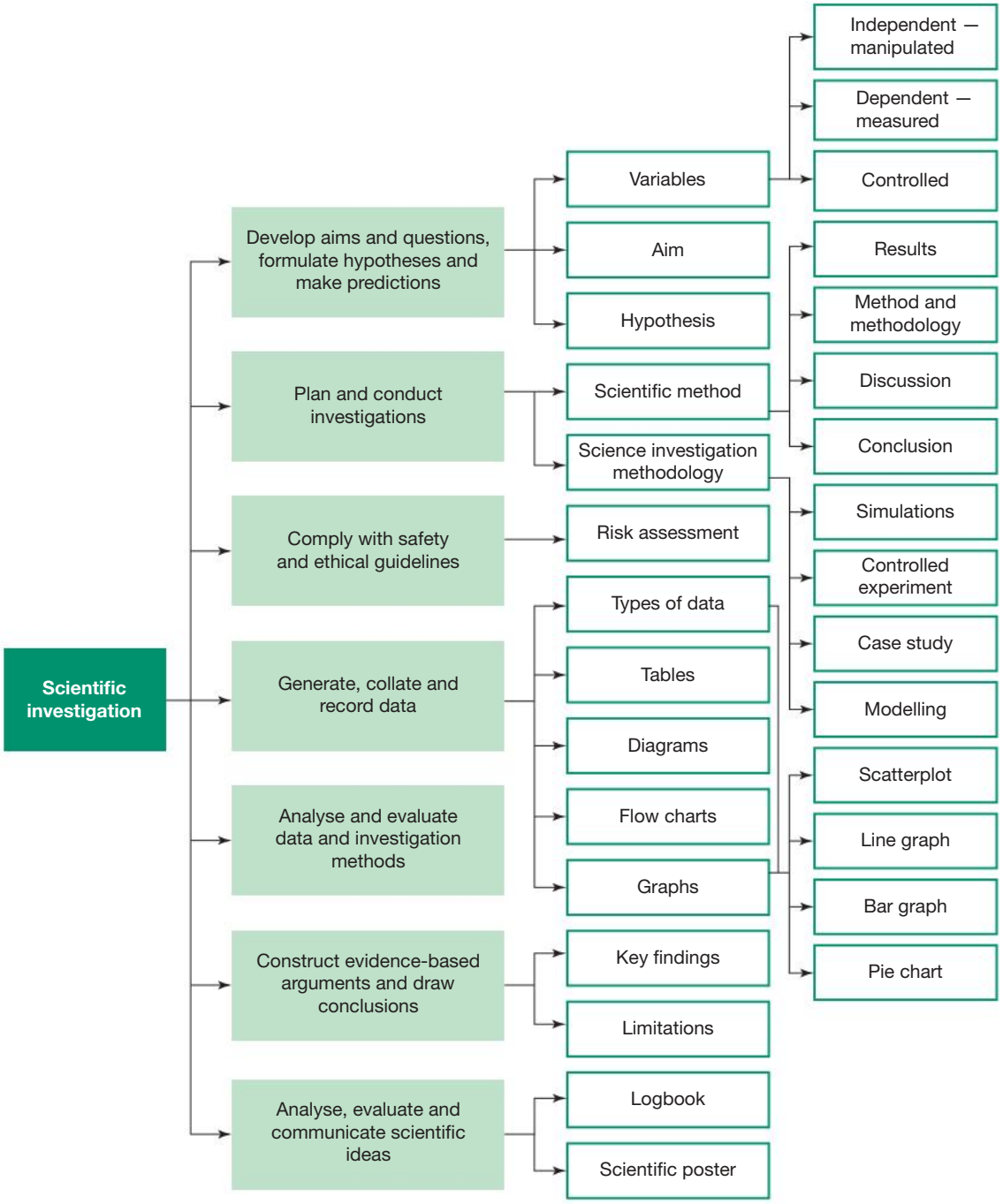
Question 4 (3 marks)

a. Why is it important to reference sources in a scientific investigation? **(1 mark)**
b. As part of referencing websites, you should reference the date the website was visited. This is not required for journal articles. Explain why the date should be included for websites but not journal articles.
c. An image is marked as Creative Commons with the abbreviation BY. Explain how you should reference this image. **(1 mark)**
d. How does a reference differ from an acknowledgement? **(1 mark)**

More exam questions are available in your learnON title.

17.11 Review

17.11.1 Topic summary



Hey students! Now that it's time to revise this topic, go online to:

Access the topic summary | Review your results | Watch teacher-led videos | Practise exam questions

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17.11.2 Key ideas summary

online only

17.11.3 Key terms glossary

online only

on Resources

Solutions	Solutions — Topic 17 (sol-0847)
Digital documents	Key science skills (doc-37066) Key terms glossary — Topic 17 (doc-37915) Key idea summary — Topic 17 (doc-37916)
Exam question booklet	Exam question booklet — Topic 17 (eqb-0126)

17.11 Activities

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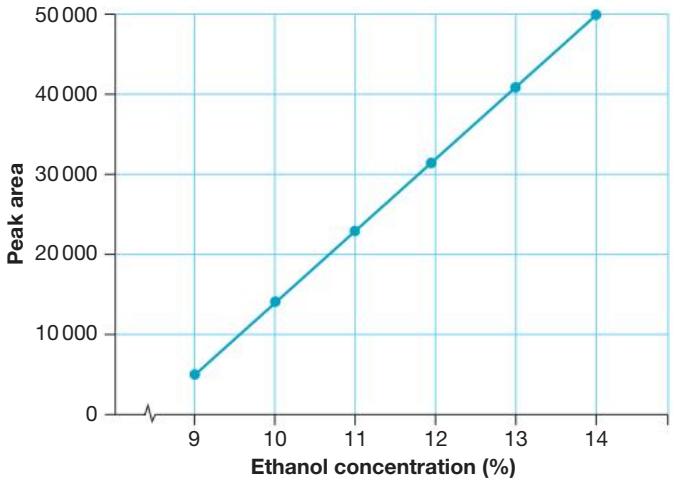
17.11 Review questions

1. An investigation was conducted to observe the differences in pH of five different swimming pools.
   a. Identify two pieces of quantitative data that you could record and measure.
   b. What instruments, if any, would you need to make these quantitative observations?
   c. Identify two pieces of qualitative data that you could observe.
   d. What instruments, if any, would you need to make these qualitative observations?
2. Identify a key difference between the members of the following pairs.
   a. Independent and dependent variables
   b. Qualitative and quantitative data
   c. Control group and experimental group
   d. Primary and secondary sources of data
   e. Uncertainty and error
3. In an investigation conducted in class, Jazz recorded a concentration of 0.0540 mol L^{-1}.
   a. How many significant figures does this recorded concentration have?
   b. Write this concentration in scientific notation.

4. Chris used gas chromatography (GC) to measure the ethanol content of some alcoholic beverages. GC works on a similar principle to high-performance liquid chromatography (HPLC). In both instruments, it is necessary to produce a calibration curve in addition to obtaining readings for the test samples. One such calibration curve is as shown.
   a. Sample A produced a reading of 36 000 from the GC. Estimate the level of ethanol in this sample. What is the name of the process used to obtain this answer?
   b. Sample B produced a reading of 50 500. Estimate the level of ethanol in this sample and comment on your answer. What is the name of the process used to obtain this estimate?
   c. Sample C produced a reading of 95 000. Is it possible to estimate the ethanol level in this sample? Explain why or why not.
   d. Write a conclusion for this investigation.



5. Ammie conducted an experiment to explore how the concentration of cordial affects the freezing rate. The results she recorded are shown in the following table.

Concentration of cordial (%)	Time to freeze (mins)
0	25
10	32
20	45
30	53
40	60
50	65

   a. Using graph paper, plot the data using the most appropriate graph type, including a scale and labels.
   b. Describe the trends and patterns in this graph.
   c. What conclusions would you make from this investigation?

17.11 Exam questions

Section A — Multiple choice questions

All correct answers are worth 1 mark each; an incorrect answer is worth 0.

Question 1

MC Belinda conducted an experiment to test the combustion of different types of fuel. In this experiment, what is the volume of fuel being tested?

A. The dependent variable
B. The independent variable
C. The controlled variable
D. The control group

Question 2

MC Students were conducting an experiment to investigate the temperature of different rooms in the school. Each measurement was taken three times using a thermometer and recorded by the student. A highly accurate digital thermometer was also used.

The following table shows the results recorded by each student.

	Recorded temperature using thermometer (°C)			Digital thermometer reading (°C)
Student	Measurement 1	Measurement 2	Measurement 3	
Mel	26	25	25.25	25.5
Josette	24	32.5	20	24.8
Fiona	28.5	28.25	28.5	24.7
Judy	29.9	31.2	32	31.2

Which of the following statements is most correct about the results recorded?

A. Mel's data is accurate but not precise.
B. Josette's data is precise but not accurate.
C. Fiona's data is the most accurate but not the most precise.
D. Judy's data is more accurate than Fiona's.

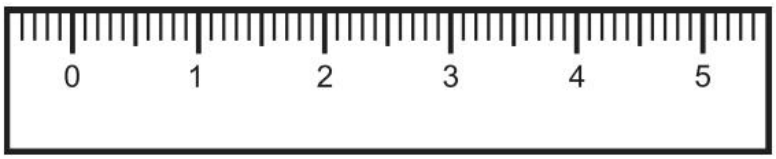
Question 3

MC What are qualitative results?

A. Results that are expected to be observed during an experiment
B. Results that are measured in an experiment
C. Results that are observed during an experiment
D. None of the above

Question 4

MC What would the measure of uncertainty be in the ruler shown?



A. 0.5 cm
B. 0.1 cm
C. 0.05 cm
D. 0.025 cm

Question 5

MC Systematic errors can

A. affect the accuracy of a reading.
B. be caused by human fault.
C. be improved by repetition of an experiment.
D. be caused by deliberate changing of results due to bias.

Question 6

MC What is the difference between an aim and a hypothesis?

A. The aim of an experiment is a statement, but the hypothesis is a question.
B. The aim includes the dependent and independent variables, but a hypothesis doesn't.
C. The aim explains the expected data, but the hypothesis explains how you will get the data.
D. The aim outlines the purpose of the investigation, but the hypothesis is a testable prediction.

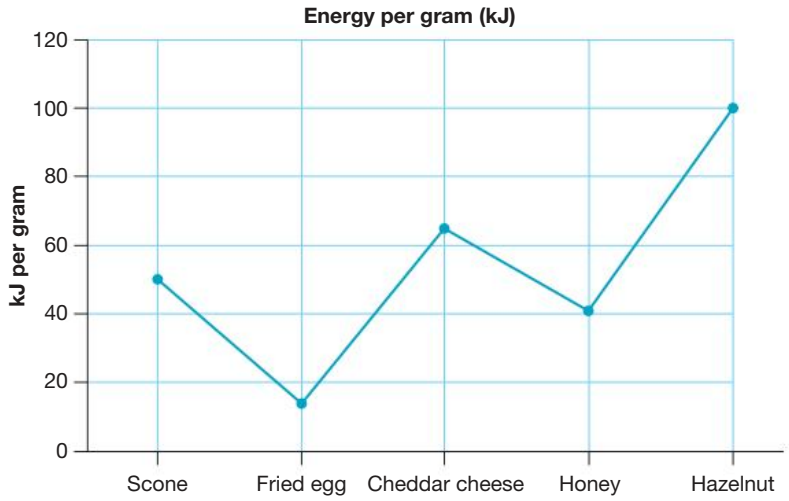
Question 7

MC Paul was investigating the energy content from different types of food through combustion. The results of his investigation are shown in the following table.

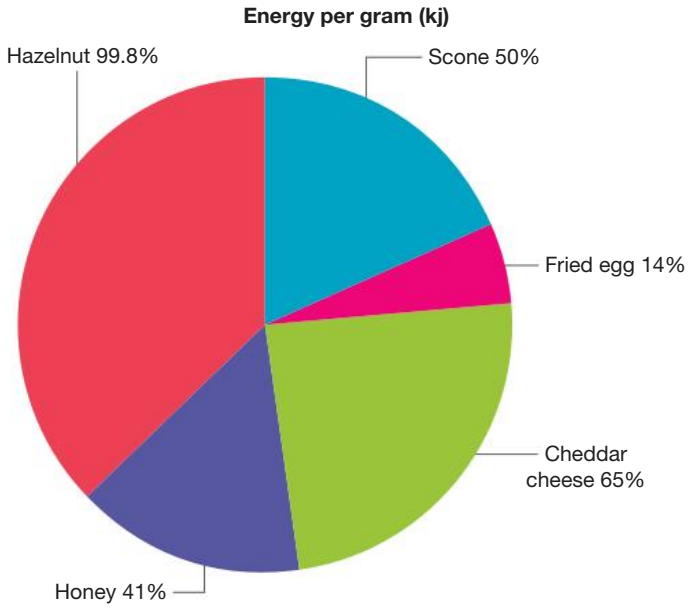
Type of food	Energy per gram (kJ)
Scone	50
Fried egg	14
Cheddar cheese	65
Honey	41
Hazelnut	99.8

What is the most appropriate graph to use to show and analyse trends in this data?

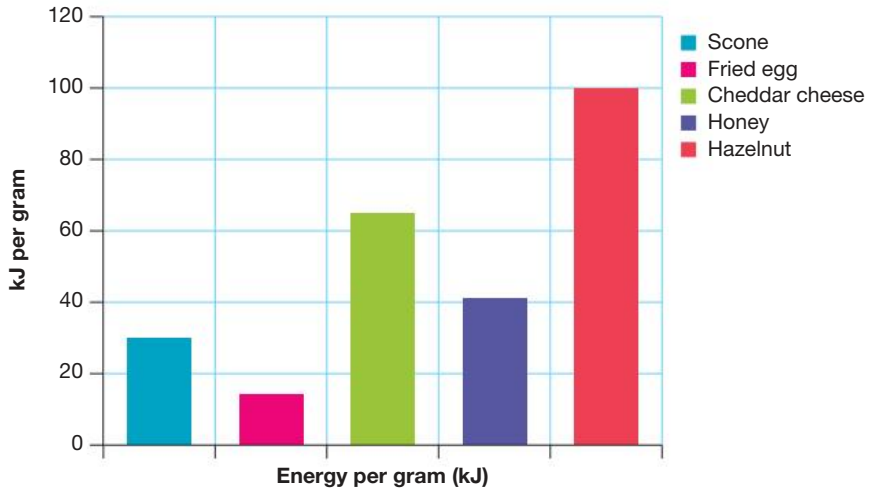
A.


Energy per gram (kJ)
kJ per gram
120
100
80
60
40
20
0
Scone
Fried egg
Cheddar cheese
Honey
Hazelnut

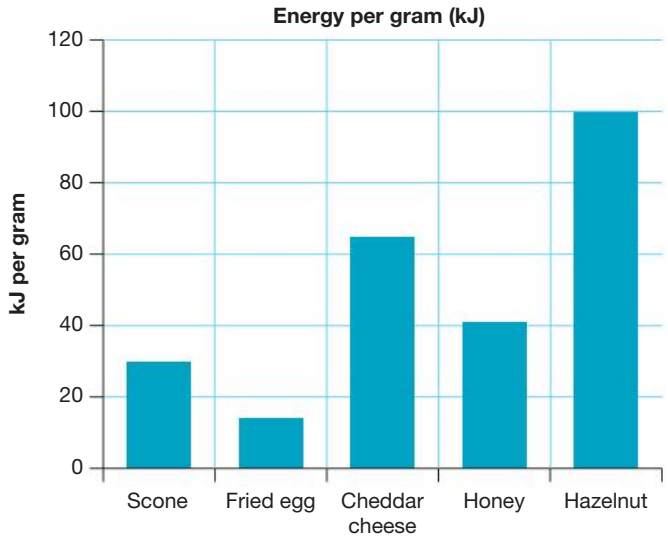
B.


Energy per gram (kj)
Hazelnut 99.8%
Scone 50%
Fried egg 14%
Cheddar cheese 65%
Honey 41%

C.


kJ per gram
120
100
80
60
40
20
0
Energy per gram (kJ)
Scone
Fried egg
Cheddar cheese
Honey
Hazelnut

D.



Question 8

MC Stacey and Bridget both swam 200 metres. On a very close finish, it was found that Stacey was 999 milliseconds faster than Bridget.

Which of the following is correct about her result?

A. It is equal to 9.99 seconds.
B. It contains two significant figures.
C. It can be written as 9.99×10^2 milliseconds.
D. It can be shortened to 999 Ms.

Question 9

MC What are theories?

A. Concepts that were once accurate, but have now been rejected
B. Diagrammatic representations of abstract concepts
C. Ideas that are unable to be rejected due to the amount of evidence they have
D. Well-supported ideas where evidence has been gained from investigations, research and observations

Question 10

MC On a risk assessment, a chemical has an image of a flame and an exclamation mark. What does this mean the chemical is?

A. An oxidiser and toxic
B. Flammable and harmful
C. Flammable and toxic
D. Harmful and explosive

Section B — Short answer questions

Question 11 (11 marks)

Vicki decided to investigate how solubility is affected by the temperature of the solution. She had a container of sodium sulfate powder. She put on some safety gloves and glasses and measured the temperature of the water in a test tube using a thermometer. She added sodium sulfate until it could no longer dissolve, and recorded how much sodium sulfate she added to reach this point. She then got a new test tube with water and heated it to 50 °C and repeated the process. She found that more sodium sulfate dissolved at 50 °C compared to that at room temperature

a. Comment on the safety precautions taken. Do you think they were sufficient? **(2 marks)**

b. Identify the dependent variable in this investigation. **(1 mark)**

c. List two variables that need to be controlled in this investigation. **(2 marks)**

d. Describe what measuring equipment should be used in this investigation and identify one factor that could affect the accuracy of this piece of equipment. **(2 marks)**

e. Write a clear experimental method for Vicki that shows how she can investigate how solubility is affected by temperature. You may need to add to her method to make sure that more data is being collected and to ensure it is reproducible. **(4 marks)**

Question 12 (10 marks)

An investigation was being conducted to examine if different brands of bottled water have different pH.

a. Write a suitable aim for this investigation. **(1 mark)**

b. Identify the hypothesis for this investigation. **(1 mark)**

c. Describe **one** piece of qualitative data and **one** piece of quantitative data that may be collected in this investigation. **(2 marks)**

d. Explain **two** factors that may lead to differences in results between different students. **(2 marks)**

e. Outline an appropriate method for this investigation. **(4 marks)**

Question 13 (8 marks)

A student decided to investigate the effect of temperature on the rate of the following reaction.

$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

Part of the student's experimental report is provided here.

Effect of temperature on the rate of production of carbon dioxide gas

Aim

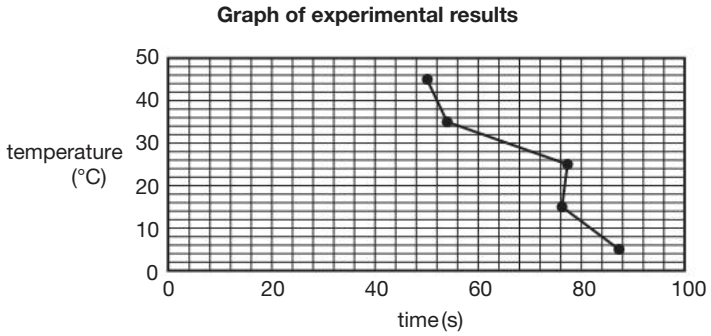
To find out how temperature affects the rate of production of carbon dioxide gas, CO_2, when a solution of hydrochloric acid, HCl, is added to chips of calcium carbonate, $CaCO_3$.

Method

1. Put 0.6 g of $CaCO_3$ chips into a conical flask.
2. Put a reagent bottle containing 2 M HCl into a water bath at 5 °C.
3. When the temperature of the HCl solution has stabilised at 5 °C, use a pipette to put 10.0 mL of the HCl solution into the conical flask containing the $CaCO_3$ chips.
4. Put a balloon over the conical flask and begin timing.
5. When the top of the balloon has inflated so that it is 10 cm over the conical flask stop timing and record the time.
6. Repeat steps 1 to 5 using temperatures of 15 °C, 25 °C. 35 °C and 45 °C.

Results

The following graph gives the experimental results.



a. **i.** Predict the relationship between the independent variable and the dependent variable. Explain your prediction. **(3 marks)**

ii. Is the graph of the student's results consistent with your prediction? Give your reasoning. **(1 mark)**

b. Identify two ways in which the graph could have been presented differently to better illustrate the relationship between the independent variable and the dependent variable. **(2 marks)**

c. Identify two changes that could be made to the experimental method to improve the precision of the results if the experiment was repeated. For each change, explain how it would improve precision. **(2 marks)**

Question 14 (7 marks)

Two students, Alan and Will, conducted an investigation to confirm that the boiling point of water is actually 100 °C, regardless of the volume of water being tested. They examined five different volumes and recorded the temperature of boiling.

Their results are shown in the table.

	Alan	Will
Volume (mL)	Recorded boiling point (°C)	Recorded boiling point (°C)
20	100.5	97
40	100	97.5
60	98.5	95.5
80	101.5	96
100	102.5	96.5

a. Describe **one** error that may have occurred that reduced the accuracy of the data for both Alan and Will, and identify how this error may be avoided. **(2 marks)**

b. Calculate the average temperature recorded by Will, recording the number to the appropriate level of precision based on the data provided. **(2 marks)**

c. Describe, with reference to the known value of the boiling point of water, why the results obtained by Alan are more accurate than those obtained by Will. **(2 marks)**

d. It was found that the density of water at the start of the investigation was 0.987 g mL^{-1}. Convert this into grams per litre. **(1 mark)**

Question 15 (12 marks)

A student conducted an investigation to determine the rate in which the temperature of water decreases when in different types of insulating material.

The student's report is as follows.

Introduction: In this experiment, three materials were examined: aluminium foil, paper and wool. This experiment was conducted over three days, with aluminium foil explored on the first day, paper on the second day and wool on the third day. All materials were wrapped around a polystyrene cup in which boiling water was placed.

Aim: To determine how different materials affect the temperature decrease of water.

Hypothesis: If aluminium foil, paper and wool are used as insulation material, then aluminium foil will work best.

Method:

1. *Collect a polystyrene cup and wrap a piece of aluminium foil around it.*
2. *Place 100 mL of boiling water in the cup and record the temperature.*
3. *Record the temperature every five minutes.*
4. *Repeat Steps 1 to 3 with paper and then with wool.*

Results:

	Aluminium	Paper	Wool
Initial	100.0 °C	98.5 °C	95.0 °C
5 minutes	95.5 °C	85.0 °C	80.0 °C
10 minutes	92.0 °C	78 °C	71.5 °C
15 minutes	85.0 °C	62.0 °C	60.0 °C
20 minutes	80.5 °C	50.0 °C	48.5 °C
25 minutes	75.0 °C	40.0 °C	39.5 °C

a. Describe an issue with the hypothesis written by the student. How would you adjust this to make it testable? **(2 marks)**

b. The experiment was conducted over three days. Explain why this may lead to errors in the data obtained. **(2 marks)**

c. Students were only able to use a thermometer in which temperature could be measured to the nearest 0.5 °C. Identify the resolution of this device and describe the uncertainty expected in the data. **(2 marks)**

d. Outline two limitations in the experimental method or data collection process that would affect the conclusions drawn. **(2 marks)**

e. Describe the most appropriate graph that the student should use to represent their data. Justify your choice. **(2 marks)**

f. Based on the student's results, write a conclusion for this investigation, linking back to the hypothesis. **(2 marks)**

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Answers

1 Elements and the periodic table

1.2 Elements

Practice problem 1

a. H

b. Each atom has one proton, the atomic number $Z = 1$

c. hydrogen-1: 0
hydrogen-2: 1
hydrogen-3: 2

d. ${}^{1}_{1}\mathrm{H}$, ${}^{2}_{1}\mathrm{H}$ (deuterium), ${}^{3}_{1}\mathrm{H}$ (tritium)

1.2 Exercise

1. C

2. B

3. a. $Z = 1$

b. $Z = 10$

c. $Z = 47$

d. $Z = 79$

4. a. $A = 13$

b. $Z = 27$

c. Aluminium-27

5. a. B

b. Mg

c. Ar

d. Ca

6. a. 22

b. 18

c. ${}^{40}_{18}\mathrm{Ar}$

7. Number of protons = 35,
Number of neutrons = 44
Number of electrons = 36

8. ${}^{2}_{1}\mathrm{H}$

9. a. See table at the bottom of the page*

b. ${}^{235}_{92}\mathrm{U}$ and ${}^{238}_{92}\mathrm{U}$

c. ${}^{238}_{92}\mathrm{U}$ has three more neutrons than ${}^{235}_{92}\mathrm{U}$.

10. The atomic number is always the same for a particular element, while the mass number can differ due to the existence of isotopes (an element with differing numbers of neutrons).

1.2 Exam questions

1. D

2. See table at the bottom of the page**

3. Isotopes

4. $Z = 17$: chlorine ${}^{37}_{17}A$, ${}^{35}_{17}D$

$Z = 12$: magnesium ${}^{26}_{12}B$, ${}^{25}_{12}E$, ${}^{24}_{12}G$

$Z = 27$: cobalt ${}^{59}_{27}C$, ${}^{60}_{27}F$

5. Tellurium atoms have more neutrons than iodine, resulting in a higher mass number and higher mass.

*9a.

Element	Number of protons	Number of electrons	Number of neutrons
${}^{12}_{6}\mathrm{C}$	6	6	6
${}^{56}_{26}\mathrm{Fe}$	26	26	30
${}^{40}_{18}\mathrm{Ar}$	18	18	22
${}^{235}_{92}\mathrm{U}$	92	92	143
${}^{238}_{92}\mathrm{U}$	92	92	146
${}^{19}_{9}\mathrm{F}$	9	9	10

**2.

Atom	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons	Name of element
${}^{23}_{11}\mathrm{Na}$	11	23	11	12	11	Sodium
${}^{19}_{9}\mathrm{F}$	9	19	9	10	9	Fluorine
${}^{28}_{14}\mathrm{Si}$	14	28	14	14	14	Silicon
${}^{56}_{26}\mathrm{Fe}$	26	56	26	30	26	Iron
${}^{197}_{79}\mathrm{Au}$	79	197	79	118	79	Gold
${}^{235}_{92}\mathrm{U}$	92	235	92	143	92	Uranium

1.3 Electrons

Practice problem 2

a. Ga^{3+}, gallium ions

b. P^{3-}, phosphide ions

Practice problem 3

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

Practice problem 4

Yes

1.3 Exercise

1. $_3Li$: $1s^2 2s^1$

$_{19}K$: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

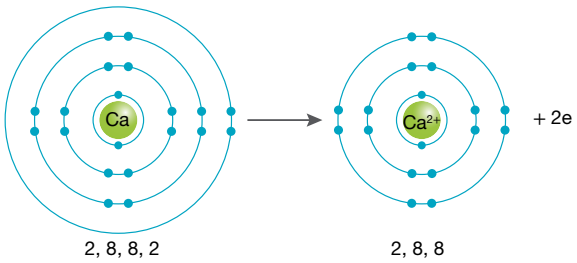
$_6C$: $1s^2 2s^2 2p^2$

$_{17}Cl$: $1s^2 2s^2 2p^6 3s^2 3p^5$

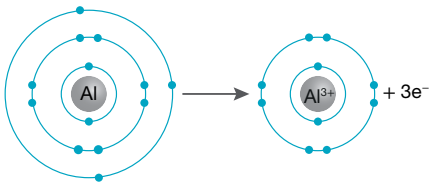
$_{18}Ar$: $1s^2 2s^2 2p^6 3s^2 3p^6$

$_7N$: $1s^2 2s^2 2p^3$

2. a.

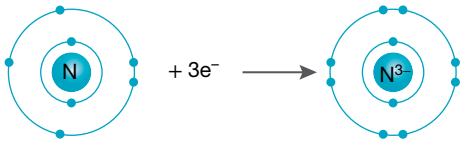


Calcium Ca $\rightarrow$ $Ca^{2+} + 2e^-$
2, 8, 8, 2 2, 8, 8

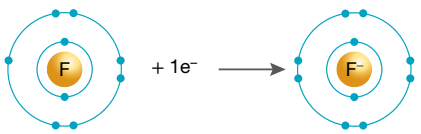


Aluminium Al $\rightarrow$ $Al^{3+} + 3e^-$
2, 8, 3 2, 8

b.



Nitrogen $N + 3e^-$ $\rightarrow$ N^{3-}(nitride)
2, 5 2, 8



Fluorine $F + e^-$ $\rightarrow$ F^-(fluoride)
2, 7 2, 8

3. a. Barium ion, Ba^{2+} **b.** Potassium ion, K^+
c. Phosphide ion, P^{3-} **d.** Chloride ion, Cl^-
e. Sulfide ion, S^{2-}

4. a. $^{32}_{15}P$: $1s^2 2s^2 2p^6 3s^2 3p^3$

b. $^{32}_{15}P$ has 17 neutrons while $^{31}_{15}P$ has only 16.

5. a. Lithium **b.** Neon
c. Sodium **d.** Aluminium
e. Argon **f.** Calcium

6. a. $Z = 20$: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

$Z = 12$: $1s^2 2s^2 2p^6 3s^2$

$Z = 4$: $1s^2 2s^2$

$Z = 9$: $1s^2 2s^2 2p^5$

b. $Z = 5$: $1s^2 2s^2 2p^1$

$Z = 6$: $1s^2 2s^2 2p^2$

$Z = 8$: $1s^2 2s^2 2p^4$

$Z = 16$: $1s^2 2s^2 2p^6 3s^2 3p^4$

7. A shell is a region in which a group of electrons of similar energies moves around the nucleus. Subshells are found within any shell and contain orbitals with the same energy level.

8. a. Ground state **b.** Excited state **c.** Ground state

9. a. The electrons are not in the lowest available energy levels.

b. When an electron changes to the ground state, energy is emitted and a photon released.

c. $1s^2 2s^2 2p^6$

10. a. Mn **b.** Ca

1.3 Exam questions

1. A

2. D

3. a. Electron transitions from higher to lower energy states.

b. 6

4. a. *Q* and *X*

b. *Q*

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

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

e. *R*

5. 4 (partially filled, or 3 completely filled)

1.4 The periodic table

Practice problem 5

This element is chromium. The atom is in period 4 in the *d*-block.

1.4 Exercise

1. D

2. Chemists study the physical and chemical properties of the elements (and their compounds) that the periodic table is based on. (Detailed responses can be found in the worked solutions in your digital resource.)

3. They have the same number of valence electrons.

4. The *d* subshell can hold 10 electrons.
5. a. 8 electrons = 8 protons = period 2, group 16, oxygen
   b. 3 electrons = 3 protons = period 2, group 1, lithium
   c. 26 electrons = 26 protons = period 4, group 8, iron
   d. 36 electrons = 36 protons = period 4, group 18, krypton
   e. 21 electrons = 21 protons = period 4, group 3, scandium
   f. 30 electrons = 30 protons = period 4, group 12, zinc
6. a. The element is manganese and belongs to period 4, group 7.
   The original ion was Mn^{7+}.
   b. The element is chlorine, and belongs to period 3, group 17.
   The original ion was Cl^- (chloride).
7. a. $1s^2 2s^2 2p^4$, period 2, group 16
   b. $1s^2 2s^2 2p^6 3s^2 3p^1$, period 3, group 13
   c. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$, period 4, group 2
   d. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4p^5 3d^{10}$, period 4, group 17
8. a. Fe, S
   b. Select from the following.

Iron	Sulfur
High melting and boiling points	Low melting and boiling points
Good conductors of heat and electricity	Poor conductors of heat and electricity
Shiny appearance	Dull colour
Ductile and malleable	Brittle when solid
Strong	Weak
Form positive ions (cations)	Form negative ions (anions)

   c. Iron is a metal; sulfur is a non-metal.
   d. Iron is found in *d*-block; sulfur is found in *p*-block.
9. Group 15
10. [Kr] $4d^7 5s^1$ OR
    $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^7 5s^1$.

1.4 Exam questions

1. B
2. a. Top, left
   b. 1, 2, 7
   c. Groups 1 and 2 metals are likely to lose electrons because atoms are more stable if they have eight electrons in their outer shell. It is easier to lose one or two electrons rather than gain seven or eight.
   d. Positive ions
   e. Alkali metals, alkaline earth metals, halogens
3. a. Barium is in the *s*-block.
   b. Copper is in the *d*-block.
   c. Uranium is in the *f*-block.
   d. Neon is in the *p*-block.
4. a. Silicon
   b. Calcium
   c. Chromium
   d. Oxygen
   e. Chlorine
   f. Iron
5. a. Elements in the first period have electrons that fill the first shell. This shell contains a maximum of two electrons.
   b. Elements in the second period have electrons that fill the second shell, two in the *s*-subshell and six in the *p*-subshell.
   c. The first series of transitions metals have electrons in the 3*d* subshell. The 3*d* subshell is filled after the 4*s* subshell and hence the first series of transition metals are in the fourth period.

1.5 Trends in the periodic table

1.5 Exercise

1. A
2. A
3. A
4. a. 1
   b. 7
5. a. More energy is required to remove the second electron. The nucleus has a greater attraction for the remaining electrons because the atom now has an overall positive charge.
   b. The lower the ionisation energy, the easier an electron is lost and so the reactivity is greater.
   c. Alkali metals because they have the lowest core charge.
   d. A high ionisation energy means it is difficult for an electron to be removed.
6. a. Peaks: H, F, Cl, Br — highest electronegativity because of highest core charge in their period.
   b. Elements 3 to 10 should be circled.
   c. Troughs: He, Ne, Ar, Kr — lowest electronegativity because they have a full valence shell and do not require any electrons.
   d. There is a similar trend of regularly increasing electronegativity (and then decreasing for noble gases) across each period (except period 1).
7. Moving down a group, metallic character increases while electronegativity decreases. This is because, as you move down a group, there are an increasing number of inner electron shells, which shield the outermost shell from the nucleus. The outer electrons are less attracted to the nucleus and more easily lost (increase in metallic character), and the nucleus is less able to attract electrons from neighbouring atoms (decrease in electronegativity).
8. a. The electronegativity increases from left to right across the periodic table.
   b. The atomic radii increase.
   c. The reducing strength decreases across a period.
9. a. Ca is more electronegative than K.
   b. Be is more electronegative than Ca.
   c. Cl is more electronegative than Br.
10. *Sample response:* Sodium is an alkali metal (and so it is very reactive and readily loses electrons to another element to form a compound); therefore, it is found as a compound. Gold is transition metal and is relatively unreactive, so can be found, in elemental form, as nuggets.

1.5 Exam questions

1. A
2. a. Electronegativity is defined as the ability of an element to gain electrons when bonding.
   b. i. F ii. K
3. B
4. a. *Q* and *J* b. *G*
   c. *X* d. *R*
   e. *J*
   f. Ionisation energy is the energy required to remove an electron from an atom.
   g. *J*
5. • Metallic character is measured by how well an atom loses electrons.
   • The stronger the core charge, the stronger the valence electrons are attracted to the nucleus, making them more difficult to lose.
   • The larger the atomic radius, the weaker the attraction between nucleus and valence electrons, making them easier to lose.

1.6 Critical elements

1.6 Exercise

1.

Description	Element(s)
Second most abundant element	He
Part of DNA	P
Used to manufacture touch screens	In
Often magnetic	REEs
Used in 'green technology'	REEs and In
Can be responsible for eutrophication	P
Used in MRI scanners	He
Can be obtained from waste organic matter	P
Stays liquid at extremely low temperatures	He
Includes elements filling the 4*f* subshell	REEs

2. C
3. B
4. a. Helium is formed from radioactive decay.
   b. Helium is a by-product of natural gas extraction.
5. B
6. D
7. a. Eutrophication is the starving of waterways of oxygen through uncontrolled plant or algal growth.
   b. Excessive use of phosphorus-based fertilisers causes the uncontrolled growth of aquatic plants, which leads to eutrophication.
8. Touch screens. Indium conducts electricity well, is transparent and adheres to glass.
9. Please see worked solutions in your online resources.
10. a. Helium is sourced as a by-product from natural gas extraction.
    b. Any three from the following:
    • Cool superconductors in MRI scanners
    • Cool and clean rocket engines
    • Cool some nuclear reactors
    • Inflate balloons
    • Add to air supply for deep sea divers

1.6 Exam questions

1. D
2. D
3. Helium is abundant in the universe but not on Earth because it is a very light gas and if it escapes after use it is lost into the atmosphere. Also, because it is inert it doesn't form compounds that might trap it on Earth.
4. $^{232}_{90}\text{Th} \rightarrow {}^{228}_{88}\text{Ra} + {}^{4}_{2}\text{He}$
5. Please see the worked solutions in your online resources.

1.7 Review

1.7 Review questions

1. See table at the bottom of the page*
2. a. A, C and D.
   b. $A = {}^{40}_{20}\text{Ca}$, $B = {}^{37}_{17}\text{Cl}$, $C = {}^{42}_{20}\text{Ca}$, $D = {}^{46}_{20}\text{Ca}$.
3. a. It needs to absorb energy.
   b. A photon equal to the difference in energy of the two levels is emitted.
4. a. Half-filled and filled subshells are more stable than other partly filled subshells.
   b. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$
   c. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$

*1.

Name of atom	Atomic number	Mass number	Protons	Neutrons	Electrons
Argon	18	36	18	36 – 18 = 18	18
Sulfur	16	34	16	34 – 16 = 18	16
Argon	18	38	18	38 – 18 = 20	18
Phosphorus	15	31	15	31 – 15 = 16	15
Lead	82	208	82	208 – 82 = 126	82
Potassium	19	39	19	39 – 19 = 20	19
Sulfur	16	36	16	36 – 16 = 20	16

5. a. Cl and I
   b. Na
   c. Nickel is a transition metal.
   d. Cl is the most reactive non-metal.
   e. Chlorine is the most electronegative.
6. a. Na and K would be the most similar because they are in the same group. Na and Cl would be the most different because Na is a metal and Cl is a non-metal.
   b. Cl and I would be the most similar because they are in the same group. Cl and Mg would be the most different because Mg is a metal and Cl is a non-metal.
7. a. $Z = 15$: $1s^2 2s^2 2p^6 3s^2 3p^3$
   b. Group 15
   c. Period 3
   d. *p*-block
   e. Phosphorus would have a lower electronegatively than chlorine.
   f. Phosphorus has a larger atomic size than nitrogen.
   g. It is essential in all life (in DNA), involved in energy transport, and used in artificial fertilisers. Supply is limited, and it has no substitutes. Solutions include recycling waste to extract phosphorus and minimising the use of fertilisers.
8. a. The nuclear charge of an element is equal to the charges on the nuclei of all atoms of that element.
   b. Calcium has the higher nuclear charge.
9. Elements with the same outer electron configuration are in the same group and so have similar properties; for example, oxygen $1s^2 2s^2 2p^4$, sulfur $1s^2 2s^2 2p^6 3s^2 3p^4$ and selenium $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$.
10. a. Strontium is lower than magnesium in the periodic table, which means that it has more shells so the valence electrons of strontium are further from the nucleus and so not as strongly attracted.
    b. Beryllium and nitrogen are both in period 2. Going across the period, the core charge is increasing so stronger attraction exists for the valence electrons, pulling them in closer to the nucleus. This means the atomic radius is smaller than expected.
    c. Fluorine is the most electronegative atom because the core charge (effective nuclear charge) increases across the periodic table, making the attraction for electrons greater. Going down the periodic table, away from fluorine, there are more electron shells and so the attraction by the nucleus is less effective.
    d. Reducing agents are good electron donors. Potassium readily loses an electron to gain a stable outer shell configuration of eight and so is a good reducing agent.

1.7 Exam questions

Section A — Multiple choice questions

1. C
2. B
3. B
4. D
5. A
6. B
7. B
8. B
9. D
10. D

Section B — Short answer questions

11. The number of neutrons is 183 – 74 = 109.
12. a. i. As = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^3$
    ii. $Ca^{(2+)} = 1s^2 2s^2 2p^6 3s^2 3p^6$
    b. The number of outer electrons in the atom determines its group number. The arsenic atom has five valence electrons in its outer fourth shell ($4s^2 4p^3$; 2 + 3 = 5) so it is in period 4, group 15. The valence electrons are in the *p* shell, so it is in the *p*-block.
13. *J*
14. a. *D* or *X*
    b. *X*
    c. *B*
    d. *J*
    e. *A*
    f. *G* or *J*
    g. *T*
15. In the circular economy, touch screens are manufactured, used, collected and recycled. They are then manufactured from the recycled material. The process is then repeated, ensuring minimum wastage of resources and energy. This results in reduced damage to human health and the environment.

2 Covalent subtances

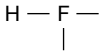
2.2 Representing molecules

Practice problem 1

Practice problem 2

Practice problem 3

a.

b.


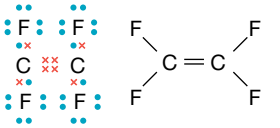
2.2 Exercise

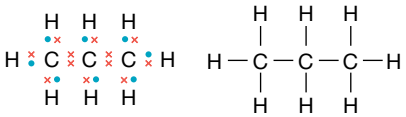
1. A
2. D
3. B
4. a.

b.

5. See table at the bottom of the page*

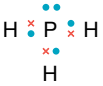
6. a.

b. 

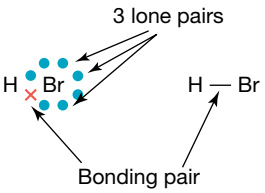
c. 

d.

7. It is a noble gas.

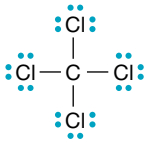
8. 

9. Advantages may include:
   - Electron dot structures can be used to work out which atoms bond together when given a molecular formula.
   - Electron dot structures show the number of bonding and non-bonding pairs of electrons.

10. 

2.2 Exam questions

1. A
2. B
3. Only iodine exists as I_2.
4. See table at the bottom of the page**
5. a. 4

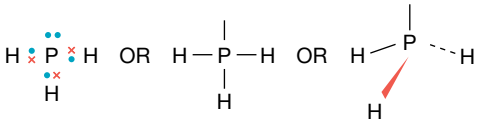
b. 

c. 4

d. 12

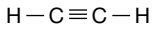
2.3 Shapes of molecules

Practice problem 4



Note: The distribution of electron pairs around a central atom and molecular shape are not necessarily the same.

Practice problem 5



Practice problem 6

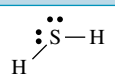
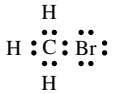
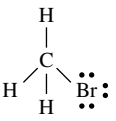
a. The bonds are non-polar.

b. The bonds are polar.

*5.

Group	Element	Valence electrons	Electron dot diagram	Lone pairs	Bonding electrons
16	S	6		2	2
14	Si	4		0	4
15	P	5		1	3
17	Cl	7		3	1
17	Br	7		3	1
16	Se	6		2	2
18	Ar	8		4	0

**4.

Molecule	Electron dot structure	Valence structure	Shape
H_2S			v-shaped
CH_3Br			tetrahedral

Practice problem 7

a. Non-polar **b.** Polar

2.3 Exercise

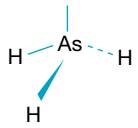
1. C

2. B

3. Non-bonding pairs repel each other.

4. HF, FCl and HCl

5.



6. a. Polar **b.** Non-polar

7. a. Non-polar **b.** Polar

8. a. $\overset{\delta+}{H}-\overset{\delta-}{F}$ **b.** $\overset{\delta-}{O}-\overset{\delta+}{H}$ **c.** $\overset{\delta-}{C}-\overset{\delta+}{H}$

d. $\overset{\delta-}{N}-\overset{\delta+}{H}$ **e.** C — C

9. Non-polar: *a*, *e*, polar: *b*, *c*, *d*.

10. No, tetrahedral molecules can be polar and non-polar.

2.3 Exam questions

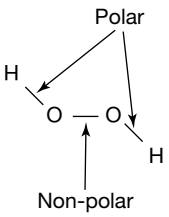
1. D

2. A

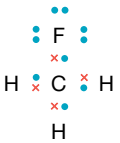
3. a.

b. Four lone pairs, two pairs per oxygen as shown in the dot diagram.

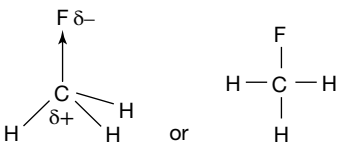
c.



4. a.



b.

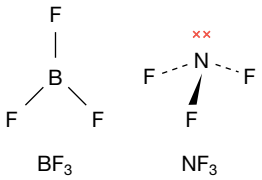


c. Tetrahedral

d. C–H are only slightly polar; C–F is very polar.

e. Polar

5. a.



b. Boron has only six electrons, not the eight required by the octet rule.

c. BF_3 is planar; NF_3 is pyramidal.

d–e. Please see online solutions document for a detailed response.

2.4 Comparing intramolecular bonding and intermolecular forces

Practice problem 8

a. CI_4 is a non-polar tetrahedral molecule. Only dispersion forces exist between the molecules.

b. CH_3NH_2 has a tetrahedral shape around the carbon atom and a trigonal planar shape around the second central atom, nitrogen, resulting in a polar molecule. It also has a highly electronegativity atom (nitrogen) bonded to hydrogen atoms. So, dispersion forces and hydrogen bonding exist between these molecules.

Practice problem 9

a. HF **b.** HI **c.** HF **d.** HF

2.4 Exercise

1. C

2. D

3. D

4. A

5. a. No **b.** CF_4 **c.** CBr_4 **d.** CBr_4

6. Boiling points should be about the same.

7. CH_3F has dispersion and dipole–dipole forces acting between its molecules. CH_3OH has dispersion forces and the stronger hydrogen bonding between molecules. CH_3OH has the higher boiling point.

2.4 Exam questions

1. C

2. B

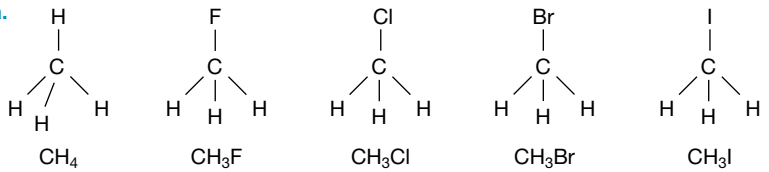
3. To boil water, the much weaker intermolecular forces are broken. However, to decompose a molecule, the strong covalent bonds within the molecule need to be broken.

4. The boiling temperature is not related to the bonds within a molecule (intramolecular) but the bonds between one molecule and another (intermolecular). Covalent bonds refer to intramolecular bonds.

5. a. See figure at the bottom of the page*

b. Covalent intramolecular bonding in each

*5. a.



c. Dipole–dipole, dispersion intermolecular forces for each

d. They are gases at room temperature, except CH_3I which is a liquid.

e. Please see worked solutions in your online resources.

2.5 Physical properties of molecular substances

2.5 Exercise

1. B
2. A
3. D
4. B
5. a. Dispersion forces
   b. Dispersion forces and hydrogen bonding
   c. Dispersion forces
   d. Dispersion forces and hydrogen bonding
   e. Dispersion forces and dipole–dipole attractions.
6. Water has an extra form of intermolecular bonding (hydrogen bonding) between its molecules that H_2S does not.
7. Although dispersion and dipole–dipole intermolecular forces exist in all three substances, as the number of electrons in each molecule increases (from 18 in HCl to 36 in HBr and 54 in HI), so does the strength of the dispersion forces. This is reflected in the increasing boiling points.
8. Br_2 is a non-polar molecule and has only dispersion forces between its molecules; ICl is a polar molecule and has dipole–dipole forces as well as dispersion forces between its molecules. This extra force causes the higher melting point because more heat energy is required to separate the molecules.
9. HF has hydrogen bonding, which is stronger than other dipole–dipole forces.
10. They are all gases because only weak dispersion forces act between the molecules.

2.5 Exam questions

1. B
2. B
3. D
4. a. C_2H_5Cl, C_2H_6O, CO_2
   b. Polar, polar, non-polar
   c. O—H
   d. Covalent bonding
   e. i. Dipole–dipole interaction and dispersion forces
      ii. Hydrogen bonding and dispersion forces
      iii. Dispersion forces
   f. Ethanol
   Hydrogen bonding is stronger than the intermolecular bonding of the other two molecules.
5. a. See table at the bottom of the page*
   b. Ammonia would have the highest boiling point because it has hydrogen bonding, which is stronger than other intermolecular bonding types.
   c. Ammonia would have the weakest dispersion forces since it is the smallest molecule of the three.
   d. None

2.6 The structure and bonding of diamond and graphite

Practice problem 10

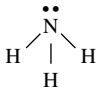
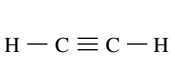
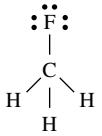
a. The bonds around each silicon atom will be covalent and tetrahedral.

b. The bonds are longer in silicon than in diamond.

2.6 Exercise

1. Every carbon atom has four outershell electrons that can be used to form strong covalent bonds with other atoms. In doing so, carbon can form molecules containing single, double and triple bonds as well as straight chain, branched chain and cyclic structures. Carbon can also form layer and network structures.

*5. a.

Compound	Ammonia, NH_3	Ethyne, C_2H_2	Fluoromethane, CH_3F
Valence structure (including non-bonding electrons)			
Shape	Pyramidal (or triangular pyramid)	Linear	Tetrahedral
Polarity: Is the molecule polar or non-polar?	Polar	Non-polar	Polar
Type/s of intramolecular bonding	Covalent	Covalent	Covalent
Type/s of intermolecular bonding	Hydrogen bonding (and dispersion forces)	Dispersion forces	Dipole–dipole interaction (and dispersion forces)

2. In diamond, each carbon atom is covalently bonded to four other carbon atoms in a three-dimensional network lattice. In graphite, each carbon atom forms covalent bonds with three other carbon atoms to form layers of hexagonal sheets. These layers are stacked on top of each other and are held together by dispersion forces. The distance between the layers is much greater than the carbon to carbon bond length in diamond. Hence, the atoms in diamond, being packed more closely than those in graphite, give diamond a greater density.

3. a. The weak dispersion forces between the layers of hexagonal sheets of carbon atoms in graphite allow the carbon atoms to be readily layered onto paper to form a mark.

   b. In diamond, each carbon atom is covalently bonded to four other carbon atoms in a three-dimensional network lattice. Hence, diamond is an exceptionally hard and strong material and is, therefore, useful for drilling.

4. In graphite, each carbon atom forms covalent bonds with three other carbon atoms to form layers of hexagonal sheets. One electron from each carbon atom is delocalised and is free to move across the sheets of carbon atoms. It can therefore conduct electricity. In diamond, each carbon atom is covalently bonded to four other carbon atoms in a three-dimensional network lattice; that is, all of carbon's outershell electrons are localised in covalent bonds and no electrons are free to conduct electricity.

5. a. Silicon atoms will be covalently bonded to four carbon atoms, which in turn will be bonded to four silicon atoms.

   This alternating pattern will continue throughout the crystal. In this way, the overall formula will be SiC (1:1 ratio). The arrangement of bonds around each carbon and each silicon atom will be tetrahedral.

   b. The bonds are longer in silicon dioxide than in diamond as silicon is in period 3 while carbon is in period 2.

2.6 Exam questions

1. D
2. $C(s) \rightarrow C(g)$
3. Due to each carbon atom bonding to only three others, the remaining unpaired electron becomes delocalised. Delocalised electrons provide movement of charge for conductivity and also interact with light to give a shiny material.
4. B
5. The properties of graphite and graphene depend on their respective structures. Because their structures vary considerably, so do their properties. In graphite, each carbon atom forms covalent bonds with three other carbon atoms to form layers of hexagonal sheets. These layers are stacked on top of each other and are held together by dispersion forces. By comparison, graphene is flat layer of carbon hexagonal sheets that is only one atom thick. There are no weak dispersion forces holding layers together, meaning that all bonding within the network is strong.

2.7 Review

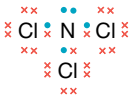
2.7 Review questions

1. Ball-and-stick
2. When bonds form, the atoms involved gain stability.
3. CaO
4. Low melting and boiling temperatures, do not conduct electricity, are soft
5. a. 14 b. 18 c. 17 d. 15
6. a.

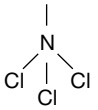
Electron dot diagram

Structural formula

b.

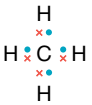


Electron dot diagram

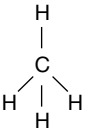


Structural formula

c.



Electron dot diagram



Structural formula

d.

Electron dot diagram

Structural formula

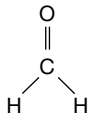
e.

Electron dot diagram

Structural formula

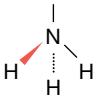
f.

Electron dot diagram



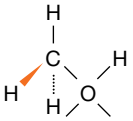
Structural formula

7. a.



N has a single lone pair, creating a pyramidal shape.

b.

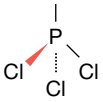


C has no lone pairs, so its structure is tetrahedral around the C atom; oxygen has two lone pairs, and hence the structure is V-shaped around the O atom.

c. $H-C\equiv N$

The triple bond between C and N creates a linear structure.

d.



P has a single lone pair, creating a pyramidal shape.

e. $-N \equiv N-$
The triple bond between the N atoms creates a linear structure.

8. Polar: NH_3, CH_3OH, HCN, PCl_3, H_2S, NF_3
Non-polar: N_2, C_2Br_4

9. Ammonia, NH_3, has a molecular dipole due to its pyramidal shape. Carbon dioxide, CO_2, is non-polar because its shape is linear and the bond dipoles cancel each other out.

10. **a** and **b** do not exhibit hydrogen bonding because the hydrogen atoms are not bonded to highly electronegative atoms. In both, hydrogen is bonded to C.

c and **d** exhibit hydrogen bonding because hydrogen atoms are bonded to the highly electronegative atoms nitrogen and fluorine respectively.

2.7 Exam questions

Section A — Multiple choice questions

1. B	**2.** A	**3.** A	**4.** A	**5.** D
6. B	**7.** C	**8.** A	**9.** A	**10.** B

Section B — Short answer questions

11. a. Iodine has more electrons and hence stronger dispersion forces.

b. Fluorine is a gas at room temperature due to its weak dispersion forces. Iodine is a solid at room temperature because it has much stronger dispersion forces.

12. a.

b. CO_2 is a linear molecule. The bond angles are 180°.

c. Four lone pairs — two on each oxygen atom.

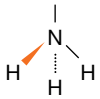
d. Polar. The bonds are polar due to the electronegativity difference between carbon and oxygen atoms.

e. Non-polar. The dipoles present in the bonds cancel out due to the linear shape (the molecule is symmetrical).

f. Dispersion forces. The molecule is non polar so can only have dispersion forces.

13. a.

The electron dot diagram shows that there are three bonding pairs and one lone pair. These four pairs will repel each other to a tetrahedral configuration. This guides the drawing of the structural formula, which indicates that the shape is pyramidal.



b. Ammonia is a polar molecule because it is asymmetrical, and the 'H' end of the molecule is slightly positive while the 'N' end of the molecule is slightly negative.

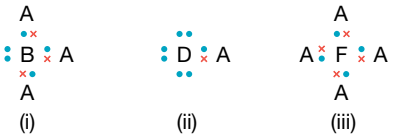
c. While dispersion forces and hydrogen bonds exist between ammonia molecules, ammonia exists as a gas because, at room temperature, these forces of attraction can be easily overcome by the kinetic energy (energy of motion) of the ammonia molecules.

14. a. Elements D and E are in same group (group 17) because they have seven electrons in their outershells.

b. C would represent any group 16 element (e.g. O) because it has six electrons in its outer shell.

c. A–C bonds will be the most polar because the largest difference in electronegativity exists between A and C.

d.



(i) (ii) (iii)

e. i. BA_3 and AD will be polar. (Bonds are polar and molecules are not symmetrical.)

ii. FA_4 will be non-polar. (The molecule is symmetrical so any dipoles will cancel.)

f. BA_3 is pyramidal. (There is a lone pair on the central atom.)

15. a. The type of bonding in each substance is covalent network bonding.

b. Germanium would be expected to be hard. This is because the strong covalent bonds between its atoms are hard to disrupt.

c. C is diamond because the dark grey colour is consistent with it turning into graphite. Of the others, the melting temperature represents the increasing strength of their covalent bonds. This is caused by the bonds becoming shorter, due to the period on the periodic table to which each belongs. Therefore, D is Ge (longest and weakest bonds), B is silicon and A is silicon carbide (shortest bonds).

3 Reactions of metals

3.2 Properties and structure of metals

3.2 Exercise

1. D

2. C

3. a. $1s^2 2s^1$.

b. Li^+, $1s^2$

c. Solid lithium contains lithium cations in a 'sea' of delocalised valence electrons. Non-directional electrostatic forces of attraction hold the lattice together.

4. a. Iron is hard because its structure is very dense and the electrostatic attraction between the cations and delocalised electrons is very strong.

b. When iron is hammered into a sheet (malleable) or drawn into a wire (ductile), layers of atoms move past one another without disrupting the force between the cations and the negative 'sea' of electrons.

c. Iron can conduct electricity because under the force of an electric field, the 'sea' of delocalised, outershell electrons is mobile whether the metal is in solid or molten form.

5. a. The metallic lattice consists of magnesium cations, Mg^{2+}, in a 'sea' of delocalised valence electrons.

b. The delocalised valence electrons within the lattice can reflect light.

6. a. Models are used to explain complex concepts in a simple way.

b. Limitations include inability to explain magnetism or the difference in hardness.

7. Aluminium is easily moulded and has an inert aluminium oxide coating which is insoluble in water.
8. Gold and silver are used for jewellery because they are soft enough to be easily worked, are lustrous and are unreactive so will not corrode easily.
9. a. Iron (steel) — it has high strength and is cheap.
   b. Aluminium — it is strong and resistant to corrosion.
   c. Copper — it is ductile and a good conductor of electricity.
   d. Lead — it is soft and dense.
10. Aluminium is cheaper and lighter than copper.

3.2 Exam questions

1. B
2. D
3. *N* and *O*
4. Any two of each of the following:
   a. Light, strong, malleable
   b. Good conductor of heat, malleable, high melting point
   c. Lustrous, light, malleable
   d. Dense, strong (corrosion resistant)
   e. Light, strong, malleable (corrosion resistant)
5. a. Lighter (three times)
   b. The sea of delocalised electrons
   c. Ductility

3.3 Reactivity of metals

Practice problem 1

$2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$

Practice problem 2

a. Hydrogen gas
b. Magnesium nitrate
c. $2HNO_3(aq) + 2Mg(s) \rightarrow 2MgNO_3(aq) + H_2(g)$
d. All of the nitric acid had been used up.

Practice problem 3

Calcium must be stored in oil to prevent it from reacting with the oxygen in the air, the oil acts as a barrier preventing a reaction from occurring. For a reaction to occur, this barrier must be removed.

3.3 Exercise

1. C
2. B
3. C
4. D
5. B
6. Electrons are lost in an oxidation reaction. In this reaction, electrons are donated to the water by the calcium metal and the calcium becomes calcium ions.
7. a. $4K(s) + O_2(g) \rightarrow 2K_2O(s)$
   b. $2Cs(s) + 2H_2O(l) \rightarrow 2Cs(OH)(aq) + H_2(g)$
   c. $Sn(s) + 2HNO_3(aq) \rightarrow Sn(NO_3)_2(aq) + H_2(g)$
8. They react with moisture and oxygen in the air.
9. $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
10. 1. Add a very small piece of a metal to a test tube or small beaker
    2. Add 20 mL of dilute hydrochloric acid using a measuring cylinder.
    3. Repeat with the other two metals
    4. Observe the rate of bubbles that form

    The most reactive metal will produce more bubbles more quickly.

3.3 Exam questions

1. B
2. C
3. B
4. a. Copper
   b. Silver
   c. Potassium
   d. Calcium
   e. Lead
5. a. Any two of immediate reaction, bubbles of hydrogen produced, alkali formed
   b. Any two of fewer bubbles for calcium, slower reaction for calcium, calcium hydroxide formed with calcium and sodium hydroxide formed with sodium. Container may feel warmer for sodium.
   c. Pop test for hydrogen, use indicator — phenolphthalein goes crimson if alkali present.

3.4 Recycling metals

3.4 Exercise

1. Metals such as silver, gold and platinum are unreactive and so are found in the uncombined state, whereas most other metals are more reactive and are found in a combined form and so have to be extracted before use.
2. Ores are rocks that are mixtures of different compounds and from which it is economically viable to separate out the metal. Copper is extracted from the ore chalcopyrite.
3. C
4. a. Advantages: For example: contribute to the nation's economy, source of jobs, use in making products and construction
   b. i. Waste rock: Material left over when metal ore is removed; can be used in road construction.
      ii. Raw materials: Raw materials are needed to manufacture the metal but are often limited so processes should be developed that minimise their use and products should be recycled where possible.
      iii. Land use: Mining, particularly if open cut mines, can be destructive to the environment. Use waste rock to refill and landscape with soil, regenerating the area when mining is complete.

5. a. Metals can be recycled because they can be repeatedly heated, melted and cooled without loss of structure or properties. In addition, because iron is magnetic, it is easier to separate from other metallic waste.

   b.
   - Collection of waste metal together with sorting processes require time and create transportation costs.
   - Limited quantities of metals in e-waste mean extra technology is required to separate the metals, which could be expensive.
   - Separating components of alloys can be difficult and expensive.
   - Health of workers in the recycling industry is at risk if contaminated or toxic waste is mixed in with other metal waste.

6. a. Bioleaching is using microbes to extract metals.

   b. It can be used to extract metals from metal ores, tailings, e-waste or contaminated water.

   c. Two from the following: It causes less damage to the environment, produces less damaging waste materials, and requires less energy than traditional methods of metal extraction because it operates at lower temperatures.

   d. Two from the following: It is slower than traditional methods, some acidic waste is produced, and the technology needs further development to operate at an industrial scale.

3.4 Exam questions

1. A
2. D
3. a. Urban mining is the process of obtaining metals and other useful substances from waste materials, particularly electronic waste.

   b. An example is obtaining gold, silver and copper from mobile phones.

   c. It is important because valuable resources are being depleted by being transformed into products that often end up in landfill and this is not environmentally sustainable.

4. a. Sample response:
   Linear:
   Extraction of ore → obtaining aluminium from ore using electrolysis → made into a can → distribution to sales outlets → use of can → put into waste

   Circular:
   Made into a can → distribution to sales outlets → use of can → recycling process

   b. Any two of the following:
   - Saves finite resources
   - Saves energy
   - Creates jobs in innovative design
   - Decreases waste
   - Reduces pollution during initial manufacturing processes
   - Can improve economy
   - Encourages innovation
   - Less damage to environment

5. a. For metals, 90 per cent is collected and recycled whereas only 15 per cent of plastics is collected and only 13 per cent of plastics is recycled. A small percentage of plastics is used to provide energy. Metals have the highest recycling rate and plastics have the lowest apart from textiles.

   b. These metals are widely used so it is more economical to collect them, or the processes to recycle these metals are well established.

3.5 Review

3.5 Review questions

1. Delocalised electrons interact with light to produce shiny materials. Their movement allows both electrical and thermal conductivity and also allows layers of the lattice to slide across each other.
2. Aluminium does not react with oxygen gas in air because its oxide forms a hard protective layer on the metal, preventing further reaction with oxygen.
3. D
4. a. (lithium) and e. (calcium)
5. *B*
6. a. Any two of the following:
   - Silver is hard, sodium is soft.
   - Silver is a very good conductor of electricity sodium is not as good.
   - Silver is unreactive and sodium is very reactive.
   - Silver is found in its natural state and sodium is obtained by electrolysis.
   - Silver has a higher melting point than sodium.
   - Silver is more dense than sodium.

   b. Both metals are lustrous and conduct electricity.

7. a. Group 2 period 4

   b. E.C.(Ca) = 2,8,8,2

   c. i. $2Ca(s) + O_2(g) \rightarrow 2CaO(s)$

   ii. $Ca(s) + 2H_2O(g) \rightarrow Ca(OH)_2(aq) + H_2(g)$

   iii. $Ca(s) + 2HCl(aq) \rightarrow Ca^{2+}(aq) + 2Cl^{-}(aq) + H_2(g)$

   d. To test for alkali, $Ca(OH)_2(aq)$, use the indicator phenolphthalein which changes colour from colourless to crimson.

   To test for hydrogen collect the gas in a test tube and place a glowing splint at the top, a pop sound will be heard if sufficient hydrogen gas is present.

8.
   - Contamination of waterways
   - Destruction of habitats
   - Air pollution

9. The actual process of recycling depends on the metal being recycled; these are general steps involved.
   1. Collection
   2. Sorting
   3. Processing
   4. Melting
   5. Purification/refining
   6. Cooling and transportation
   7. New products

10. a. Elements to include in flow chart are as follows:
   1. Mineral waste is crushed.
   2. Crushed waste is placed in a bioreactor with specific bacteria.
   3. Water and nutrients are added.

4. Bacteria transform the metal into a solution while gaining energy.
5. The acidic solution is treated to remove the metal.
6. The remaining acid is retained so as not to escape into the environment.

b. The e-waste could be collected and crushed. It is placed in a bioreactor with water and nutrients, along with the appropriate bacteria to catalyse the breakdown of gold. The bacteria cause the gold to go into solution. The solution is treated to extract the gold, which is then used to manufacture parts for new mobile phones.

3.5 Exam questions

Section A — Multiple choice questions

1. C	**2.** B	**3.** A	**4.** C	**5.** D
6. C	**7.** B	**8.** A	**9.** C	**10.** D

Section B — Short answer questions

11. a. An ore is a naturally occurring rock that contains metal in sufficient quantity to be economically mined.

b. *C*, then *A*, then *B*.

c. Metal *B*

d. Metals can be recycled without losing any of their properties.

e. Any two of the following:
- Recycling uses less energy to re-manufacture the metal
- Less waste generated to put in landfill
- Less damage to environment from mining and saves resources.

12. a. Silver, nickel, manganese, barium

b. $Mn(s) + 2HCl(aq) \rightarrow MnCl_2(aq) + H_2(g)$

c. Barium would be more reactive than calcium as it is lower in group 2, the valence electrons are further from the nucleus and so more easily lost.

d. $Ba(s) + 2H_2O(aq) \rightarrow Ba(OH)_2(aq) + H_2(g)$

e. Silver is unreactive so would most like to be found uncombined.

13. a. The reaction of metal with acids would be preferable because these metals are not group 1 metals, and so don't react with water, and they are not unreactive metals such as silver and gold. They are of medium reactivity, so adding acid may result in the production of bubbles, meaning the rate of bubble production or temperature change could be used.

b. Similar size, mass and shape pieces of zinc and lead, dilute hydrochloric acid, test tubes, test tube stand, spatula, measuring cylinder, electronic balance

c. A qualitative way would be to observe the rate of bubble formation. Alternatively, the temperature of the acid before and after could be measured for a quantitative measurement.

d. Wear lab coat and safety glasses when using acids, use a spatula so as not to handle metals that could be toxic.

14. a. Type of metal

b. Rate of bubble production

c. Same mass and shape of metal, the same concentration of sulfuric acid and the same volume of acid, same size containers and same temperature.

d. Sodium is a group 1 metal and would react violently with acid, which was not observed.

e. Metal *O* reacted most vigorously and so was the one that donated electrons most easily and so was most readily oxidised.

f. Metal *O* was the most reactive and so most likely to react with steam.

15. a. The microbes could be used to separate metals from the ore in the rocks on Mars instead of using the more energy-demanding and environmentally hazardous smelting processes. Nutrients and a bioreactor would also be necessary. (Forming the metals into useful materials would be your next problem to solve.)

b. Magnetic separation could be used to separate iron particles from the leachate.

c. Native iron refers to iron metal and it is less likely to corrode on Mars because there is so little oxygen (and water) to oxidise the metal.

d. Because the gravity on Mars is far lower than on Earth, the weight (but not the mass) of iron would be less and so it would be easier to work with.

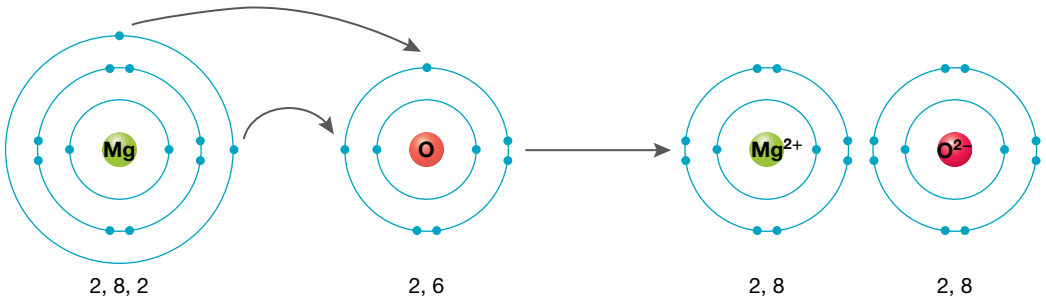
e. Aluminium is a more reactive metal than iron, so it requires more energy to separate it from its ore, if using electrolysis, than separating iron from its ore by smelting. This also depends on the availability of other raw materials required in the processes.

4 Reactions of ionic compounds

4.2 Structure and properties of ionic substances

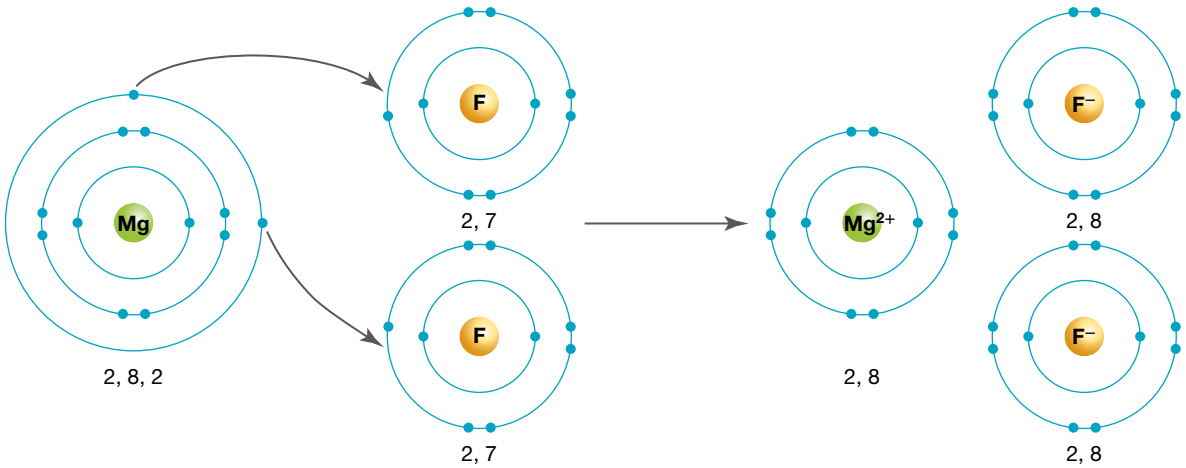
Practice problem 1

a.



$Mg^{2+} + O^{2-} \rightarrow MgO$

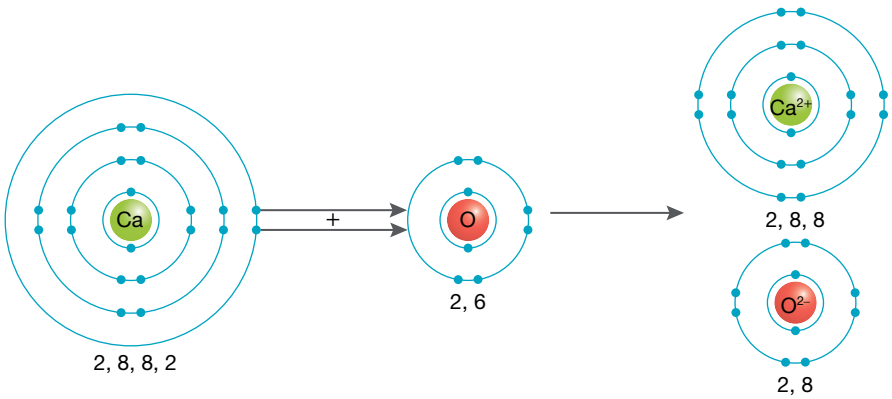
b.



$Mg^{2+} + 2F^{-} \rightarrow MgF^{2}$

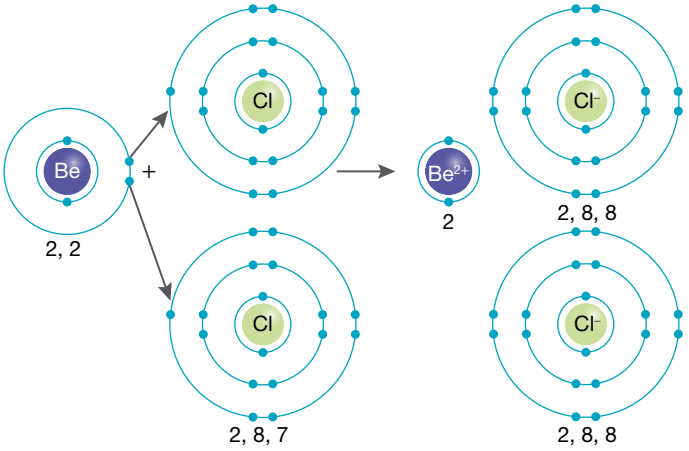
4.2 Exercise

1. Ca and Al each form an ion in order to attain a noble gas configuration and, hence, stability.
2. **a.**



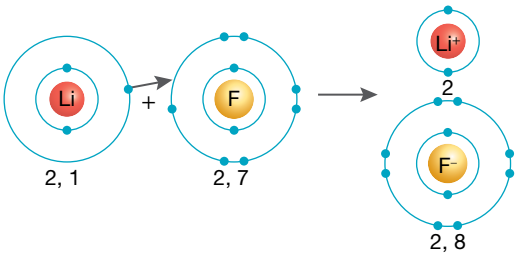
$Ca^{2+} + O^{2-} \rightarrow CaO$

b.



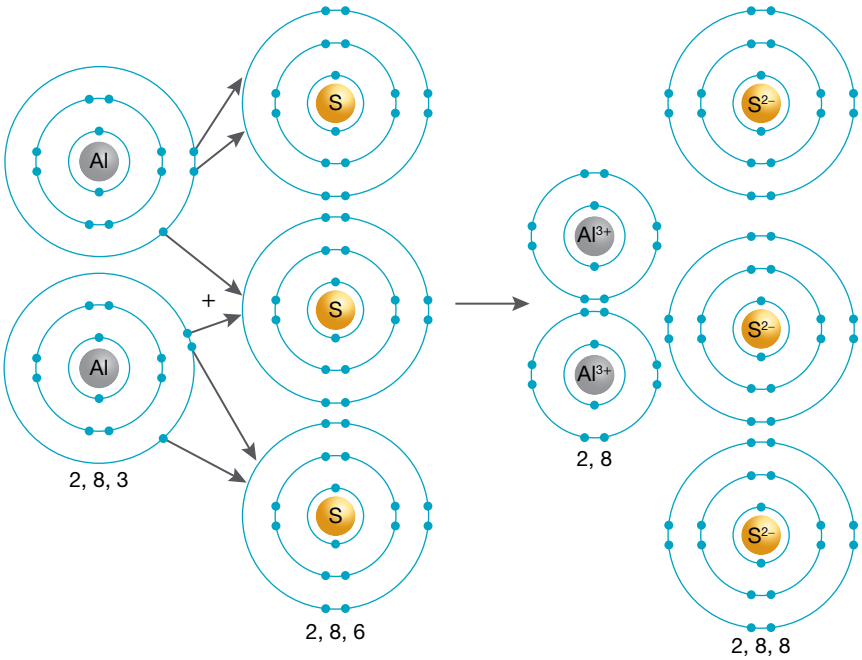
$$Be^{2+} + 2Cl^- \rightarrow BeCl_2$$

c.



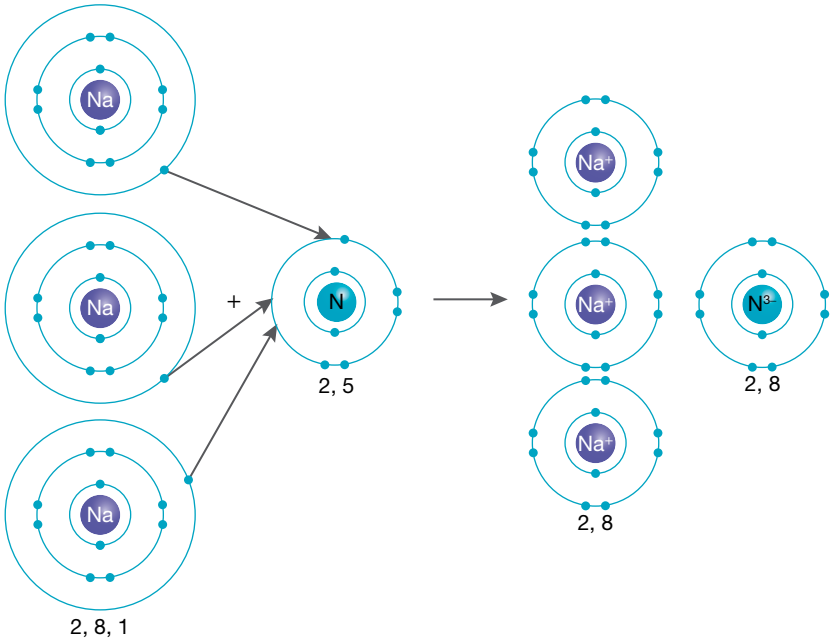
$$Li^+ + F^- \rightarrow LiF$$

d.



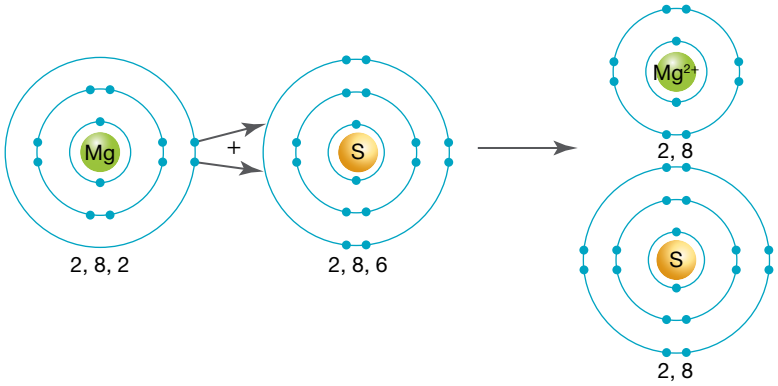
$$2Al^{3+} + 3S^{2-} \rightarrow Al_2S_3$$

e.



$3Na^+ + N^{3-} \rightarrow Na_3N$

f.



$Mg^{2+} + S^{2-} \rightarrow MgS$

3. a. The ions Mg^{2+} and $2Cl^-$ make up $MgCl_2$.
The ions K^+ and Cl^- make up KCl.

b. Two Cl^- ions are needed to balance the one Mg^{2+} ion.
One K^+ ion balances one Cl^- ion.

c. Ionic bonds

4. a. Calcium chloride has a high melting point because it is an ionic compound with strong electrostatic attractive forces holding the ions together in a three-dimensional lattice.

b. Calcium chloride will shatter when pressure is applied because ions of like charge repel when aligned with one another.

c. Calcium chloride will conduct electricity only in the liquid state where the ions are free to move. In the solid state, the ions are held strongly in place by the strong ionic bond and so will not conduct electricity as there are no mobile charged particles.

5. a. i. KCl has one K^+ ion for each Cl^- ion; therefore, the ratio is 1 : 1.

ii. CaO has one Ca^{2+} ion for each O^{2-} ion; therefore, the ratio is 1 : 1.

b. MgO has a higher melting point because the electrostatic attraction is greater between its ions due to their higher charges.

6. The substance is likely to be ionically bonded. Detailed responses can be found in the worked solutions in your digital formats.

4.2 Exam questions

1. B
2. C
3. A
4. • Relative size of ions
   • Ratio of cations to anions in lattice
5. The hammer blow causes a layer of ions to move out of alignment in the lattice, leading to repulsion between similarly charged ions.

4.3 Formation of ionic compounds

Practice problem 2

Strontium fluoride

Practice problem 3

a. Strontium iodide

b. V_2O_5

Practice problem 4

$Na_2Cr_2O_7$

4.3 Exercise

1. a. $AlCl_3$ b. BaO
   c. Na_2S d. Mg_3P_2
2. a.

Ions	K^+	Ca^{2+}	Al^{3+}
F^-	KF	CaF_2	AlF_3
O^{2-}	K_2O	CaO	Al_2O_3
N^{3-}	K_3N	Ca_3N_2	AlN

   b. KF: potassium fluoride
   CaF_2: calcium fluoride
   AlF_3: aluminium fluoride
   K_2O: potassium oxide
   CaO: calcium oxide
   Al_2O_3: aluminium oxide
   K_3N: potassium nitride
   Ca_3N_2: calcium nitride
   AlN: aluminium nitride
3. a. Potassium chloride
   b. Silver sulfide
   c. Sodium nitride
4. a. Iron(II) sulfide
   b. Iron(III) sulfide
   c. Copper(II) chloride
5. a. Tin(II) oxide
   b. Copper(I) phosphide
   c. Lead(II) bromide
6. a. SnF_4 b. PbS c. HgO
   d. FeN e. Cu_2S f. SnO
7. HSO_4^-, hydrogen sulfate
   HSO_3^-, hydrogen sulfite
   SO_4^{2-}, sulfate
   SO_3^{2-}, sulfite
   $S_2O_3^{2-}$, thiosulfate
8.

Cation	Anion	Empirical formula
Ag^+	Cl^-	AgCl
K^+	S^{2-}	K_2S
Mg^{2+}	O^{2-}	MgO
Al^{3+}	Br^-	$AlBr_3$
Fe^{3+}	CO_3^{2-}	$Fe_2(CO_3)_3$
Ba^{2+}	PO_4^{3-}	$Ba_3(PO_4)_2$
NH_4^+	SO_4^{2-}	$(NH_4)_2SO_4$

9. a. Na_2SO_3 b. NH_4NO_2 c. $Cu(HCO_3)_2$
10. a. $Al_2(CO_3)_3$, aluminium carbonate
    b. $NaNO_3$, sodium nitrate
    c. $Hg_3(PO_4)_2$, mercury(II) phosphate
    d. $PbSO_4$, lead(II) sulfate

4.3 Exam questions

1. C
2. C
3. D
4. a. i. Na_2CO_3
   ii. $BaSO_4$
   iii. $Fe_2(CO_3)_3$
   iv. NH_4NO_3
   v. Cu_2S
   b. i. Calcium phosphate
   ii. Vanadium(III) bromide
   iii. Iron(II) sulfate
   Note that the charge on transition metals is shown in roman numerals.
5. a. Ca^{2+} ions and F^- ions
   b. i. The nearest neighbours to calcium ions are fluoride ions. The nearest neighbours to fluoride ions are calcium ions.
   ii. The second closest neighbour to calcium is calcium. The second closest neighbour to fluoride is fluoride.

4.4 Precipitation reactions

Practice problem 5

$$RbBr(s) \xrightarrow{H_2O} Rb^+(aq) + Br^-(aq)$$

Practice problem 6

$MgCl_2$ is soluble.

Practice problem 7

$$2KOH(aq) + Ca(NO_3)_2(aq) \rightarrow 2KNO_3(aq) + Ca(OH)_2(s)$$

Practice problem 8

$$Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$$

Spectator ions: Na^+, Cl^-

4.4 Exercise

1. The solute is the dissolved substance in solution. The solvent is the substance that the solute dissolves in. For example, salt (solute) dissolves in water (solvent).
2. Water molecules have a polar arrangement of the oxygen and hydrogen atoms. One side (hydrogen) has a positive electrical charge and the other side (oxygen) has a negative charge. This can disrupt attractive forces that hold the ionic compounds together and thus dissolve it. Detailed responses can be found in the worked solutions in your digital formats.
3. The ionic bonds between the Li^+ and Br^- ions must be broken, and the hydrogen bonds between the water molecules must be broken. New ion–dipole interactions are formed between the ions and water.
4. a. When two solutions are mixed and a solid is formed.
   b. Two solutions are mixed together and a cloudy solid forms.
5. a. Insoluble
   b. Soluble
   c. Insoluble
   d. Soluble
6. a. Not a precipitation reaction because no precipitate formed.
   b. Not a precipitation reaction because two solutions have not been mixed to form this solid.
   c. This is a precipitate reaction because two solutions have been combined and a solid product has been formed.
7. a. There will be a precipitate of iron (lll) hydroxide.
   b. There will be a precipitate of barium sulfate.
   c. There will be no precipitate.
8. a. $2LiOH(aq) + Ca(NO_3)_2(aq) \rightarrow Ca(OH)_2(s) + 2LiNO_3(aq)$
   b. $Na_2S(aq) + (CH_3COO)_2Pb(aq) \rightarrow PbS(s) + 2(CH_3COO)Na(aq)$
   c. $2(NH_4)_3PO_4(aq) + 3CaCl_2(aq) \rightarrow Ca_3(PO_4)_2(s) + 6NH_4Cl(aq)$
9. a. $BaCl_2\,(aq) + MgSO_4\,(aq) \rightarrow BaSO_4\,(s) + MgCl_2\,(aq)$
   $Ba^{2+}\,(aq) + SO_4^{2-}\,(aq) \rightarrow BaSO_4\,(s)$
   Spectator ions: Mg^{2+}, Cl^-
   b. $Ca(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow 2NaNO_3(aq) + CaCO_3(s)$
   $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$
   Spectator ions: Na^+, NO_3^-
10. Silver nitrate solution
    $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

4.4 Exam questions

1. AgCl, $BaSO_4$, $MgCO_3$
2. C
3. a. $CoCl_2(aq) + 2NaOH(aq) \rightarrow Co(OH)_2(s) + 2NaCl(aq)$
   b. $Co^{2+}(aq) + 2OH^-(aq) \rightarrow Co(OH)_2(s)$
4. a. Dissociation is the breaking up or separation of an ionic compound into its component ions.
   b. Ionisation is the reaction of a molecular substance with water to produce ions.
   c. When sodium chloride is added to water, it dissociates into $Na^+(aq)$ and $Cl^-(aq)$ ions; i.e.
   $NaCl(s) \xrightarrow{H_2O(l)} Na^+(aq) + Cl^-(aq)$.
   Such solutions are good conductors of electricity because the ions are free to move.
   d. When hydrochloric acid is added to water, it is ionised into $H_3O^+(aq)$ and $Cl^-(aq)$ ions; i.e.
   $HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$.
   Such solutions are good conductors of electricity because the ions are free to move.
5. a. Test tube B (barium chloride)
   b. $BaCl_2(aq) + K_2CO_3(aq) \rightarrow BaCO_3(s) + 2KCl(aq)$
   c. $Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$
   d. Spectator ions are present in the reaction but do not take part in the formation of the product $BaCO_3$. Spectator ions are $Cl^-(aq)$ and K^+.
   e. KCl(s): As the water is evaporated, the solution will form solid crystals of potassium chloride.

4.5 A review of bonding

Practice problem 3

a. Substance *C* is metallic.
b. Substance *D* is covalent layer.

Practice problem 4

a. Copper: metallic bonding
b. Copper(II) oxide: ionic bonding
c. Oxygen gas: covalent bonding within its molecules and dispersion forces between them

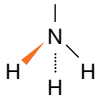
4.5 Exercise

1. *E*: covalent molecular bonding.
   F: covalent lattice.
2. a. Covalent molecular, due to its poor conductivity in its solid and liquid state.
   b. Metallic, due to its conductivity in solid and liquid form.
   c. Covalent molecular, due to its poor conductivity of electricity and heat.
   d. Ionic, due to its conductivity of electricity in its liquid, but not its solid state.
3. a. Substance *O*
   b. Substance *M*
   c. Substance *L*
   d. Substance *N*
4. a. Metallic bonding
   b. Ionic bonding
   c. Dispersion forces between atoms
   d. Covalent bonding
   e. Covalent bonding
5. a. Metallic bonding
   b. Covalent network lattice
   c. Covalent layer lattice
   d. Hydrogen bonding
   e. Ionic bonding

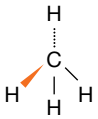
6. a. Covalent molecular bonding
   b. Covalent network
   c. Ionic network lattice
   d. Metallic lattice
   e. Metallic lattice
   f. Covalent molecular bonding

7. i. a. Carbon dioxide is a covalent molecular substance with weak dispersion forces between its molecules.
   b. O=C=O

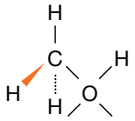
   ii. a. Ammonia is a covalent molecular substance with hydrogen bonding and dispersion forces between its molecules.
   b.



   iii. a. Methane is a covalent molecular substance with weak dispersion forces between its molecules.
   b.



   iv. a. Methanol is a covalent molecular substance with hydrogen bonding and dispersion forces between its molecules.
   b.



8. D. H–F has the strongest dipole because the electronegativity difference between hydrogen and fluorine is greater than any of the other electronegativity differences (it is a hydrogen bond).

9. a. Sodium chloride has ionic bonding, which is characterised by strong electrostatic attractions between oppositely charged ions extend throughout the lattice.
   b. Salt crystals are brittle because of their structure. Ionic substances are made of lattices of cations and anions. When pressure forces like charges to align, repulsion occurs and the crystal shatters.
   c. Salt is useful as a flavouring because it is soluble in water. Water molecules can move between the ions forming ion–dipole bonds with the ions, hence freeing them by disrupting the rigid ionic lattice.

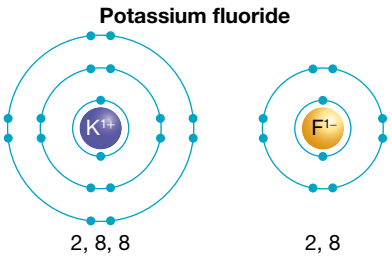
10. a. Gold is used in jewellery because it is shiny, durable, highly ductile and malleable. These properties are due to its metallic lattice. The non-directional nature of the metallic bond allows distortion without disrupting the lattice, while the lustre is due to light being reflected from the delocalised, outershell electrons.
    b. Graphite has a covalent layer lattice with weak dispersion forces between the layers. This accounts for its use as a lubricant; it is slippery because the layers slide readily over one another.

4.6 Review

4.6 Review questions

1. KCl is composed of K^+ and Cl^-. Potassium oxide is composed of K^+ and O^{2-}. Because two K^+ ions are necessary to balance the O^{2-} charge, the ratio of the different ions in potassium oxide is 2 : 1.

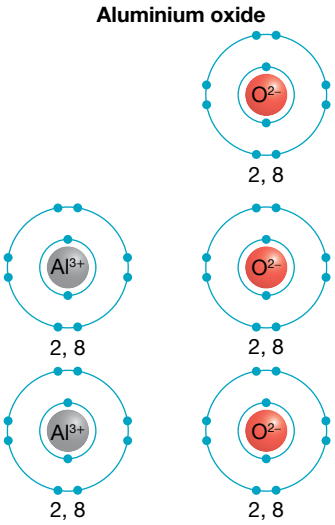
2. a.



$K^+ + F^- \rightarrow KF$

$2, 8, 8 + 2, 8$

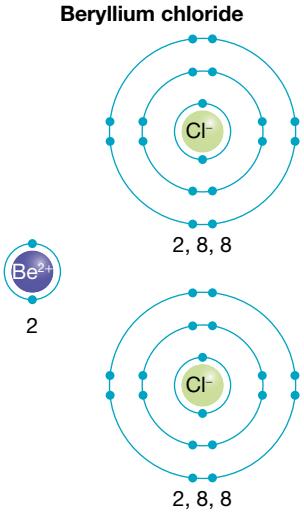
   b.



$2Al^{3+} + 3O^{2-} \rightarrow Al_2O_3$

$2, 8 + 2, 8$

   c.

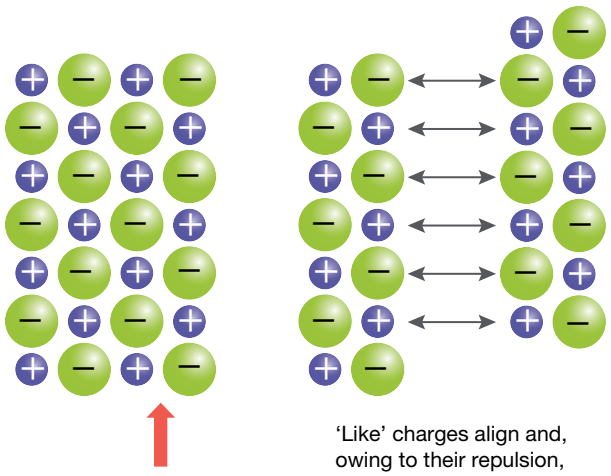


$Be^{2+} + 2Cl^- \rightarrow BeCl_2$

$2 + 2, 8, 8$

3. Ionic salts conduct electricity in the molten and aqueous states because the ions have dissociated and are now mobile. In the solid state, they are held rigidly by strong ionic bonds and, therefore, cannot conduct electricity.

4.



5. a. XY
   b. The bonding between X and Y must be ionic.
   c. $X^{2+} + Y^{2-} \rightarrow XY$
6. a. NaF
   b. $Ba(NO_3)_2$
   c. $Fe(OH)_3$
   d. NH_4OH
   e. Cr_2O_3
   f. Na_2HPO_4
7. a. Zinc chloride
   b. Aluminium carbonate
   c. Sodium sulfate
   d. Silver nitrate
   e. Sodium hydroxide
8. a. Yes, AgCl will precipitate.
   b. Yes, $BaSO_4$ will precipitate.
   c. No, the products would be NaCl and KNO_3 which are both soluble and so all of the ions would be dissociated.
   d. Yes, $Ca_3(PO_4)_2$ will precipitate.
   e. Yes, PbS will precipitate.
   f. No, all nitrate and ammonium compounds are soluble.
   g. No, most chlorides and bromides are soluble.
9. AgCl: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

   $BaSO_4$: $Ba^{2+}(aq) + SO_4{}^{2-}(aq) \rightarrow BaSO_4(s)$

   $Ca_3(PO_4)_2$: $3Ca^{2+}(aq) + 2PO_4{}^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$

   PbS: $Pb^{2+}(aq) + S^{2-}(aq) \rightarrow PbS(s)$
10. a. $K_2SO_4(aq) + Ca(NO_3)_2(aq) \rightarrow 2KNO_3(aq) + CaSO_4(s)$
    $SO_4{}^{2-}(aq) + Ca^{2+}(aq) \rightarrow CaSO_4(s)$
    b. $Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow 2KNO_3(aq) + PbI_2(s)$
    $Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$
    c. $3MgCl_2(aq) + 2Na_3PO_4(aq) \rightarrow 6NaCl(aq) + Mg_3(PO_4)_2(s)$
    $3Mg^{2+}(aq) + 2PO_4{}^{3-} \rightarrow Mg_3(PO_4)_2(s)$
    d. $K_2S(aq) + ZnCl_2(aq) \rightarrow ZnS(s) + 2KCl(aq)$
    $S^{2-}(aq) + Zn^{2+}(aq) \rightarrow ZnS(s)$
    e. No reaction

4.6 Exam questions

Section A — Multiple choice questions

1. A
2. C
3. D
4. B
5. A
6. D
7. B
8. A
9. B
10. D

Section B — Short answer questions

11. • An ionic lattice must maintain an alternating positive and negative ion arrangement; therefore, movement of layers cannot happen without repulsion breaking the lattice.
    • Metal lattices have the same charges throughout the lattice, allowing for flexibility.
12. a. $Mg(HCO_3)_2$
    b. $(NH_4)_2SO_4$
    c. $Ca(OH)_2$
13. Sodium chloride.
    Although potassium and rubidium ions have the same charge, the sodium ion is smaller, allowing greater electrostatic attraction between the ions and, hence, a higher melting temperature.
14. a. $AgNO_3(aq) + NH_4I(aq) \rightarrow AgI(s) + NH_4NO_3(aq)$
    b. $Zn(NO_3)_2(aq) + 2KOH(aq) \rightarrow Zn(OH)_2(s) + 2KNO_3(aq)$
15. a. $Na_2CO_3(aq) + CaI_2(aq) \rightarrow CaCO_3(s) + 2NaI(aq)$
    b. $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$

5 Separation and identification of the components of mixtures

5.2 Solutions, solvents and chromatography

Practice problem 1

Hexane, xylene or toluene

Practice problem 2

Compound *D*

Practice problem 3

Component *B*: 0.59

Component *C*: 0.41

5.2 Exercise

1. A
2. B

3. A solvent front can never travel slower than a component from the mixture.
4. a. Mobile phase: The mixture of solvent (or carrier gas) and components that move over the stationary phase.
   b. Stationary phase: The surface over which the mixture is made to move. It forms temporary bonds with the components of the mixture.
   c. Solvent: In paper chromatography or TLC, it is the liquid that dissolves the mixture to be separated.
5. Because it is through adsorption and desorption onto and from the surface that separation occurs.
6. a. Because pen ink contains dyes that might interfere with those being separated.
   b. So that all the mixture will be 'swept' up the paper; the sample does not just dissolve in the solvent without travelling up the paper.
   c. A different solvent could be used such that the spot dissolves.

5.2 Exam questions

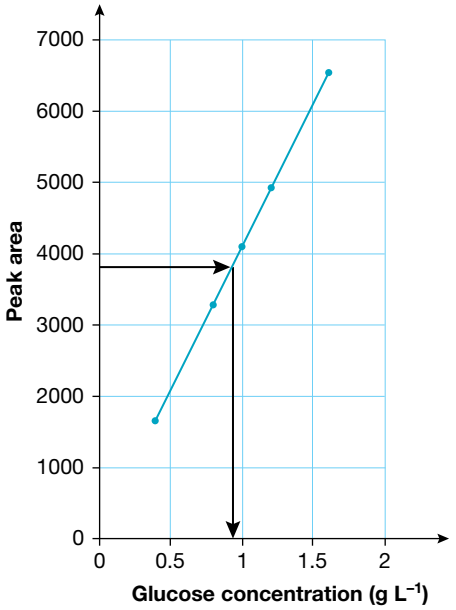
1. D
2. a. A: 0.33; B: 0.69; C: 0.60; D: 0.60.
   b. Lolly C and lolly D contain the same dye.
3. a. Two bands were expected but three were obtained.
   b. Calculate Rf values for each of the impurities and look for a match with Rf values calculated from the sample.
   c. Rf values for bands 1, 2 and 3 are 0.34. 0.50, 0.81 respectively. Rf values for A and B are 0.52 and 0.59 respectively. Therefore, impurity A is likely to be present.
4. a. Different colours experience different strengths of attraction to the stationary phase.
   b. i. The yellow components have different Rf values so they are different substances.
   ii. The blue components have the same Rf values so they are the same substance.
   iii. This is unlikely because the components are different.
5. 12.5 cm

5.3 Review

5.3 Review questions

1. The polarity of their molecules
2. Not possible to determine
3. a. Blue, pink, yellow, black. (Note that separation at this stage is poor.)
   b. Colours would be further separated
4. Paper:
   - Advantages: cheap, readily available, quick to set up.
   - Disadvantages: not as precise as TLC, cannot use corrosive solvents.

   TLC:
   - Advantages: quicker, greater range of stationary phases available, able to use solvents of a more corrosive nature.
   - Disadvantages: not as cheap as paper, more specialised equipment required (glass/metal slide rather than paper).
5. a.



   b. ≈ 0.9 g L^{-1}
   c. The glucose is within normal range.

5.3 Exam questions

Section A — Multiple choice questions

1. D	2. B	3. A	4. D	5. A
6. A	7. C	8. B	9. C	10. B

Section B — Short answer questions

11. The mobile phase contains the mixture to be analysed. This is the water containing the dye mixture.
12. The chromatogram for the red jelly beans should have erythrosine above green; the chromatogram for the purple jelly beans should have croceine orange above carmoisine.
13. a. A: 0.20; B: 0.50; C: 0.65; D: 0.80; E: 0.50; F: 0.63; G: 0.81.
   b. A
   c. G
   d. B and E (Note that D and G are also very close, suggesting further analysis is required.)
   e. The solvent front does not have to travel the same distance for comparison.
14. a. Isoleucine
   b. Taurine
   c. Arganine

15. a, b. See table at the bottom of the page*

b. It appears that chlorophyll a and chlorophyll b are replaced by anthocyanin.

c. Phaeophytin and xanthophyll and chlorophyll a and anthocyanin

d. Chlorophyll a and anthocyanin because they have different colours.

Unit 1 | Area of Study 1 review

Section A — Multiple choice questions

1. D	**2.** C	**3.** A	**4.** A	**5.** C
6. C	**7.** A	**8.** B	**9.** B	**10.** A
11. B	**12.** C	**13.** D	**14.** B	**15.** D
16. C	**17.** C	**18.** B	**19.** B	**20.** C

Section B — Short answer questions

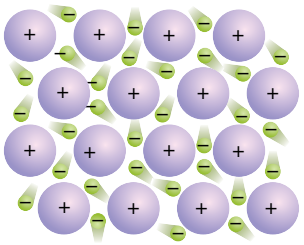
21. a. O **b.** At **c.** Sr **d.** Cl

22. a. Electronegativity is defined as the tendency of an atom to attract electron cloud density towards itself in a covalent bond.

b. Electronegativity increases across a period as the effective nuclear charge increases, while the separation between the nucleus and valence shell remains similar (or decreases), causing a greater attraction to electrons from other atoms.

c. The greater the separation between the nucleus and the valence shell (due to large atomic radius), the weaker the attraction, leading to lower electronegativity. A response may also refer to the increased shielding as more shells are occupied, reducing the overall attraction of the nucleus to valence electrons.

23. a.



b. Metals are malleable because they have non-directional bonding within the lattice, allowing for the layers to slide over each other. Ionic solids have a specific alternating arrangement of ions so any disruption to the alignment brings about repulsion, destroying the lattice structure.

c. Both metals and ionic compounds are held together by the electrostatic forces of attraction between oppositely charged particles.

24. See table at the bottom of the page**

25.

Bond type	Process
Covalent	Burning methane
Hydrogen bonding	Evaporating ethonal

26. a. i. Full equation: $Fe(NO_3)_2(aq) + K_2CO_3(aq) \rightarrow FeCO_3(s) + 2KNO_3(aq)$ Ionic equation: $Fe^{2+}(aq) + CO_3^{2-}(aq) \rightarrow FeCO_3(s)$

ii. Spectator ions: NO_3^- and K^+

b. i. Full equation: $2FeCl_3(aq) + 3K_2CO_3(aq) \rightarrow Fe_2(CO_3)_3(s) + 6KCl(aq)$

ii. Ionic equation: $2Fe^{3+}(aq) + 3CO_3^{2-}(aq) \rightarrow Fe_2(CO_3)_3(s)$

iii. Spectator ions: K^+ and Cl^-

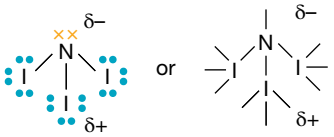
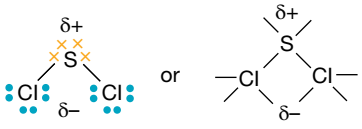
27. a. The R_f values are different because the mobile and stationary phases are different, resulting in different attractions for the components in the mixture.

***15. a, b**

Leaf colour	Colour of separated pigment	Distance of pigment from origin (cm)	Distance of solvent front from origin (cm)	Rf value	Pigment
Ggreen	Pale yellow	5.1	5.3	0.96	Carotene
	Dark yellow	4.1	5.2	0.79	Xanthophyll*
	Light green	3.2	5.3	0.60	Chlorophyll a
	Dark green	2.6	5.4	0.48	Chlorophyll *
Red	Pale yellow	7.2	7.5	0.96	Carotene
	Dark yellow	5.8	7.4	0.78	Xanthophyll*
	Red-blue	3.4	7.4	0.46	Anthocyanin*

*see parts c and d.

****24.**

Molecule	Nitrogen tri-iodide, NI_3	Sulfur dichloride, SCl_2
Structure of molecule showing shape and non-bonding electrons		
Polar or non-polar	Polar	Polar
Shape	Pyramidal	Angular/V-shaped/bent

b. Spot 1 could be either or both of *W* and *Y* because their R_f values are so similar.

c. Spot 2 is *Z*, and spot 3 is *X*. (This done by calculating and comparing R_f values).

d. Spot 1 did consist of two components, *W* and *Y*. Although not able to be separated in trial 1, they were able to be separated in trial 2.

6 Quantifying atoms and compounds

6.2 Relative isotopic mass and the carbon-12 scale

Practice problem 1

$A_r = 63.5$

Practice problem 2

a. 310.3 **b.** 195.8 **c.** 262.0

6.2 Exercise

1. See table at the bottom of the page*

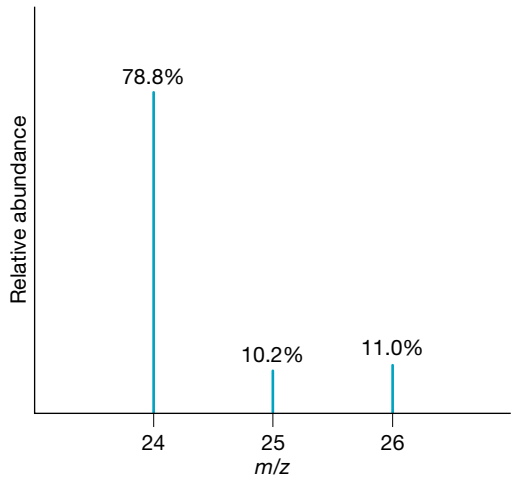
2. a. Indium isotope 1 = 112.9
Indium isotope 2 = 115.0

b. $^{113}_{49}\text{In}$

$^{115}_{49}\text{In}$

3. 108

4. a.



b. $A_r(\text{mg}) = 24.3$

5. a. 44.0 **b.** 58.5 **c.** 34.0
d. 98.1 **e.** 180.0

6. 71 (two significant figures)

7. The occurrence of heavier isotopes of carbon makes the relative atomic mass higher than 12.

8. The mass number of 63 is the sum of the number of protons and neutrons in the nucleus of the isotope of ^{63}Cu. The relative atomic mass of 62.9296 is the true mass of the isotope, as measured on the relative isotopic mass scale.

9. $A_r(\text{B}) = 11$ (two significant figures)

6.2 Exam questions

1.
- Mass number is the sum of protons and neutrons in an atom and must, therefore, be an integer.
- The relative atomic mass is both a comparison against the mass of ^{12}C as well as a weighted average of isotopes of that element and is, therefore, unlikely to be a whole number.

2. B

3. a. The three isotopes and their percentage abundances are 226 (65.0%), 228 (15.0%) and 230 (20.0%).

b. Relative atomic mass

$$= \frac{(226.0 \times 65.0) + (228 \times 15.0) + (230 \times 20.0)}{100}$$

$$= \frac{22\,710}{100}$$

$$= 227$$

4. Let x = % of lighter isotope.
$\therefore (100 - x)$ = % of heavier isotope.

$$\text{RAM} = \frac{(\text{RIM}_1 \times \%_1) + (\text{RIM}_2 \times \%_2)}{100}$$

$$\Rightarrow 121.8 = \frac{120.90x + 122.90(100 - x)}{100}$$

$$\Rightarrow 12180 = 120.90x + 12290 - 122.90x$$
$$\Rightarrow 12180 - 12290 = (120.90 - 122.90)x$$
$$\Rightarrow -110 = -2x$$
$$\Rightarrow x = 55.0\%$$

$\therefore$ % of lighter isotope 55.0%.

5. a. 7 isotopes

b. $^{152}_{62}\text{Sm}$, 27%

c. 144.0

6.3 Avogadro's constant and the mole

Practice problem 3

$M_r = 239.1\ \text{g mol}^{-1}$

*1.

Isotope	Element	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
$^{16}_{8}\text{O}$	Oxygen	8	16	8	8	8
$^{17}_{8}\text{O}$	Oxygen	8	17	8	9	8
$^{56}_{26}\text{Fe}$	Iron	26	56	26	30	26

Practice problem 4

1.2×10^{24} atoms of oxygen

6.3 Exercise

1. **a.** 48.0 g mol^{-1} **b.** 253.8 g mol^{-1}
   c. 56.1 g mol^{-1} **d.** 36.5 g mol^{-1}
2. C
3. **a.** 0.01000 mol **b.** 0.200 mol
   c. 0.100 mol **d.** 0.00500 mol
4. **a.** 0.500 mol **b.** 0.0100 mol
   c. 0.150 mol **d.** 0.166 mol
5. B
6. D
7. **a.** 1.57×10^{3} g **b.** 0.140 g
   c. 2.60×10^{-8} g **d.** 5.23×10^{6} g
   e. 5.53×10^{2} g
8. A
9. **a.** 1.008 g mol^{-1}
   b. 1.67×10^{-24} g
   c. This is an average mass because the relative atomic mass of hydrogen includes the mass of all the isotopes of hydrogen.
10. See table at the bottom of the page*

6.3 Exam questions

1. A
2. B
3. D
4. **a.** $N(CO_2) = n \times N_A = 4.50 \times 10^{-3} \times 6.02 \times 10^{23}$
   $= 2.71 \times 10^{21}$ CO_2 molecules
   b. $N(C) = 2.71 \times 10^{21}$ atoms, since one C atom is in each CO_2 molecule.
5. B

6.4 Using the mole concept

Practice problem 5

28%

Practice problem 6

K_2CO_3

Practice problem 7

a. $C_9H_{18}O$

b. $C_{18}H_{36}O_2$

6.4 Exercise

1. **a.** 84.2% **b.** 70.3%
   c. 32.8% **d.** 87.5%
2. VCl_3
3. **a.** By definition, an empirical formula is a ratio of atoms. It cannot be a ratio of masses because different types of atoms have different masses.
   b. A mole is just a counting unit for atoms. The ratio of moles will, therefore, be the same as the ratio of atoms.
   c. C_4H_{10}
4. C_3H_3O
5. C_8H_{18} (0.84 g C per g of fuel)
6. $Na_2B_4O_7$
7. $Fe_2(SO_4)_3$

*10.

	Substance	Amount (g)	Molar mass (M)	Number of atoms in the molecule	Number of moles (n) of substance	Number of molecules	Total number of atoms
1	Water, H_2O	3.2	18.0	3	0.18	1.1×10^{23}	3.3×10^{23}
2	Methane, CH_4	2.7	16.0	5	0.17	1.0×10^{23}	5.1×10^{23}
3	Ammonia, NH_3	0.056	17.0	4	0.0033	2.0×10^{21}	7.9×10^{21}
4	Ethanoic acid, CH_3COOH	1.64×10^{3}	60.0	8	27.3	1.64×10^{25}	1.31×10^{26}
5	Benzene, C_6H_6	44	78.0	12	0.56	3.4×10^{23}	4.0×10^{24}
6	Octane, C_8H_{18}	267	114.0	26	2.34	1.41×10^{24}	3.66×10^{25}
7	Ethanol, CH_3CH_2OH	4.6×10^{2}	46.0	9	10	6.0×10^{24}	5.4×10^{25}
8	Ozone, O_3	0.101	48.0	3	0.002 11	1.27×10^{21}	3.81×10^{21}
9	Sulfuric acid, H_2SO_4	7.0×10^{3}	98.0	7	71	4.3×10^{25}	3.0×10^{26}
10	Carbon dioxide, CO_2	1.8×10^{3}	44.0	3	42	2.5×10^{25}	7.5×10^{25}

8. a. $M((NH_4)_2SO_4) = (14.0 \times 2) + (1.0 \times 8) + (32.1 \times 1) + (16.0 \times 4)$
$= 132.1 \text{ g mol}^{-1}$
$\therefore$ 132.1 g will contain 28.0 g of N
$\therefore \frac{50 \times 132.1}{28.0} = 2.4 \times 10^2$ g (236 g) will be required to obtain 50 g of nitrogen

b. $M(NH_4NO_3) = (14.0 \times 2) + (1.0 \times 4) + (16.0 \times 3)$
$= 80.0 \text{ g mol}^{-1}$
$\therefore$ 80.0 g will contain 28.0 g of N
$\therefore \frac{50 \times 80.0}{28.0} = 1.4 \times 10^2$ g (143 g) will be required to obtain 50 g of nitrogen

c. $M((NH_4)_3PO_4) = (14.0 \times 3) + (1.0 \times 12) + (31.0 \times 1) + (16.0 \times 4)$
$= 149.0 \text{ g mol}^{-1}$
$\therefore$ 149.0 g will contain 42.0 g of N
$\therefore \frac{50 \times 149.0}{42.0} = 1.8 \times 10^2$ g (177 g) will be required to obtain 50 g of nitrogen

d. $M(CO(NH_2)_2) = (12.0 \times 1) + (16.0 \times 1) + (14.0 \times 2) + (1.0 \times 4)$
$= 60.0 \text{ g mol}^{-1}$
$\therefore$ 60.0 g will contain 28.0 g of N
$\therefore \frac{50 \times 60.0}{28} = 1.1 \times 10^2$ g (107 g) will be required to obtain 50 g of nitrogen

9. CH_4O

10. a. C_5H_{12} b. C_6H_{10} c. C_6H_8 d. C_6H_{12} e. $C_{10}H_{20}$ f. $C_6H_{12}O_6$ g. $C_2H_4Cl_2$

6.4 Exam questions

1. D

2. If %S – 23.44, %Gd = 100 – 23.44 = 76.56%.

	Gd	:	S
mass	76.56	:	23.44
mole	$\frac{76.56}{157.3}$	:	$\frac{23.44}{32.1}$
	= 0.4867	:	0.7302 (÷ by smallest number)
	= 1	:	1.5 ⇒ 2 : 3

$\therefore$ empirical formula = Gd_2S_3

3.

	Fe	:	O	:	H
mass	62.14	:	35.63	:	2.23
mole	$\frac{62.14}{55.8}$	:	$\frac{35.63}{16.0}$	:	$\frac{2.23}{1.0}$
	= 1.114	:	2.227	:	2.23 (÷ by smallest number)
	= 1	:	2	:	2

$\therefore$ empirical formula = $Fe(OH)_2$

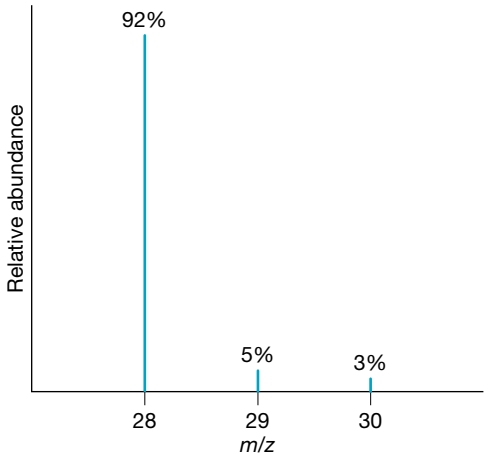
4. $C_4H_8O_2$

5. mass (Pb) = 1 tonne
$= 1 \times 10^6$ g
$= \frac{78.4}{100} \times$ mass (ore)
$\therefore$ mass (ore) $= 1 \times 10^6 \times \frac{100}{78.4}$ g
$= 1.28 \times 10^6$ g

6.5 Review

6.5 Review questions

1. a.



b. $Z = 14$
c. Approximately 28, 29 and 30
d. A_r (Si) = 28 (to 2 significant figures)

2. a. 6.0 mol nitrogen atoms
b. 24.0 mol hydrogen atoms
c. 8.0 mol oxygen atoms
d. 2.0 mol phosphorus atoms
e. 3.6×10^{24} nitrogen atoms
f. 84.0 g of nitrogen atoms

3. See table at the bottom of the page*

*3.

Name	Formula	Molar mass (M)	Mole (n)	Mass (m)
Sodium hydroxide	NaOH	40.0 g mol^{-1}	0.0853	3.41 g
Carbon tetrachloride	CCl_4	154.0 g mol^{-1}	1.40	216 g
Sodium carbonate	Na_2CO_3	106.0 g mol^{-1}	1.00	106 g
Potassium chloride	KCl	74.6 g mol^{-1}	0.25	9 g
Ammonium phosphate	$(NH_4)_3PO_4$	149.0 g mol^{-1}	0.0568	8.46 g

4. **a.** 8.93×10^3 g **b.** 4.5 g **c.** 194 g

5. **a.** 0.000 31 mol **b.** 1.9×10^{20} ions

6. **a.** 1.00×10^{22} atoms
   b. 3.16×10^{22} atoms

7. 73% C
   5.4% H
   22% O

8. U_3O_8

9. The abundance of the lighter isotope is 55%.

10. Molecular formula: $C_8H_{10}N_4O_2$

Section A — Multiple choice questions

1. B
2. D
3. D
4. A
5. A
6. D
7. C
8. B
9. A
10. D

Section B — Short answer questions

11. **a.** 3.4×10^{-4} mol
    b. 2.1×10^{20} molecules
    c. 5.0×10^4 g (50 kg)

12. **a.** **i.** $M\left(H_2C_2O_4\right) = 90.0\ \text{g mol}^{-1}$

$$n\left(H_2C_2O_4\right) = \frac{m}{M}$$
$$= \frac{6.53}{90.0}\ \text{mol}$$
$$= 7.26 \times 10^{-2}\ (0.0726)\ \text{mol}$$

ii. $N\left(H_2C_2O_4\right) = n \times N_A$
$= 7.26 \times 10^{-2} \times 6.02 \times 10^{23}$ molecules
$= 4.37 \times 10^{22}$ molecules

$N(O) = 4 \times N(H_2C_2O_4)$, since there are 4 O per $H_2C_2O_4$ molecule
$= 4 \times 4.37 \times 10^{22}$ atoms
$= 1.75 \times 10^{23}$ atoms

b. $M(X) = \frac{m}{n} = \frac{1.827}{0.0150}\ \text{g mol}^{-1}$
$= 121.8\ \text{g mol}^{-1}$
$\therefore$ antimony, Sb

13. Mass of H = 0.282 – (0.01600 + 0.228) = 0.038 g
See table at the bottom of the page*
The empirical formula is $C_{19}H_{38}O$.

14. Cu_2O

15. **a.** %(O) = 100 – (37.5 + 12.5) = 50.0%

	C	:	H	:	O
Mass	37.5	:	12.5	:	50.0
Mole	$\frac{37.5}{12.0}$	:	$\frac{12.5}{1.0}$	:	$\frac{50.0}{16.0}$
	= 3.125	:	12.5	:	3.125 (÷ by the smallest number)
	= 1	:	4	:	1

so CH_4O = empirical formula

b. $\frac{M(\text{compound})}{M(\text{empirical formula})} = \frac{128}{32} = 4$
$\therefore$ molecular formula = $C_4H_{16}O_4$

7 Families of organic compounds

7.2 Hydrocarbon families

Practice problem 1

$C_{16}H_{34}$

Practice problem 2

C_6H_{12}

7.2 Exercise

1. C
2. B
3. A
4. **a.** Alkanes have the general formula C_nH_{2n+2}. If 10 carbons are present, there will be (2 × 10) + 2 = 22 hydrogen atoms. The formula is $C_{10}H_{22}$.
   b. Alkenes have the general formula C_nH_{2n}. If 13 carbons are present, there will be 2 × 13 = 26 hydrogen atoms. The formula is $C_{13}H_{26}$.
5. **a.** $CH_2CHCH_2CH_2CH_3$ or $CH_2CH(CH_2)_2CH_3$
   b. $CH_3CHCHCH_3$
   c. $CH_3CH(CH_3)CH_2CH_3$
6. **a.** Carbon dioxide and water
   b. Carbon monoxide or carbon, and water

*13.

Step 1	C	H	O
Step 2	0.228 g	0.038	0.016 00
Step 3	$\frac{0.228}{12.0} = 0.0190$	$\frac{0.038}{1.0} = 0.038$	$\frac{0.01600}{16.0} = 0.00100$
Step 4	$\frac{0.0190}{0.00100} = 19$	$\frac{0.038}{0.00100} = 38$	$\frac{0.00100}{0.00100} = 1$
Step 5	19	38	1

7. a. Combustion and substitution reactions
   b. Combustion and addition reactions
8. As the intermolecular forces between molecules increase, the melting and boiling points increase. The intermolecular forces between alkanes are dispersion forces. Dispersion forces increase with molecular size (with an increase in the number of electrons in a molecule). So, as the molecular size of an alkane increases, the forces between molecules increases, meaning that the alkane will be more solid at room temperature than a smaller alkane, which has weaker dispersion forces.
9. Butane has a higher boiling point than propane because it is a larger molecule (contains more electrons) and, hence, stronger dispersion forces will exist between butane molecules.
10. Although both molecules have dispersion forces of about the same magnitude, molecules of hexane can pack closer together than molecules of 2,3-dimethylbutane. They, therefore, have a stronger effect in the former due to the molecules being closer. This means hexane has the higher boiling point.

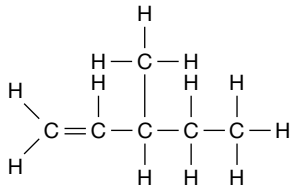
7.2 Exam questions

1. A
2. B
3. B
4. B
5. a. $CH_2C(CH_3)_2$ or $CH_2C(CH_3)CH_3$
   b. Alkenes
   c. 10
   Each successive member of a homologous series increases by CH_2 only.
   d. Unsaturated. It has a multiple bond between carbon atoms.
   e. This is a non-polar molecule, because little difference exists between the electronegativity of carbon and hydrogen, so the bonds are not polar, and the molecule does not have a positive and negative end.
   f. Dispersion forces only

7.3 Naming organic compounds and isomers

2,2,3-trimethylbutane

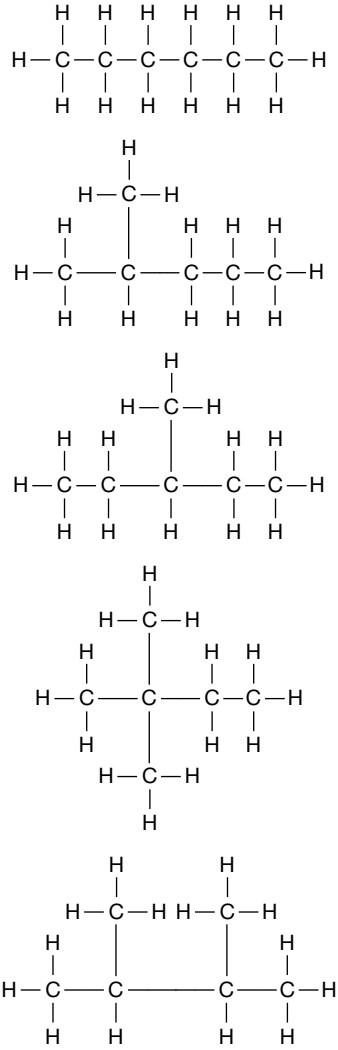
Practice problem 4



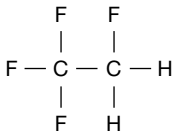
Practice problem 5

5-methylhept-2-ene

Practice problem 6



7.3 Exercise

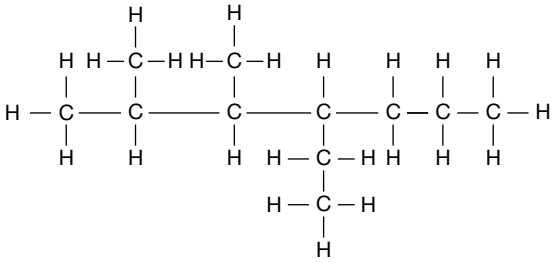
1. D
2. C
3. 
4. C

5. See figure at the bottom of the page*

6. a. The *only* position for the methyl group is carbon 2. Hence, numbers are not required.

   b. Propan-3-ol is incorrect because it has been numbered from the wrong end. It should be propan-1-ol.

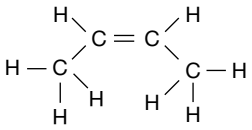
7. 3,3 dimethylhexane

8. 

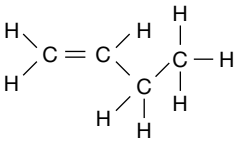
7.3 Exam questions

1. D

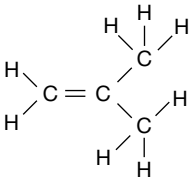
2. B

3. 

   but-2-ene (or 2-butene)



but-1-ene (or 1-butene)



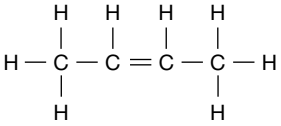
methylpropene (or 2-methylpropene)

4. a. Numbers must be assigned from the end that gives the lowest number to the double bond.

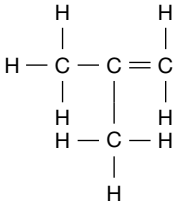
   b. But-1-ene

   c. $CH_3CH_2CHCH_2$

   d. Any one of the following.



2-butene



2-methylpropene

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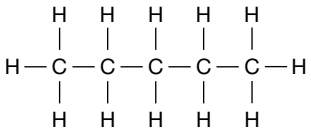
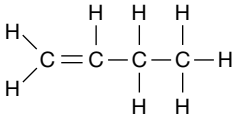
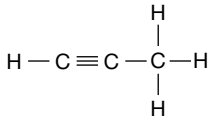
*5.

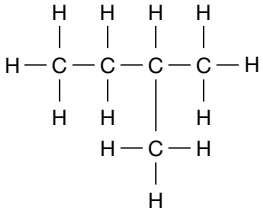
	(i) pentane	(ii) butene	(iii) propyne
(a)	3 isomers	3 isomers	1 isomer

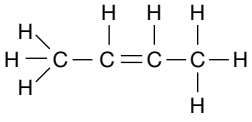
(b)

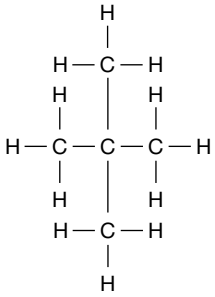


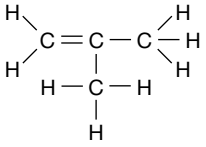


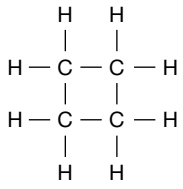




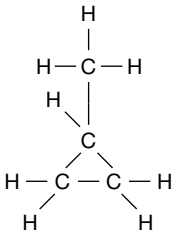








Cyclobutane



Methylcyclopropane

5. **a.** See figure at the bottom of the page*

   b. i. 2,2,4,4-tetramethylheptane

   ii. Octane

   c. a. i. $CH_3CHCHCH_2CH_2CH_3$

   ii. $CH_2C(CH_3)CH_3$

   iii. $CH_3CCCH_2CH_3$

   b. i. $CH_3C(CH_3)_2CH_2C(CH_3)_2CH_2CH_2CH_3$

   ii. $CH_3(CH_2)_6CH_3$

7.4 Functional groups — alcohols and carboxylic acids

Practice problem 7

a. Octan-1-ol

b. Propan-1-ol

Practice problem 8

a. 2-methylpentan-3-ol

b. $CH_3CH(CH_3)CH(OH)CH_2CH_3$

Practice problem 9

a. The longest carbon chain is five carbons long, with the carboxyl group attached to the first carbon. A methyl group is also attached to the third carbon, so the name of the molecule is 3-methylpentanoic acid.

b. $CH_3CH_2CH(CH_3)CH_2COOH$

Practice problem 10

a. Methyl propanoate: $CH_3CH_2COOH + CH_3OH \rightarrow CH_3CH_2COOCH_3 + H_2O$; $C_4H_8O_2$

b. Propyl methanoate: $HCOOH + CH_3CH_2CH_2OH \rightarrow HCOOCH_2CH_2CH_3 + H_2O$; $C_4H_8O_2$

c. Ethyl ethanoate: $CH_3COOH + CH_3CH_2OH \rightarrow CH_3COOCH_2CH_3 + H_2O$; $C_4H_8O_2$

d. Isomers are molecules that have the same molecular formula but different structures. All of the esters listed are isomers with the molecular formula $C_4H_8O_2$.

7.4 Exercise

1. $CH_3CH_2CH_2CH_2CH_2CH_2OH$ or $CH_3(CH_2)_6OH$. Alternately, the hydroxyl functional group (OH) could be swapped with any other hydrogen along the chain.
2. $CH_3CH_2CH_2CH_2CH_2COOH$ or $CH_3(CH_2)_4COOH$
3. A functional group is an atom, group of atoms or a bond that, when inserted into a molecule, gives that molecule a characteristic set of properties. Two examples of functional groups are the hydroxyl group (—OH) and the carboxyl group (—COOH).
4. **a.** Methylpropan-2-ol

   b. $CH_3C(CH_3)(OH)CH_3$

   c. The CH_3 group does not need a number because it can only go on carbon 2. (If it is placed on carbon 1 or carbon 3, it makes a butanol.) However, the —OH group can go on either carbon 1 or carbon 2; hence, the number must be specified.
5. **a.** 3,4-dimethylpentanoic acid

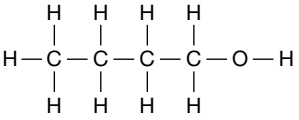
   b. $CH_3CH(CH_3)CH(CH_3)CH_2COOH$
6. **a.** A homologous series is a series of similar compounds, containing the same functional group and general formula, where the formulas differ by CH_2 for successive members.

   b. The alcohols are a homologous series because the formulas of successive members differ by CH_2. For example, the first three members are methanol, ethanol and propanol. The respective formulas of these are CH_3OH, CH_3CH_2OH and $CH_3CH_2CH_2OH$.

   c. The carboxylic acids are a homologous series because the formulas of successive members differ by CH_2. For example, the first three members are methanoic acid, ethanoic acid and propanoic acid. The respective formulas of these are HCOOH, CH_3COOH, CH_3CH_2COOH.

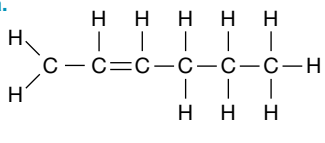
   d. The alkanes are a homologous series because the formulas of successive members differ by CH_2. For example, the first three members are methane, ethane and propane. The respective formulas of these are CH_4, C_2H_6 and C_3H_8. They have the general formula C_nH_{2n+2}.
7. **a.** Molecular formula: $C_4H_{10}O$

   Structural formula:

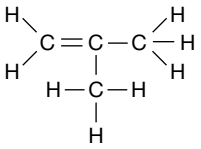


   Condensed formula: $CH_3CH_2CH_2CH_2OH$

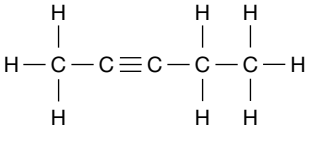
*5. a.



(i)



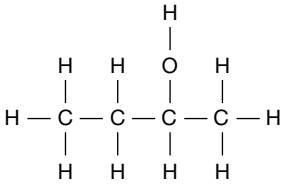
(ii)



(iii)

b. Molecular formula: $C_4H_{10}O$

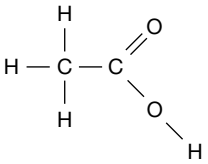
Structural formula:



Condensed formula: $CH_3CH_2CH(OH)CH_3$

8. a. Molecular formula: $C_2H_4O_2$

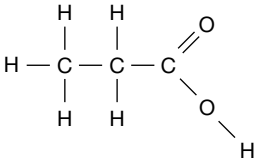
Structural formula:



Condensed formula: CH_3COOH

b. Molecular formula: $C_3H_6O_2$

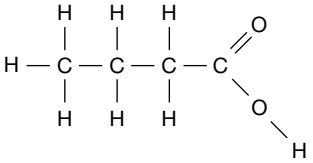
Structural formula:



Condensed formula: CH_3CH_2COOH

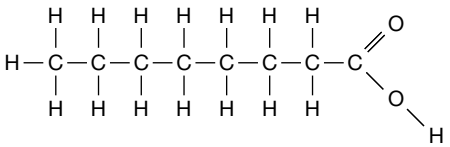
c. Molecular formula: $C_4H_8O_2$

Structural formula:



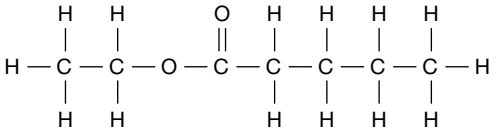
Condensed: $CH_3CH_2CH_2COOH$

9. a.



$CH_3CH_2CH_2CH_2CH_2CH_2COOH$ or $CH_3(CH_2)_6COOH$

b.



$CH_3CH_2CH_2CH_2COOCH_2CH_3$

Note: Drawing/writing this in the opposite direction (from right to left) is equally correct.

10. Pentanoic acid exhibits hydrogen bonding between its molecules while ethyl propanoate does not. This results in stronger intermolecular bonding and, hence, a higher boiling point because the molecules are harder to separate.

7.4 Exam questions

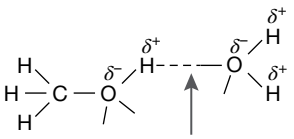
1. B

2. C

3. a. Due to the electronegativity difference between O and H (as well as O and C), the molecule contains polar bonds, which makes the molecule polar.

b. The —COOH functional group is highly polar so the O and H can hydrogen bond to water in smaller carboxylic acids.

4.

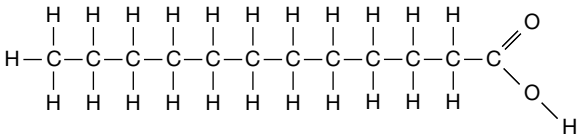


hydrogen bond between methanol and water

5. a. Carboxylic acids homologous series

b. 6-methyl-5-propyloctanoic acid

c. An example is shown. *Note*: Many other isomers are valid answers.



7.5 Sources and uses of organic chemicals

7.5 Exercise

1. D

2. C

3. B

4.

Metal	Symbol	Atomic number	Group	Period
Ruthenium	Ru	44	8	5
Cobalt	Co	27	9	4
Rhodium	Rh	45	9	5

5. a. Any of the following:

- as a basis of flavouring to replace chemical food additives
- as plastics, adhesives, surfactants and lubricants — to replace fossil fuels as the feedstock
- in the production of ethanol — to reduce amount of fossil fuels required (e.g. E10 fuel), as a feedstock for solvents and other chemicals manufactured using fossil fuels.

b. Any of the following:

- in furans (to synthesise lycra) to replace petrochemical-based components
- in disposable coffee cups, detergents, paint stabilisers to replace toxic synthetic oil-based materials, in packaging and as a natural emulsifier in cosmetics
- in the production of ethanol — to reduce amount of fossil fuels required (e.g. E10 fuel), as a feedstock for solvents and other chemicals manufactured using fossil fuels.

6. **a.** Nylon is a polymer derived from oil.
   b. *Benefits:* Strong, lightweight, durable, elastic, waterproof, resistant to sunlight, requires less water than natural fibres to produce, long-lasting so reusable.
   c. *Disadvantages:* Non-biodegradable so lasts indefinitely, recycling (if possible) is energy intensive, production uses considerable energy and creates nitrous oxide (a greenhouse gas), uses fossil fuel resources, source of microfibres (which cause pollution), and can be hazardous to workers in production.

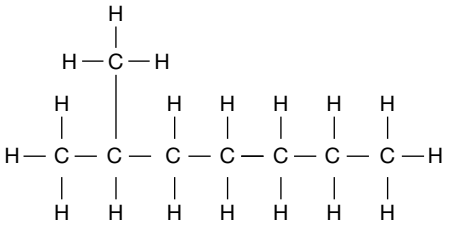
7.5 Exam questions

1. Any of the following:
   Land traditionally used for crops may be used for biomass instead, limiting food supplies.
   Challenges exist in converting lignocellulose into a useable feedstock.
2. Plant cells contain a cell wall made of cellulose. Animal cells do not have a cell wall.
3. **a.** Lignocellulose is a complex tissue consisting of cellulose, which gives strength to the cell walls, hemicellulose, which helps strengthen and connect the cellulose, and lignin, which binds these compounds together.
   b. It is obtained from the cell walls of woody plants.
   c. Lignocellulose is insoluble in water and separating its parts is difficult.
4. **a.** Both are polymers, both have sugar monomers, both are part of lignocellulose.
   b. Cellulose: long unbranched polymer molecules with glucose monomers.
   Hemicellulose: short branched polymer molecules with different monomers.
   c. Cellulose can be made into paper, textiles, pharmaceuticals, cosmetics, emulsifiers, detergents and explosives.
   Hemicellulose can be made into lycra, takeaway coffee cups, detergents, paint stabilisers, emulsifiers and ethanol.
5. **a.** Phthalates
   b. Nylon
   c. Cochineal
   d. Volatile organic compounds
   e. Monosodium glutamate (MSG)

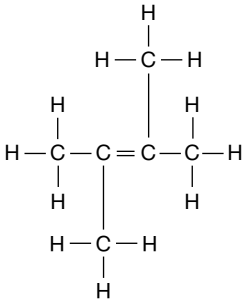
7.6 Review

7.6 Review questions

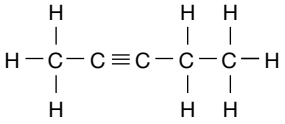
1. Alkenes and alkynes need at least two carbons to have a double or triple bond between them. This is not possible with just one carbon.
2. **a.** Saturated hydrocarbons contain only single carbon–carbon bonds. Unsaturated hydrocarbons contain one or more carbon–carbon double (or triple) bonds.
   b. Unsaturated hydrocarbons
3. **a.** $C_{22}H_{46}$ **b.** $C_{17}H_{34}$ **c.** $C_{13}H_{24}$
4. **a. i.** HBr
   ii. CH_2CH_2
   iii. CH_2Cl_2
   iv. $C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4H_2O$
   b. i. Substitution **ii.** Addition
   iii. Substitution **iv.** Combustion/oxidation
5. **a.** $CH_2C(CH_3)CH_2CH(CH_3)CH_2CH_2CH_3$
   b. $CHCCH_2CH_2CH_3$
   c. $CH_3CH_2CHCHCH_2CH_2CH_2CH_3$
   d. $CH_3CH(CH_3)CH(CH_3)CHCH_2$
   e. $CH_3CH(CH_3)CH(CH_3)CH_3$
   f. $CH_3CH_2CH(CH_3)CH(CH_2CH_3)CH_2CH_2CH_3$
6. **a.** Methylbutane **b.** 3-methylpentane
   c. 3-ethylhexane **d.** 3,3-dimethylhexane
   e. 2,4-dimethylhexane **f.** 2,2,4,4-tetramethylpentane
7. **a.**



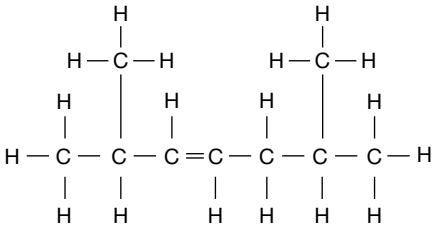
   b.



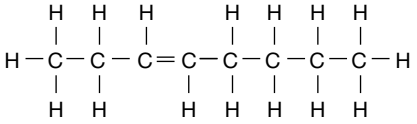
   c.



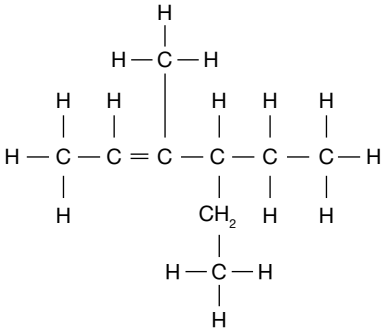
   d.



   e.



   f.



8. I, II and IV

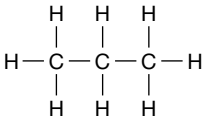
9. Isomers are molecules with the same molecular formula but different structures. For example, C_5H_{12} has three isomers — pentane, 2-methylbutane and 2,2-dimethylpropane — and C_4H_8 has three isomers — but-1-ene, but-2-ene and 2-methylpropene.

10. a. A hydrocarbon is a compound consisting of carbon and hydrogen

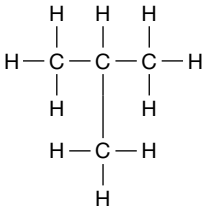
b. Alkanes

c. C_nH_{2n+2}

d. C_5H_{12}

e. 

Propane



2-methylpropane

7.6 Exam questions

Section A — Multiple choice questions

1. C 2. D 3. A 4. A 5. C
6. C 7. C 8. D 9. D 10. D

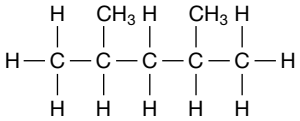
Section B — Short answer questions

11. a. Boiling points reflect the strength of the intermolecular forces operating. The intermolecular forces operating are dispersion forces, which increase with molecule size.

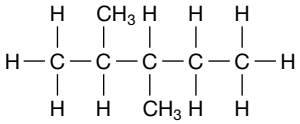
b. The easier a hydrocarbon is to vaporise, the more flammable it is. Petrol has smaller molecules than paraffin wax and, hence, has weaker dispersion forces between its molecules, making it easier to vaporise.

c. Methane, ethane, chloroethane, because chloroethane has the strongest intermolecular forces (dipole–dipole and dispersion). Methane and ethane only have dispersion forces, but ethane has a longer chain so they are stronger between ethane molecules than methane molecules.

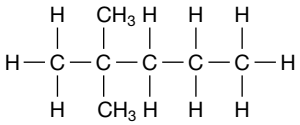
12. The longest chain has seven carbon atoms. The compound should have been named 3,3-dimethylheptane.

13. 

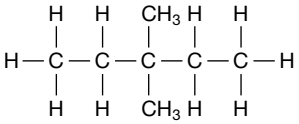
2,4-dimethylpentane



2,3-dimethylpentane



2,2-dimethylpentane



3,3-dimethylpentane

14. a. Unsaturated molecules have a double or triple carbon-to-carbon bond.
A homologous series is a series of compounds with the same functional group and differ by CH_2.

b. $CH_2CH_2(g) + H_2(g) \rightarrow CH_3CH_3(g)$
$CH_3CH_3(g) + Cl_2(g) \rightarrow CH_3CH_2Cl(g) + HCl(g)$

c. $CH_2CH_2(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$ or (l)

d.
- Yeast is used for fermenting the sugars in the plant biomass into ethanol.
  $C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$
- The ethanol is distilled to separate it from the water, and dehydrated to remove water, leaving ethene.

15. a. Carboxylic acid

b. Alcohols (and esters)

c. Butanoic acid

d. $C_4H_8O_2$

e. %C = 54.5%
%H = 9.1%
%O = 36.4%

8 Polymers and society

8.2 Polymers

Practice problem 1

7.00×10^5 g mol^{-1}

8.2 Exercise

1. Monomers for addition polymerisation require a carbon–carbon double bond. Alkanes do not possess a double carbon–carbon bond but alkenes do.

2. a. Examples of natural polymers include proteins, starch, cellulose, wool, cotton and many others.

b. Examples of synthetic polymers include polyethene, polypropene, polyvinyl chloride, nylon, PTFE, PVA, PET, polystyrene and many others.

3. Approximate molar mass = 7.00×10^6 g mol^{-1}

4. Like most addition monomers, special conditions are required before polymerisation can occur. These often involve the use of a catalyst and elevated temperatures and pressures. If these conditions are not present, polymerisation will not take place.

5. a. 1.60×10^4

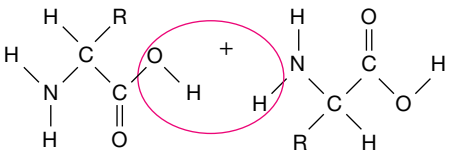
b. 3.20×10^4

6. Alkynes can act as addition monomers because they have a triple carbon–carbon bond. The polymer produced uses electrons from one of the triple bonds to link the monomers together. The resulting polymer would contain residual carbon–carbon double bonds.

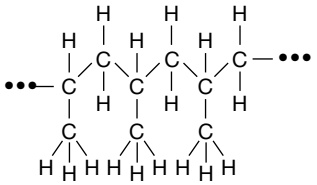
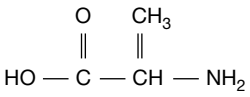
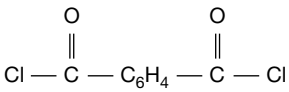
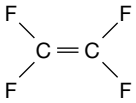
7.
- Both types are made up of smaller units called monomers.
- Both form very long molecules called polymers as a result of the reaction.

8. • Addition polymerisation: breaking of a double carbon–carbon bond.
Condensation polymerisation: the reaction of functional group on separate molecules.
• Addition polymerisation: one product is formed.
Condensation polymerisation: a reaction where a small molecule is produced for each monomer that joins to the polymer chain.

9. a. Natural polymers are formed by living organisms.
b. Condensation polymers are formed by the reaction of functional groups at the ends of each monomer with the loss of a small molecule.
c. Monomer refers to the small molecules that link together to form a polymer.

10. 

8.2 Exam questions

1. C
2. D
3. 
4. a. i. 
ii. Condensation polymerisation
b. i. 
ii. Condensation polymerisation
5. a. Tetrafluoroethylene 
b. C_2F_4
c. Addition reaction

8.3 Linear and cross-linked polymers

8.3 Exercise

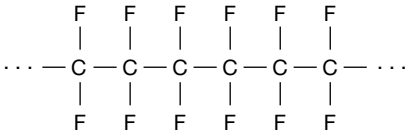
1. D
2. C
3. Thermosetting polymers:
• have extensive covalent cross-linking between chains
• are rigid and decompose or char rather than melt.
4. A
5. B
6. The bonding between chains in thermosetting polymers is covalent bonding. This is much stronger than the weak intermolecular forces that exist between the chains in thermoplastic polymers.
7. a. Thermoplastic polymers only have weak forces, most often dispersion forces, between their chains.
Elastomers have a small amount of covalent cross-linking between their chains.
Thermosetting polymers have a large number of covalent cross-links between their chains.
b. Thermoplastic polymers are able to be softened and melted and then re-moulded.
Elastomers are characterised by their ability to be stretched and to then return to their original shape once the stretching force is removed.
Thermosetting polymers are not able to be softened by heating and many will instead char.
8. a. Making an article from a thermosetting polymer involves two main steps. The first is to place the preliminary linear polymers (called the resin) into the mould. These polymers have functional groups attached to the long polymer backbone. The second step is to induce cross-linking by making the functional groups from one chain react with functional groups from a neighbouring chain. This may be done by heat and pressure, or by adding a chemical to cause this reaction.
b. A plastic is defined as a substance that can be moulded at some stage in its production. Thermosetting polymers are moulded at the stage before the cross-linking is initiated. Hence, by definition, they can be called plastics.
9. A thermoplastic article is able to be softened and melted. After processing (such as moulding and extruding), it will retain its new shape upon cooling. Thermosetting polymers cannot do this because they cannot be melted and re-shaped.
10. It would be expected that polymer *A* would be harder, less flexible and more heat-resistant.

8.3 Exam questions

1. C
2. a. III b. II c. V
d. III or IV e. III or IV f. VI
3. A
4. a. Thermosetting plastics: covalent cross-linking between the polymer chains, chars when heated.
Thermoplastics: often linear polymers, can be repeatably melted, shaped and hardened by cooling.
b. Sample response: Polyvinyl chloride used for insulating wires because it is a non-conductor and resists water.
c. Sample response: Polyethene used for milk bottles because it is water resistant to water and chemically inert.
5. a. Polymer A — within chains: covalent bonding. Between chains: possibly weak dispersion forces.
Polymer B — within chains: covalent bonding. Between chains: covalent bonding.
b. Polymer A. Only has weak forces between the carbon chains.
c. Neither polymer (no free moving ions or electrons are present).

8.4 Polymer selection

Practice problem 2



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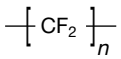
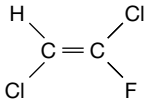


Practice problem 3



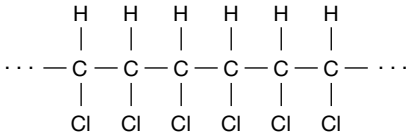
8.4 Exercise

1. Chain length: in a linear polymer, longer chains will be harder, less flexible and more heat resistant.
   Degree of branching: as the amount of branching increases, polymers will be more flexible, softer and not as dense.
   (Detailed responses can be found in the worked solutions in your digital formats.)
2. Plasticisers are small molecules that fit between the chains of a linear polymer and push them further apart. This makes the dispersion forces acting between these chains less effective. As a result, the polymer will be softer and more flexible.
3. Polyethene is water repellent and polyvinyl alcohol is water-soluble due to the monomers from which each is made. Polyethene is non-polar due to its C–H bonds and its symmetry. Polyethene is thus insoluble in water and water repellent. Vinyl alcohol and, subsequently, polyvinyl alcohol, have O–H groups, which make them polar and soluble in water due to the formation of hydrogen bonds.
4. Crystallinity is caused by polymer chains being able to pack together in an orderly pattern. This is more likely to occur when the side groups attached to these chains are in an orderly or predictable pattern. If the pattern of these side groups is more random, it is more difficult for the chains to pack together.
5. Because the required polymer needs to soften under moderate heat, it cannot be thermosetting. Hence, the polymer cannot be cross-linked. It needs to be made from a monomer that is non-polar so that it will be water repellent. For it to be rigid, the movement of its long polymer chains needs to be restricted in the solid state. This is possible if the chains are long and able to tangle due to their length. Alternatively, if the chains can be encouraged to align more often with each other so that the degree of crystallinity is increased, this will also improve the polymer's hardness. However, if the use specifies continual cycles of melting and solidifying, this last effect may not be reproducible.
6. Advantages: polymers are lighter than metal, making them more fuel efficient, and are easier to produce. Disadvantage: polymers are not as strong as metal, and hence are more easily damaged.
7. Detailed responses can be found in the worked solutions in your digital formats.
8. PVC is not recycled to the same extent for two main reasons:
   - it is often used for a long time (e.g. pipes, gutters), and is not replaced often, which reduces the amount of recycling.
   - it contains chlorine and additives (plasticisers) that make the recycling process more complicated.

8.4 Exam questions

1. **a.** Monomer II is the correct monomer.

   b.



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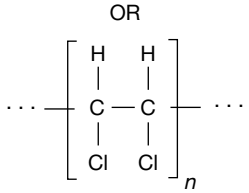
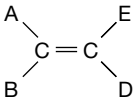
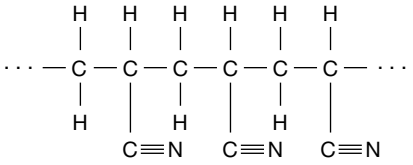


2. Polypropene, polystyrene and polytetrafluoroethylene are all made from non-polar monomers. Because water is polar, the polymers will be water insoluble and water repellent.

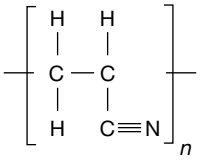
3. 

4. The monomer to choose would be ethene, because a food container needs to be water-proof and vinyl alcohol is water-soluble due to the presence of polar hydroxyl groups. Ethene is non-polar and so water-repellent. The ethene would be used to make low-density polyethylene because it is flexible and transparent due to the high degree of branching, which keeps the molecules apart and so creates lower density. The carbon chains would be shorter so they are not too tangled and can move past each other. So low-density polyethene would be the preferable polymer to produce.

5. **a.**



   b.



   c. Polyacrylonitrile

8.5 Plastic recycling and innovations in design

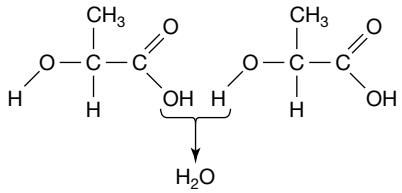
8.5 Exercise

1. a. Bioplastic is a polymer obtained from renewable biomass; for example, from agricultural, cellulose and corn starch waste.
   b. Polylactic acid, biopolyethene and biopolypropene
2. Degradable: will eventually break down by any means.
   Biodegradable: can be broken down by microorganisms but not necessarily in a short period of time.
   Compostable: will break down using microorganisms to natural substances in months rather than years.
3. Advantages (any three of the following): plants are renewable and sustainable unlike fossil fuels; plants use carbon dioxide when growing; process can use inedible plant matter; less energy and fewer hazardous chemicals are required in production of plant-based plastics.
   Disadvantages (any three of the following): Need land for plants to grow on, which may be scarce in some countries; use of land may be needed to grow food crops; these plastics may not be biodegradable and are currently more expensive to manufacture; process requires water, fertilisers and pesticides, which could impact on the environment.
4. a. Polylactic acid (PLA) is a thermoplastic, biodegradable polymer obtained from plant products such as corn starch, maize or sugar cane.
   b. 1. The raw material, starch, is extracted from the plant material.
      2. The sugar dextrose is processed from the starch.
      3. Dextrose is converted to lactic acid by fermentation using microorganisms.
      4. Either of two methods of condensation polymerisation:
         a. producing lactide as an intermediate
         b. direct condensation polymerisation of lactic acid.
   c. Any three of the following:
      - biodegradable
      - low heat resistance
      - soluble in some organic solvents
      - easy to melt and mould
      - can be hydrolysed into its monomers
      - solid form is non-toxic.
5. Both are polymers made from biomass.
   PLA is a biodegradable condensation polymer and bio-PE is an addition polymer and must be recycled at the end of life.
6. Any four of the following:
   - Plastics are long lasting materials.
   - Plastics are often only used once and then discarded so the amount of waste is considerable.
   - Very little plastic waste is collected and recycled.
   - Recycling often requires sorting, which can be complex.
   - Some plastic waste is incinerated as an energy source, and this can contribute to air pollution.
   - A significant amount of plastic collects in the ocean and it is difficult to retrieve.
   - The health of humans and other organisms is being adversely affected by plastic waste.
7. a. Mechanical, chemical and organic recycling
   b. In mechanical recycling, the waste plastic is collected, shredded, washed, dried and melted and then converted into pellets to be manufactured into new products. It is beneficial to recycle plastics because it conserves energy and decreases the use of fossil fuels.
   Chemical recycling has the advantage that a wider range of plastics can be recycled, and the process breaks the plastics into monomers so they can be combined to produce new plastics that are not degraded. Catalysts and heat energy are required to heat and treat the waste; ideally, renewable energy would be used.
   Organic recycling, to be effective, needs to be carried out on an industrial scale and used for compostable plastics. It requires particular microorganisms and the right conditions. The plastics are broken down to carbon dioxide, water vapour and heat in a few months, so they do not have to go to landfills.
8. Sample response:
   - Do not purchase goods that have unnecessary plastic packaging or contents.
   - Ensure that plastics are disposed of in recycling bins.
   - Reuse plastic containers or bags.
   - Use compostable plastic where possible, and dispose of properly.
9. a. Household: avoid single-use plastics, purchase less, purchase recyclable plastics and dispose of properly.
   b. Business: avoid plastic packaging where possible, and invest in recyclable/compostable products.
   c. Industry: use renewable energy to recycle or manufacture necessary plastic, invest in bioplastic development and use.
   d. Government: legislate to limit the amount of plastic in society (e.g. ban single use plastics that are not compostable), provide funding to support bioplastic innovation and design.
10. i. Choose three types of thin plastic, including a biodegradable plastic, a compostable plastic and a non-renewable plastic.
    ii. Cut three samples of the same size from each of the three plastics.
    iii. Bury the nine samples in the same area of soil at the same depth with markers indicating the plastic type.
    iv. Remove one of each type from the soil and make observations after one week, two weeks and four weeks. Take photos and record the appearance.
    v. Discuss the degree of break down for each sample, and comment on the validity and reliability of the data obtained.

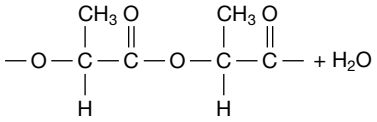
8.5 Exam questions

1. A
2. A
3. a. PLA is a condensation polymer because two functional groups are reacting, resulting in the formation of a larger molecule and a smaller molecule (water).

b.



Condensation reaction



4. a. 3.7×10^5 tonnes

b. 1.3×10^5 tonnes

5. a. Sample response — any three of the following:

- shopping bags
- packaging for dry foods
- packaging for dry materials that have to be dissolved in water after purchase (e.g. fertiliser, dishwashing powder)
- packaging for products that need protection from the environment
- pet waste bags
- laundry bags.

b. Sample response:

- They do not contribute to landfills.
- They are not detrimental to the environment because they are non-toxic.

c. Sample response — one of following:

- They would break down if left in the rain for a period of time, so useful for indoor use only.
- Not suitable for storing liquids.

8.6 Review

8.6 Review questions

1. Advantages: can be moulded into any shape, don't corrode, are cheap to produce and lighter than metallic options, are good insulators of heat and electricity, some are recyclable, most are chemically inert, are strong and relatively cheap to produce. (Detailed responses can be found in the worked solutions in your digital formats.)
Disadvantages: non-renewable/finite resource (made from petroleum), take a long time to decompose (if at all) and contribute to landfill. (Detailed responses can be found in the worked solutions in your digital formats.)

2. Alkanes are represented by the formula, C_nH_{2n+2}.
Options (a), (c) and (g) follow this relationship and, therefore, cannot act as monomers in a polymerisation reaction. Options (b) and (e) are alkenes, because they follow the formula C_nH_{2n}. Because these molecules contain double bonds, they can act as monomers. Options (d) and (f) can also act as monomers in polymerisation reactions because they are alkynes and follow the formula C_nH_{2n-2}.

3. Thermoplastic polymers, such as polyethene, polypropene and polystyrene, can be repeatedly melted, reshaped and hardened by cooling. Thermosetting polymers, such as Bakelite, Formica and rubber, do not melt when heated but instead char, and cannot be reshaped through melting.

4. a. Long polymer chains made by polymerising monomers.

b. Has cross-links between its long polymer chains (but not as many).

5. a. Condensation polymer

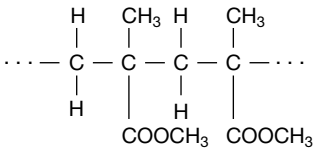
b. HCl

6. In LDPE, the extensive branching results in the polymer chains being kept further apart, resulting in less effective dispersion forces. In HDPE, minimal branching is present so the polymer chains can pack together much more closely, resulting in more effective and, hence, stronger dispersion forces.

7. Recycling is important because it reduces both the use of valuable finite resources to produce plastics and the quantity of non-biodegradable material going into landfill.

8. Many recycling methods are dependent on plastic being sorted according to type. The recycling code both assists and reduces errors in this procedure.

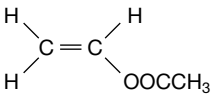
9. a.



b. Perspex is stronger and more impact-resistant than glass.

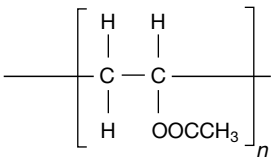
c. Perspex is more expensive than glass and melts when exposed to high temperatures; it can also scratch and may not be as clear as glass.

10. a.

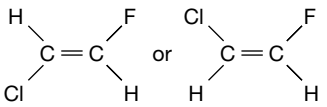


b. Dispersion forces exist between these chains as well as dipole–dipole forces, due to the polar carbon–oxygen bonds.

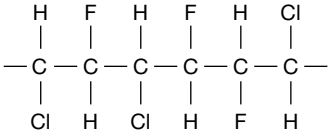
c.



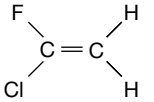
11. a.



b.



c.



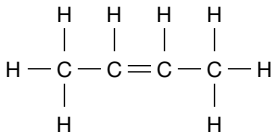
8.6 Exam questions

Section A — Multiple choice questions

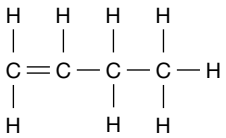
1. D	**2.** D	**3.** B	**4.** A	**5.** D
6. A	**7.** C	**8.** B	**9.** A	**10.** B

Section B — Short answer questions

11. a. i.

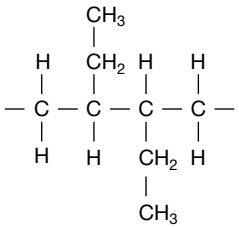


ii.

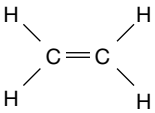


b. This polymer is made from the monomer but-2-ene.

c.

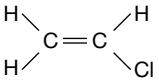


12. a. Monomer for polyethene (HDPE)



ethene

Monomer for polyvinyl chloride (PVC)



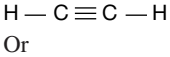
chloroethene

vinyl chloride

b. Both chains are about the same length and have roughly equal amounts of branching. They would, therefore, be expected to show the same degree of entanglement. Differences could be caused by the presence of chlorine atoms in place of some hydrogen atoms. This would increase the dispersion forces between the chains. Acting against this might be the fact that chlorine atoms are larger than hydrogen atoms, so the chains might not be able to pack as closely together. This would decrease the effect of the dispersion forces.

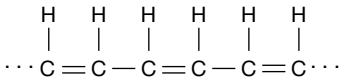
c. Plasticisers push the polymer chains a little further apart. This means the dispersion forces between the chains are not quite as effective. This results in a polymer that is softer and more flexible.

d. The main difference between LDPE and HDPE is the amount of branching that exists in the polymer chains. This results in chains that cannot pack as effectively together in LDPE as they can in HDPE, resulting in LDPE being less dense, softer and more flexible than HDPE.

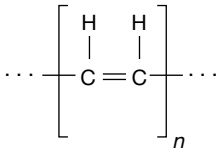
13. a. 

Or

b.

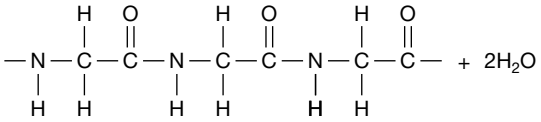


c.



d. Polyethyne would be more reactive than polyethene due to the presence of double bonds.

14. a.



b. It is a condensation reaction, because a larger molecule is formed from three smaller molecules and water is produced as a by-product.

15. Please see the worked solutions in your online resources.

Unit 1 | Area of Study 2 review

Section A — Multiple choice questions

1. C	**2.** C	**3.** B	**4.** D	**5.** B
6. C	**7.** C	**8.** D	**9.** D	**10.** A
11. B	**12.** B	**13.** A	**14.** C	**15.** D
16. A	**17.** D	**18.** A	**19.** B	**20.** C

Section B — Short answer questions

21. a.

	Isotope 1	Isotope 2	Isotope 3
Relative isotopic mass	273	274	275
Relative abundance (%)	55	12	33

b. $A_r = 2.7 \times 10^2$ (2 significant figures)

22. a. i. 0.343 mol

ii. 0.687 mol

iii. 8.27×10^{23} atoms

b. 30 mg

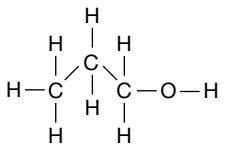
23. $C_2H_6Cl_2$

24. a. $C_3H_6 + Cl_2 \rightarrow C_3H_6Cl_2$

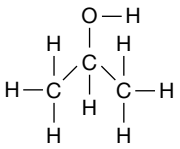
b. $C_2H_6 + Cl_2 \rightarrow C_2H_5Cl + HCl$

c. $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$

25.

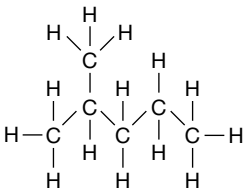
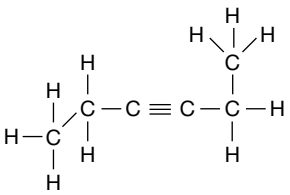
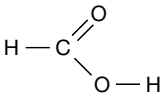


propan-1-ol or 1-propanol

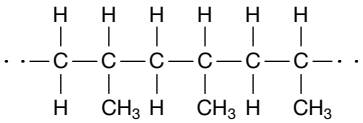


propan-2-ol or 2-propanol

26. a.

Structural formula	Systematic name
	**2-methylpentane** Five carbons are in the main chain, giving pent-. As all C–C bonds are single, this is a pent-ane. The methyl group sites on carbon-2, resulting in **2-methylpentane**
	**Hex-3-yne** Six carbons are in the main chain, giving hex-. Because there is one C–C triple bond, this is a hex-yne. The triple bond is in position C3, resulting in **hex-3-yne**.
 The molecule is a carboxylic acid, hence it has the functional group COOH. The meth-prefix indicates one carbon atom.	Methanoic acid

b. $CH_3CH_2CH_2COOCH_2CH_3$

27. a. 

Addition polymerisation breaks the double carbon–carbon bonds, allowing long chains to form.
See figure at the bottom of the page*

b. Due to the closer packing of the less-branched structure, the high density form will have greater opportunities for dispersion forces between its chains. This will increase its melting point and make it less flexible (or more opaque) than the low density form.

28. a. Plants or plant-based products contain cellulose, so any two out of corn, paper, sawdust or recycled cotton would be acceptable.

b. Advantages would include:
- removal of carbon dioxide from atmosphere
- lower energy costs due to solar energy being used
- less reliance on hazardous chemicals.

Other reasonable answers are also acceptable.

10 Water as a unique chemical

10.2 Water on Earth

10.2 Exercise

1. Australia has the highest per capita water storage rate in the world due to geography and climate. Australia does not have many natural lakes to use as water reservoirs and so must actively manage its water storage capacities. Furthermore, Australia is the driest inhabited continent in the world, and has very variable rainfall. The unpredictability of rainfall is exacerbated by heavy rains. This leads to significant run-off that needs to be carefully managed to ensure as much water as possible can be stored. Careful water management tries to ensure drinking water is available from a range of sources, rather than just one.

2. C

3. a. 1.82% **b.** 98.1791%

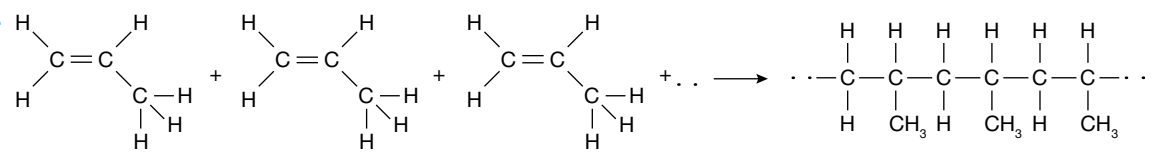
4. The United Nations Sustainable Development Goals were adopted by all United Nations Member States in 2015. There are 17 goals encompassing 169 urgent call for action targets, with the overall goal of a sustainable future for all, to be met by 2030. Within the field of chemistry, nine goals are particularly relevant:
- Goal 2: Zero Hunger
- Goal 6: Clean water and sanitation
- Goal 7: Affordable and clean energy
- Goal 9: Industry, innovation and infrastructure
- Goal 11: Sustainable cities and communities
- Goal 12: Responsible consumption and production
- Goal 13: Climate action
- Goal 14: Life below water
- Goal 15: Life on land.

5. a. Goal 6 (Clean water and sanitation) aims to 'ensure availability and sustainable management of water and sanitation for all', and encompasses eight targets.

b. Goal 6 aims to improve access to safely managed drinking water, safely managed sanitation and basic hygiene.

c. i. Globally, safely managed drinking water has been improved by approximately 4 per cent, safely managed sanitation by approximately 7 per cent and basic hygiene by approximately 3.5 per cent

ii. Australia only reported against two of the eight targets in Goal 6, making identifying an overall improvement or not difficult. However, water consumption per capita has significantly reduced and is below the global benchmark and is, therefore, rated as on-track. The average household expenditure on water, sewage and wastewater, on the other hand, can be improved.

*27. a. 

10.2 Exam questions

1. C
2. a. Only 2.6 per cent of the world's water is fresh because water is such a good solvent. Many substances that water comes in contact with will dissolve in it, and these include naturally occurring minerals or introduced contaminants.
   b. The largest source of the world's freshwater is in glaciers and icecaps.
   c. i. Australia's drinking water is sourced from:
      - water catchments
      - lakes
      - rivers and creeks
      - groundwater
      - desalinated water
      - recycled water.

      ii. Australia is protecting its water security through the development of:
      - desalination plants
      - water storage in underground aquifers
      - development of catchment reservoirs
      - ongoing testing of drinking water sources to monitor potential contamination
      - government protocols for the treatment and disposal of industrial waste products to prevent contamination of water sources.
3. - Gas — atmosphere as water vapour.
   - Liquid — as water in lake.
   - Solid — as ground ice.
4. D
5. - Increase in atmospheric CO_2 levels.
   - Atmospheric CO_2 dissolves in ocean water to produce the weak acid carbonic acid.
   - Carbonic acid dissociates into bicarbonate ions and hydrogen ions (H^+).
   - Increase in H^+ ions in oceans.
   - Decrease in pH, increase in ocean acidity.

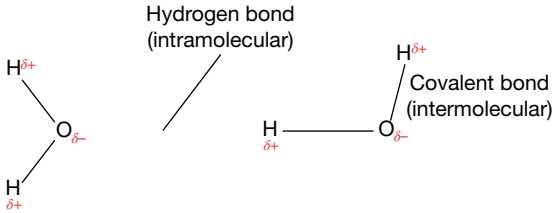
10.3 Properties of water

Practice problem 1

$m(H_2O) = 150$ g (3 significant figures and correct units)

10.3 Exercise

1. Water contains hydrogen–oxygen bonds, which allow for the formation of comparatively strong permanent dipoles called hydrogen bonds.
2. Weakest to strongest: dispersion forces, dipole–dipole forces, hydrogen bonding, covalent bonding.
3. a. A water molecule consists of two hydrogen atoms covalently bonded to a single oxygen atom. Two lone pairs are also present and the repulsion of the bonding pairs and the lone pairs results in the water molecule being bent.
   b. Oxygen is more highly electronegative than hydrogen. The shared electrons will, therefore, be more closely attracted to the oxygen, meaning the oxygen atom becomes slightly negatively charged and the hydrogen slightly positively charged. This is a polar covalent bond. The molecule can also be described as polar because the bent shape results in an uneven charge distribution overall.
   c.



4. Gas particles have higher kinetic energy than liquid particles so energy needs to be removed, meaning changing state from gas to liquid (such as in condensation) is an exothermic process.
5. When liquid water evaporates into steam the intermolecular bonds, hydrogen bonds, are broken. The covalent bonds between the atoms in the molecule are still intact.
6. a. Mass and volume are required to calculate density.

$$d = \frac{m}{V}$$

   b. Units of density: g mL^{-1} or kg m^{-3}
   c. Higher
7. $m(H_2O) = 65.3$ g
8. When water freezes, the density decreases because as the molecules lose energy they form a hexagonal crystalline lattice, which occupies more volume than the random arrangement of molecules in the liquid form.
9. Polonium's melting point can be approximated as –25 °C to –35 °C; the boiling point would be approximately 25 °C to 35 °C.
10. The hydrogen bonds between the H_2O molecules are stronger than the dipole–dipole forces between the H_2S molecules. (The electronegativity difference between S and H is much smaller than that between O and H so H_2S is considerably less polar.)

10.3 Exam questions

1. D
2. B
3. Hydrogen bonding is the attraction between a hydrogen atom covalently bonded to a fluorine, oxygen or nitrogen atom and another very electronegative atom (F, O, N).
4. C
5. a. When substances boil, the intermolecular forces are broken. H_2S has a much lower boiling point than water because it does not have hydrogen bonding between its molecules, as water does.
   b. The boiling points for H_2Se and H_2Te increase due to the increasing size of the molecules, causing an increase in dispersion forces.

10.4 Heat capacity and latent heat

Practice problem 2

$q = 87.7$ kJ (3 significant figures and correct units)

Practice problem 3

$q(H_2O) = 226\ kJ$ (3 significant figures and correct units)

10.4 Exercise

1. Latent heat of vapourisation is the energy needed to change the state of a substance from liquid to gas and vice versa; the temperature of the substance remains constant during this process.
2. The heat energy is used to break the intermolecular hydrogen bonds between the water molecules while the temperature remains constant.
3. The energy (measured in joules) needed to raise the temperature of 1 g of water by 1 °C is 4.18 joules.
4. Water has a higher specific heat capacity than sand so it requires more heat energy to raise its temperature. This means the temperature of the sand increases at a faster rate than that of the water.
5. It would require more heat to raise the temperature of copper because it has higher specific heat capacity. (0.39 J g^{-1} $°C^{-1}$ for copper and 0.13 J g^{-1} $°C^{-1}$ for gold.)
6. $m(H_2O) = 300\ g$
   $c(H_2O) = 4.18\ J\ g^{-1}\ °C^{-1}$
   $\Delta T(H_2O) = 30.2\ °C$
   $q(H_2O) = ?$
   $q = mc\Delta T$
   $q(H_2O) = 300 \times 4.18 \times 30.2$
   $= 3.79 \times 10^4\ J$
   $= 37.9\ kJ$ (3 significant figures and correct units)
7. $m(H_2O) = 50.0\ g$
   $c(H_2O) = 4.18\ J\ g^{-1}\ °C^{-1}$
   $\Delta T(H_2O) = ?$
   $q(H_2O) = 6274\ J$
   $q = mc\Delta T$
   $\Delta T(H_2O) = \dfrac{q}{mc}$
   $\Delta T = \dfrac{6274}{50.0 \times 4.18}$
   $= 30.0\ °C$
   Final temperature = 25.0 + 30.0 = 55.0 °C (3 significant figures and correct units)
8. $m(Cu) = 5100\ g$
   $\Delta T(Cu) = 15.0\ °C$
   $q(Cu) = 30.0\ kJ$
   $q = mc\Delta T$
   $c(Cu) = \dfrac{q}{m\Delta T}$
   $= \dfrac{30 \times 1000}{5100 \times 15.0}$
   $= 0.392\ J\ g^{-1}\ °C^{-1}$
   (3 significant figures and correct units)
9. $m(Al) = 250\ g$
   $c(Al) = 0.90\ J\ g^{-1}\ °C^{-1}$
   $\Delta T(Al) = 45.0\ °C - 20.0\ °C = 25\ °C$
   $q(Al) = ?$
   $q = mc\Delta T$
   $q(Al) = 250 \times 0.90 \times 25$
   $\Delta T = 5.63 \times 10^3\ J$
   $= 5.63\ kJ$ (3 significant figures and correct units)
10. $m(H_2O) = 152\ g$
    $L_f(H_2O) = 6.01\ kJ\ mol^{-1}$
    $q(H_2O) = ?$
    $q = nL_f$
    $n = \dfrac{m}{M}$
    $= \dfrac{152}{18.0}$
    $= 8.44\ mol$
    $q = 8.44 \times 6.01$
    5.07 kJ (3 significant figures and correct units)
11. $q = mc\Delta T$
    $\Rightarrow 1.20 \times 10^5\ J = 500 \times 4.18 \times \Delta T$
    $\Rightarrow \Delta T = \dfrac{1.20 \times 10^5}{500 \times 4.18}$
    $\Rightarrow \Delta T = 57.4\ °C$
    $\therefore T_{final} = 23.0 + 57.4 = 80.4\ °C$
12. $C_{H_2O} \times m_{H_2O} \times \Delta T_{H_2O} = 4.18 \times 50.0 \times (51.4 - 20.0)\ J$
    $= 6563\ J$
    $=$ energy supplied to water
    Since oil receives the same amount of energy,
    $6563\ J = C_{oil} \times m_{oil} \times \Delta T_{oil}$
    $= C_{oil} \times 50.0 \times (85.6 - 20.0)$
    $= 3280 \times C_{oil}$
    $\therefore C_{oil} = \dfrac{6563}{3280}$
    $= 2.00\ J\ °C^{-1}\ g^{-1}$
    Alternative method:
    $\Delta T_{H_2O} = (51.4 - 20.0)\ °C = 31.4\ °C$
    $\Delta T_{oil} = (85.6 - 20.0)\ °C = 65.6\ °C$
    Since the same energy is applied to both,
    $q_{oil} = qH_2O$
    $\Rightarrow m_{oil} \times c_{oil} \times \Delta T_{oil} = m_{H_2O} \times c_{H_2O} \times \Delta T_{H_2O}$
    $\Rightarrow 50.0 \times c_{oil} \times 65.6 = 50.0 \times 4.18 \times 31.4$
    $\Rightarrow c_{oil} = \dfrac{50.0 \times 4.18 \times 31.4}{50.0 \times 65.6}\ J\ g^{-1}\ °C^{-1}$
    $\Rightarrow 2.00\ J\ g^{-1}\ °C^{-1}$

10.4 Exam questions

1. D
2. C
3. B

4. a. Due to its high heat capacity, water absorbs a lot of heat energy, resulting in cooling of the surrounding land and air during the heat of the day, and loses it slowly, resulting in warming during the cool night. As a result, the temperature in coastal regions tends to be more moderate.

   b. The latent heat of vaporisation is the amount of energy needed to change a fixed amount of water from a liquid to a gas. As sea surface temperatures rise, the sea water has more kinetic energy and, therefore, the amount of water able to be evaporated (that is, changed from a liquid to a gas) increases. This increases the amount of water vapour in the atmosphere, leading to heavier precipitation and more severe storms.

5. a.
   - The two liquids are poured into the same type of container.
   - The *mass* of each liquid is recorded.
   - The *initial temperature* of each liquid is measured.
   - The two containers are heated simultaneously on the same heating appliance (or heated for equal times on the same heat setting) to ensure that both received the same amount of energy.
   - The *final temperature* of each liquid is measured.

   Note: Using the equation $q = c \times m \times \Delta T$, the energy delivered to each liquid can be calculated from water's known specific heat capacity, its recorded mass and temperature change. Hence, the unknown heat capacity of the oil can be found using q, m and ΔT.

   b. Protective clothing, particularly safety glasses, should be worn when heating liquids.

10.5 Review

10.5 Review questions

1. A large amount of energy is required to evaporate the water from clothing because of water's high latent heat of vapourisation value.

2. a. Specific heat capacity is the energy needed to raise the temperature of 1 gram of a substance by 1 °C.

   b. Water has a relatively high specific heat capacity compared to other substances because it is capable of absorbing a relatively large amount of heat before its temperature rises significantly.

   c. Water is used as a coolant in car radiators to keep the engine from overheating.

3. Water has the highest melting and boiling temperatures because of the hydrogen bonding between the molecules. Other group 16 hydrides depend on dispersion forces, which are relatively weak but increase as the size of the molecule increases.

4. a.
$$m(H_2O) = 435\,g$$
$$c(H_2O) = 4.18\,J\,g^{-1}\,°C^{-1}$$
$$\Delta T(H_2O) = 86.4 - 33.5 = 52.9\,°C$$
$$q = mc\Delta T$$
$$q(H_2O) = 435 \times 4.18 \times 52.9 = 9.62 \times 10^4\,J = 96.2\,kJ$$

   b.
$$m(H_2O) = ?$$
$$c(H_2O) = 4.18\,J\,g^{-1}\,°C^{-1}$$
$$\Delta T(H_2O) = 5.00\,°C$$
$$q(H_2O) = 925\,J$$
$$q = mc\Delta T$$
$$m = \frac{q}{c\Delta T}$$
$$m = \frac{925}{4.18 \times 5.00}$$
$$m = 44.3\,g \text{ (3 significant figures and correct units)}$$

5.
$$n(H_2O) = 6.10\,mol$$
$$m(H_2O) = n \times M = 6.10 \times 18.0 = 110\,g$$
$$c(H_2O) = 4.18\,J\,g^{-1}\,°C^{-1}$$
$$\Delta T(H_2O) = 25.0\,°C$$
$$q = mc\Delta T$$
$$q(H_2O) = 110 \times 4.18 \times 25.0 = 1.15 \times 10^4\,J = 11.5\,kJ \quad \text{(3 significant figures and correct units)}$$

6. Cells are filled with water (and other substances) so when a lettuce leaf is frozen the water expands and breaks the cell walls, allowing the water to escape. This means the cell loses its shape, leaving the lettuce limp when defrosted.

7. 70.3 °C

8. Silver, copper, aluminium.

9. a.
$$m(H_2O) = 220\,g$$
$$n(H_2O) = \frac{m}{M} = \frac{220}{18.0} = 12.2\,mol$$
$$L_v(H_2O) = 40.7\,kJ\,mol^{-1}$$
$$q = nL_v = 12.2 \times 40.7 = 4.97 \times 10^2\,kJ \quad (3 \text{ significant figures and correct units})$$

   b. A large amount of additional heat energy is released when the steam condenses into a liquid.

10. a. Latent heat refers to the energy that is required to change the state of a substance.

    b. The latent heat of vapourisation for water is much greater than the latent heat of fusion because when water is boiled the molecules have to break all of the hydrogen bonds and separate completely as they turn into a gas. The process of melting only partially separates the molecules and so requires less energy.

    c. i. The high latent heat of vaporisation of water means the oceans can absorb significant amounts of heat (energy) without the average sea temperatures being affected, and oceans can release heat over long periods of time.

    ii. Any two of the following:
    - Hot spots in the oceans develop, changing rainfall patterns, which affects both marine life and rainfall patterns on land.

- Storms can increase in severity due to increased amounts of evaporation (atmospheric water vapour) over the oceans, leading to increased risk of heavier rainfall.
- Impacts on ecosystems and biological diversity because different aquatic organisms have different optimal temperatures for survival, meaning any changes to sea temperatures will affect ecosystems.

10.5 Exam questions

Section A — Multiple choice questions

1. C **2.** D **3.** A **4.** A **5.** C
6. C **7.** C **8.** D **9.** C **10.** D

Section B — Short answer questions

11. a. $m\left(C_2H_5OH\right) = ?$
$d\left(C_2H_5OH\right) = 0.789\ g\ mL^{-1}$
$V\left(C_2H_5OH\right) = 250.0\ mL$
$d = \frac{m}{V}$
$m\left(C_2H_5OH\right) = d \times V$
$= 0.789 \times 250$
$= 197\ g$ (3 significant figures and correct units)

b. Specific heat capacity is the energy (measured in joules) needed to raise the temperature of 1 g of a pure substance by 1 °C.

c. Water has the higher capacity, which means that it can absorb more heat energy with a lesser increase in temperature than ethanol and so the ethanol would show the higher temperature.

d. $m\left(C_2H_5OH\right) = 180\ g$
$c\left(C_2H_5OH\right) = 2.5\ J\ g^{-1}\ °C^{-1}$
$\Delta T\left(C_2H_5OH\right) = 17.0\ °C$
$q = mc\Delta T$
$q\left(C_2H_5OH\right) = mc\Delta T$
$= 180 \times 2.5 \times 17.0$
$= 7.65 \times 10^3\ J$
$= 7.7\ kJ$ (2 significant figures and correct units)

12. a. $C_{H_2O} \times m_{H_2O} \times \Delta T_{H_2O} = 4.18 \times 50.0 \times (51.4 - 20.0)\ J$
$= 6563\ J$
$=$ energy supplied to water

Since oil receives the same amount of energy,

$$6563\ J = C_{oil} \times m_{oil} \times \Delta T_{oil}$$
$$= C_{oil} \times 50.0 \times (85.6 - 20.0)$$
$$= 3280 \times C_{oil}$$
$$\therefore C_{oil} = \frac{6563}{3280}$$
$$= 2.00\ J\ °C^{-1}\ g^{-1}$$

Alternative method:

$$\Delta T_{H_2O} = (51.4 - 20.0)\ °C = 31.4\ °C$$
$$\Delta T_{oil} = (85.6 - 20.0)\ °C = 65.6\ °C$$

Since the same energy is applied to both,

$$q_{oil} = q_{H_2O}$$
$$\Rightarrow m_{oil} \times c_{oil} \times \Delta T_{oil} = m_{H_2O} \times c_{H_2O} \times \Delta T_{H_2O}$$
$$\Rightarrow 50.0 \times c_{oil} \times 65.6 = 50.0 \times 4.18 \times 31.4$$
$$\Rightarrow c_{oil} = \frac{50.0 \times 4.18 \times 31.4}{50.0 \times 65.6}\ J\ g^{-1}\ °C^{-1}$$
$$= 2.00\ J\ g^{-1}\ °C^{-1}$$

b. IV: time spent on hotplate/temperature of hotplate.
DV: temperature of the oil and water.

13. a. The density of water increases as the temperature decreases until maximum density at 4 °C because the water molecules start to form a hexagonal crystalline lattice in ice held together by hydrogen bonds instead of the closer random arrangement found in liquid water.

b. Because ice is less dense than water, it floats on top of the water, insulating the water and trapping oxygen below so organisms can survive below the ice. If ice were denser than water, it would freeze from the bottom up and sea creatures would not survive.

c. Place 9.0 mL of water in a measuring cylinder. Place 9.0 mL of oil in a measuring cylinder. Place both in a freezer overnight. Observe the volume of each the next day. If the volume has increased, this means that the density has decreased.

14. a. The damp cloth absorbs heat energy from the water bottle to evaporate/change state, and this keeps the bottle and water inside cool.

b. The water vapour from the air condenses when it contacts the cold can. Energy goes into the can from the water in the condensation process.

c. The fish will remain at 0 °C while ice is present. If the temperature of the fish were higher than 0 °C, heat energy would be transferred to the ice to melt the ice, and so cooling the fish.

15. a. E **b.** E **c.** B **d.** B **e.** D **f.** A

11 Acid–base (proton transfer) reactions

11.2 Acids and bases

Practice problem 1

Conjugate pairs are HNO_3/ NO_3^- and NH_4^+/NH_3.

11.2 Exercise

1. Acid: H_2SO_4, base: H_2O

2. a. HBr **b.** HSO_4^- **c.** HSO_4^- **d.** H_3O^+

3. (b), (c), (e)

4. a. Base **b.** Amphiprotic
c. Amphiprotic **d.** Acid
e. Amphiprotic

5. a. $H_3PO_4(aq) + H_2O(l) \rightleftharpoons H_2PO_4^-(aq) + H_3O^+(aq)$
$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$
$HPO_4^{2-}(aq) + H_2O(l) \rightleftharpoons PO_4^{3-}(aq) + H_3O^+(aq)$

b. $H_3PO_4(aq)/\ H_2PO_4^-(aq)$, $H_3O^+(aq)/H_2O(l)$
$H_2PO_4^-(aq)/\ HPO_4^{2-}(aq)$, $H_3O^+(aq)/H_2O(l)$
$HPO_4^{2-}(aq)/\ PO_4^{3-}(aq)$, $H_3O^+(aq)/H_2O(l)$

6. **a.** $HS^-(aq) + H_2O(l) \rightleftharpoons S^{2-}(aq) + H_3O^+(aq)$
$HS^-(aq) + H_2O(l) \rightleftharpoons H_2S(aq) + OH^-(aq)$
b. $HSO_4^-(aq) + H_2O(l) \rightleftharpoons SO_4^{2-}(aq) + H_3O^+(aq)$
$HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_2SO_4(aq) + OH^-(aq)$
7. **a.** Triprotic **b.** Monoprotic **c.** Monoprotic
d. Diprotic **e.** Monoprotic **f.** Diprotic

11.2 Exam questions

1. $H_2S(aq) + H_2O(l) \rightleftharpoons HS^-(aq) + H_3O^+(aq)$
$HS^-(aq) + H_2O(l) \rightleftharpoons S^{2-}(aq) + H_3O^+(aq)$
2. Acidic protons are bonded to electronegative atoms. In sulfuric acid, both hydrogen atoms are bonded to oxygen atoms and are, therefore, able to be donated. In ethanoic acid, only one hydrogen atom is bonded to an oxygen atom.
3. HSO_4^-
because
$$HSO_4^- + H^+ \rightarrow H_2SO_4$$
$$HSO_4^- \rightarrow SO_4^{2+} + H^+$$
4. C
5. D

11.3 Concentration and strength of acids and bases

Practice problem 2

$n(NaOH) = 0.135\ mol$

11.3 Exercise

1. **a.** 0.130 mol **b.** 0.240 mol **c.** 0.0475 mol
2. 0.882 M
3. **a.** 0.714 L **b.** 5.00 L **c.** 10.4 L
4. 0.00189 mol
5. A weak acid only partially ionises in water (only a small number of the acid molecules become ions).
6. $HN(CH_3)_2(aq) + H_2O(l) \rightleftharpoons H_2N^+(CH_3)_2(aq) + OH^-(aq)$
7. 0.5 M HNO_3, since it is both a strong acid and has the higher concentration.

11.3 Exam questions

1. D
2. A
3. $C_2H_5COOH(aq) + H_2O(l) \rightleftharpoons C_2H_5COO^-(aq) + H_3O^+(aq)$
4. Strong acids completely ionise to produce solutions that are good conductors of electricity;
e.g. $HBr\ (aq) + H_2O\ (l) \rightarrow Br^-\ (aq) + H_3O^+\ (aq)$.
Weak acids only partially ionise to produce solutions that are poor conductors of electricity;
e.g. $HS^-\ (aq) + H_2O\ (l) \rightleftharpoons S^{2-}\ (aq) + H_3O^+\ (aq)$.
5. **a.** HNO_3/NO_3^- and HCl/Cl^-
b. Wear protective clothing, safety glasses and safety gloves, and work in a well-ventilated area.

11.4 The pH scale

Practice problem 3

pH = 2

Practice problem 4

pH = 1.4

Practice problem 5

$[HNO_3] = 1.6 \times 10^{-2}\ M$

Practice problem 6

pH = 12.4

Practice problem 7

[KOH] = 0.0063 M

Practice problem 8

$c_2 = 1.02\ M$

Practice problem 9

450 mL

Practice problem 10

pH = 1.1

11.4 Exercise

1. **a.** $[OH^-] = 10^{-9}\ M$
b. $[H_3O^+] > [OH^-]$, so the solution is acidic.
2. The pH scale provides a measure of the acidity of a substance. The formula for calculating pH is $pH = -\log_{10}[H_3O^+]$.
3. **a.** A weak base should be added, such as sodium carbonate.
b. Alkaline
4. **a.** pH = 5 **b.** pH = 1
5. pH = 2.82
6. **a.** pH = 10 **b.** pH = 13
7. **a.** pH = 3.04 **b.** pH = 9.25
8. In a neutral solution, $[H_3O^+] = [OH^-] = 10^{-7}\ M$.
$$pH = 7$$
9. If the HCl acid solution has a low enough concentration to compensate for the smaller hydronium concentration, the pH of the two acids could be the same.
10. **a.** **i.** $pH = -\log[H_3O^+] = -\log(0.001) = 3.0$
ii. $pH = -\log[H_3O^+] = -\log(0.0005) = 3.3$

b. i. $[H_3O^+] \times [OH^-] = 10^{-14}\ M^2$.

$$\text{Therefore, } [H_3O^+] = \frac{10^{-14}}{[OH^-]}$$
$$= \frac{10^{-14}}{0.001}$$
$$= 10^{-11}\ M$$
$$\text{And pH} = -\log[H_3O^+]$$
$$= -\log(10^{-11})$$
$$= 11$$

ii. $[H_3O^+] \times [OH^-] = 10^{-14}\ M^2$.

$$\text{Therefore, } [H_3O^+] = \frac{10^{-14}}{[OH^-]}$$
$$= \frac{10^{-14}}{0.000500}$$
$$= 2.0 \times 10^{-11} M$$
$$\text{pH} = -\log[H_3O^+]$$
$$= -\log(2.0 \times 10^{-11})$$
$$= 10.7$$

11. 31.3 mL

11.4 Exam questions

1. a. i. $[H_3O^+] = 10^{-8}$ M

ii. pH = 8

b. i. $[H_3O^+] = 4 \times 10^{-13}$ M

ii. pH = 12.4

2. a. 4 **b.** 4.82 **c.** –0.30

3. 0.0025 M

4. D

5. D

11.5 Measuring pH

Practice problem 11

1. a. i. *Tap water:* Expected pH: 6–7 (relatively neutral – slightly acidic).
Expected colour: blue/purple.

ii. *Carbonated water:* Expected pH: 4–5.
Expected colour: pink-purple.

iii. *Bicarb soda:* $NaHCO_3$; expected pH: 8–9.
Expected colour: bright blue.

iv. *Vinegar:* Ethanoic acid, CH_3COOH; expected pH: 4–5.
Expected colour: pink-purple.

v. *Drain cleaner:* NaOH; expected pH: 12–14.
Expected colour: green-yellow.

vi. *Pool Clear:* HCl; expected pH: 1–2.
Expected colour: bright red-pink.

b. Anthocyanin molecules.
Anthocyanins are in the most protonated state in high acidic environments. As the pH increases, anthocyanins will start donating their hydrogen ions, becoming less protonated, and the colour will change from red to pink, purple, blue, green and yellow as the pH becomes highly alkaline.
The most protonated state is in highly acidic environments, predicted to be in the Pool Clear solution, which contains the strong acid hydrochloric acid.

Practice problem 12

a. Student 3

b. Student 2

c. Student 1 had the most accurate data while Student 3 had the most precise data.

11.5 Exercise

1. A

2. C

3. pH meters need to be calibrated to ensure results are accurate (reducing systematic errors).
pH is temperature dependent, and ideally should be recorded at 25 °C.

4. Any two of:
- pH is temperature dependent, so you should try to control the air temperature surrounding your testing zone to 25°C.
- If a solution being tested has a colour of its own, the colour of the indicator may be distorted.
- Colour is subjective. This can be a significant issue for people who are colour-blind.

5. a. Purple-blue

b. Clear; using many samples to confirm

c. Yellow

d. Accurate and precise results of 7.40.

11.5 Exam questions

1. C

2. Orange

3. C

4. a. Anthocyanins

b. pH 3 – 4

c. Sodium hydrogen carbonate ($NaHCO_3$); pH 8.0.

d. The anthocyanin molecule begins to deprotonate from its acidic conformation to another alkaline conformation. This change in structure results in the molecule absorbing orange wavelength light and appearing blue in colour.

5. a. $[H_3O^+] < [OH^-]$

b. i. A weak base or any solution around pH 8.

ii. Example: bicarb soda (sodium hydrogen carbonate).

c. i. Starting as blue, you would expect the colour to change to purple.

ii. More acidic.

d. As you blow carbon dioxide, CO_2, into the solution, a small percentage of the CO_2 gas molecules react and dissolve with the OH^- ion from water, resulting in aqueous carbonic acidic, H_2CO_3. This is a weak acid and will release a H^+ ion, creating the bicarbonate ion, HCO_3^-. This results in the protonation of the anthocyanin molecule, causing the colour observed to change from blue to purple.

e. The dissolved CO_2 is much greater in soda water than by blowing through a straw; therefore, the solution would become even more acidic (pH ~3.5), further protonating the anthocyanin molecule, causing the colour to change to pink.

f. Simply by adding a more alkaline component to your solution will deprotonate the anthocyanin molecule. This will cause the hydronium ion concentration to decrease, increasing the pH and shifting the colour observed back towards blue.

11.6 Neutralisation reactions to produce salts

Practice problem 13

a. $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

b. $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

Practice problem 14

a. $2CH_3COOH(aq) + Na_2CO_3(s) \rightarrow 2CH_3COONa(aq) + CO_2(g) + H_2O(l)$

b. $2H^+(aq) + Na_2CO_3(s) \rightarrow 2Na^+(aq) + CO_2(g) + H_2O(l)$

Practice problem 15

a. $2HCl(aq) + Mg(OH)_2(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$

b. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

11.6 Exercise

1. See table at the bottom of the page*

2. a. Acid + metal → salt and hydrogen
   b. Acid + metal carbonate → salt + water + carbon dioxide
   c. Acid + metal hydroxide → salt + water

3. a. $Zn(s) + 2HBr(aq) \rightarrow ZnBr_2(aq) + H_2(g)$
   b. $Ca(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + H_2(g)$
   c. $2K(s) + 2HNO_3(aq) \rightarrow 2KNO_3(aq) + H_2(g)$

4. a. $MgCO_3(s) + 2CH_3COOH(aq) \rightarrow Mg(CH_3COO)_2 + CO_2(g) + H_2O(l)$
   b. $Na_2CO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$
   c. $MnCO_3(s) + 2HNO_3(aq) \rightarrow Mn(NO_3)_2(aq) + CO_2(g) + H_2O(l)$

5. a. $2NaOH + H_2SO_4 \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
   b. $Ca(OH)_2 + 2HCl \rightarrow CaCl_2(aq) + 2H_2O(l)$
   c. $Mg(OH)_2 + 2HNO_3 \rightarrow Mg(NO_3)_2(aq) + 2H_2O(l)$

6. a. $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$; spectator ion Br^-
   b. $MgCO_3(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + CO_2(g) + H_2O(l)$; spectator ion CH_3COO^-
   c. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$; spectator ions Na^+, SO_4^{2-}

7. a. $Ca(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + H_2(g)$
   $Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$
   b. $ZnCO_3(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + CO_2(g) + H_2O(l)$
   $ZnCO_3(s) + 2H^+(aq) \rightarrow Zn^{2+} + CO_2(g) + H_2O(l)$
   c. $Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$
   $CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$

8. a. Collect the gas from above the reaction flask in a test tube. Place a lit match at the mouth of the test tube. If a pop is heard, the gas is hydrogen.
   b. $2H_2(g) + O_2(g) \rightarrow H_2O(l)$
   c. No, there may not be enough gas produced or collected.

9. a. Direct the gas downwards into a test tube containing limewater, $Ca(OH)_2$. If a white precipitate is observed, the gas is likely to be carbon dioxide.
   b. $CO_2(g) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + 2H_2O(l)$

10. a. $Ni(s) + H_2SO_4(aq) \rightarrow NiSO_4(aq) + H_2(g)$
    $Ni(s) + 2H^+(aq) \rightarrow Ni^{2+}(aq) + H_2(g)$
    Spectator ions SO_4^{2-}
    b. $2Al(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$
    $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$
    Spectator ions SO_4^{2-}
    c. $CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$
    $CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+} + CO_2(g) + H_2O(l)$
    Spectator ions NO_3^-

11.6 Exam questions

1. B
2. C
3. $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$
4. $2HNO_3(aq) + CaCO_3(s) \rightarrow Ca(NO_3)_2(aq) + H_2O(l) + CO_2(g)$
5. a. Full equation: $2HCl(aq) + Mg(OH)_2(s) \rightarrow MgCl_2(aq) + 2H_2O(l)$
   Ionic equation: $2H^+(aq) + Mg(OH)_3(s) \rightarrow Mg^{2+}(aq) + 2H_2O(l)$
   b. Full equation: $3HCl(aq) + Al(OH)_3(s) \rightarrow AlCl_3(aq) + 3H_2O(l)$
   Ionic equation: $3H^+(aq) + Al(OH)_3(s) \rightarrow Al^{3+}(aq) + 3H_2O(l)$
   c. Full equation: $2HCl(aq) + MgCO_3(s) \rightarrow MgCl_2(aq) + H_2O(l) + CO_2(g)$
   Ionic equation: $2H^+(aq) + MgCO_3(s) \rightarrow Mg^{2+}(aq) + H_2O(l) + CO_2(g)$

*1.

Acid	Formula of acid	Type of salt formed from the acid	Anion present
Hydrochloric	HCl	Chloride	Cl^-
Nitric	HNO_3	Nitrate	NO_3^-
Sulfuric	H_2SO_4	Sulfate	SO_4^{2-}
Ethanoic acid	CH_3COOH	Ethanoate	CH_3COO^-

11.7 Applications of acid–base reactions in society

11.7 Exercise

1. Acid rain is rain made acidic by mixing with atmospheric pollution, resulting a pH of less than 7.
2. Carbon dioxide and water produces carbonic acid.
   $CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$
3. Carbonic acid is a weak acid and partially ionises.
   $H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$
4. $SO_2(g) + H_2O(l) \rightleftharpoons H_2SO_3(aq)$
   $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$
   $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$
5. Burning of coal and other fossil fuels together with active volcanoes produce SO_2.
6. Possible control methods include:
   - using natural gas because it produces less SO_2
   - using coal with a lower SO_2 content
   - removing SO_2 before emitting waste gases and using it in other industries.
7. Oceans are slightly alkaline but the added CO_2 being absorbed by the oceans is lowering the pH.
8. 1.3×10^{-6} M
9. Problems that might occur as oceans become less basic include:
   - coral bleaching
   - interference with marine organisms
   - reproductive disorders in some organisms
   - interference with shell building.
10. Calcium carbonate is a constituent of corals and the shells of shellfish. The presence of carbon dioxide decreases the already small concentration of carbonate ions present:
    $CO_2(g) + H_2O(l) + CO_3^{2-}(aq) \rightarrow 2HCO_3^-(aq)$
    If the carbonate ion concentration decreases too much, the shells of marine organisms may be affected.

11.7 Exam questions

1. C
2. B
3. B
4. a. $[H_3O^+] = 0.5$ M
   b. Acidic gases released underground by the volcano; including hydrochloric and sulfuric acids.
   c. Maybe. For an explanation, see the worked solution in your online resources.
5. a. Increases in atmospheric CO_2 levels lead to the following flow-on effects:
   - atmospheric CO_2 dissolves in ocean water to produce the weak acid carbonic acid, H_2CO_3
   - carbonic acid, H_2CO_3, dissociates into bicarbonate ions, HCO_3^-, and hydrogen ions, H^+
   - this increases the H^+ ion levels in oceans
   - this decreases the pH in the ocean, leading to ocean acidity.

   $$CO_2 + H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons HCO_3^- + H^+$$

   b. Any two of the following:
   - coral bleaching
   - interference with marine organisms, particularly those at the lower end of the food chain
   - reproductive disorders in certain marine organisms
   - interference with shell building in shellfish and other similar organism
   - changes to food chains in ecosystems, potentially leading to dead zones.

   c. Sample response:
   - Impacts on livelihoods for people who rely on the ocean for food and employment.
   - Reduction in tourism due to coral bleaching.

11.8 Review

11.8 Review questions

1. a. A Brønsted–Lowry acid is a proton donor and a Brønsted–Lowry base is a proton acceptor.

   $HCl(aq) + H_2O \rightarrow H_3O^+(aq) + Cl^-(aq)$

   acid → H^+ → base

   b. Ethanoic acid is a weak acid, while hydrochloric acid is a strong acid. Hydrochloric acid will completely ionise and produce a significant amount of hydronium ions in comparison to ethanoic acid, which only partially ionises and will produce significantly fewer hydronium ions. Therefore, hydrochloric acid would be expected to turn red-pink using the red cabbage indicator, while ethanoic acid should be more purplely-pink.

   c. Red cabbage contains a variant of the anthocyanin molecule. In acidic environments, it appears red. However, as the pH increases with a decrease in hydronium ion concentration, the anthocyanin molecule starts donating its hydrogen ions, changing the molecular conformation, resulting in the appearance of different colours from purple (slightly acidic) to blue (slightly alkaline) to green and yellow (alkaline).

   d. Using a calibrated pH meter would enable more accurate and precise pH measurements.

   e. $CH_3COOH(aq) + H_2O(l) \rightarrow CH_3COO^-(aq) + H_3O^+(aq)$
2. $Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$
   $Al(OH)_3(s) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$
   $NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$
3. $2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$
4. $H_2SO_4(aq) + H_2O(l) \rightarrow HSO_4^-(aq) + H_3O^+(aq)$
   $HSO_4^-(aq) + H_2O(l) \rightleftharpoons SO_4^{2-}(aq) + H_3O^+(aq)$
5. a. A weak base
   b. Alkaline
6. a. Conjugate acid–base pairs differ by one proton. The conjugate base of an acid is formed in a reaction where the acid loses a proton. The conjugate acid of a base is formed in a reaction where the base gains a proton.

   b. i. HCl ii. HCO_3^- iii. H_2SO_4
   iv. H_3O^+ v. H_2O

   c. i. F^- ii. CO_3^{2-} iii. SO_4^{2-}
   iv. OH^- v. HS^-

7. a. $HCO_3^-(aq) + H_2O(l) \rightleftharpoons CO_3^{2-}(aq) + H_3O^+(aq)$

$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$

b. $H_2S(aq) + H_2O(l) \rightleftharpoons HS^-(aq) + H_3O^+(aq)$

$HS^-(aq) + H_2O(l) \rightleftharpoons S^{2-}(aq) + H_3O^+(aq)$

c. $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

or $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

8. a. $3Na_2S(s) + 2H_3PO_4(aq) \rightarrow 2Na_3PO_4(aq) + 3H_2S(g)$
$2H^+(aq) + Na_2S(s) \rightarrow 2Na^+(aq) + H_2S(g)$

b. $2HCl(aq) + CuO(s) \rightarrow CuCl_2(aq) + H_2O(l)$
$2H^+(aq) + CuO(s) \rightarrow Cu^{2+}(aq) + H_2O(l)$

c. $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$
$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

9. a. i. $n(\text{NaOH}) = \dfrac{m}{M} = \dfrac{0.120}{40.0}\text{mol} = 3.00 \times 10^{-3}\text{mol}$

$$c(\text{NaOH}) = \frac{n}{V} = \frac{3.00 \times 10^{-3}}{0.2500}\text{M} = 1.20 \times 10^{-2}\text{M}$$

ii. $c(\text{NaOH}) = 1.20 \times 10^{-2}\text{M} = [OH^-]$

$$[H_3O^+] = \frac{10^{-14}\text{M}^2}{[OH^-]} = \frac{10^{-14}\text{M}^2}{1.20 \times 10^{-2}\text{M}} = 8.33 \times 10^{-13}\text{M}$$

iii. $\text{pH} = -\log[H_3O^+] = -\log[8.33 \times 10^{-13}] = 12.08$

b. i. Less basic

ii. pH = 11.08
Note: Diluting by a factor of 10 will reduce the pH by one unit for bases (and increase by one unit for acids).

c. i. HCl (or HNO_3) and ethanoic acid. (Other acceptable strong acids are HCl and HNO_3, but not H_2SO_4. For the weak acid, any carboxylic acid could be named.)

ii. The strong acid has the higher acidity constant.

10. a. i. Accuracy refers to how close the data points are to the true value of the measurement.

ii. Precision refers to how close the repeated measurements are to each other, such that your data has a low range of values.

b. When you are using a pH meter, if it is not calibrated, you would still expect your pH values to be precise; however, the accuracy may be wrong. If you use a calibrated pH meter, you would expect a highly accurate, highly precise pH reading. If you were to use commercial pH indicators, these would be less accurate than using a pH meter purely because they provide a qualitative measurement. Natural pH indicators would also be less accurate than commercial pH indicators because they have not been aligned to quantitative pH values.

11.8 Exam questions

Section A — Multiple choice questions

1. D	2. C	3. B	4. C	5. B
6. B	7. D	8. C	9. D	10. A

Section B — Short answer questions

11. a. Monoprotic means it can donate only one proton.

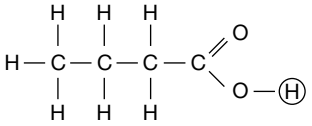
b. $C_3H_7COOH(aq) + H_2O(l) \rightleftharpoons C_3H_7COO^-(s) + H_3O^+(aq)$

c. $C_3H_7COOH(aq) + NaOH(aq) \rightarrow C_3H_7COO^-(aq) + H_2O(l) + Na^+(aq)$

d. The equation in (b) does not react completely; butanoic acid is only partially ionised in water.

e. Measure the conductivity. Weak acids only partly ionise, so there will not be a significant increase in conductivity.

f.



12. a. Amphiprotic means it can donate or receive a proton.

b. $H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3PO_4(aq) + OH^-(aq)$
$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$

c. The second reaction occurs to a greater extent than the first

13. a. Ammonium sulfate could be produced by combining ammonia and sulfuric acid.

b. $2NH_3(g) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq)$

c. Full ionic equation: $NH_3(g) + H^+(aq) \rightarrow NH_4^+$

d. Ammonium sulfate would be obtained by evaporating the water from the solution.

e. $(NH_4)_2SO_4(s) \rightarrow 2NH_4^+(aq) + SO_4^{2-}(aq)$

14. a. Iron, and some boron, manganese

b. Calcium, magnesium, nitrogen, potassium

c. Phosphorus

d. Boron, iron, manganese

e. Boron, iron, manganese and potassium

f. Phosphorus

g. The soil for the lettuces may be lacking a particular nutrient. To test this, set up nine identical pots of seedlings using the gardener's soil and add a specified volume of pure water daily to pot 1 (as the control). For the remaining pots, add the same volume of solutions containing the same concentration of each of the elements listed. Maintain the same conditions of light and temperature for all pots. Observe over a specified period of time and measure the heights of the seedlings to see which has grown successfully.

15. a. $CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$

b. $H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+ + HCO_3^-$
$H_2CO_3(aq)$ is a weak acid that reacts with water to produce the hydronium ion and the hydrogen carbonate ion. The acidity is due to the presence of the hydronium ion.

c. $H_2SO_4(aq) + CaCO_3(s) \rightarrow CaSO_4(aq) + H_2O(l) + CO_2(g)$

d. Sulfur dioxide, SO_2, sulfur trioxide, SO_3, and nitrogen dioxide, NO_2.

12 Redox (electron transfer) reactions

12.2 Redox reactions

Practice problem 1

$2Br^-$ has lost two electrons to become Br_2.

Cl_2 has gained two electrons to become $2Cl^-$.

Br^- is oxidised, while Cl_2 is reduced; that is, the chlorine molecule is reduced.

Practice problem 2

a. $2Ag^+(aq) + Cu(s) \rightarrow Cu^{2+}(aq) + 2Ag(s)$

b. $Ag^+(aq) + e^- \rightarrow Ag(s)$
$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

c. Conjugate pairs: $Ag^+(aq)/Ag(s)$ and $Cu^{2+}(aq)/Cu(s)$

Practice problem 3

$SO_2(g) + 2H_2O(l) \rightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$

12.2 Exercise

1. Sample answer: fireworks, explosions, combustion, photosynthesis, respiration and metal corrosion.
2. **a.** Oxidation **b.** Reduction **c.** Oxidation
   d. Oxidation **e.** Reduction
3. **a.** $Pb^{2+}(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + Pb(s)$
   b. $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$ and $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
   c. Zinc metal
   d. Lead ions
   e. Lead ions are the oxidising agents.
   f. Zinc metal is the reducing agent.
4. This reaction is not an acid–base reaction because no protons have been transferred. Instead, it is a redox reaction because electrons have been transferred. Magnesium atoms have been oxidised and the hydrogen ions have been reduced.
   $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$ and $2H^+(aq) + 2e^- \rightarrow H_2(g)$
5. **a.** Reduction: $Ag^+(aq) + e^- \rightarrow Ag(s)$.
   Oxidation: $Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$.
   b. Silver ions accept electrons to form silver metal.
   c. Lead metal, Pb(s), is oxidised to lead ions, $Pb^{2+}(aq)$.
   d. Lead metal
   e. Lead metal
   f. $Ag^+(aq)/Ag(s)$ and $Pb^{2+}/Pb(s)$
6. **a.** $2H^+$ accepts electrons to form $H_2(g)$.
   b. H^+ is reduced to H_2 because it is the species accepting electrons.
   c. Al(s) is the reducing agent.
   d. Al(s). The reducing agent has been oxidised.
   e. Al is the reducing agent because it donates electrons, $Al \rightarrow Al^{3+}$.
   f. The oxidising agent is H^+ because it accepts electrons.
   g. $Al^{3+}(aq)/Al(s)$ and $H^+(aq)/H_2(g)$
   h. $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$
7. **a.** $Ca(s) \rightarrow Ca^{2+}(s) + 2e^-$
   $Cl_2(g) + 2e^- \rightarrow 2Cl^-(s)$
   $Ca^{2+}(s)/Ca(s)$ and $Cl_2(g)/Cl^-(s)$
   b. $Mg(s) \rightarrow Mg^{2+}(s) + 2e^-$
   $S(s) + 2e^- \rightarrow S^{2-}(s)$
   $Mg^{2+}(s)/Mg(s)$ and $S(s)/S^{2-}(s)$
   c. $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$
   $Ag^+(aq) + e^- \rightarrow Ag(s)$
   $Al^{3+}(aq)/Al(s)$ and $Ag^+(aq)/Ag(s)$
8. **a.** $NO_3^-(aq) + 2H^+(aq) + e^- \rightarrow NO_2(g) + H_2O(l)$
   b. $ClO^-(aq) + 2H^+(aq) + 2e^- \rightarrow Cl^-(aq) + H_2O(l)$
   c. $BrO_3^-(aq) + 2H^+(aq) + e^- \rightarrow BrO_2(aq) + H_2O(l)$
9. Responses should include the steps towards these half-equations. Fully worked solutions are available in your digital formats.
   $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$
   $NO_2^-(aq) + H_2O(l) \rightarrow NO_3^-(aq) + 2H^+(aq) + 2e^-$
10. **a.** $I_2(s) + H_2S(g) \rightarrow 2I^-(aq) + S(s) + 2H^+(aq)$
    b. $5SO_3^{2-}(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(l)$
    c. $3Cu(s) + 2NO_3^-(aq) + 8H^+(aq) \rightarrow 3Cu^{2+}(aq) + 2NO(g) + 4H_2O(l)$

12.2 Exam questions

1. A
2. C
3. **a.** **i.** $Br_2(I) + 2e^- \rightarrow 2Br^-(aq)$
   ii. $Co(s) \rightarrow Co^{2+}(aq) + 2e^-$
   b. Reaction (i) is the reduction half-equation.
   c. $Br_2(1) + Co(s) \rightarrow Co^{2+}(aq) + 2Br^-(aq)$
   or $Br_2(1) + Co(s) \rightarrow CoBr_2(aq)$
   d. Co is the reducing agent.
   e.
   - The mixture would initially be deep orange due to bromine with the grey metal cobalt present. As time progresses, the gray metal would dissolve.
   - The orange colour due to bromine would fade and the cobalt(II) ions would from a pink-red solution.
4. **a.** $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$
   b. $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$
   c. $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$: Oxidation
   $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$: Reduction
5. **a.** $2Br^-(aq) + SO_4^{2-}(aq) + 4H^+(aq) \rightarrow Br_2(l) + SO_2(g) + 2H_2O(l)$
   b. $I_2(s) + H_2S(g) \rightarrow I^-(aq) + S(s) + H^+(aq)$

12.3 EXTENSION: Oxidation numbers

Practice problem 4

a. The oxidation number of phosphorus is +5.

b. The oxidation number of nitrogen is +5.

c. The oxidation number of Cl^- is –1.

d. The oxidation number of K is 0.

Practice problem 5

Oxidation numbers: $\overset{0}{Pb}$, $\overset{0}{Br}$, $\overset{2+}{Pb}$, $\overset{-1}{Br}$

Lead is the reducing agent.

Br_2 is the oxidising agent.

Electrons have been transferred from the lead to the bromine and so this is a redox reaction.

12.3 Exercise

1. **a.** Oxidation number P = 0
   b. Oxidation number $F_2 = 0$
   c. Oxidation number is $Zn^{2+} = 2+$.
   d. Oxidation number of H = +1, O = –2, C = +4

e. Oxidation number of Na = +1 (ion), O = −2, N = +5

f. Oxidation number of O = −2, Be = +2

g. Oxidation number of Ag^+ = +1, S = −2

2. a. i. 0 ii. +5 iii. −3 iv. +4

b. NH_4^+

3. NH_4^+, N_2, NO_2, NO_3^-

4. The maximum oxidation number sulfur can have is six.

5. a. $\overset{0}{Br_2}(l) + \overset{+4-2}{SO_2}(g) \rightarrow \overset{-1}{2Br^-}(aq) + \overset{+6-2}{SO_4^{2-}}(aq)$
is a redox reaction.

b. $\overset{+1-2+1}{KOH}(aq) + \overset{+1+5-2}{HNO_3}(aq) \rightarrow \overset{+1+5-2}{KNO_3}(aq) + \overset{+1\ -2}{H_2O}(aq)$
is not a redox reaction.

c. $\overset{+3}{Fe^{3+}}(aq) + 3\overset{-2+1}{OH^-}(aq) \rightarrow \overset{+3\ -2+1}{Fe(OH)_3}(s)$
is not a redox reaction.

d. $\overset{0}{5Zn}(s) + 2\overset{+5-2}{NO_3^-}(aq) + 12\overset{+1}{H^+}(aq) \rightarrow 5\overset{+2}{Zn^{2+}}(aq) + \overset{0}{N_2}(g) + 6\overset{+1\ -2}{H_2O}(l)$ is a redox reaction.

12.3 Exam questions

1. B

2. C

3. a. i. $\overset{0}{Ca}(s) + 2\overset{+1\ -2}{H_2O}(l) \rightarrow \overset{+2\ -2+1}{Ca(OH)_2}(aq) + \overset{0}{H_2}(g)$

ii. Calcium has been oxidised and water has been reduced.

iii. Oxidising agent: H_2O; reducing agent: Ca.

b. i. $\overset{+4-2}{PbO_2}(s) + \overset{-1}{I^-} \rightarrow \overset{+2}{Pb^{2+}}(aq) + \overset{0}{I_2}(g)$

ii. Lead oxide has been reduced and iodide ions have been oxidised.

iii. Oxidising agent: PbO_2; reducing agent: I^-.

c. i. $\overset{-1}{I^-}(aq) + \overset{+3\ -2}{NO_2^-}(aq) \rightarrow \overset{0}{I_2}(s) + \overset{+2-2}{NO}(g)$

ii. The nitrite ions, NO_2^-, have been reduced and iodide ions have been oxidised.

iii. Oxidising agent: NO_2^-; reducing agent: I^-.

d. i. $\overset{+1\ -1}{H_2O_2}(aq) + \overset{+7\ -2}{MnO_4^-} \rightarrow \overset{+2}{Mn^{2+}}(aq) + \overset{0}{O_2}(g)$

ii. The permanganate ions, MnO_4^-, have been reduced and the hydrogen peroxide molecules have been oxidised.

iii. Oxidising agent: MnO_4^-; reducing agent: H_2O_2.

4. a. −1 b. Sn^{2+}
c. H_2O_2 d. Hydrogen peroxide
e. Tin II ion (Sn^{2+})

5.

Change	Oxidation	Reduction	Neither
$O_2 \rightarrow O_3$			✓
$MnO_2 \rightarrow MnO_4^-$	✓		
$P_2O_5 \rightarrow P_4H_{10}$			✓
$NH_3 \rightarrow NO_2$	✓		
$Fe^{2+} \rightarrow Fe^{3+}$	✓		
$S_2O_3^{2-} \rightarrow S_2O_4^{2-}$	✓		

12.4 Reactivity series of metals

Practice problem 6

a. $Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$

b. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$

c. Zn is the reducing agent.
Pb^{2+} is the oxidising agent.

12.4 Exercise

1. a. Group 1 metals, they only need to lose one electron to achieve a stable full outer shell

b. The reactivity increases as you go down a group.

2. a. Silver (Ag), gold (Au)

b. Iron (Fe) will displace tin ions (Sn^{2+}) but will be displaced by zinc (Zn)

c. Potassium (K), sodium (Na), calcium (Ca)

d. Tin (Sn) and lead (Pb)

e. Calcium (Ca)

f. Copper (Cu) and mercury (Hg)

3. a. The blue colour, $CuSO_4$, of the solution fades as the copper metal forms. This displacement reaction occurs because iron, Fe, is more reactive than copper, Cu. The iron dissolves and brown copper metal is formed.

b. The products are copper and iron sulfate solution.

4. a. $Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$
Reducing agent: Mg; oxidising agent: Cu^{2+}

b. $2Al(s) + 3Pb^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Pb(s)$.
Reducing agent: Al; oxidising agent: Pb^{2+}

c. No reaction because copper is less reactive than iron.

d. No reaction because silver is less reactive than zinc.

5. a. $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$
$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
$Pb^{2+}(aq) + Fe(s) \rightarrow Pb(s) + Fe^{2+}(aq)$

b. $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$
$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
$Sn^{2+}(aq) + Zn(s) \rightarrow Sn(s) + Zn^{2+}(aq)$

c. No reaction because iron is less reactive than aluminium and so will not displace aluminium from solution.

d. $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$
$Al(s) \rightarrow Al^{3+}(aq) + 3e^-$
$3Zn^{2+}(aq) + 2Al(s) \rightarrow 3Zn(s) + 2Al^{3+}(aq)$

6. a. The aluminium metal may have had an unreactive oxide coating.

b. Use sand paper or emery paper to strip off the coating to expose the metal before repeating the experiment.

7. $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$
Oxidising agent: H_2O
Reducing agent: Na

8. Aluminium oxide
$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

9. a. $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
Magnesium chloride solution and hydrogen gas

b. $Mg(s) + 2H_2O(g) \rightarrow Mg(OH)_2(aq) + H_2(g)$
Magnesium hydroxide and hydrogen

c. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$
Magnesium oxide

10. Gold would be too expensive and too soft, and magnesium would be too reactive. Copper would be a better choice or perhaps an alloy of copper. Factors to be considered

include toxicity, availability and cost of metal, lack of reactivity, hardness, ability to be shaped, and ease and cost of manufacture.

12.4 Exam questions

1. C
2. B
3. D
4. **a.** Overall: $3Sn^{2+}(aq) + 2Al(s) \rightarrow 3Sn(s) + 2Al^{3+}(aq)$
   b. Overall: $Pb(s) + 2Ag^{+}(aq) \rightarrow Pb^{2+}(aq) + 2Ag(s)$
5. **a.** *P* since it reacted with all other solutions.
   b. $Q^{2+} > S^{2+} > R^{2+} > P^{2+}$

12.5 Applications of redox reactions

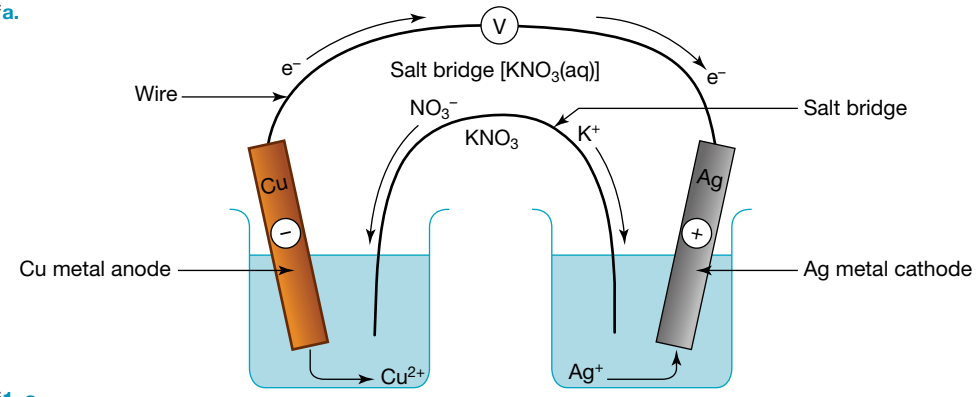
Practice problem 7

a. See figure at the bottom of the page*
b. Oxidation: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^{-}$
Reduction: $Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$
c. Overall ionic redox equation: $Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + Ag(s)$

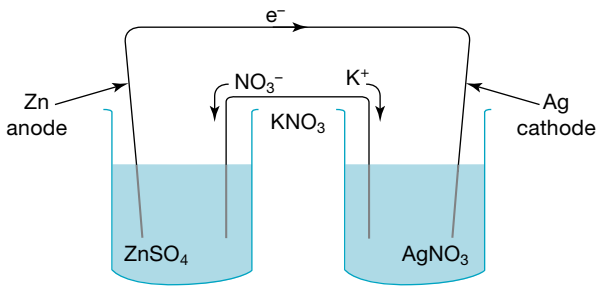
12.5 Exercise

1. **a.** See figure at the bottom of the page**
   b. Oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-}$
   Reduction: $Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$
   c. $Zn(s) + 2Ag^{+}(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$
2. **a.** See figure at the bottom of the page#
   b. and **c.** Oxidation (at anode): $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^{-}$
   Reduction (at cathode): $Zn^{2+}(aq) + 2e^{-} \rightarrow Zn(s)$
   d. Overall reaction: $Mg(s) + Zn^{2+}(aq) \rightarrow Mg^{2+}(aq) + Zn(s)$
   Zn^{2+} is the oxidising agent, Mg is the reducing agent.
3. **a.** The species will be reduced.
   b. Cathode
4. **a.** Corrosion is the oxidation of metals by materials in their environment.
   b. Corrosion is a redox reaction.
5. Dry corrosion occurs when metals react directly with oxygen. For example, when sodium reacts directly with oxygen and so must be stored in oil. Wet corrosion involves the presence of water. For example, iron in a damp environment.

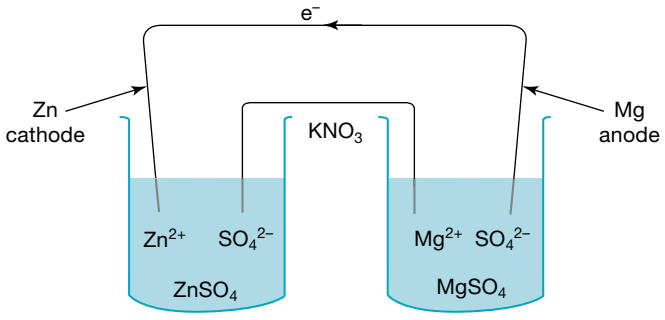
*a.



**1. a.



#2. a



6. Rusting is the wet corrosion of iron. Corrosion refers to the oxidation of metals in general.

7. Water and oxygen are necessary for rusting to occur, and this process is speeded up in the presence of salty water.
$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$
$2Fe(s) + O_2(g) + 2H_2O(l) \rightarrow 2Fe(OH)_2(s)$

8. Set up two identical test tubes with the same volume of each type of water at the same temperature and place identical iron nails in each. Observe over the same period of time.

9. Iron is versatile, relatively abundant and relatively economically extracted; it is strong and as a steel alloy is less susceptible to corrosion.

10. a. A bike chain is protected by grease.
b. A metal dish-draining rack is protected by a waterproof plastic coating.
c. A metal roof is protected by painting.
d. A ship's hull is protected by sacrificial protection.

12.5 Exam questions

1. B

2. a. See figure at the bottom of the page*
i. Pb is oxidised
Cu^{2+} is reduced
ii. Anode: Pb
Cathode: Cu^{2+}
iii. From anode to cathode

b. $Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

c. $Pb(s) + Cu^{2+}(aq) \rightarrow Pb^{2+}(aq) + Cu(s)$

3. a. $Pb^{2+}(aq)/Pb$ b. $Mg^{2+}(aq)/Mg$
c. $Pb^{2+}(aq)/Pb$ d. $Zn^{2+}(aq)/Zn$

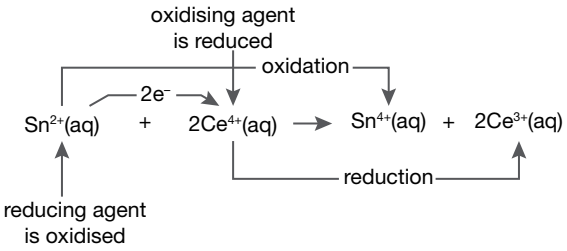
4. a. Metal, oxygen and water
b. Salt, acids and pollutants
c. i. Very little water ii. No oxygen
iii. Very little oxygen

5. a. The block of zinc will corrode instead of the iron.
b. No protection would be provided, because lead is less reactive than iron.
c. Corrosion could also be prevented through painting or coating the statue with an unreactive material.

12.6 Review

12.6 Review questions

1. a. Oxidation is the gain of oxygen; reduction is the loss of oxygen.
b. Oxidation is the loss of electrons; reduction is the gain of electrons. (OIL RIG)
c. An oxidising agent is an electron acceptor; it is reduced.

2. a.


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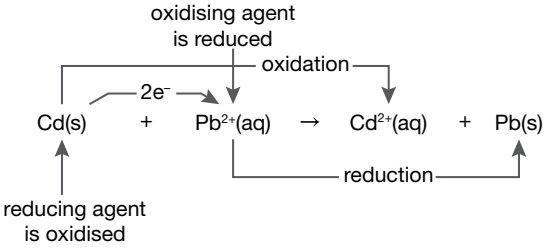

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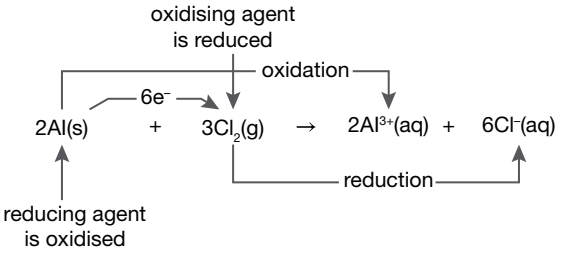

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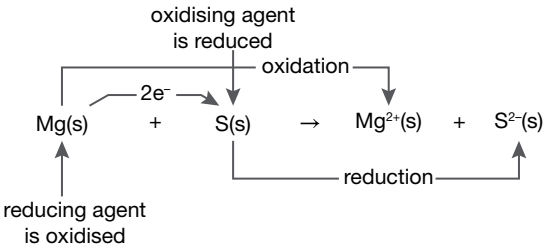
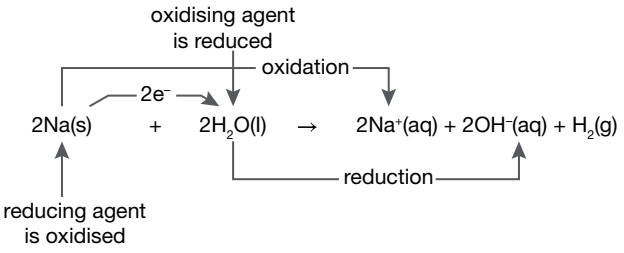
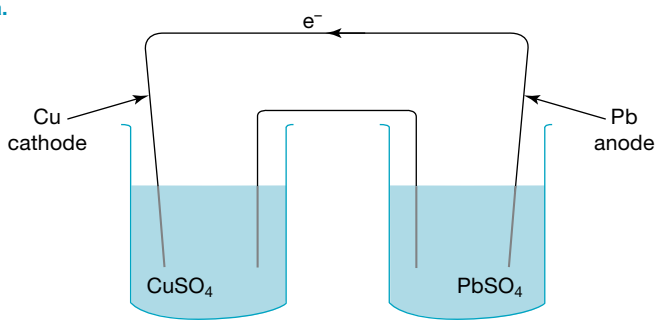


e.


*2. a.


3. $Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$

4. a. $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

   b. $Fe(s) + Sn^{2+}(aq) \rightarrow Fe^{2+}(aq) + Sn(s)$

   c. $Pb(s) + 2Ag^{+}(aq) \rightarrow Pb^{2+}(aq) + 2Ag(s)$

   d. $2Al(s) + 3Zn^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Zn(s)$
   In each case, the reducing agent has been placed first in the equation and the oxidising agent is second.

5. a. $Sn(s) + Cu^{2+}(aq) \rightarrow Sn^{2+}(aq) + Cu(s)$
   $Cd(s) + Sn^{2+}(aq) \rightarrow Cd^{2+}(aq) + Sn(s)$

   b. Cd > Sn > Cu (in order of decreasing reactivity)

   c. Only reaction (iv) will occur.

6. a. $E > A > F > C > D > B$.

   b. Sample answer: Mg, Zn, Fe, Sn, Pb, Cu

7. a. Galvanic cell: electrochemical cell that converts chemical energy into electrical energy.

   b. Electrode: a solid conductor of electricity in an electrochemical cell; can be made of inert graphite or be a metal; connects the electrolyte solution in the beakers to the wire.

   c. Electrolyte: the solution within a cell that conducts electricity.

   d. Anode: the electrode at which oxidation (loss of electrons) occurs.

   e. Cathode: the electrode at which reduction (gain of electrons) occurs.

8. a. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-}$

   b. $Fe^{2+}(aq) + 2e^{-} \rightarrow Fe(s)$

   c. $Zn(s) + Fe^{2+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s)$

9. Balanced reaction:
   $4Fe(OH)_2(s) + O_2(g) + 2H_2O(l) \rightarrow 4Fe(OH)_3(s)$
   The corrosion of iron metal is a redox reaction because it involves the transfer of electrons. Iron loses electrons to form iron ions and oxygen in the presence of water gains electrons to form hydroxide ions.

10. a. The steel will rust underneath the plastic.

    b. The steel will rust underneath the paint.

    c. The steel will rust underneath the tin.

    d. No effect will occur — the remaining zinc will continue to protect the steel.

12.6 Exam questions

Section A — Multiple choice questions

1. D	2. D	3. B	4. D	5. A
6. D	7. A	8. B	9. C	10. A

Section B — Short answer questions

11. a. Redox reaction b. Not a redox reaction
    c. Redox reaction d. Redox reaction

12. a. Reduction: $Fe_2O_3(s) + 6H^{+}(aq) + 6e^{-} \rightarrow 2Fe(s) + 3H_2O(l)$
    Oxidation: $(CO(g) + H_2O(l) \rightarrow CO_2(g) + 2H^{+}(aq) + 2e^{-}) \times 3$
    Balanced equation: $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

    b. Oxidation: $(P_4(s) + 16H_2O(l) \rightarrow 4H_3PO_4(aq) + 20\,H^{+}(aq) + 20e^{-}) \times 3$
    Reduction: $(HNO_3(aq) + 3H^{+}(aq) + 3e^{-} \rightarrow NO(g) + 2H_2O(l)) \times 20$
    Balanced equation: $3P_4(s) + 20HNO_3(aq) + 8H_2O(l) \rightarrow 12H_3PO_4(aq) + 20NO(g)$

    c. Reduction: $Cr_2O_7^{2-}(aq) + 14H^{+}(aq) + 6e^{-} \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$
    Oxidation: $(H_2C_2O_4(aq) \rightarrow 2CO_2(g) + 2H^{+}(aq) + 2e^{-}) \times 3$
    Balanced equation: $Cr_2\,O_7^{2-}(aq) + 3H_2C_2O_4(aq) + 8H^{+}(aq) \rightarrow 2Cr^{3+}(aq) + 6CO_2(g) + 7H_2O(l)$.

    d. Oxidation: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^{-}$
    Reduction: $(HNO_3(aq) + H^{+}(aq) + e^{-} \rightarrow NO_2(g) + H_2O(l)) \times 2$
    Balanced equation: $Cu(s) + 2HNO_3(aq) + 2H^{+}(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$

    e. Oxidation: $PbS(s) + 4H_2O(l) \rightarrow PbSO_4(s) + 8H^{+}(aq) + 8e^{-}$
    Reduction: $(H_2O_2(l) + 2H^{+}(aq) + 2e^{-} \rightarrow 2H_2O(l)) \times 4$
    Balanced equation: $PbS(s) + 4H_2O_2(l) \rightarrow PbSO_4(s) + 4H_2O(l)$

13. a. $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^{-}$

    b. $Pb^{2+}(aq) + 2e^{-} \rightarrow Pb(s)$

    c. Pb^{2+}

    d. Sn

    e. No. A further redox reaction will not occur because Mg^{2+} is a weaker oxidising agent than Sn^{2+}.

14. a. A^{+}, B^{2+}, C^{+}, D^{3+}

    b. $B > D > A > C$

    c. $B > D > A > C$

    d. B^{2+}

15. Only statements I and II are correct.

Unit 2 | Area of Study 1 review

Section A — Multiple choice questions

1. B	2. A	3. D	4. B	5. D
6. B	7. C	8. A	9. C	10. C
11. D	12. C	13. C	14. A	15. D
16. B	17. B	18. D	19. A	20. C

Section B — Short answer questions

21. a. Below 0 °C, water forms ice, where each water molecule is hydrogen bonded to four other water molecules in a tetrahedral arrangement. This is more spread out than the temporary hydrogen bonds formed by moving water molecules, reducing the density of ice compared to liquid water.

    b. 9.97×10^3 g

    c. 3.13×10^6 J

    d. As the temperature of the water increases, the molecules gain kinetic energy, causing them to become more widespread and thus reducing the overall density.

22. a. $HNO_3(aq) + H_2O(1) \rightarrow H_3O^{+}(aq) + NO_3^{-}(aq)$

    b. $H_2S(aq) + H_2O(1) \rightleftharpoons H_3O^{+}(aq) + HS^{-}(aq)$
    $HS^{-}(aq) + H_2O(1) \rightleftharpoons H_3O^{+}(aq) + S^{2-}(aq)$

    c. $HSO_3^{-}(aq) + H_2O(1) \rightleftharpoons H_3O^{+}(aq) + SO_3^{2-}(aq)$
    $HSO_3^{-}(aq) + H_2O(1) \rightleftharpoons OH^{-}(aq) + H_2SO_3^{2}(aq)$

23. a. When placed in water, HCOOH forms ions (ionises) according to $HCOOH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCOO^-(aq)$, and these ions enable the solution to conduct electricity (current is defined as flow of charge).

b. Citric acid is weak acid that only partially ionises in water, hence only a slight reaction occurs when consumed.
HCl is a strong acid and will completely react with water, producing a corrosive reaction with skin.

24. a. $Ca(OH)_2(aq) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2H_2O(l)$

b. $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

c. $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

25. a. i. 0.30 **ii.** 0.083 M

b. 11

26. a. i. $HNO_3(aq) + H^+(aq) + e^- \rightarrow NO_2(aq) + H_2O(1)$
$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

ii. Reduction reaction: $HNO_3(aq) + H^+(aq) + e^- \rightarrow NO_2(aq) + H_2O(1)$

iii. Overall reaction: $Cu(s) + 2HNO^3(aq) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(aq) + 2H_2O(1)$

b. i. +5 **ii.** +4

c. $\mathbf{2}HNO_3(aq) + \mathbf{3}V(s) + \mathbf{6}H^+(aq) \rightarrow \mathbf{2}NO(g) + \mathbf{3}V^{2+}(aq) + \mathbf{4}H_2O$

27. a. i. Nickel metal, tin metal and Fe^{2+} ions.

ii. $Fe(s) + Ni^{2+}(aq) \rightarrow Fe^{2+}(aq) + Ni(s)$ or
$Fe(s) + Sn^{2+}(aq) \rightarrow Fe^{2+}(aq) + Sn(s)$

b. Any one of Pb, Cu, Ag or Au

28. a. Iron is higher on the reactivity series than tin. This means that it is more reactive. If the area around the scratch is exposed to typical corrosion conditions (such as water and salt), the more reactive iron will react (rust) in preference to the less reactive tin.

b. Zinc is higher on the activity series than iron. This means that it is more reactive. If the area around the scratch is exposed to typical corrosion conditions (such as water and salt), the more reactive zinc will react in preference to the less reactive iron. The iron thus stays protected.

c. A freshly produced aluminium surface reacts almost instantaneously with atmospheric oxygen to produce a thin but strong and continuous layer of aluminium oxide. This then insulates the surface underneath from further corrosion by environmental agents. It is the initial reactivity of aluminium that prevents it from further corrosion.

13 Measuring solubility and concentration

13.2 Measuring solution concentration

Practice problem 1

0.020 g

Practice problem 2

4.25 L

Practice problem 3

1.5×10^{-4} g L^{-1} = 1.5×10^{-1} ppm

Practice problem 4

49%(m/m)

Practice problem 5

0.4%(m/v)

Practice problem 6

The label is incorrect. The label states a 35%(v/v) alcohol content, whereas the sample found a concentration of 32%(v/v).

Practice problem 7

0.8 mol L^{-1}

Practice problem 8

30.3 g

Practice problem 9

562 mL

Practice problem 10

0.017 M

13.2 Exercise

1. a. 1.70×10^4 mg L^{-1} or 17.0 g L^{-1}

b. 0.480 g L^{-1} or 480 mg L^{-1}

2. 0.0950 g

3. 4.00 mL

4. a. 25.6%(m/m) **b.** 0.50%(m/m) **c.** 0.0720%(m/m)

5. a. 20%(m/v) **b.** 0.0050%(m/v)

6. a. 37.5 mL **b.** 500 mL

7. a. 0.50 mol L^{-1} **b.** 4.29 mol L^{-1} **c.** 1.6 mol L^{-1}

8. a. 2.0 mol L^{-1} **b.** $c\left(Na^+\right) = 2.0\,\text{mol}\,L^{-1}$
$c\,(OH^-) = 2.0\,\text{mol}\,L^{-1}$

9. 0.015 mol L^{-1}

10. a. 11.6 M means that 11.6 mol of hydrogen chloride is dissolved in every litre of water to make the hydrochloric acid solution.

b. Use n = cV to calculate the number of moles the final dilute solution will contain.

$$n(HCl) = cV$$
$$= 1 \times 2$$
$$= 2\,\text{mol}$$

This has to come from the large bottle. Calculate the volume required to obtain this many moles.

$$V(conc.HCl) = \frac{n}{c}$$
$$= 2 \div 11.6$$
$$= 0.172\,\text{L}$$
$$= 172\,\text{mL}.$$

To give a final volume of 2 L (2000 mL), 2000 – 172 = 1828 mL of water will be required.
The 172 mL of concentrated acid will then be slowly added to the 1828 mL of water with constant stirring — never the other way around!

13.2 Exam questions

1. C
2. 6.93%
3. 0.300 M
4. C
5. $n(AgNO_3) = 7.36 \times 10^{-3}$ mol

   $c(AgNO_3) = 0.0490$ M

13.3 Factors that influence solubility

13.3 Exercise

1. Higher temperatures mean solvent particles have more energy, which allows the solute structure to be more easily disrupted, and the particles kept apart once they have dissolved.
2. Oil is insoluble in water because oil consists of non-polar molecules with weak dispersion forces, while water is polar with strong hydrogen bonds between its molecules. This hydrogen bonding holds the water molecules together and so they do not interact with the oil molecules.
3. **a.** Insoluble
   b. Insoluble
   c. Insoluble (although at this higher temperature, it might be described as sparingly soluble)
4. Add hot water to the mixture and stir. The small amount of lead(II) chloride will dissolve because it is soluble in hot water, while the lead(II) sulfate will remain insoluble. Filter before the solution cools, thus collecting only lead(II) sulfate in the filter paper.
5. **a.** Ion–dipole interactions
   b. Dipole–dipole interactions
   c. Hydrogen bonding

13.3 Exam questions

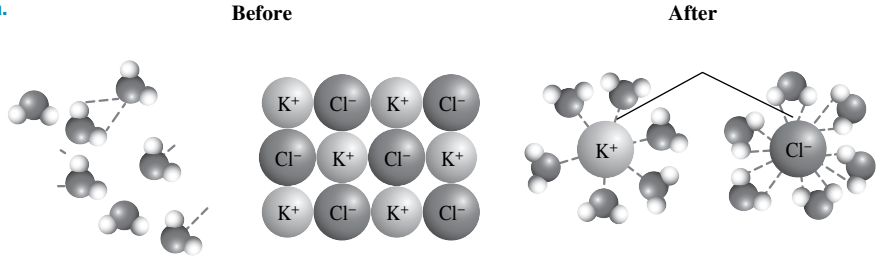
1. **a, b** See figures at the bottom of the page*
2. D
3. A
4. B
5. C

13.4 Solubility graphs

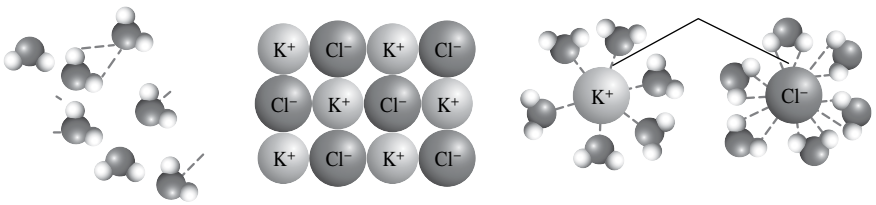
Practice problem 11

a. See figure at the bottom of the page**
b. From the graph, at 35 °C, solubility is around 36.4 g/100 g water.
c. 53.7 g

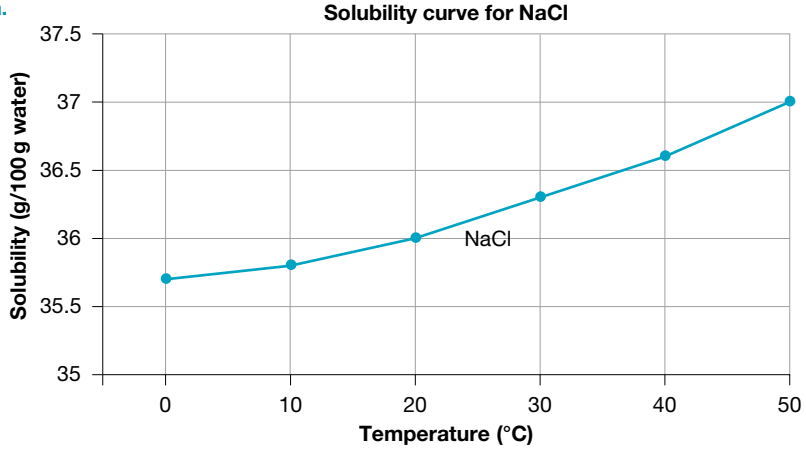
*1. a.



b.



**a.



Practice problem 12

Step 1 Find the mass of KCl in 0700 mols.
$m = 52.2$ g

Step 2 At SLC 25 °C, the solubility of KCl is around 35 g/100 g. Therefore, a sample of 52.2 g will form a supersaturated solution.

Step 3 Obtain an unsaturated solution:

a. Increase temperature of the solution to a minimum of 80 °C, or

b. Continue working at SLC, but dilute the solution. At 25 °C, solubility is 35 g/100 g; we have 52.2 g, thus we need:
$52.2 \times \frac{100}{35} = 149.14$ g solution or 150 g solution.
Therefore a minimum of 50 g (50 mL) of water will need to be added.

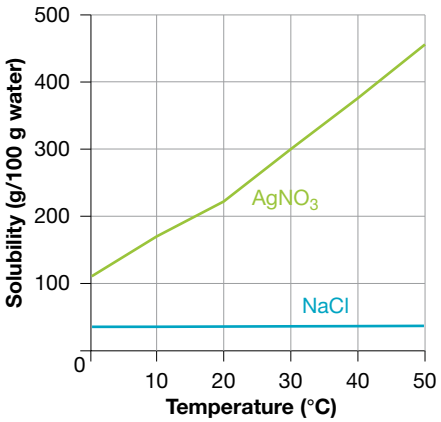
Practice problem 13

a. Both substances are fully dissolved.

b. All the potassium sulfate is still dissolved, and no crystals form. Predict that 14 g of pure potassium chloride crystals are formed.

c. Below 20 °C, potassium sulfate crystals would then start to form as well, and the potassium chloride crystals would no longer be pure.

13.4 Exercise

1. **a.** Unsaturated
   b. 390 g
   c. 119 g
   d. Supersaturated
2. 

3. 30 °C.
4. 240 g water
5. 60 °C
6. 0.23 mols
7. 12 g
8. 24 g
9. Sodium sulfate
10. **a.** Yes. 31 g will dissolve in 100 g at 70 °C. It will certainly dissolve in more water and at a higher temperature.
    b. 52 °C
    c. 20 g
    d. 11 g
    e. 0.25 mol

13.4 Exam questions

1. 21 g/100 g H_2O
2. D
3. A
4. 136 g
5. **a.** 8.69 mL **b.** 6.9 g

13.5 Review

13.5 Review questions

1. **a.** 4×10^{-4} g L^{-1} **b.** 0.4 ppm
2. **a. i.** 48 °C **ii.** 58 °C **iii.** 24 °C
   b. 65 °C
   c. i. 36 g/100 g **ii.** 99 g/100 g **iii.** 55 g/100 g
3. An unsaturated solution contains less than the maximum mass of solute for a given volume of solution at a given temperature. A saturated solution contains the maximum mass of solute for a given volume of solution at a given temperature. A supersaturated solution contains more than the theoretical maximum mass of solute for a given volume of solution at a given temperature.
4. 40 g
5. 2.9%(v/v)
6. **a.** 18 g **b.** 49 g **c.** 2.8×10^3 g
   d. 1.3×10^2 g
7. **a.** 39 g **b.** 105 g **c.** 128 g
   d. 13.9 g **e.** 1.38 g **f.** 12.3 g
8. 16.4 g
9. **a.** 2.14 M
   b. $c(Al^{3+}) = 2.14$ M and $c(Cl^-) = 6.42$ M
10. 0.083 M

13.5 Exam questions

Section A — Multiple choice questions

1. D 2. D 3. A 4. D 5. C
6. A 7. D 8. C 9. B 10. D

Section B — Short answer questions

11. $n(\text{NaCl}) = \frac{m}{M} = \frac{5.00}{58.5}$ mol $= 0.0855$ mol

$V(\text{NaCl}) = \frac{n}{c} = \frac{0.0855}{0.110}$ L $= 0.777$ L

12. **a.** V, W **b.** T, X, Z **c.** U, Y
13. **a. i.** 3×10^1 g L^{-1} **ii.** 3×10^4 mg L^{-1}
    b. i. 0.005 g L^{-1} **ii.** 5 mg L^{-1}
    c. i. 2.50 g L^{-1} **ii.** 2.50×10^3 g L^{-1}

14. a. 17 g/100 g H_2O
    b. 65 °C
    c. 100 g
    d. 19 g
    e. $KClO_3(s) \xrightarrow{H_2O} K^+(aq) + ClO_3^-(aq)$
15. a. 0.1 mg/tonne = 1×10^{-4} ppm
    2.0 mg/tonne = 2×10^{-3} ppm
    b. 1.0×10^9 L
    c. The energy cost required to extract 1 kg of gold from 1.0 billion litres of sea water would be well in excess of the value of the gold.

14 Analysis for acids and bases

14.2 Solution stoichiometry (volume–volume stoichiometry)

Practice problem 1

Moles of water = 1.33

Practice problem 2

200 mL

Practice problem 3

10 mol L^{-1}

Practice problem 4

13.25 g

Practice problem 5

0.0499 mol L^{-1}

14.2 Exercise

1. a. $Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$
   b. 0.181 mol L^{-1}
2. a. i. 30.0 mL ii. 60.0 mL iii. 90.0 mL
   b. Hydrochloric acid is monoprotic and can donate one proton, sulfuric acid is diprotic and can donate two protons, and phosphoric acid is triprotic and can donate three protons. Therefore, it is calculated to take 30 mL of KOH for each acidic proton the acid molecule has.
3. 0.560 mol L^{-1}
4. 45 mL
5. 1.6 mol L^{-1}
6. 0.12 mol L^{-1}
7. 0.075 L or 75 mL
8. 1.88 mol L^{-1}
9. a. Responses should include all steps for making a standard solution. Detailed responses can be found in the worked solutions in your digital formats.
   b. 0.095 01 mol L^{-1}
10. a. Standard solutions have accurately known concentrations.
    b. 0.0551 mol L^{-1}

14.2 Exam questions

1. C
2. a. $Na_2CO_3(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$
   b. 0.2255 M
   c. 0.2565 M
3. 11.1 mL
4. a. 0.09 mol L^{-1} b. 0.16 mol L^{-1} c. 1.25 mol L^{-1}
5. 0.0750 mol L^{-1}

14.3 Acid–base titrations

Practice problem 6

0.340 mol L^{-1}

Practice problem 7

6.67 mol L^{-1}

14.3 Exercise

1. B
2. Orange
3. a. $2HNO_3(aq) + Na_2CO_3(aq) \rightarrow 2NaNO_3(aq) + CO_2(g) + H_2O(l)$
   b. 0.099 50 mol L^{-1}
4. The equivalence point is where the reactants have reacted and reached their correct mole ratio. The end point in a titration is when the indicator changes colour.
5. Methyl orange, thymol blue, bromphenol blue, congo red or bromcresol green
6. a. Litmus or bromothymol blue
   b. Litmus or bromothymol blue
   c. Phenolphthalein
   d. Methyl orange
7. 0.015 L
8. a. 0.0018 mol L^{-1} b. 64 ppm
9. 75 mg mol
   Because 1 L was the volume analysed, the total alkalinity was 75 mg L^{-1}.
10. 0.262 M

14.3 Exam questions

1. C
2. B
3. The funnel may allow drops of solution to fall during the titration, altering the volume measurement in the burette.
4. 32.62
5. a. i. 0.82
   ii. A solution of accurately known concentration
   iii. $n(Na_2CO_3) = 0.0250$ mol
   $m(Na_2CO_3) = 2.65$ g
   iv. $Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$
   v. Sodium carbonate

vi. Average titre = 27.97
$n(Na_2CO_3) = 2.00 \times 10^{-3}$ mol
$n(HCl) = 4.00 \times 10^{-3}$ mol
$c(HCl) = 0.143$ M

b. i. $[OH^-] = 10^{-5}$ M

ii. 495 L

14.4 Review

14.4 Review questions

1. a. An aliquot is the volume delivered by a pipette.
   b. Concordant titres are volumes delivered by the burette that are very close in value (within 0.1 mL).
   c. Equivalence point is the point where two reactants are present in their stoichiometric ratio.
2. a. 21 mL b. 33 mL c. 3.1 mL
3. 149 mL
4. 16.2 mL
5. a. 200 mL
   b. 325 mL
   c. 1.60×10^3 mL (1.60 L)
6. 100 mL
7. a. 0.184 mol L^{-1} b. 0.105 mol L^{-1}
8. 4.0 mol L^{-1}
9. 526 mL
10. a. 0.004 10 mol L^{-1} b. 0.164 g L^{-1}

14.4 Exam questions

Section A — Multiple choice questions

1. D 2. B 3. B 4. B 5. D
6. B 7. B 8. A 9. A 10. B

Section B — Short answer questions

11. a. $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
    b. 0.001 85 mol
    c. 0.000 925 mol
    d. 0.0370 mol L^{-1}
    e. 0.370 mol L^{-1}
    f. 0.0102 mol L^{-1}
    g. 35.3 L
12. a. 2.65 g b. 0.106 mol L^{-1}
13. a. 0.1046 M
    b. i. $2HCl(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$
    ii. 0.2137 M
    c. i. Distilled water
    ii. Hydrochloric acid
    iii. Distilled water
    d. If the pipette was rinsed with water, a smaller number of moles of base would be delivered to the conical flask, resulting in a smaller titre. This would result in a lower concentration than it should be for the acid.
14. a. i. Sodium hydroxide solution
    ii. Distilled water
    b. i. It would lower the concentration of the solution in the pipette, leading to an underestimate of its concentration.
    ii. One
    c. Distilled water
15. a. $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
    b. The function of the methyl orange indicator is to detect the end point of the titration.
    c. i. 21.48 mL ii. 19.16 (+/– 0.01 mL) mL
    d. 0.001 832 mol
    e. 0.1466 M

15 Measuring gases

15.2 Gases and the enhanced greenhouse effect

15.2 Exercise

1. The natural greenhouse effect is a result of infrared radiation from the Sun striking the Earth and warming its surface, which then radiates heat energy back into space. The enhanced greenhouse effect is caused by greenhouse gases absorbing some of this heat radiation, so the air warms up.
2. Any two from:
   - deforestation
   - cows producing large amounts of methane
   - transport of food requiring use of petroleum-based fuels.

   Possible reductions:
   - adopting a more plant-based diet to reduce methane emissions from cows
   - buying local, in-season produce to reduce emissions from transportation.

   Other ideas also possible.
3. Electricity can be generated by wind, hydro or solar power, reducing the greenhouse gas emissions related to its production and consumption in comparison to natural gas (a fossil fuel), which would release carbon dioxide into the atmosphere.
4. Global warming is a gradual increase in the overall temperature of the Earth's atmosphere, probably due to an increase in greenhouse gas emissions that cause the enhanced greenhouse effect.
   Climate change refers to changes in various measures of climate such as rainfall, humidity and wind speed over a long period of time.
5. Cattle release methane, which is the most potent greenhouse gas. In addition, the plants grown on the land will take in carbon dioxide during photosynthesis, likely to a greater extent than the grass the cattle graze on because the crops will be allowed to grow to maturity.

15.2 Exam questions

1. CO_2, CH_4, H_2O
2. A
3. D
4. A
5. Demand for energy, land, resources and food production all have contributed to greenhouse gas production and imbalance in the environment.
   As countries become more industrial and manufacturing increases, more greenhouse gases will be produced.

15.3 Gases at standard laboratory conditions (SLC)

Practice problem 1

a. −146 °C

b. 0.065 L or 6.5×10^{-2} L

Practice problem 2

a. $M = 44.0$ g mol^{-1}

b. Carbon dioxide

15.3 Exercise

1. a. 27 °C b. 154 °C c. −100 °C
   d. 392 K e. 73 K
2. Particles collide with the walls of their container, creating pressure.
3. a. 3.0×10^3 mol b. 0.103 mol c. 0.0101 mol
4. a. 37.9 L b. 21.1 L c. 1.0×10^8 L
5. a. 0.194 g b. 7.9 g

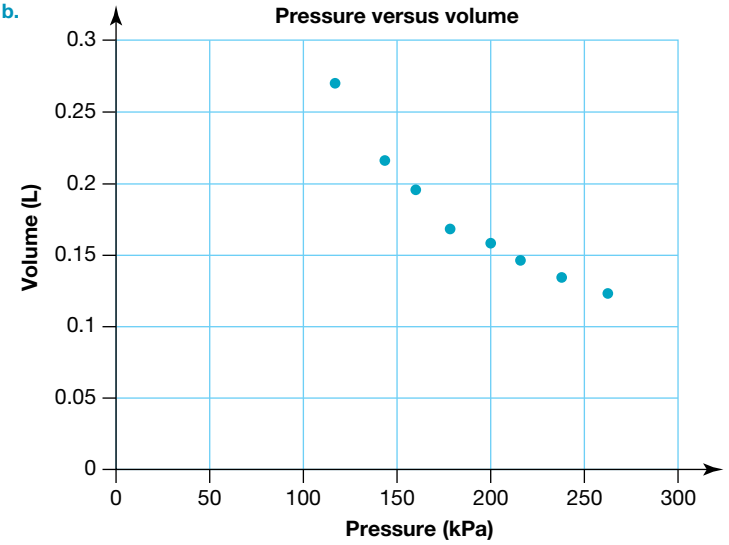
15.3 Exam questions

1. B
2. C
3. 0.960 kg
4. Krypton (Kr), molar mass of 83.8 g mol^{-1}
5. a.

Pressure (kPa)	$\frac{1}{P}$
120	8.33×10^{-3}
145	6.90×10^{-3}
162	6.17×10^{-3}
180	5.56×10^{-3}
200	5.00×10^{-3}
216	4.63×10^{-3}
240	4.17×10^{-3}
258	3.88×10^{-3}

b. See figure at the bottom of the page*

*5. b.



c. See figure at the bottom of the page*

d. The first graph is non linear, with volume decreasing as pressure increased. The second graph shows a linear proportionality, where volume increases linearly as $\frac{1}{P}$ increases.

15.4 Calculations using the ideal gas equation and stoichiometry

Practice problem 3

35 L

Practice problem 4

735 000 Pa

Practice problem 5

5.7×10^3 mL

Practice problem 6

a. $V(O_2) = 50$ mL

b. $V(CO_2) = 25$ mL

c. $V(H_2O) = 50$ mL

Practice problem 7

$m(C_6H_{14}) = 57.8$ g

15.4 Exercise

1. a. 21.8 L b. 88.0 °C

2. a. 94 L b. 10 L c. 6.1×10^{-3} L
   d. 0.010 L e. 0.24 L

3. a. 40 mL b. 60 mL c. 40 mL
   d. 20 mL

4. 0.059 g

5. 101 g

15.4 Exam questions

1. A 2. A 3. 4.2 L
4. –165 °C 5. 36.6 °C

15.5 Review

15.5 Review questions

1. The effects of raising prices of non-renewable fuels are uncertain.

   Some arguments for the proposal are:
   - Raising the prices will encourage people to look for alternative fuel options that are more environmentally friendly.
   - Raising prices will make these fuel sources less accessible and possibly reduce their use.

   Some arguments against the proposal are:
   - Raising the prices would unfairly limit poorer people's access to fuels in general, creating a class of people who could no longer afford to travel by car.
   - Raising the prices would increase the income made by the big oil and gas companies without necessarily reducing the use of the fuels.

2. a. Answers will depend on electricity usage. As an example, for a household that uses 5000 kWh per year: 1.44 × 5000 = 7200 kg of carbon dioxide.

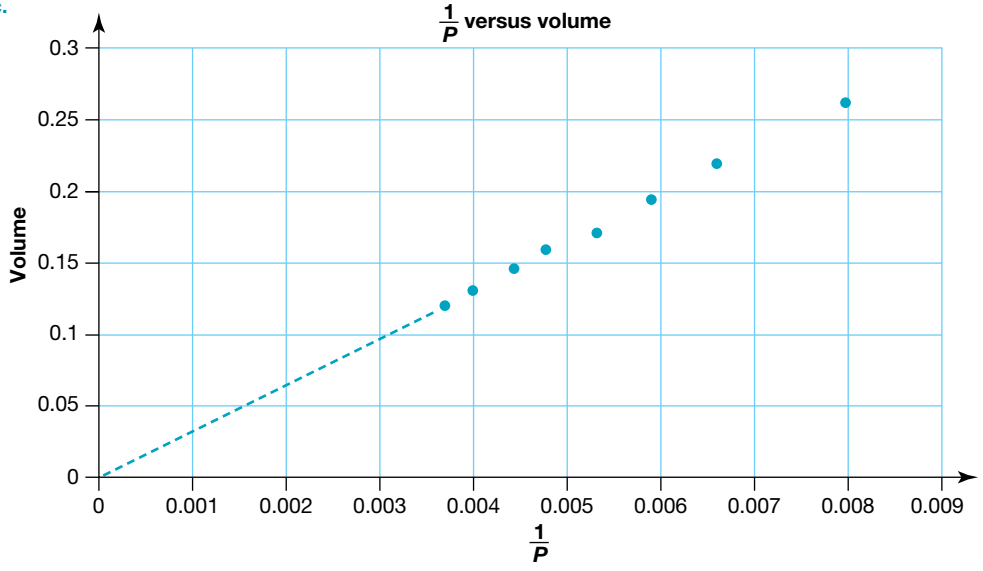
   b. Carbon dioxide is a greenhouse gas that contributes to the enhanced greenhouse effect. The enhanced greenhouse effect causes more heat to be trapped in the Earth's atmosphere, leading to environmental issues such as global warming.

   c. Sample answer:
   - installing solar panels or switching to any other form of sustainable energy
   - limiting electricity use
   - using public transport instead of driving.

3. a. 119 °C b. 300 K c. 0.056 L

4. a. –267 °C b. –93 °C

*5. c.



5. The mass is 28 mol^{-1}; the gas is most likely to be nitrogen.
6. a. 1.4 mol b. 13 mol

15.5 Exam questions

Section A — Multiple choice questions

1. C 2. B 3. D 4. D 5. A
6. A 7. D 8. C 9. A 10. B

Section B — Short answer questions

11. a. Gas particles move with continual random motion and, when doing so, collide with the walls of any container, producing the effect we know as pressure.
    b. As temperature increases, gas particles move more rapidly. This results in an increase in the frequency of collisions of the gas particles with the walls of any container.
12. a. 0.043 mol b. 2.6×10^{22} c. 60 °C
13. 1.9×10^{20}
14. a. i. 9.9 g ii. 11 L iii. 14 L
    b. 0.092 mL
15. 4.41×10^{5} L

16 Analysis for salts

16.2 The sources of salts in soil and water

16.2 Exercise

1. A chemical substance obtained from the reaction between an acid and a base
2. Sodium chloride
3. Sample answer: Sodium, potassium, calcium, magnesium, chloride, sulfate, nitrate, iron or aluminium
4. Despite $CaCO_3$ being almost insoluble in water, rainwater saturated with CO_2 is slightly acidic and forms calcium hydrogen carbonate, $Ca(HCO_3)_2$, which readily dissociates into its ions.
5. Industry discharge, corroded pipes, pollution
6. Both are toxic to humans and animals.
7. An organometallic substance is formed when a heavy metal combines with an organic substance. An example is methyl mercury.
8. All salt ions conduct electricity, meaning that particular salts can't be differentiated with electricity.
9. EC units
10. Na_2CO_3
11. Substances that undergo biomagnification are persistent in the environment, fat-soluble, and difficult to eliminate because they are stored in the fatty tissue of the animals. This results in concentration as the substance moves up the food chain.

16.2 Exam questions

1. D
2. Above the limit
3. A
4. Tap water contains dissolved ions that conduct electricity.
5. $NH_4NO_3(s) \xrightarrow{H_2O} NH_4^+(aq) + NO_3^-(aq)$

16.3 Quantitative analysis of salts — stoichiometry and molar ratios

Practice problem 1

a. $CuSO_4 \cdot 5H_2O$
b. Copper(II) sulfate pentahydrate

Practice problem 2

3.976 g

Practice problem 3

9.44 g

Practice problem 4

0.600 M

Practice problem 5

5.0%

Practice problem 6

9.95% or 9.99%

Practice problem 7

1.252 g L^{-1}

16.3 Exercise

1. a. Salt crystals that can incorporate water molecules within their structure, produced as an aqueous solution crystallises
   b. A salt that has no water molecules
   c. Please see the worked solutions in your online resources.
2. B
3. $SrCl_2 \cdot 2H_2O$
4. $CaSO_4 \cdot 2H_2O$
5. $MgSO_4 \cdot 7H_2O$
6. a. The barium carbonate is a solid.
   b. 4.837 g
7. a. 6.04 g
   b. All of the Na_2S will react.
8. a. All the HCl had reacted.
   b. 0.91 M
   c. Weighing the strip of magnesium and measuring the volume of hydrochloric acid
   d. Precautions include wearing a safety coat, safety glasses and gloves, and working in a well-ventilated environment.
9. 96.7%
10. B

16.3 Exam questions

1. The degree of hydration is 2.
2. $NaHCO_3$
3. 7
4. A
5. 3.93% (w/w)

16.4 Quantitative analysis of salts — colorimetry and UV-visible spectroscopy

Practice problem 8

0.05 mol L^{-1}

Practice problem 9

0.04 mol L^{-1}

16.4 Exercise

1. A calibration curve is a graph of concentration against absorbance. It is created using a series of standards of known concentrations and a line of best fit, and is used to be able to determine the concentration of an unknown sample.
2. A UV–visible spectroscopy machine is required because both 260 nm and 280 nm are outside the visible spectrum range and within the ultraviolet region of the electromagnetic spectrum.
3. a. See figure at the bottom of the page*
   b. 2.0 mM
   c. Yes, the patient has a calcium blood level below the normal range, classifying the patient as being hypocalcemic.
   d. To ensure precision, ensure you have replicates (at least triplicates) of the standards and samples you are testing. To ensure your results are accurate, make sure you are measuring your standards and test samples at the same time in the same environmental conditions.
4. See table at the bottom of the page**
5. B

16.4 Exam questions

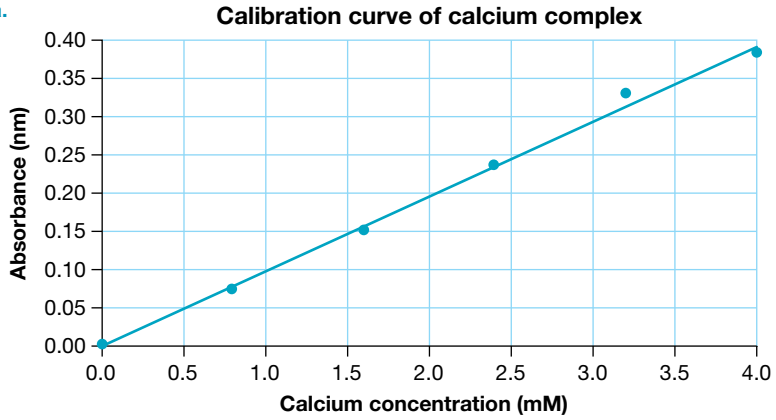
1. D
2. D
3. 0.83 mg of phospate per litre
4. 11.25 mg/L
5. A

16.5 Review

16.5 Review questions

1. a. $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow 2NaNO_3(aq) + BaSO_4(s)$
   Precipitate is barium sulfate.
   b. 5.973 g
2. a. 0.15 mol b. 13 g
3. a. 9.47 g b. 13.02 g
4. 0.029 g
5. a. 159 g b. 5.2 g
6. 0.794 g
7. a. 0.920 kg
   b. $2NaOH(s) + CO_2(g) \rightarrow Na_2CO_3(aq) + H_2O(1)$
   c. 550 g
   d. When planning space travel, experts focus on minimising the mass of material that needs to be lifted into space.

*3. a.



**4.

Instrumental colorimeter	UV–visible spectrometer
Only measures absorbance within the visible spectrum	Can measure absorbance within the ultraviolet and visible spectrum
Uses a coloured filter complementary to the colour of the sample being tested for optimal absorbance readings	Uses a specific wavelength to measure the standards and samples
Relatively cheap	Relatively more expensive
Can produce results within 1–2% accuracy	Accuracy is relatively better than colorimetry
Can result in interference from other coloured species in the solution	More selective, decreasing interference from similarly coloured species
A calibration curve of concentration versus absorbance of the standards is needed to be able to determine the concentration of the sample	A calibration curve of concentration versus absorbance of the standards is needed to be able to determine the concentration of the sample

Lithium hydroxide would be preferred because a lower mass of lithium hydroxide would be needed to remove the same mass of carbon dioxide.

e. $2OH^-(s) + CO_2(g) \rightarrow CO_3^{2-}(aq) + H_2O(l)$ for both reactions mentioned.

8. a. 46.8 tonnes

b. The term 'burning' usually means the combustion of a substance with the oxygen in air.

9. a. Ions dissolved in water are required for it to conduct electricity. The more ions from dissolved salts, the greater the conductivity. Because salts form ions when dissolved in water, electrical conductivity can be used to measure salt levels in river water.

b. Before the early 1980s, the trend was increasing. Since the early 1980s, the trend has been decreasing.

c. These results represent total ion content; ions from salts other than NaCl are probably present also.

d. Conductivity is related to the total amount of all ions present. It does not depend on which ions are present and individual types of ions cannot be identified through electrical conductivity measurements.

10. a. 13 ppm b. 1.3% c. Orange

11. a. Lead b. Silver
c. Barium d. Hydroxide

12. A systematic error. For sample responses, see the worked solutions in your digital formats.

13. C, E, H, A, F, D

16.5 Exam questions

Section A — Multiple choice questions

1. D	2. B	3. D	4. D	5. C
6. D	7. D	8. A	9. D	10. D

Section B — Short answer questions

11. a. Carbon dioxide

b. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

c. 82.8%

d. Response should state one assumption. Detailed responses can be found in the worked solutions in your digital formats.

12. a. 1.14 g

b. 7.97%

c. Response should state one assumption. Detailed responses can be found in the worked solutions in your digital formats.

13. a. $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(S) + NaNO_3(aq)$

b. 1.82 g

c. 0.182 g L^{-1}

d. It is far less likely that any of the AgCl precipitate will dissolve; i.e. it is far more likely that an accurate result will be obtained.

e. Response should state one assumption. Detailed responses can be found in the worked solutions in your digital formats.

14. $x = 7$

15. Degree of hydration, x, = 2

Unit 2 | Area of Study 2 review

Section A — Multiple choice questions

1. A	2. B	3. B	4. C	5. C
6. C	7. A	8. A	9. D	10. B
11. C	12. D	13. C	14. B	15. A
16. B	17. B	18. A	19. C	20. D

Section B — Short answer questions

21. a. 10.8 g L^{-1}

b. i. 3.8×10^{-4} g

ii. 7.9×10^{-5} mol L^{-1}

iii. 1.5 ppm

22. a. 800 mg L^{-1}

b. i. 5500 mg L^{-1}

ii. 1.9×10^4 kg

c. 450 mL

23. a. 20 g/100 g H_2O

b. 55 °C

c. 90 g

24. a. Distilled water

b. 18.14 mL

c. $2HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$

d. 0.0227 mol L^{-1}

e. Blue → green or blue → green → yellow (if the end point was passed)

f. Concentration of barium hydroxide would be overestimated.

25. a. 8.77×10^{-3} mol

b. 24.6 L

c. The molar volume of a gas is independent of the nature of the gas. The two results would, therefore, be expected to be similar.

26. a. $3Zn^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Zn_3(PO_4)_2(s)$

b. i. 0.004 775 mol

ii. 0.057 30 mol L^{-1}

c. The precipitate should be dried overnight in an oven (110 °C), cooled and weighed, and then this procedure should be repeated until a constant mass is obtained.

d. i. It is assumed that only Zn^{2+} ions formed a precipitate with the phosphate ions. *or* It is assumed that all the Zn^{2+} ions formed a precipitate with the phosphate ions.

ii. Since sodium phosphate was in excess, no further increase in the mass of precipitate would occur, so the actual Zn^{2+} concentration would be unaffected.

GLOSSARY

%(m/m) percentage mass per mass
%(m/v) percentage mass per volume
%(v/v) percentage volume per volume
absolute zero the lowest temperature that is theoretically possible; 0 K
abundance the relative amount of a substance
accuracy how close an experimental measurement is to a known value
accurate when an experimental measurement is close to a known value
acid substance that can donate a proton to a base
acid rain rain made acidic by mixing with atmospheric pollution resulting a pH less than 5
acidic oxide oxide of a non-metal, such as CO_2 or SO_2
acidic protons the hydrogen atoms that will readily dissociate from the acidic molecule
acyclic hydrocarbons based on straight and branched chains of carbon atoms
addition polymerisation the process in which monomers with at least one double carbon–carbon bond react together to form a polymer by addition reactions
addition reaction a reaction in which one molecule bonds covalently with another molecule without losing any other atoms; only one product is formed
adsorption the adhesion of atoms, ions or molecules from a gas, liquid or dissolved solid to a surface
aim a statement outlining the purpose of an investigation, linking the dependent and independent variables
aliquot known, fixed volume delivered by a pipette
alkali soluble base
alkane a compound of carbon and hydrogen containing only single bonds, no rings, and having the general formula C_nH_{2n+2}
alkene a hydrocarbon that contains one carbon–carbon double bond and the general formula C_nH_{2n}
alkyne a hydrocarbon that contains one carbon–carbon triple bond and the general formula C_nH_{2n-2}
allotrope different forms of an element
alloying forming a substance with metallic properties that consists of two or more elements, at least one of which is a metal
amorphous polymer polymers that do not have a crystalline structure; their atoms are held together loosely and their arrangement is not ordered or predictable
amphiprotic substance that can both accept or donate protons
analyte the solution of unknown concentration to go into the conical flask and be determined by titration
anecdote an individual's story based on personal experience rather than strong evidence
anhydrous contains no water molecules in its chemical structure
anion a negatively charged atom or group of atoms
anode site of oxidation in an electrochemical cell
anthocyanin pH-sensitive pigment molecule found in plants that will change colour in response to different acidic and alkaline environments
aqueous solution with water as the solvent
aqueous solutions mixtures in which substances are dissolved in water
assumptions ideas that are accepted as true without evidence in order to overcome limitations in experiments
atom a neutral particle with a nucleus; the smallest sample of an element
atomic emission spectrum a spectrum emitted as distinct bands of light of diagnostic frequencies by elements or compounds
atomic number the number of protons in the nucleus of an atom of a particular element
average titre titre determined by averaging the results of several titrations
Avogadro constant the number of elementary particles (atoms) in exactly 12 g of carbon-12; this number is equal to 6.02×10^{23}

ball-and-stick models representation of a molecule in which the atoms are shown as balls and the bonds as sticks

bar graph graph in which data is represented by a series of bars; usually used when one variable is quantitative and the other is qualitative

base substance that can accept a proton from an acid

bias the intentional or unintentional influence on a research investigation

binary ionic compound describes ionic compounds made up of only two elements

biodegradable can be broken down or degraded by natural means, including by microbes

bioethene chemical identical to ethene but produced from biomass

bioleaching (biomining) the process of extracting metals from ores or waste by using microorganisms

biomagnification process by which a compound (such as a pollutant or organometal) increases its concentration in the tissues of organisms as it travels up the food chain

biomass renewable organic matter that can be obtained from living or recently living plants and animals

bioplastic a polymer obtained from renewable biomass; for example, agricultural, cellulose and corn starch waste

biopropene chemically identical to propene but produced from biomass

bioreactor an industrial device that allows a biological process to occur

biotechnology the use of biological processes, systems and resources in industrial production

boiling temperature temperature at which a liquid boils and turns to vapour

bond dipole separation of charge in a polar covalent bond

bonding electrons the electrons shared as a pair to form a covalent bond

Boyle's law describes the relationship between pressure and volume of a gas, such that for a fixed amount of gas at constant temperature, pressure is inversely proportional to volume; $P_1V_1 = P_2V_2$

branched molecule has alkyl groups attached along the carbon chain

calibrated to adjust an instrument using standards of known measurements to ensure an instrument's accuracy

carboxylic acids homologous series containing the —COOH functional group

case study an investigation of a particular event or problem, involving the discussion of knowledge learned from a situation

cathode site of reduction in an electrochemical cell

cathodic protection method of protecting a metal from corrosion by connecting it to a power source and another metal that is less active; see *electrochemical protection*

cation a positively charged atom or group of atoms

causation when one factor or variable directly influences the results of another factor or variable

cellulose the most common carbohydrate and a condensation polymer of D-glucose; humans cannot hydrolyse cellulose, so it is not a source of energy

central atom the atom in a molecule with the most bonding electrons

Charles' law describes the relationship between temperature and volume of a gas, such that for a given amount of gas at constant pressure, volume is directly proportional to the absolute temperature; $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

chemical recycling recycling using heat or chemicals in which the chemical structure of the polymer is converted into monomers that can then be reused in chemical processes

chromatography a laboratory technique for the separation of a mixture by passing it in solution (mobile phase) over stationary phase; separation occurs as components move at different rates

circular economy economy in which waste and pollution is eliminated, products at the end of life are reused, repaired or recycled, and the environment is not harmed

climate change long-term shifts in temperatures and weather patterns

combustion reaction between a fuel and oxygen (usually) releasing heat energy

compostable a product that can be degraded by microorganisms in a moist, warm environment to produce matter that can support plant life in a relatively small amount of time

compound substance consisting of two or more elements

computer-generated model representation of a molecule that is produced by a computer

concentrated a large quantity of solute in a known volume of solution

concentration the amount of solute that is dissolved in a known volume of solution
conclusion section at the end of a scientific report that relates back to the question, sums up key findings and states whether the hypothesis was supported or rejected
concordant titres results of repeated titrations showing precision to be within ± 0.10 mL of each other
condensation polymerisation polymerisation in which two monomers combine and a smaller molecule is eliminated
conjugate acid the product that a base forms when it has accepted a proton from an acid
conjugate base the product that an acid forms when it has donated a proton to a base
conjugate redox pair two species that differ only by a certain number of electrons
continuous data quantitative data that can take any continuous value
control group a group that is not affected by the independent variables and is used as a baseline for comparison
controlled variable variable that is kept constant across different experimental groups
copolymer polymer formed from the polymerisation of two monomers
correlation measure of a relationship between two or more variables
corrosion oxidation, generally of metals, by materials in their environment
covalent molecules which share electrons between nuclei to form the bond that holds them together in a molecule
covalent bonding sharing of electrons between nuclei that bonds them together in a molecule
covalent layer lattice a substance that displays covalent bonding in three dimensions without the formation of discrete molecules
covalent molecular compounds a molecular compound in which atoms of different elements share electrons with each other
covalent molecular element element made up of identical atoms held together by covalent bonds
critical elements elements considered vital for the world's economy but in short supply
cross-linking the bonding between two polymers
crude oil naturally occurring mixture of compounds that are mainly hydrocarbons; also known as petroleum
crystalline lattice structure of particles that are held together in an ordered, three-dimensional arrangement
cyclic compound where the atoms are connected in a ring
degradable can be broken into smaller and smaller fragments or simpler chemical structure
delocalised describes electrons that are not bound to any one atom but are free to move throughout a lattice
density physical property found by dividing the mass of an object by its volume
dependent variable variable that is influenced by the independent variable; the variable that is measured
deprotonation the donation or dissociation of a proton, H^+, from the acidic molecule
desorption the opposite of adsorption, where molecules are released from the surface of other particles
diatomic molecules substance containing two atoms only
difunctional a molecule that contains two functional groups
dilute not concentrated; has a small quantity of solute in a known volume of solution
dipole unequal sharing of electrons between atoms in a molecule
dipole–dipole bond weak bonding caused by the positive end of one dipole attracting the negative end of another dipole
diprotic (of an acid) can donate two protons
discrete separate, distinct; not in an infinite array or lattice
discrete data quantitative data that can only take on set values
discussion detailed area of a scientific report in which results are discussed, analysed and evaluated, relationships to concepts are made, errors, limitations and uncertainties are assessed and suggestions for future improvements are outlined
dispersion force the bond between adjacent molecules formed by instantaneous dipoles; this weak non-directional bonding is also known as van der Waals force
displacement reaction chemical reaction in which a more reactive element displaces a less reactive element from its compound
dissociate the process by which ions separate when an ionic compound dissolves
double arrow used to show a chemical equation is reversible

double bond strong bond between two atoms formed by two pairs of electrons that are shared by the two nuclei
dry corrosion corrosion of a metal in the air when water is not a reactant
ductile able to be drawn out into a thin wire
economic factors factors that are related to the financial implications of an action or decision; these may be on an individual level or on a larger community or national level
elastomer polymer that can be stretched and can return to its original shape
electrochemical protection protecting a metal from corrosion by connecting it to a power source and another metal that is less active; the electrons feed onto the metal being protected, making it a cathode so that it cannot be oxidised
electrode a solid conductor of electricity in an electrochemical cell
electrode polarity the positive or negative sign on the electrodes in a galvanic cell
electrolysis the decomposition of a chemical substance (in solution or the molten state) by the application of electrical energy
electrolyte solution or liquid that can conduct electricity
electron configuration the arrangement of electrons in the shells of an atom
electron dot diagrams representation where the atom's nucleus and all innershell electrons are replaced by its element symbol and the outershell electrons are represented by dots around the symbol in a square arrangement
electronegativity the electron-attracting power of an atom
electrostatic attraction strong force of attraction between positive and negative charges
electrostatic force attractive or repulsive force between particles that is due to their electric charges
element a pure chemical species consisting of atoms of a single type
empirical formula chemical formula that shows the simplest numerical ratio in which atoms are combined
end point the point at which the indicator just changes colour or the pH curve becomes vertical
endothermic chemical process or reaction that absorbs energy
enhanced greenhouse effect the effect of increasing concentrations of greenhouse gases in the atmosphere as the result of human activity
equivalence point point at which two reactants have reacted in their stoichiometric molar ratios in a titration as determined by a balanced equation
ester an organic compound formed from a condensation reaction between an alcohol and a carboxylic acid
ethics acceptable and moral conduct determining what is 'right' and 'wrong'
eutrophication form of water pollution involving excess nutrients, such as nitrogen and phosphorus, leaching from soils, typically resulting in excessive growth of algae
evidence reliable and valid data used to support or refute a hypothesis, model or theory
excess reactant reactant left over when the limiting reagent has been used up
excited state raised to a higher than ground-state energy level
exothermic chemical process or reaction that releases energy through light or heat
experimental bias a type of influence on results in which an investigator either intentionally or unintentionally manipulates results to get a desired outcome
experimental group test group that is exposed to the independent variable
external circuit the wire joining the two half-cells
extrapolation estimation of a value outside the range of data points tested
falsifiable able to be proven false using evidence
feedstock raw material used in industrial processes to produce a product
filtrate liquid that has passed through a filter
first ionisation energy the energy required to remove one valence electron from a gaseous atom
full equation equation that includes all of the reactants and products that are present
functional group an atom or group of atoms attached to or part of a hydrocarbon chain that influence the physical and chemical properties of the molecule
galvanic cell electrochemical cell that converts chemical energy into electrical energy
global warming a gradual increase in the overall temperature of the Earth's atmosphere

global warming potential (GWP) the energy one tonne of gas will absorb compared to one tonne of carbon dioxide
gravimetric analysis analysis in which the amount of a substance is established through the measurement of a mass
green chemistry the design of chemical products and processes that decrease or eliminate the use or creation of hazardous wastes and the impact on the environment
greenhouse effect a natural process that warms the Earth's surface; when the Sun's energy reaches the Earth's atmosphere, some of it is reflected back to space and the rest is absorbed and re-radiated by greenhouse gases
greenhouse gases gases that contribute to the greenhouse effect by absorbing infrared radiation
ground state the least excited state of an atom, where the electrons are occupying the lowest possible energy levels
group vertical column of the periodic table
half-cell one of two containers containing the electrode and electrolyte in an galvanic cell
half-equation chemical equation showing either oxidation or reduction in a redox reaction
haloalkane compound with the same structure as an alkane, except that one or more of the hydrogen atoms has been replaced with halogen atoms
heavy metal metal with high density or of high relative atomic weight, and that is a risk to health or environment
hemicellulose branched polysaccharide found in plant cell walls that have a simpler structure than cellulose
high-density polyethene (HDPE) thermoplastic polymer with very few side branches produced from the monomer ethene
histogram graph in which data is sorted in intervals and frequency is examined; used when both pieces are data are quantitative
homologous series series of organic compounds with the same general forumla, in which the formula of one molecule differs from the next by a —CH_2 group
host rock the rock surrounding or containing the ore minerals
human error mistakes made by individuals conducting an investigation, such as mistakes in entering data or the use of the wrong type of chemical in a reaction
hydrate any compound that absorbs water in the form of H_2O molecules into their chemical structure
hydride binary compound containing hydrogen
hydrocarbon compound containing only carbon and hydrogen
hydrogen bond the bond between a hydrogen atom covalently bonded to an atom of F, O or N and another molecule that also contains an atom of H, F, O or N
hydrolysis a reaction where a molecule is broken into smaller molecules, with water as a reactant
hydronium ion H_3O^+, combination of a proton with a water molecule
hygiene conditions or practices beneficial to maintaining health and preventing disease, especially through cleanliness
hypothesis a tentative, testable and falsifiable statement for an observed phenomenon that acts as a prediction for the investigation
ideal gas equation $PV = nRT$, where pressure is in kilopascals, volume is in litres and temperature is in Kelvin
immiscible when two substances are unable to form a homologous solution
independent variable variable that is changed or manipulated by an investigator
indicator substances that undergo an observable change in an acidic or basic solution
inorganic salt hydrates inorganic salts where water molecules have been incorporated into their crystal structure
insoluble will not dissolve
instrumental colorimeter instrument used to determine the concentration of coloured compounds in solution through comparison of a test sample against samples of known concentration through production of a calibrations curve (absorbance versus concentration)
intermolecular bond bond between molecules
intermolecular forces forces that exist between molecules
internal circuit the movement of ions in the electrolyte solutions

interpolation estimation of a value within the range of data points tested
intramolecular bond internal bonds within a molecule
investigation question the focus of a scientific investigation in which experiments act to provide an answer
ion an atom that has lost or gained electrons and so has a charge
ion–dipole bond bond between an ion and a polar molecule
ionic bond bond between cations and anions in an ionic compound
ionic bonding bonding that forms between cations and anions
ionic compound compound containing cations and anions in an ionic lattice structure or an ionic liquid, depending on ion size
ionic concentrations concentrations of the anions and cations resulting from the dissolution of an ionic lattice; designated by square brackets
ionic equation equation that only includes the species that have reacted, without the spectator ions
ionisation a process in which ions are formed
ionise reaction in which a substance reacts with water to form ions
isomer have the same molecular formula but a different arrangement of the atoms in space
isoprene volatile liquid hydrocarbon obtained from petroleum, the molecule of which forms the basic structural unit of natural and synthetic rubbers
isotopes forms of an element with the same number of protons but different numbers of neutrons in the nucleus, resulting in elements with different masses
isotopes forms of an element with the same number of protons but different numbers of neutrons in the nucleus
isotopic symbol representation of an element as ${}^{A}_{Z}E$, where E is the symbol for the element, A is the mass number and Z is the atomic number
IUPAC International Union of Pure and Applied Chemistry
Kelvin the SI base unit of thermodynamic temperature, equal in magnitude to the degree Celsius
kinetic molecular theory of gases gas particles are in continuous, random motion; collisions between gas particles are completely elastic
latent heat energy needed to change state
latent heat of fusion, (L_f) energy needed to change a fixed amount from the solid to the liquid state
latent heat of vaporisation, (L_v) energy needed to change a fixed mass from the liquid to the gaseous state
legal factors factors that are related to the legality of an action
Lewis structure diagram that shows how outer shell electrons are arranged in a molecule using dashes for bonding pairs and dots for lone pairs
lignin complex organic polymer that provides rigidity to plants
lignocellulose a complex formed from lignin and cellulose
like dissolves like rule polar solutes are more likely to dissolve in polar solvents, while non-polar solutes are more likely to dissolve in non-polar solvents
limitations factors that affect the interpretation and/or collection of findings in a practical investigation
limiting reactant the reactant that is completely used up in a reaction
line graph graph in which points of data are joined by a connecting line; used when both pieces of data are quantitative (numerical)
line of best fit trend line added to a scatterplot to best express the data shown; these are straight lines, and are not required to pass through all points
linear economy raw materials are converted into products that are used and then discarded as waste
linear polymer polymer that does not have side chains
literature review using and analysing secondary evidence and data from a group of scientists' research experiments
logarithm quantity representing the power to which a fixed number (the base) must be raised to produce a given number
logbook a record containing all the details of progress through the steps of a scientific investigation
lone pairs pairs of outershell electrons which are not involved in bonding (non-bonding electrons)
low-density polyethene (LDPE) thermoplastic polymer with a large number of side branches produced from the monomer ethene

lustrous having a shiny surface
malleable able to be hammered or pressed into shape without breaking or cracking
mass number the total number of protons and neutrons in the nucleus of a particular isotope of an element
mass spectrometer analytical instrument that determines the relative isotopic masses of the different isotopes of an element and abundance
measurement bias a type of influence on results in which an experiment manipulates results to get a desired outcome; may be unintentional (i.e. through the placebo effect) or intentional
measurement result a final result; usually the average of several measurement values
mechanical recycling physical recycling of plastics where the plastic is shredded and melted, and converted into pellets
melting temperature temperature at which state changes from solid to liquid
meniscus the curved liquid surface in a tube caused by the hydrostatic properties of fluid with the walls of the glassware
metal ductile, malleable, conducting element
metallic bonding positively charged metal cations arranged in a lattice with delocalised valence electrons being able to flow around them
metallic lattice structure crystal lattice made up of an array of cations
minerals naturally occurring inorganic elements or salts
mining the process of extracting metal ores from the earth
mistakes human errors or personal errors that can impact results, but should not be included in analysis; instead, the experiment should be repeated correctly
mobile phase the mixture to be analysed in the chromatography process
model representation of ideas, phenomena or scientific processes; can be physical models, mathematical models or conceptual models
Mohs scale a scale used to measure the hardness of substances
molar gas constant R, the constant of the universal gas equation; $R = 8.31\ \text{J mol}^{-1}\ \text{K}^{-1}$ when pressure is measured in kPa, volume is measured in L, temperature is measured in K and the quantity of the gas is measured in moles (n)
molar gas volume the volume occupied by a mole of a substance at a given temperature and pressure; at SLC, one mole of gas occupies 24.8 litres
molar mass mass of one mole of a compound expressed in grams per mol
molarity concentration of moles of solute per litre of solution, with units M or mol L^{-1}
mole 6.02×10^{23} particles of a substance
mole ratio represents the relative number of moles of each substance involved in the reaction
molecular formula actual number of atoms of each element in a molecule of a substance
molecule group of atoms bonded together covalently
monomer molecule that links together to form a polymer
monoprotic (of an acid) can donate one proton
multiple bond bond formed when two atoms share two or more pairs of electrons
natural polymers naturally occurring polymers in plants and animals
network lattice a substance that displays covalent bonding in three dimensions without the formation of discrete molecules
neutralisation chemical reaction in which an acid and a base react to form a salt and water
nominal data qualitative data that has no logical sequence
non-bonding electrons electrons that are not involved in bonding
non-metal non-ductile, non-malleable, non-conducting element
non-polar covalent bonds formed between atoms with the same electronegativity
non-polar molecules molecule that does not have permanent dipoles or is symmetrical
non-renewable resources that are depleted or consumed faster than they can be replaced
ocean acidification the decrease in the pH of Earth's oceans, due to increased absorption of carbon dioxide (CO_2) from the atmosphere

octet rule a generalisation that works for many (but not all) atoms, stating that atoms will donate or share electrons in order to achieve eight electrons in their outer shells
open-cut a surface mining technique that removes ore and waste rock from an open pit from the surface
opinions statements based on an individual's personal belief and experience
orbital describes a region where an electron is likely to be found, with a maximum capacity of two electrons
ore a rock containing minerals from which a valuable metal can be removed for profit
organic chemistry the study of carbon-containing compounds and their properties
organic recycling waste material is treated using microorganisms in humid, warm conditions
organometallic compounds a compound containing a metal bonded to a carbon atom of an organic molecule
origin the position on either chromatography paper or a TLC slide that the sample is applied
outlier result that is a long way from other results and seen as unusual
oxidant the substance that causes oxidation of another substance; it accepts electrons and is reduced; see *oxidising agent*
oxidation an increase in the oxidation number; a loss of one or more electrons
oxidation number a theoretical number that corresponds to the charge that atom would have if the compound was composed of ions
oxidising agent the substance that causes oxidation of another substance; it accepts electrons and is reduced
oxidising strength how readily an element gains electrons
partially ionise only a small fraction of the species will break to produce ions
parts per million (ppm) number of a particular component present within one million parts
pascal (Pa) unit of pressure, measured in newtons per square metre (NM^{-2})
percentage composition the percentage contribution that each element makes to the mass of a compound
period horizontal row of the periodic table
personal errors human errors or mistakes that can affect results but should not be included in analysis
PET thermoplastic polymer produced from the monomer ethylene terephthalate; typically used as synthetic fibres (polyester) and packaging (PET); commonly recycled
pH scale used to measure acidity; pH can be calculated using $pH = -\log_{10}[H_3O^+]$
phase the physical form in which a substance is present; common phases include solid, liquid, gas and aqueous solution, but other phases exist as well
photon particle of light
plastic polymer that can be moulded when hot and retains its shape when cooled
plasticisers small molecules that are added to some polymers to improve their flexibility
polar molecules that have an uneven distribution of electrons due to differences in electronegativity of the atoms
polar bond bond formed when two atoms that have different electronegativities share electrons unequally
polar covalent bond formed when two atoms that have different electronegativities share electrons unevenly
polar molecule a molecule which, due to its polar bonds *and* its asymmetric shape, has an overall imbalance in the distribution of its electrons
polarity localised imbalances in electric charges within a molecule resulting in a negatively charged end and positively charged end
political factors factors that are related to the actions of government
polyatomic ion charged ion composed of two or more atoms
polyester synthetic polymer in which the structural units are linked by ester bonds
polyethene polymer consisting of monomers of ethene
polylactic acid (PLA) a versatile, thermoplastic, biodegradable polymer obtained from plant products such as corn starch, maize or sugar cane
polymer molecule made up of thousands of units (monomers) strung together into very long chains
polymerisation the formation of giant molecules by repeated monomers that are joined by covalent bonds
polyprotic (of an acid) able to donate more than one proton
'pop' test a test where a lighted wooden splint is placed in the vicinity of captured hydrogen gas, causing a 'pop' sound in its presence
precipitate solid with low solubility formed in a liquid

precipitation reaction reaction where two solutions are mixed and an insoluble solid is formed
precise how close multiple measurements of the same investigation are to each other, a measure of repeatability or reproducibility
pressure the force per unit area that one region of a gas, liquid or solid exerts on another
primary cell a cell that cannot easily be recharged after use
primary data direct or firsthand evidence about some phenomenon, obtained from investigations or observations
primary source a document which is a record of direct or firsthand evidence about some phenomenon
primary standard substance used in volumetric analysis that is of such high purity and stability that it can be used to prepare a solution of accurately known concentration
principal quantum number, *n* a number related to the energy of an electron and distance from the nucleus
product chemical species obtained as the result of a chemical reaction
qualitative analysis use of experimental procedures to determine which elements are in a substance
qualitative data categorical data that examines the quality of something (e.g. colour or gender) rather than numerical values
qualitatively measurement or observation that describes the qualities of something; labels or categories rather than numbers
quantitative analysis use of experimental procedures to determine the percentage composition of a compound or the percentage of a component in a mixture
quantitative data numerical data that examines the quantity of something (e.g. length, time); also known as numerical data
quantitatively measurement of the quantity of something (e.g. length or time); always numerical
random errors chance variations in measurements; result in a spread of readings
randomised assignment of individuals to an experimental or a control group at random, not influenced by external factors
reactivity series of metals an ordered list of how readily the metals react with oxygen, water, steam, dilute acids and salt solutions
recyclable materials which can be reprocessed into new materials
redox a reaction that involves the transfer of one or more electrons between chemical species
reducing agent the substance that causes reduction of another substance; it donates electrons and is oxidised
reducing strength how readily an element loses electrons
reductant the substance that causes reduction of another substance; it donates electrons and is oxidised; see *reducing agent*
reduction a gain of electrons (decrease in oxidation number)
refining metals purifying a metal
refractive index a measure of how much a substance refracts (bends) light when it enters it
relative atomic mass (A_r) mean mass of the naturally occurring mixture of the isotopes of an element on the relative atomic mass scale, on which the masses of particles are compared with the mass of the carbon-12 isotope
relative formula mass (M_r) the sum of the relative atomic masses of the atoms in the formula of an ionic compound
relative isotopic mass mass of an individual isotope of an element on the relative atomic mass scale, on which the masses of particles are compared with the mass of the carbon-12 isotope
relative molecular mass (M_r) the sum of the relative atomic masses of the atoms in a molecule of a covalent molecular compound
reliability the consistency of a measurement across multiple trials
renewable energy energy sources that can be produced faster than they are used
repeatability how close the results of successive measurements are to each other in the exact same conditions
repeated measurements precision between the results of successive measurements of the same quantity being measured, carried out under the same conditions of measurement
reproducible how close the results of measuring the same quantity are to each other in changed conditions (different observer, method, instrument, time etc.)

residue substance remaining in filter paper after the process of filtration
resolution the smallest change of measurement that a particular equipment can detect
response bias a type of influence on results in which only certain members of the target population respond to an invitation to participate in the clinical trial, resulting in an unrepresentative sample of the larger population
results section in a scientific report in which all data obtained is recorded, usually in the form of tables and graphs
retention time (Rt) time taken for a substance to travel through a chromatography column and exit at the end
reticulated water water provided through a network of pipes
R_f in chromatography, the retardation factor; the ratio of the distance travelled by the solute to the distance travelled by the solvent
risk assessment a document that examines the different hazards in an investigation and suggested safety precautions
rubber an elastic polymer formed from the latex of tropical rubber trees and plants
rusting wet corrosion of iron
sacrificial anode a reactive metal used to prevent a less active metal from corroding
sacrificial protection protection of a metal from corrosion by attaching another more active metal to it; the more active metal reacts preferentially with corrosive elements and protects the less reactive metal
salt ionic compound consisting of a metal ion and a non-metal ion, except oxides and hydroxides
salt bridge provides ions to balance ions consumed or produced in each half-cell
sample size the number of trials in an investigation
sampling bias a type of influence on results in which participants chosen for a study are not representative of the target population
sanitation conditions of public health, specifically the delivery of clean drinking water and adequate sewage disposal
saturated hydrocarbon compound composed of carbon and hydrogen and containing only single covalent bonds
saturated solution solution in which the maximum amount of solvent has been dissolved
scatterplots graph in which two quantitative variables are plotted as a series of dots
scientific investigation methodology the principles of research based on the scientific method
scientific methodology the type of investigation being conducted to answer a question and resolve a hypothesis
secondary cell a cell that can be recharged
secondary data researchers' comments on or summaries and interpretations of primary data
secondary fuel a fuel that is produced from another energy source
secondary source a document that comments on, summarises or interprets primary data
secondary standard a standard prepared for a specific volumetric analysis that has been standardised against a primary standard
selection bias a type of influence on results in which test subjects are not equally and randomly assigned to experimental and control groups
self-ionisation constant *Kw*; in all aqueous solutions this value is constant and results from the self-ionisation of water
semi-structural formula derived from the structural formula that shows the atoms present in a molecule, carbon by carbon, and is written on a single line; also called a condensed formula
single bond formed by one pair of electrons shared between two atoms
skeletal formula structural representation where a molecule is represented by lines and vertices; carbons and hydrogens are not shown but assumed to be at each vertex and at the ends unless otherwise indicated
slag waste material from the smelting or refining of ore
social factors factors that influence both a whole society and the individuals within the society
socioeconomic the interaction of social and economic factors
solubility grams of a substance that will dissolve in an amount of solvent at a particular temperature
solubility curves graphs of grams of a substance that will dissolve in 100 g of solvent at a particular temperature
soluble will dissolve

solute dissolved substance in a solution
solution solute dissolved in a solvent
solvent substance in which the solute is dissolved
space-filling models three-dimensional representation of a molecule that shows the relative sizes of atoms within the molecule and the distances between them
species general term that could refer to atoms, molecules, molecular fragments or ions
specific heat capacity energy (measured in joules) needed to raise the temperature of 1 g of a pure substance by 1 ^{o}C
spectator ion ion that is present but takes no part in a reaction
standard solution solution that has a precisely known concentration
stationary phase a material with a high surface area, used in the chromatography process
stock solution concentrated solution that will be diluted as required for later use
stoichiometry the amount and ratio of reactants and products in a chemical reaction
straight chain carbon atoms connected in one continuous line without any branches
strong acid acid that fully ionises in water
strong base base that fully dissociates in water
structural formula a diagrammatic representation of a molecule showing every bond, but not showing lone pairs of electrons
structural isomers compounds that have the same molecular formula but a different semi-structural formula due to differences in the arrangement of the atoms
subatomic particles particles in atoms: electrons, protons and neutrons
substitution reaction reaction in which an atom (or group of atoms) is removed and replaced by a different atom (or group of atoms)
supersaturated solution solution in which a greater amount of solute is dissolved at a particular temperature than is predicted by a solubility curve temperature
sustainability development that meets current needs without compromising the needs of future generations
Sustainable Development Goals adopted by all United Nations Member States in 2015, consists of 17 goals that address 169 urgent calls for action to global challenges with the overall goal of a sustainable future for all
symbol simplified representation of an element consisting of one or two letters
synthetic polymers polymers made from fossil fuels or biomass
systematic errors errors that affect the accuracy of a measurement that cannot be improved by repeating an experiment; usually due to equipment or system errors
tailings waste material remaining after the removal of the ore from the host rock
tentative not fixed or certain; may be changed with new information
testable able to be supported or proven false through the use of observations and investigation
theory a well-supported explanation of a phenomena, based on facts that have been obtained through investigations, research and observations
thermoplastic polymers that soften on heating; also known as thermosoftening
thermosetting polymers that do not soften on heating and char if heated strongly
titrant the standard solution of known concentration to go into the burette
titration process used to determine the concentration of a substance by using a pipette to deliver one substance and a burette to deliver another substance until they have reacted in exactly their mole ratio
titration curve results of a titration showing the pH change versus the volume of the titre added during titration
titre volume of liquid measured by a burette during titration
transition elements metallic elements in groups 3 to 12 with the *d* subshell being filled
triple bonds very strong bond between two atoms formed by three pairs of electrons that are shared by the two nuclei
triprotic (of an acid) can donate three protons
true value the value that would be found if the quantity could be measured perfectly
uncertainty a limit to the precision of data obtained; a range within which a measurement lies
underground a variety of mining techniques that use tunnels and shafts to extract ore and waste rock from under the ground; this process minimises waste rock removal when the ore body is deep below the surface

unified atomic mass units (u) standard unit of mass based on the mass of the C-12 isotope, where mass of an atom of C-12 equals 12 u; it is approximately equal to one nucleon (proton or neutron)
United Nations Declaration on the Rights of Indigenous Peoples a universal framework of minimum standards for the survival, dignity and wellbeing of the Indigenous peoples of the world
unsaturated hydrocarbon compound containing carbon and hydrogen but possessing one or more double or triple bonds
unsaturated solution solution containing less than the maximum amount of solute that can dissolve
urban mining obtaining metals from e-waste and other materials already in our cities
UV–visible spectroscopy technique used to study compounds that absorb light in the UV–visible region
valence electrons electrons in the outermost shell of an atom; largely determine chemical properties of an element and contribute to chemical bond formation
valence shell electron pair repulsion (VSEPR) theory model used to predict the shape of a molecule based on minimising electron repulsions around a given atom in a molecule
valence structure a diagrammatic representation of the outershell electrons in a molecule; similar to an electron dot diagram but replaces each pair of electrons with a dash (–)
valid describes an experiment that investigates what it sets out to investigate (via appropriately controlling variables, eliminating bias etc.)
validity how accurately an experiment investigates the claim it is intended to investigate
volumetric analysis measurement of volume of one solution of accurately known concentration to determine the concentration of a second solution using stoichiometry
vulcanisation the process in which sulfur is added to rubber and heated to cause cross-linking of polymer chains, increasing the strength of rubber
water of hydration water molecules present in the crystal lattice of hydrated compounds; can also be referred to as ‘water of crystallisation’
weak acid acid that partially ionises in water
weak base base that partially dissociates in water
wet corrosion corrosion of a metal by reaction with oxygen in the presence of water

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PERIODIC TABLE OF THE ELEMENTS

	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9
	Alkali metals ↓	**Alkaline earth metals** ↓					**Transition metals**		
Period 1	**1** Hydrogen **H** 1.0								
Period 2	**3** Lithium **Li** 6.9	**4** Beryllium **Be** 9.0							
Period 3	**11** Sodium **Na** 23.0	**12** Magnesium **Mg** 24.3							
Period 4	**19** Potassium **K** 39.1	**20** Calcium **Ca** 40.1	**21** Scandium **Sc** 45.0	**22** Titanium **Ti** 47.9	**23** Vanadium **V** 50.9	**24** Chromium **Cr** 52.0	**25** Manganese **Mn** 54.9	**26** Iron **Fe** 55.8	**27** Cobalt **Co** 58.9
Period 5	**37** Rubidium **Rb** 85.5	**38** Strontium **Sr** 87.6	**39** Yttrium **Y** 88.9	**40** Zirconium **Zr** 91.2	**41** Niobium **Nb** 92.9	**42** Molybdenum **Mo** 96.0	**43** Technetium **Tc** (98)	**44** Ruthenium **Ru** 101.1	**45** Rhodium **Rh** 102.9
Period 6	**55** Caesium **Cs** 132.9	**56** Barium **Ba** 137.3	**57–71** Lanthanoids	**72** Hafnium **Hf** 178.5	**73** Tantalum **Ta** 180.9	**74** Tungsten **W** 183.8	**75** Rhenium **Re** 186.2	**76** Osmium **Os** 190.2	**77** Iridium **Ir** 192.2
Period 7	**87** Francium **Fr** (223)	**88** Radium **Ra** (226)	**89–103** Actinoids	**104** Rutherfordium **Rf** (261)	**105** Dubnium **Db** (262)	**106** Seaborgium **Sg** (266)	**107** Bohrium **Bh** (264)	**108** Hassium **Hs** (267)	**109** Meitnerium **Mt** (268)

Period 1

1 Hydrogen **H** 1.0	**2** Helium **He** 4.0

Key
- **Atomic number**
- Name
- **Symbol**
- Relative atomic mass

- Alkali metals
- Alkaline earth metals
- Transition metals
- Lathanoids
- Actinoids
- Unknown chemical properties
- Post-transition metals
- Metalloids
- Non-metals
- Halides
- Noble gases

Lanthanoids

57 Lanthanum **La** 138.9	**58** Cerium **Ce** 140.1	**59** Praseodymium **Pr** 140.9	**60** Neodymium **Nd** 144.2	**61** Promethium **Pm** (145)	**62** Samarium **Sm** 150.4	**63** Europium **Eu** 152.0

Actinoids

89 Actinium **Ac** (227)	**90** Thorium **Th** 232.0	**91** Protactinium **Pa** 231.0	**92** Uranium **U** 238.0	**93** Neptunium **Np** (237)	**94** Plutonium **Pu** (244)	**95** Americium **Am** (243)

Noble gases

Group 18

Non-metals

Group 10	Group 11	Group 12	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
			5 Boron **B** 10.8	**6** Carbon **C** 12.0	**7** Nitrogen **N** 14.0	**8** Oxygen **O** 16.0	**9** Fluorine **F** 19.0	**2** Helium **He** 4.0
			13 Aluminium **Al** 27.0	**14** Silicon **Si** 28.1	**15** Phosphorus **P** 31.0	**16** Sulfur **S** 32.1	**17** Chlorine **Cl** 35.5	**10** Neon **Ne** 20.2
								18 Argon **Ar** 39.9
28 Nickel **Ni** 58.7	**29** Copper **Cu** 63.5	**30** Zinc **Zn** 65.4	**31** Gallium **Ga** 69.7	**32** Germanium **Ge** 72.6	**33** Arsenic **As** 74.9	**34** Selenium **Se** 79.0	**35** Bromine **Br** 79.9	**36** Krypton **Kr** 83.8
46 Palladium **Pd** 106.4	**47** Silver **Ag** 107.9	**48** Cadmium **Cd** 112.4	**49** Indium **In** 114.8	**50** Tin **Sn** 118.7	**51** Antimony **Sb** 121.8	**52** Tellurium **Te** 127.6	**53** Iodine **I** 126.9	**54** Xenon **Xe** 131.3
78 Platinum **Pt** 195.1	**79** Gold **Au** 197.0	**80** Mercury **Hg** 200.6	**81** Thallium **Tl** 204.4	**82** Lead **Pb** 207.2	**83** Bismuth **Bi** 209.0	**84** Polonium **Po** (210)	**85** Astatine **At** (210)	**86** Radon **Rn** (222)
110 Darmstadtium **Ds** (271)	**111** Roentgenium **Rg** (272)	**112** Copernicium **Cn** (285)	**113** Nihonium **Nh** (280)	**114** Flerovium **Fl** (289)	**115** Moscovium **Mc** (289)	**116** Livermorium **Lv** (292)	**117** Tennessine **Ts** (294)	**118** Oganesson **Og** (294)

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64 Gadolinium **Gd** 157.3	**65** Terbium **Tb** 158.9	**66** Dysprosium **Dy** 162.5	**67** Holmium **Ho** 164.9	**68** Erbium **Er** 167.3	**69** Thulium **Tm** 168.9	**70** Ytterbium **Yb** 173.1	**71** Lutetium **Lu** 175.0
96 Curium **Cm** (247)	**97** Berkelium **Bk** (247)	**98** Californium **Cf** (251)	**99** Einsteinium **Es** (252)	**100** Fermium **Fm** (257)	**101** Mendelevium **Md** (258)	**102** Nobelium **No** (259)	**103** Lawrencium **Lr** (262)