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Getting the best from the book

Welcome to Collins *Cambridge IGCSE Chemistry*.

This textbook has been designed to help you understand all of the requirements needed to succeed in the Cambridge IGCSE Chemistry course. The book has taken the Cambridge IGCSE syllabus and split each of the topics across four main sections of study: Principles of chemistry, Physical chemistry, Inorganic chemistry and Organic chemistry.

Each topic in the textbook covers the essential knowledge and skills you need. The textbook also has some very useful features which have been designed to really help you understand all the aspects of Chemistry which you will need to know for this syllabus.

SAFETY IN THE SCIENCE LESSON

This book is a textbook, not a laboratory or practical manual. As such, you should not interpret any information in this book that relates to practical work as including comprehensive safety instructions. Your teachers will provide full guidance for practical work and cover rules that are specific to your school.

A brief introduction to the section to give context to the science covered in the section.

Starting points will help you to revise previous learning and see what you already know about the ideas to be covered in the section.

The section contents shows the separate topics to be studied matching the syllabus order.


2

Physical chemistry

Modern physical chemistry originated in the 19th century. It is not as clearly defined a category as organic chemistry, but it is still a useful description of this branch of science. Physical chemistry focuses on chemical processes at the 'macro level' (where properties can be observed) more than at the 'micro level' (too small to see) of individual atoms, molecules and ions. However, observed physical properties can still be explained in terms of what the atoms, molecules or ions are doing.

In this section you will explore the chemical reactions that can be caused by using electricity, a process known as electrolysis. You will then investigate some chemical reactions that produce significant amounts of heat energy, as well as some strange ones that seem to absorb energy and make everything cooler. The speed or rate of chemical reactions will also be explored, together with chemists' strategies to try to control them. You will learn about reactions that go from reactants to products and then back again. These are a particular challenge when chemists want to make a product that could turn back into the reactants that made it! You will learn about redox reactions, which are reactions involving reduction and oxidation, as well as about acids, bases and salts. Finally, you will look at some of the simple analytical techniques that can be used to identify ions and gases.

STARTING POINTS

1. How many non-renewable fuels can you name? What products do they form when they burn?
2. Give an example of a very rapid, almost instantaneous, chemical reaction. Now give an example of a very slow one.
3. Explain how you can easily distinguish between an acid and an alkali.
4. What is a catalyst? Name two examples where catalysts are used in everyday life.
5. Name a process that can be reversed easily.
6. Acids react with alkalis in neutralisation reactions. What is meant by neutralisation?

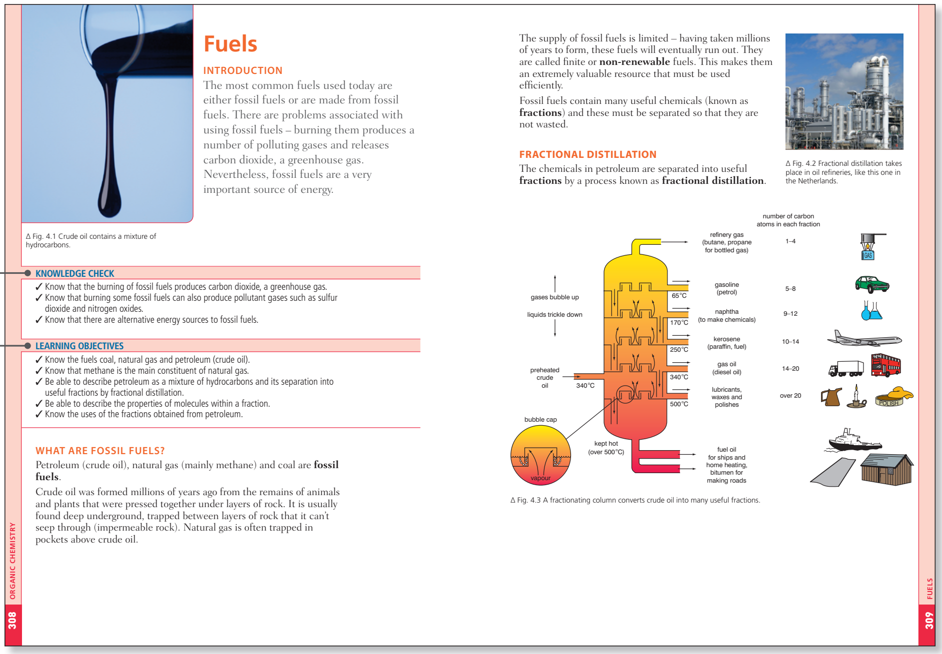
SECTION CONTENTS

a) Electricity and chemistry
b) Chemical energetics
c) Rate of reaction
d) Reversible reactions
e) Redox reactions
f) Acids, bases and salts
g) Identification of ions and gases
h) Exam-style questions

△ Physical chemistry deals with properties that can be observed.

Knowledge check shows the ideas you should have already encountered in previous work before starting the topic.

Learning objectives cover what you need to learn in this topic.



Fuels

INTRODUCTION

The most common fuels used today are either fossil fuels or are made from fossil fuels. There are problems associated with using fossil fuels – burning them produces a number of polluting gases and releases carbon dioxide, a greenhouse gas. Nevertheless, fossil fuels are a very important source of energy.

Δ Fig. 4.1 Crude oil contains a mixture of hydrocarbons.

KNOWLEDGE CHECK

✓ Know that the burning of fossil fuels produces carbon dioxide, a greenhouse gas.
✓ Know that burning some fossil fuels can also produce pollutant gases such as sulfur dioxide and nitrogen oxides.
✓ Know that there are alternative energy sources to fossil fuels.

LEARNING OBJECTIVES

✓ Know the fuels coal, natural gas and petroleum (crude oil).
✓ Know that methane is the main constituent of natural gas.
✓ Be able to describe petroleum as a mixture of hydrocarbons and its separation into useful fractions by fractional distillation.
✓ Be able to describe the properties of molecules within a fraction.
✓ Know the uses of the fractions obtained from petroleum.

WHAT ARE FOSSIL FUELS?

Petroleum (crude oil), natural gas (mainly methane) and coal are **fossil fuels**.

Crude oil was formed millions of years ago from the remains of animals and plants that were pressed together under layers of rock. It is usually found deep underground, trapped between layers of rock that it can't seep through (impermeable rock). Natural gas is often trapped in pockets above crude oil.

The supply of fossil fuels is limited – having taken millions of years to form, these fuels will eventually run out. They are called finite or **non-renewable** fuels. This makes them an extremely valuable resource that must be used efficiently.

Fossil fuels contain many useful chemicals (known as **fractions**) and these must be separated so that they are not wasted.

Δ Fig. 4.2 Fractional distillation takes place in oil refineries, like this one in the Netherlands.

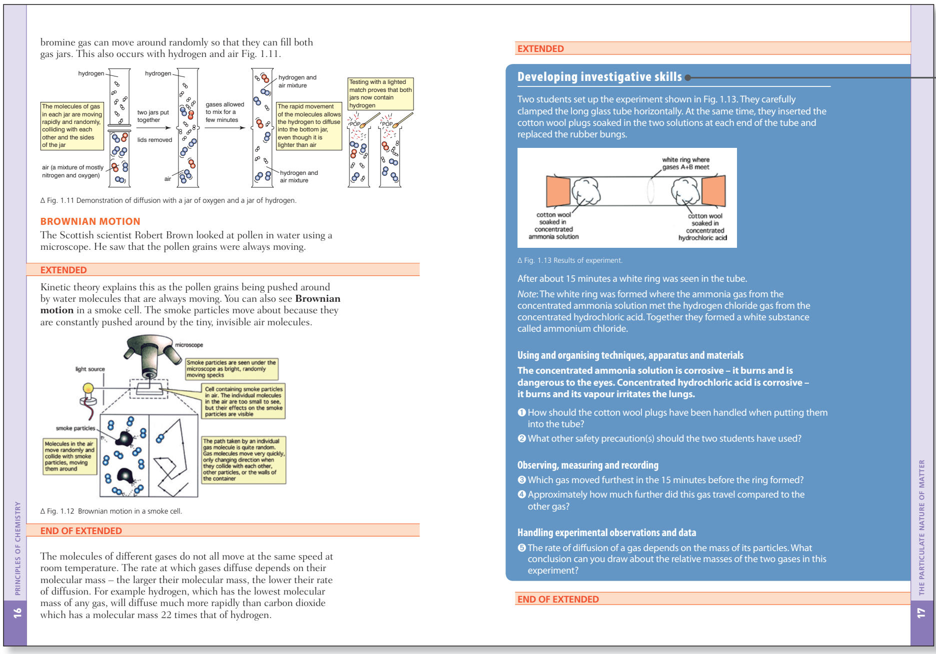
FRACTIONAL DISTILLATION

The chemicals in petroleum are separated into useful **fractions** by a process known as **fractional distillation**.

Δ Fig. 4.3 A fractionating column converts crude oil into many useful fractions.

308 ORGANIC CHEMISTRY

FUELS 309



bromine gas can move around randomly so that they can fill both gas jars. This also occurs with hydrogen and air Fig. 1.11.

Δ Fig. 1.11 Demonstration of diffusion with a jar of oxygen and a jar of hydrogen.

BROWNIAN MOTION

The Scottish scientist Robert Brown looked at pollen in water using a microscope. He saw that the pollen grains were always moving.

EXTENDED

Kinetic theory explains this as the pollen grains being pushed around by water molecules that are always moving. You can also see **Brownian motion** in a smoke cell. The smoke particles move about because they are constantly pushed around by the tiny, invisible air molecules.

Δ Fig. 1.12 Brownian motion in a smoke cell.

END OF EXTENDED

The molecules of different gases do not all move at the same speed at room temperature. The rate at which gases diffuse depends on their molecular mass – the larger their molecular mass, the lower their rate of diffusion. For example hydrogen, which has the lowest molecular mass of any gas, will diffuse much more rapidly than carbon dioxide which has a molecular mass 22 times that of hydrogen.

16 PRINCIPLES OF CHEMISTRY

EXTENDED

Developing investigative skills

Two students set up the experiment shown in Fig. 1.13. They carefully clamped the long glass tube horizontally. At the same time, they inserted the cotton wool plugs soaked in the two solutions at each end of the tube and replaced the rubber bungs.

Δ Fig. 1.13 Results of experiment.

After about 15 minutes a white ring was seen in the tube.

Note: The white ring was formed where the ammonia gas from the concentrated ammonia solution met the hydrogen chloride gas from the concentrated hydrochloric acid. Together they formed a white substance called ammonium chloride.

Using and organising techniques, apparatus and materials

The concentrated ammonia solution is corrosive – it burns and is dangerous to the eyes. Concentrated hydrochloric acid is corrosive – it burns and its vapour irritates the lungs.

❶ How should the cotton wool plugs have been handled when putting them into the tube?
❷ What other safety precaution(s) should the two students have used?

Observing, measuring and recording

❸ Which gas moved furthest in the 15 minutes before the ring formed?
❹ Approximately how much further did this gas travel compared to the other gas?

Handling experimental observations and data

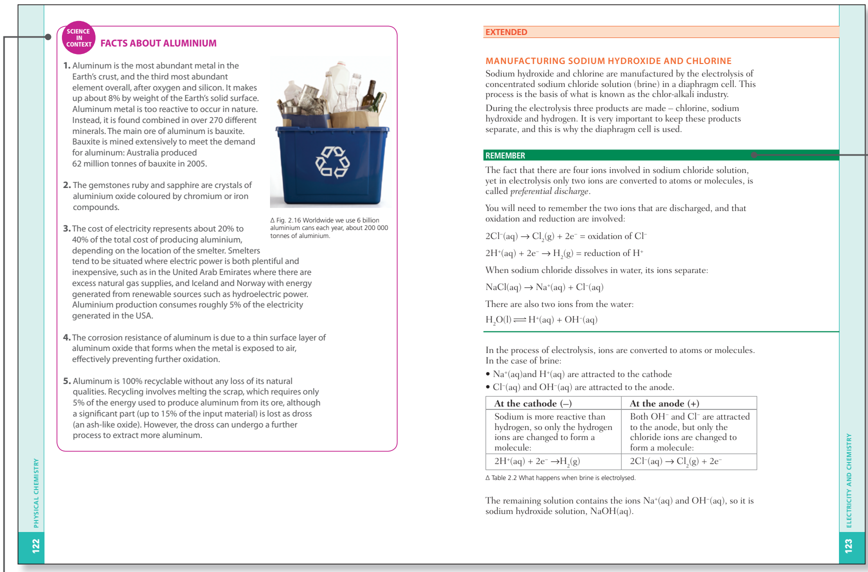
❺ The rate of diffusion of a gas depends on the mass of its particles. What conclusion can you draw about the relative masses of the two gases in this experiment?

END OF EXTENDED

THE PARTICULATE NATURE OF MATTER 17

Examples of investigations are included with questions matched to the investigative skills you will need to learn.

Getting the best from the book *continued*



SCIENCE IN CONTEXT **FACTS ABOUT ALUMINIUM**

1. Aluminum is the most abundant metal in the Earth's crust, and the third most abundant element overall, after oxygen and silicon. It makes up about 8% by weight of the Earth's solid surface. Aluminum metal is too reactive to occur in nature. Instead, it is found combined in over 270 different minerals. The main ore of aluminum is bauxite. Bauxite is mined extensively to meet the demand for aluminum: Australia produced 62 million tonnes of bauxite in 2005.
2. The gemstones ruby and sapphire are crystals of aluminium oxide coloured by chromium or iron compounds.
3. The cost of electricity represents about 20% to 40% of the total cost of producing aluminium, depending on the location of the smelter. Smelters tend to be situated where electric power is both plentiful and inexpensive, such as in the United Arab Emirates where there are excess natural gas supplies, and Iceland and Norway with energy generated from renewable sources such as hydroelectric power. Aluminium production consumes roughly 5% of the electricity generated in the USA.
4. The corrosion resistance of aluminum is due to a thin surface layer of aluminum oxide that forms when the metal is exposed to air, effectively preventing further oxidation.
5. Aluminum is 100% recyclable without any loss of its natural qualities. Recycling involves melting the scrap, which requires only 5% of the energy used to produce aluminum from its ore, although a significant part (up to 15% of the input material) is lost as dross (an ash-like oxide). However, the dross can undergo a further process to extract more aluminum.

Δ Fig. 2.16 Worldwide we use 6 billion aluminium cans each year, about 200 000 tonnes of aluminium.

122 PHYSICAL CHEMISTRY

EXTENDED

MANUFACTURING SODIUM HYDROXIDE AND CHLORINE

Sodium hydroxide and chlorine are manufactured by the electrolysis of concentrated sodium chloride solution (brine) in a diaphragm cell. This process is the basis of what is known as the chlor-alkali industry.

During the electrolysis three products are made – chlorine, sodium hydroxide and hydrogen. It is very important to keep these products separate, and this is why the diaphragm cell is used.

REMEMBER

The fact that there are four ions involved in sodium chloride solution, yet in electrolysis only two ions are converted to atoms or molecules, is called *preferential discharge*.

You will need to remember the two ions that are discharged, and that oxidation and reduction are involved:

$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$ = oxidation of Cl^-

$2H^+(aq) + 2e^- \rightarrow H_2(g)$ = reduction of H^+

When sodium chloride dissolves in water, its ions separate:

$NaCl(aq) \rightarrow Na^+(aq) + Cl^-(aq)$

There are also two ions from the water:

$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$

In the process of electrolysis, ions are converted to atoms or molecules. In the case of brine:

- $Na^+(aq)$and $H^+(aq)$ are attracted to the cathode
- $Cl^-(aq)$ and $OH^-(aq)$ are attracted to the anode.

At the cathode (–)	At the anode (+)
Sodium is more reactive than hydrogen, so only the hydrogen ions are changed to form a molecule:	Both OH^- and Cl^- are attracted to the anode, but only the chloride ions are changed to form a molecule:
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

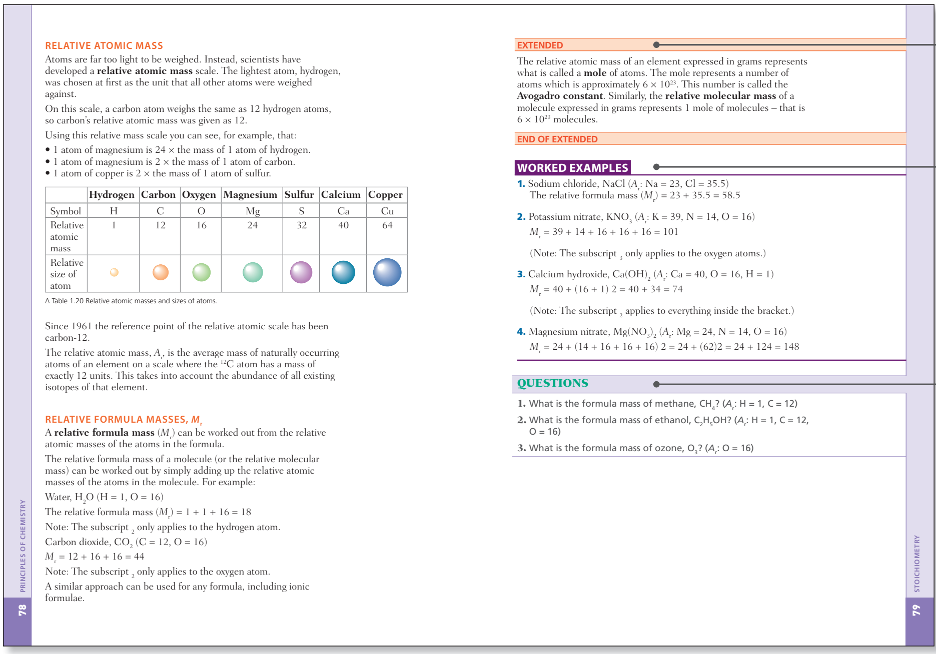
Δ Table 2.2 What happens when brine is electrolysed.

The remaining solution contains the ions $Na^+(aq)$ and $OH^-(aq)$, so it is sodium hydroxide solution, $NaOH(aq)$.

ELECTRICITY AND CHEMISTRY 123

Remember boxes provide tips and guidance to help you during your course and to prepare for examination.

Science in context boxes put the ideas you are learning into real-life context. It is not necessary for you to learn the content of these boxes as they do not form part of the syllabus. However, they do provide interesting examples of scientific application that are designed to enhance your understanding.



RELATIVE ATOMIC MASS

Atoms are far too light to be weighed. Instead, scientists have developed a **relative atomic mass** scale. The lightest atom, hydrogen, was chosen at first as the unit that all other atoms were weighed against.

On this scale, a carbon atom weighs the same as 12 hydrogen atoms, so carbon's relative atomic mass was given as 12.

Using this relative mass scale you can see, for example, that:

- 1 atom of magnesium is 24 × the mass of 1 atom of hydrogen.
- 1 atom of magnesium is 2 × the mass of 1 atom of carbon.
- 1 atom of copper is 2 × the mass of 1 atom of sulfur.

	Hydrogen	Carbon	Oxygen	Magnesium	Sulfur	Calcium	Copper
Symbol	H	C	O	Mg	S	Ca	Cu
Relative atomic mass	1	12	16	24	32	40	64
Relative size of atom							

Δ Table 1.20 Relative atomic masses and sizes of atoms.

Since 1961 the reference point of the relative atomic scale has been carbon-12.

The relative atomic mass, A_r, is the average mass of naturally occurring atoms of an element on a scale where the ^{12}C atom has a mass of exactly 12 units. This takes into account the abundance of all existing isotopes of that element.

RELATIVE FORMULA MASSES, M_r

A **relative formula mass** (M_r) can be worked out from the relative atomic masses of the atoms in the formula.

The relative formula mass of a molecule (or the relative molecular mass) can be worked out by simply adding up the relative atomic masses of the atoms in the molecule. For example:

Water, H_2O (H = 1, O = 16)

The relative formula mass (M_r) = 1 + 1 + 16 = 18

Note: The subscript $_2$ only applies to the hydrogen atom.

Carbon dioxide, CO_2 (C = 12, O = 16)

M_r = 12 + 16 + 16 = 44

Note: The subscript $_2$ only applies to the oxygen atom.

A similar approach can be used for any formula, including ionic formulae.

78 PRINCIPLES OF CHEMISTRY

EXTENDED

The relative atomic mass of an element expressed in grams represents what is called a **mole** of atoms. The mole represents a number of atoms which is approximately 6×10^{23}. This number is called the **Avogadro constant**. Similarly, the **relative molecular mass** of a molecule expressed in grams represents 1 mole of molecules – that is 6×10^{23} molecules.

END OF EXTENDED

WORKED EXAMPLES

1. Sodium chloride, NaCl (A_r: Na = 23, Cl = 35.5)
   The relative formula mass (M_r) = 23 + 35.5 = 58.5
2. Potassium nitrate, KNO_3 (A_r: K = 39, N = 14, O = 16)
   M_r = 39 + 14 + 16 + 16 + 16 = 101
   (Note: The subscript $_3$ only applies to the oxygen atoms.)
3. Calcium hydroxide, $Ca(OH)_2$ (A_r: Ca = 40, O = 16, H = 1)
   M_r = 40 + (16 + 1) 2 = 40 + 34 = 74
   (Note: The subscript $_2$ applies to everything inside the bracket.)
4. Magnesium nitrate, $Mg(NO_3)_2$ (A_r: Mg = 24, N = 14, O = 16)
   M_r = 24 + (14 + 16 + 16 + 16) 2 = 24 + (62)2 = 24 + 124 = 148

QUESTIONS

1. What is the formula mass of methane, CH_4? (A_r: H = 1, C = 12)
2. What is the formula mass of ethanol, C_2H_5OH? (A_r: H = 1, C = 12, O = 16)
3. What is the formula mass of ozone, O_3? (A_r: O = 16)

STOICHIOMETRY 79

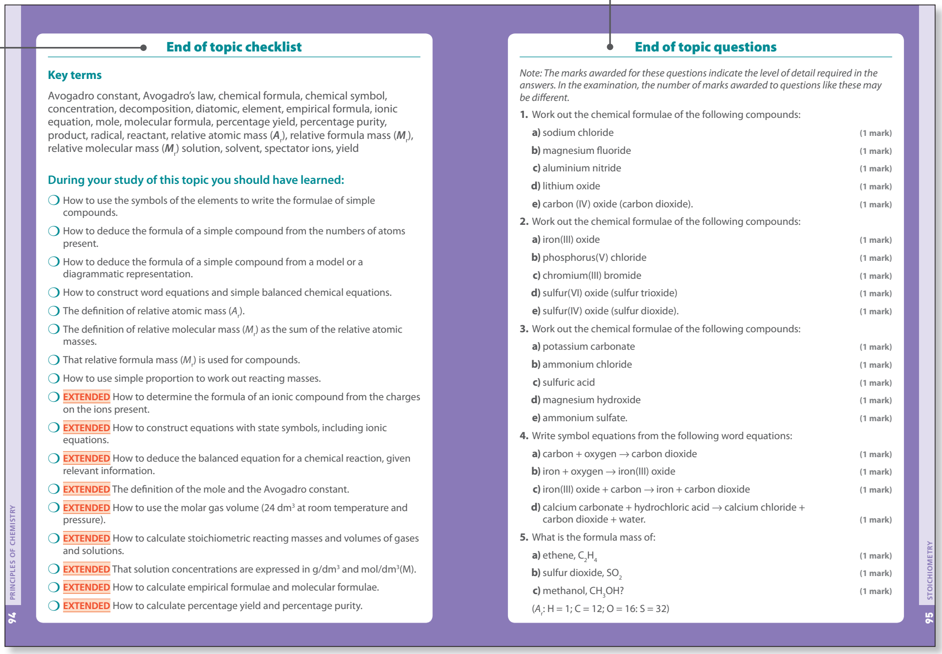
Clearly differentiated Extended material takes your learning even further.

Learn to apply formulae through worked examples.

Questions to check your understanding.

End of topic questions allow you to apply the knowledge and understanding you have learned in the topic to answer the questions.

A full checklist of all the information you need to cover the complete syllabus requirements for each topic.



94 PRINCIPLES OF CHEMISTRY

End of topic checklist

Key terms

Avogadro constant, Avogadro's law, chemical formula, chemical symbol, concentration, decomposition, diatomic, element, empirical formula, ionic equation, mole, molecular formula, percentage yield, percentage purity, product, radical, reactant, relative atomic mass (A_r), relative formula mass (M_r), relative molecular mass (M_r) solution, solvent, spectator ions, yield

During your study of this topic you should have learned:

- How to use the symbols of the elements to write the formulae of simple compounds.
- How to deduce the formula of a simple compound from the numbers of atoms present.
- How to deduce the formula of a simple compound from a model or a diagrammatic representation.
- How to construct word equations and simple balanced chemical equations.
- The definition of relative atomic mass (A_r).
- The definition of relative molecular mass (M_r) as the sum of the relative atomic masses.
- That relative formula mass (M_r) is used for compounds.
- How to use simple proportion to work out reacting masses.
- EXTENDED How to determine the formula of an ionic compound from the charges on the ions present.
- EXTENDED How to construct equations with state symbols, including ionic equations.
- EXTENDED How to deduce the balanced equation for a chemical reaction, given relevant information.
- EXTENDED The definition of the mole and the Avogadro constant.
- EXTENDED How to use the molar gas volume (24 dm³ at room temperature and pressure).
- EXTENDED How to calculate stoichiometric reacting masses and volumes of gases and solutions.
- EXTENDED That solution concentrations are expressed in g/dm³ and mol/dm³(M).
- EXTENDED How to calculate empirical formulae and molecular formulae.
- EXTENDED How to calculate percentage yield and percentage purity.

End of topic questions

Note: The marks awarded for these questions indicate the level of detail required in the answers. In the examination, the number of marks awarded to questions like these may be different.

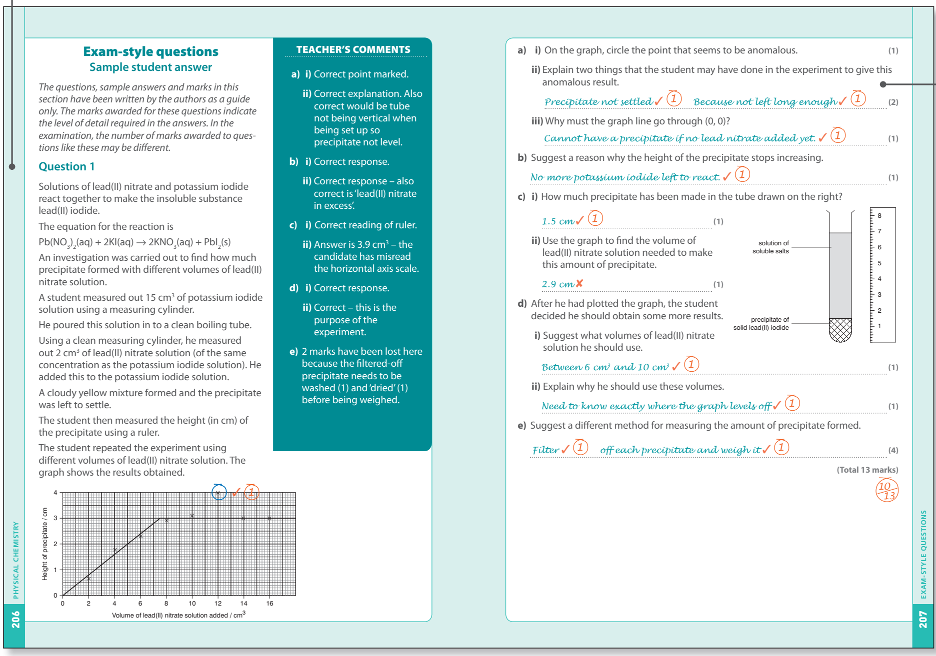
1. Work out the chemical formulae of the following compounds:
   - a) sodium chloride (1 mark)
   - b) magnesium fluoride (1 mark)
   - c) aluminium nitride (1 mark)
   - d) lithium oxide (1 mark)
   - e) carbon (IV) oxide (carbon dioxide). (1 mark)
2. Work out the chemical formulae of the following compounds:
   - a) iron(III) oxide (1 mark)
   - b) phosphorus(V) chloride (1 mark)
   - c) chromium(III) bromide (1 mark)
   - d) sulfur(VI) oxide (sulfur trioxide) (1 mark)
   - e) sulfur(IV) oxide (sulfur dioxide). (1 mark)
3. Work out the chemical formulae of the following compounds:
   - a) potassium carbonate (1 mark)
   - b) ammonium chloride (1 mark)
   - c) sulfuric acid (1 mark)
   - d) magnesium hydroxide (1 mark)
   - e) ammonium sulfate. (1 mark)
4. Write symbol equations from the following word equations:
   - a) carbon + oxygen → carbon dioxide (1 mark)
   - b) iron + oxygen → iron(III) oxide (1 mark)
   - c) iron(III) oxide + carbon → iron + carbon dioxide (1 mark)
   - d) calcium carbonate + hydrochloric acid → calcium chloride + carbon dioxide + water. (1 mark)
5. What is the formula mass of:
   - a) ethene, C_2H_4 (1 mark)
   - b) sulfur dioxide, SO_2 (1 mark)
   - c) methanol, CH_3OH? (1 mark)

(A_r: H = 1; C = 12; O = 16: S = 32)

STOICHIOMETRY 95

Each section includes exam-style questions to help you prepare for your exam in a focussed way and get the best results.

The first question is a student sample with teacher's comments to show best practice.



206 PHYSICAL CHEMISTRY

Exam-style questions

Sample student answer

The questions, sample answers and marks in this section have been written by the authors as a guide only. The marks awarded for these questions indicate the level of detail required in the answers. In the examination, the number of marks awarded to questions like these may be different.

Question 1

Solutions of lead(II) nitrate and potassium iodide react together to make the insoluble substance lead(II) iodide.

The equation for the reaction is

$Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow 2KNO_3(aq) + PbI_2(s)$

An investigation was carried out to find how much precipitate formed with different volumes of lead(II) nitrate solution.

A student measured out 15 cm³ of potassium iodide solution using a measuring cylinder.

He poured this solution in to a clean boiling tube.

Using a clean measuring cylinder, he measured out 2 cm³ of lead(II) nitrate solution (of the same concentration as the potassium iodide solution). He added this to the potassium iodide solution.

A cloudy yellow mixture formed and the precipitate was left to settle.

The student then measured the height (in cm) of the precipitate using a ruler.

The student repeated the experiment using different volumes of lead(II) nitrate solution. The graph shows the results obtained.

TEACHER'S COMMENTS

a) i) Correct point marked.

ii) Correct explanation. Also correct would be tube not being vertical when being set up so precipitate not level.

b) i) Correct response.

ii) Correct response – also correct is 'lead(II) nitrate in excess'.

c) i) Correct reading of ruler.

ii) Answer is 3.9 cm³ – the candidate has misread the horizontal axis scale.

d) i) Correct response.

ii) Correct – this is the purpose of the experiment.

e) 2 marks have been lost here because the filtered-off precipitate needs to be washed (1) and 'dried' (1) before being weighed.

a) i) On the graph, circle the point that seems to be anomalous. (1)

ii) Explain two things that the student may have done in the experiment to give this anomalous result.

Precipitate not settled ✓ 1 *Because not left long enough* ✓ 1 (2)

iii) Why must the graph line go through (0, 0)?

Cannot have a precipitate if no lead nitrate added yet. ✓ 1 (1)

b) Suggest a reason why the height of the precipitate stops increasing.

No more potassium iodide left to react. ✓ 1 (1)

c) i) How much precipitate has been made in the tube drawn on the right?

1.5 cm ✓ 1 (1)

ii) Use the graph to find the volume of lead(II) nitrate solution needed to make this amount of precipitate.

2.9 cm ✗ (1)

d) After he had plotted the graph, the student decided he should obtain some more results.

i) Suggest what volumes of lead(II) nitrate solution he should use.

Between 6 cm³ and 10 cm³ ✓ 1 (1)

ii) Explain why he should use these volumes.

Need to know exactly where the graph levels off ✓ 1 (1)

e) Suggest a different method for measuring the amount of precipitate formed.

Filter ✓ 1 *off each precipitate and weigh it* ✓ 1 (4)

(Total 13 marks)

10/13

EXAM-STYLE QUESTIONS 207

This section provides the basic ideas that the rest of your course is built on. You may have covered some aspects in your previous work, but it is important to understand the key principles thoroughly before seeing how these can be applied across all the other sections. The section covers some of the experimental techniques you will meet in your course.

First you will look at the existing evidence for the particulate nature of matter. Next, you will consider the structure of an atom and why the atoms of different elements have different properties. You will look at the different ways that atoms of elements join together when they form compounds, and how the method of combination will determine the properties of the compound formed. You will develop your skills in writing word and symbol equations; and, as well as being able to use an equation to work out the products of a reaction, you will be able to calculate how much of the product can be made in the reaction. These quantitative aspects of chemistry are crucially important in the chemical industry.

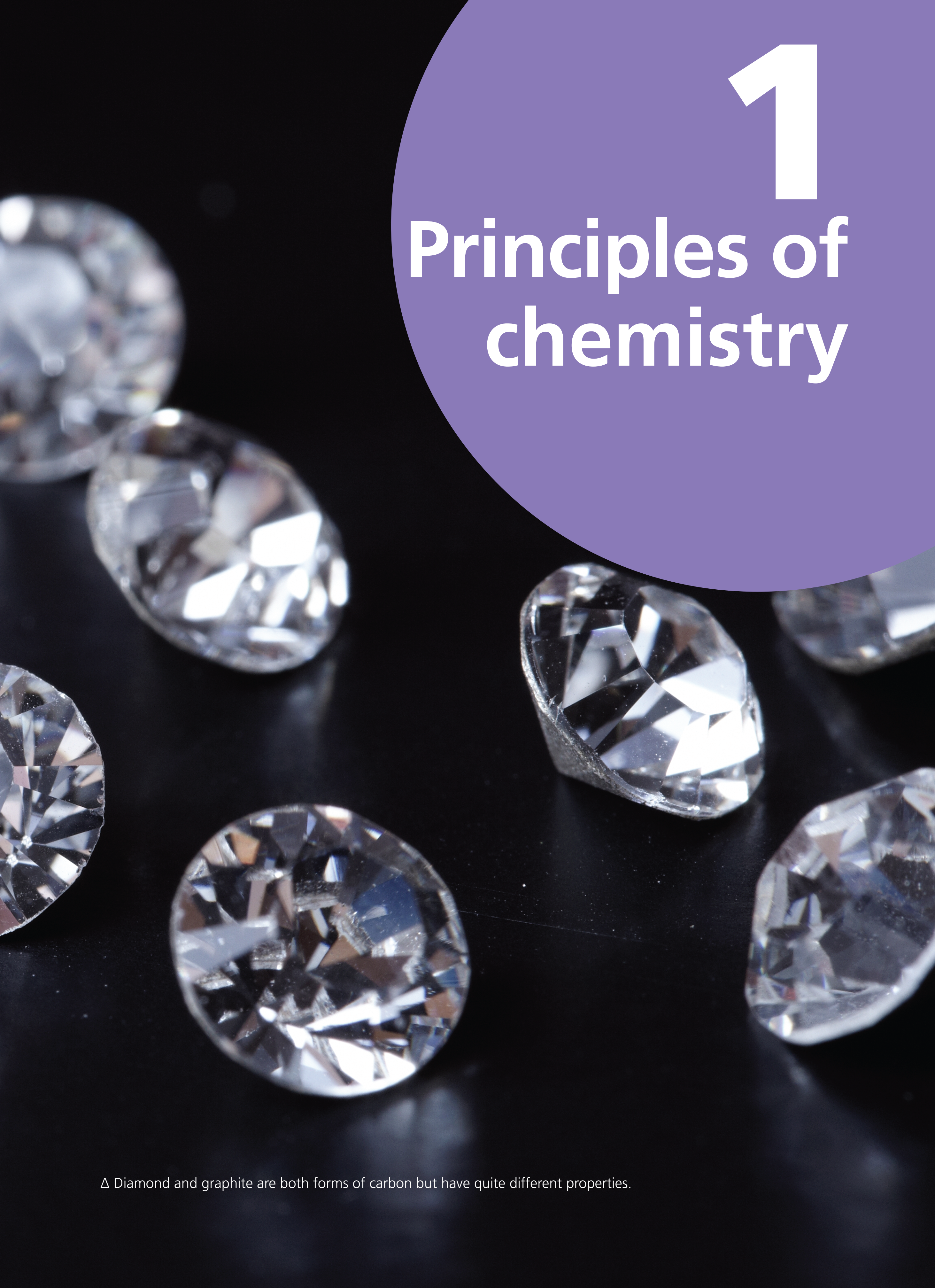
STARTING POINTS

1. What is an atom?
2. Name some of the particles that are found in an atom.
3. What name is given to a particle formed when two atoms combine together?
4. You will be learning about the states of matter. Do you know what these states are?
5. One type of chemical bonding you will study is called ionic bonding. Find out what an ion is.
6. Diamond and graphite are both covalent substances. They contain the same atoms but have very different structures and properties. What do you know about what diamond and graphite are used for?

SECTION CONTENTS

a) The particulate nature of matter
b) Experimental techniques
c) Atomic structure and the Periodic Table
d) Ions and ionic bonds
e) Molecules and covalent bonds
f) Metallic bonding
g) Stoichiometry
h) Exam-style questions

1 Principles of chemistry



Δ Diamond and graphite are both forms of carbon but have quite different properties.

Δ Fig. 1.1 Water in all its states of matter.

The particulate nature of matter

INTRODUCTION

Nearly all substances may be classified as solid, liquid or gas – the states of matter. In science these states are often shown in shorthand as (s), (l) and (g) after the formula or symbol (these are called **state symbols**). The kinetic particle theory is based on the idea that all substances are made up of extremely tiny particles. The particles in these three states are arranged differently and have different types of movement and different energies. In many cases, matter changes into different states quite easily. The names of many of these processes are in everyday use, such as melting and condensing. Using simple models of the particles in solids, liquids and gases can help to explain what happens when a substance changes state.

KNOWLEDGE CHECK

✓ Be able to classify substances as solid, liquid or gas.
✓ Be familiar with some of the simple properties of solids, liquids and gases.
✓ Know that all substances are made up of particles.

LEARNING OBJECTIVES

✓ Be able to state the distinguishing properties of solids, liquids and gases.
✓ Be able to describe the structure of solids, liquids and gases in terms of particle separation, arrangement and types of motion.
✓ Be able to describe the changes of state in terms of melting, boiling, evaporation, freezing, condensation and sublimation.
✓ Be able to describe qualitatively the pressure and temperature of a gas in terms of the motion of its particles.
✓ Be able to show an understanding of the random motion of particles in a suspension (sometimes known as Brownian motion) as evidence for the kinetic particle (atoms, molecules or ions) model of matter.
✓ Be able to describe and explain diffusion.
✓ EXTENDED Be able to explain changes of state in terms of the kinetic theory.
✓ EXTENDED Be able to describe and explain Brownian motion in terms of random molecular bombardment.
✓ EXTENDED Be able to state evidence for Brownian motion.
✓ EXTENDED Be able to describe the dependence of rate of diffusion on molecular mass.

Δ Fig. 1.2 Water covers nearly four-fifths of the Earth's surface. In this photo you can see that all three states of matter can exist together: solid water (the ice) is floating in liquid water (the ocean), and the surrounding air contains water vapour (clouds).

HOW DO SOLIDS, LIQUIDS AND GASES DIFFER?

The three states of matter each have different properties, depending on how strongly the particles are held together.

- **Solids** have a fixed volume and shape.
- **Liquids** have a fixed volume but no definite shape. They take up the shape of the container in which they are held.
- **Gases** have no fixed volume or shape. They spread out to fill whatever container or space they are in.

Substances don't always exist in the same state; depending on the physical conditions, they change from one state to another (interconvert).

Some substances can exist in all three states in the natural world. A good example of this is water.

QUESTIONS

1. What is the state symbol for a liquid?
2. Which is the only state of matter that has a fixed shape?
3. In what ways does fine sand behave like a liquid?

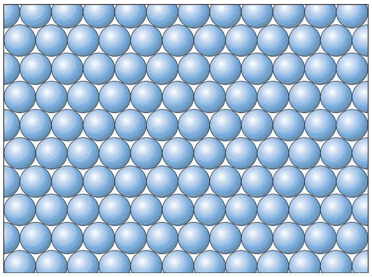
Why do solids, liquids and gases behave differently?

The behaviour of solids, liquids and gases can be explained if we think of all matter as being made up of very small particles that are in constant motion. This idea has been summarised in the **kinetic theory** of matter.

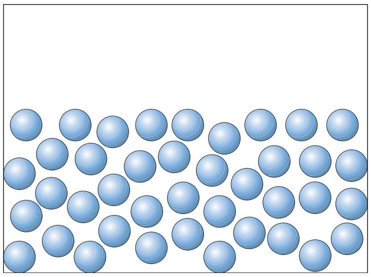
In solids, the particles are held tightly together in a fixed position, so solids have a definite shape. However, the particles are vibrating about their fixed positions because they have energy.

In liquids, the particles are held tightly together but have enough energy to move around. Liquids have no definite shape and will take on the shape of the container they are in.

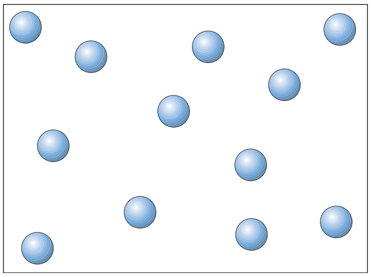
In gases, the particles are further apart with enough energy to move apart from each other and are constantly moving. Gas particles can spread apart to fill the container they are in.



Δ Fig. 1.3 Particles in a solid.

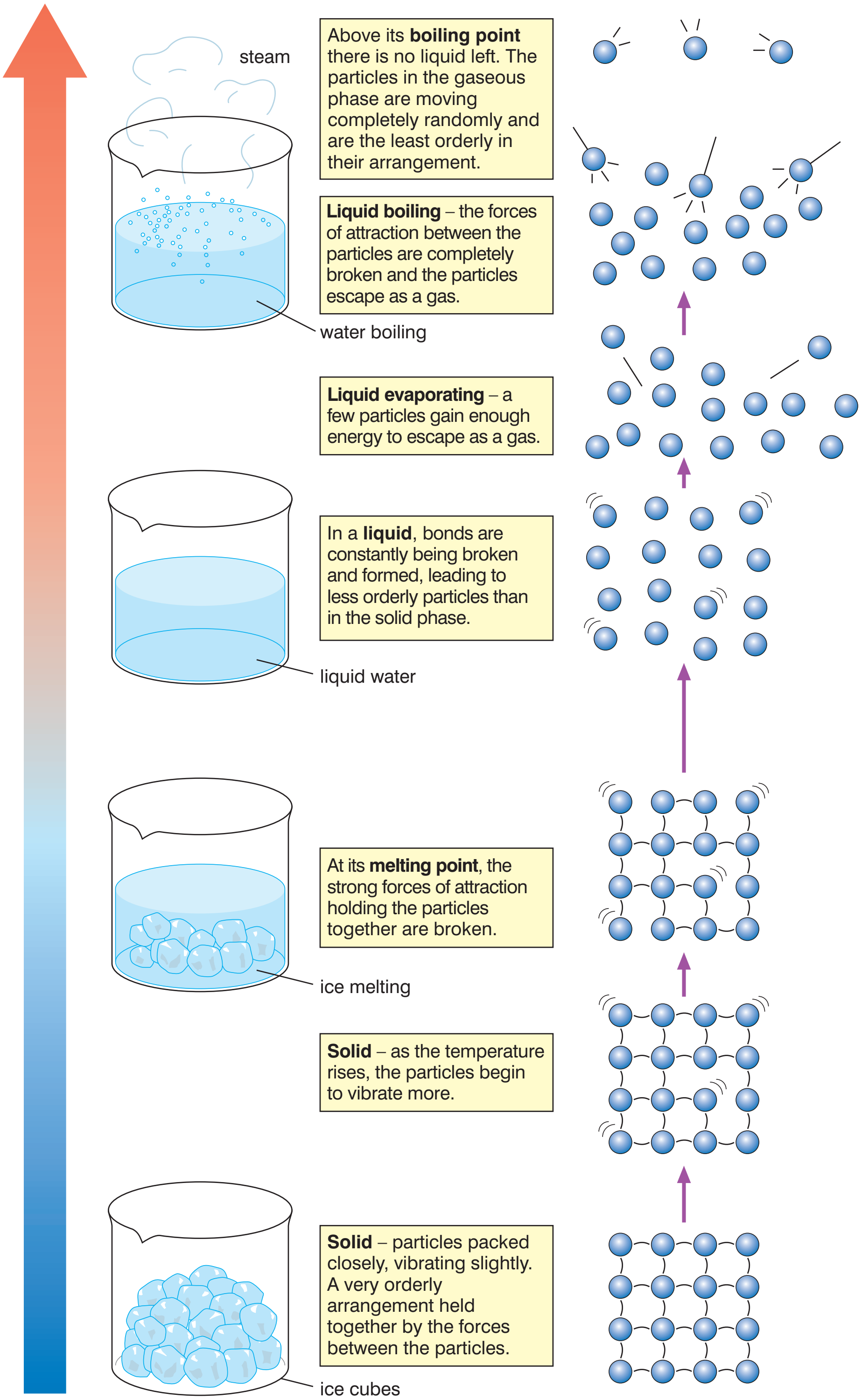


Δ Fig. 1.4 Particles in a liquid.



Δ Fig. 1.5 Particles in a gas.

EXTENDED



Δ Fig. 1.6 Particles in the different states of matter.

END OF EXTENDED

HOW DO SUBSTANCES CHANGE FROM ONE STATE TO ANOTHER?

To change solids into liquids and then into gases, heat energy must be put in. The heat provides the particles with enough energy to overcome the forces holding them together.

To change gases into liquids and then into solids involves cooling, so removing heat energy. This makes the particles come closer together as the substance changes from gas to liquid and the particles bond together as the liquid becomes a solid.

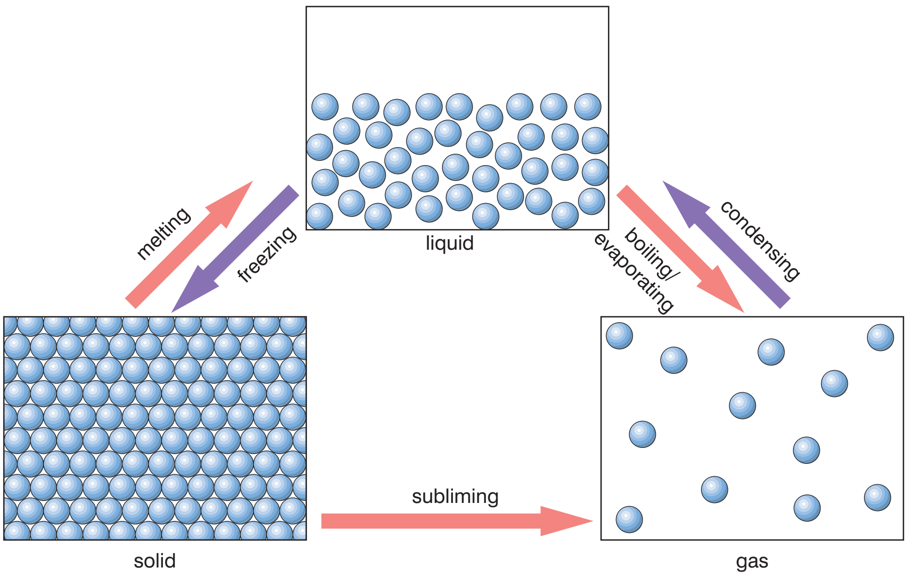
The temperatures at which one state changes to another have specific names:

Name of temperature	Change of state
Melting point	Solid to liquid
Boiling point	Liquid to gas
Freezing point	Liquid to solid
Condensation point	Gas to liquid

Δ Table 1.1 Changes of state.

The particles in a liquid can move around. They have different energies, so some are moving faster than others. The faster particles have enough energy to escape from the surface of the liquid and it changes into the gas state (also called **vapour** particles). This process is **evaporation**. The rate of evaporation increases with increasing temperature because heat gives more particles the energy to be able to escape from the surface.

Fig. 1.7 summarises the changes in states of matter. Note that melting and freezing happen at the same temperature – as do boiling and condensing.



Δ Fig. 1.7 Changes of state. Note that melting and freezing happen at the same temperature – as do boiling and condensing.

SCIENCE IN CONTEXT

THE STATES OF MATTER

There are three states of matter – or are there? To complicate this simple idea, some substances show the properties of two different states of matter. Some examples are given below.

Liquid crystals

Liquid crystals are commonly used in displays in computers and televisions. Within particular temperature ranges the particles of the liquid crystal can flow like a liquid, but remain arranged in a pattern in which the particles cannot rotate.

Δ Fig. 1.8 An LCD (liquid crystal display) television.

Superfluids

When some liquids are cooled to very low temperatures they form a second liquid state, described as a *superfluid* state. Liquid helium at just above absolute zero has infinite fluidity and will 'climb out' of its container when left undisturbed – at this temperature the liquid has zero viscosity. (You may like to look up 'fluidity' and 'viscosity'.)

Plasma

Plasmas, or ionised gases, can exist at temperatures of several thousand degrees Celsius. An example of a plasma is the charged air produced by lightning. Stars like our Sun also produce plasma. Like a gas, a plasma does not have a definite shape or volume but the strong forces between its particles give it unusual properties, such as conducting electricity. Because of this combination of properties, plasma is sometimes called the fourth state of matter.

The **pressure** exerted by a gas is caused by the gas particles bombarding the sides of the container the gas is in. If a gas is compressed into a smaller volume, the number of gas particles hitting the sides of the container every second will increase (there is less space for them to move in). The increase in the number of collisions per second causes the increase in pressure. If the **temperature** of a gas is increased, the gas particles have more energy and will move faster. Again there will be more collisions with the sides of the container each second and so the gas will exert a greater pressure.

QUESTIONS

1. What type of movement do the particles in a solid have?
2. In which state are the particles held together more strongly: in solid water, liquid water or water vapour?
3. What is the name of the process that occurs when the faster-moving particles in a liquid escape from its surface?
4. What name is given to the temperature at which a solid changes into a liquid?

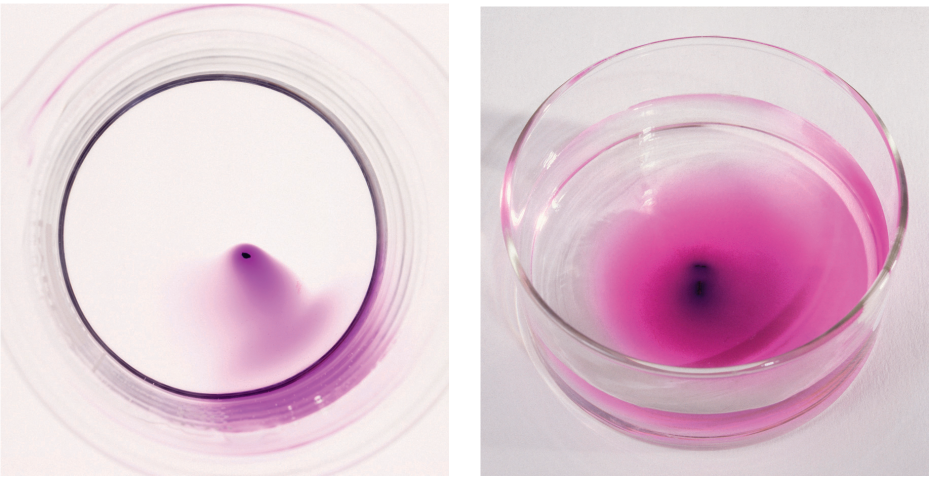
DIFFUSION EXPERIMENTS

Scientists have confidence in the kinetic theory because of the evidence from simple experiments.

The random mixing and moving of particles in liquids and gases is known as **diffusion**. The examples given below show the effects of diffusion.

Dissolving crystals in water

Fig. 1.9 shows purple crystals of potassium manganate(VII) dissolving in water.



Δ Fig. 1.9 Crystals of potassium manganate(VII) dissolving in water. The picture on the left shows the water immediately after the crystal was added; the picture on the right shows the water 1 hour later.

There are no water currents, so only the kinetic theory can explain this. The particles of the crystal gradually move into the water and mix with the water particles.

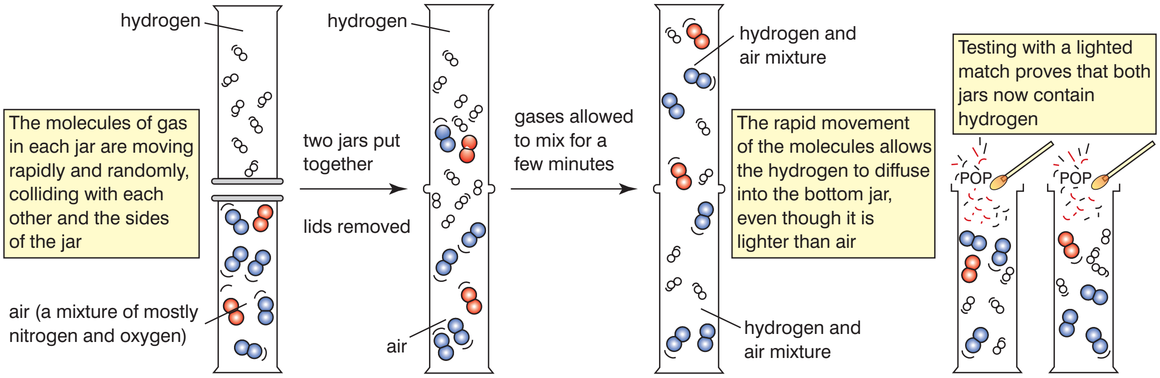
Mixing gases

These photos show a jar of air and a jar of bromine gas. Bromine gas is red-brown and heavier than air. The jar of air has been placed on top of the jar of bromine and the lids removed so the gases can mix (left-hand part of Fig.1.10).

After about 24 hours the bromine gas and the air have spread throughout both jars. Kinetic theory says that the particles of

Δ Fig. 1.10 Diffusion of bromine.

bromine gas can move around randomly so that they can fill both gas jars. This also occurs with hydrogen and air Fig. 1.11.



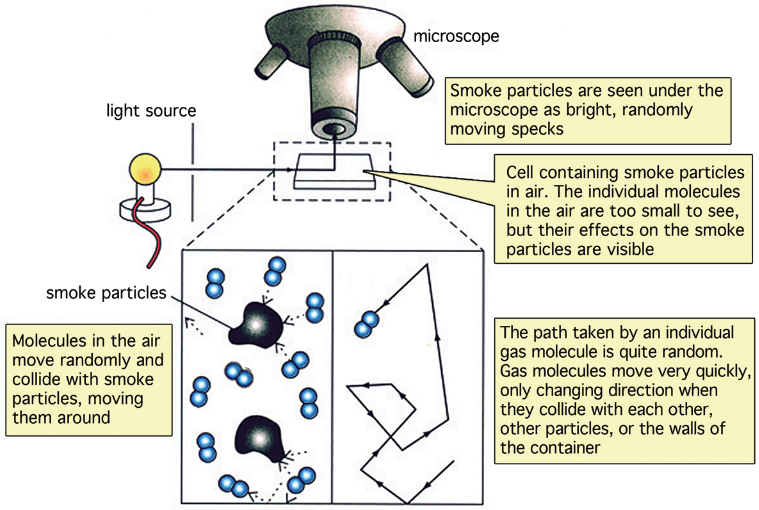
∆ Fig. 1.11 Demonstration of diffusion with a jar of oxygen and a jar of hydrogen.

BROWNIAN MOTION

The Scottish scientist Robert Brown looked at pollen in water using a microscope. He saw that the pollen grains were always moving.

EXTENDED

Kinetic theory explains this as the pollen grains being pushed around by water molecules that are always moving. You can also see **Brownian motion** in a smoke cell. The smoke particles move about because they are constantly pushed around by the tiny, invisible air molecules.



∆ Fig. 1.12 Brownian motion in a smoke cell.

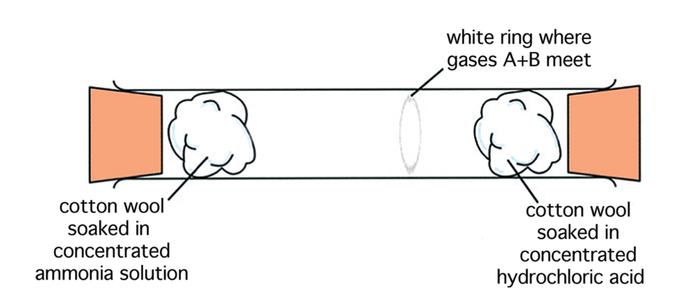
END OF EXTENDED

The molecules of different gases do not all move at the same speed at room temperature. The rate at which gases diffuse depends on their molecular mass – the larger their molecular mass, the lower their rate of diffusion. For example hydrogen, which has the lowest molecular mass of any gas, will diffuse much more rapidly than carbon dioxide which has a molecular mass 22 times that of hydrogen.

EXTENDED

Developing investigative skills

Two students set up the experiment shown in Fig. 1.13. They carefully clamped the long glass tube horizontally. At the same time, they inserted the cotton wool plugs soaked in the two solutions at each end of the tube and replaced the rubber bungs.



Δ Fig. 1.13 Results of experiment.

After about 15 minutes a white ring was seen in the tube.

Note: The white ring was formed where the ammonia gas from the concentrated ammonia solution met the hydrogen chloride gas from the concentrated hydrochloric acid. Together they formed a white substance called ammonium chloride.

Using and organising techniques, apparatus and materials

The concentrated ammonia solution is corrosive – it burns and is dangerous to the eyes. Concentrated hydrochloric acid is corrosive – it burns and its vapour irritates the lungs.

1. How should the cotton wool plugs have been handled when putting them into the tube?
2. What other safety precaution(s) should the two students have used?

Observing, measuring and recording

3. Which gas moved furthest in the 15 minutes before the ring formed?
4. Approximately how much further did this gas travel compared to the other gas?

Handling experimental observations and data

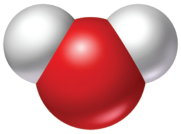
5. The rate of diffusion of a gas depends on the mass of its particles. What conclusion can you draw about the relative masses of the two gases in this experiment?

END OF EXTENDED

ELEMENTS, ATOMS AND MOLECULES

All matter is made from **elements**. Elements are substances that cannot be broken down into anything simpler, because they are made up of only one kind of the same small particle. These small particles are called **atoms**.

Almost always, the atoms in an element combine with other atoms to form molecules. For example, the particles in water are molecules containing two hydrogen atoms joined up with one oxygen atom. The formula is therefore H_2O. There are also particles called ions (see page 43).



Δ Fig. 1.14 Model of a water molecule.

QUESTIONS

1. What is *diffusion*?

2. Explain how the purple colour of the potassium manganate(VII) shown in Fig. 1.9 spreads through the water.

3. A bottle of perfume is broken at one end of a room. Explain why the perfume can soon be smelled all over the room.

4. What is the name of the particle that is found in all elements?

5. EXTENDED Brownian motion can be observed in a laboratory using a smoke cell:

a) In what situation was Brownian motion first observed?

b) In the smoke cell, what causes the random motion of the smoke particles?

End of topic checklist

Key terms

atom, boiling, Brownian motion, condensation, diffusion, element, evaporation, freezing, gas, kinetic theory, liquid, melting, pressure, solid, state symbols, temperature, vapour

During your study of this topic you should have learned:

- ❍ About the different properties of solids, liquids and gases.
- ❍ How to describe the structure of solids, liquids and gases in terms of particle separation, arrangement and types of motion.
- ❍ How to describe changes of state in terms of melting, boiling, evaporation, freezing, condensation and sublimation.
- ❍ How to describe the pressure and temperature of a gas in terms of the motion of its particles.
- ❍ An understanding of the random motion of particles in a suspension (Brownian motion) as evidence for the kinetic particle (atoms, molecules or ions) model of matter.
- ❍ How to describe and explain diffusion.
- ❍ EXTENDED How to explain changes of state in terms of the kinetic theory.
- ❍ EXTENDED How to describe and explain Brownian motion in terms of random molecular bombardment
  - about evidence for Brownian motion
  - how to describe how the rate of diffusion depends on molecular mass.

End of topic questions

Note: The marks awarded for these questions indicate the level of detail required in the answers. In the examination, the number of marks awarded to questions like these may be different.

1. In which of the three states of matter are the particles moving fastest? **(1 mark)**
2. Describe the arrangement and movement of the particles in a liquid. **(2 marks)**
3. In which state of matter do the particles just vibrate about a fixed point? **(1 mark)**
4. Sodium (melting point 98 °C) and aluminium (melting point 660 °C) are both solids at room temperature. From their melting points, what can you conclude about the forces of attraction between the particles in the two metals? **(1 mark)**
5. What is the name of the process involved in each of the following changes of state:

   a) $Fe(s) \rightarrow Fe(l)$? **(1 mark)**

   b) $H_2O(l) \rightarrow H_2O(g)$? **(1 mark)**

   c) $H_2O(g) \rightarrow H_2O(l)$? **(1 mark)**

   d) $H_2O(l) \rightarrow H_2O(s)$? **(1 mark)**
6. Ethanol liquid turns into ethanol vapour at 78 °C. What is the name of this temperature? **(1 mark)**
7. Explain how water in the Earth's polar regions can produce water vapour even when the temperature is very low. **(2 marks)**
8. A student wrote in her exercise book, 'The particle arrangement in a liquid is more like the arrangement in a solid than in a gas'. Do you agree with this statement? Explain your reasoning. **(2 marks)**
9. What word is used to describe the rapid mixing and moving of particles in gases? **(1 mark)**
10. Look at Fig. 1.10 showing gas jars of air and bromine. Explain why bromine gas fills the top gas jar even though it is denser than air. **(2 marks)**
11. EXTENDED **a)** What would you see if you observed illuminated smoke particles in a smoke cell? **(1 mark)**

    b) What causes what you observe? **(1 mark)**
12. EXTENDED The molecular masses of some gases are shown in the table:

    Which gas would diffuse at the greatest rate? Explain your answer. **(2 marks)**

Gas	Molecular mass
Oxygen	32
Nitrogen	28
Chlorine	71

Experimental techniques

INTRODUCTION

Practical work is a very important part of studying chemistry. In your practical work you will need to develop your skills so that you can safely, correctly and methodically use and organise techniques, apparatus and materials. This involves being able to use appropriate apparatus for measurement to give readings to the required degree of accuracy. It is important to be able to use techniques that will determine the purity of a substance and, if necessary, techniques that can be used to purify mixtures of substances.

Δ Fig. 1.15 Using the neutralisation method for a titration.

KNOWLEDGE CHECK

✓ Be familiar with some simple equipment for measuring time, temperature, mass and volume.
✓ Know that some substances are mixtures of a number of different components.

LEARNING OBJECTIVES

✓ Be able to name appropriate apparatus for accurate measurement of time, temperature, mass and volume.
✓ Be able to describe paper chromatography.
✓ Be able to interpret simple chromatograms.
✓ Be able to identify substances and assess their purity from the melting point and boiling point.
✓ Understand the importance of purity in substances in everyday life.
✓ **EXTENDED** Be able to use R_f values in interpreting simple chromatograms.
✓ **EXTENDED** Be able to outline how chromatography techniques can be applied to colourless substances by exposing chromatograms to substances called locating agents.
✓ Be able to describe methods of purification by the use of a suitable solvent, filtration, crystallisation and distillation.
✓ Be able to suggest suitable purification techniques, given information about the substances involved.

MEASUREMENT

In your study of chemistry you will carry out practical work. It is essential to use the right apparatus for the task.

Time is measured with clocks, such as a wall clock. The clock should be accurate to about 1 second. You may be able to use your own wristwatch or a stopclock.

Temperature is measured using a thermometer. The range of the thermometer is commonly –10 °C to +110 °C with intervals of 1 °C.

Mass is measured with a balance or scales.

Volume of liquids can be measured with burettes, pipettes and measuring cylinders.

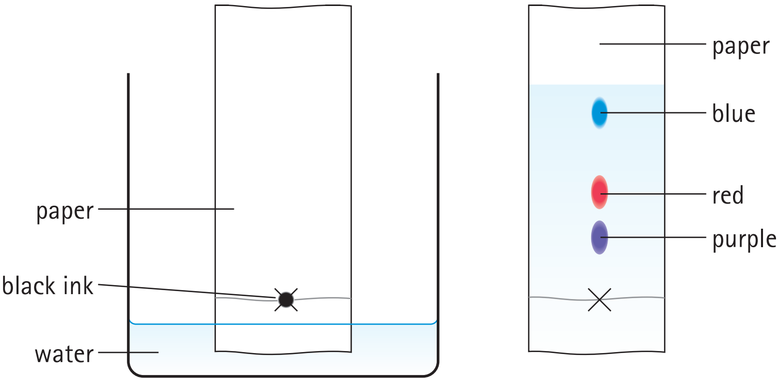
Δ Fig. 1.16 Measuring equipment.

CRITERIA OF PURITY

Paper chromatography

Paper **chromatography** is a way of separating solutions or liquids that are mixed together.

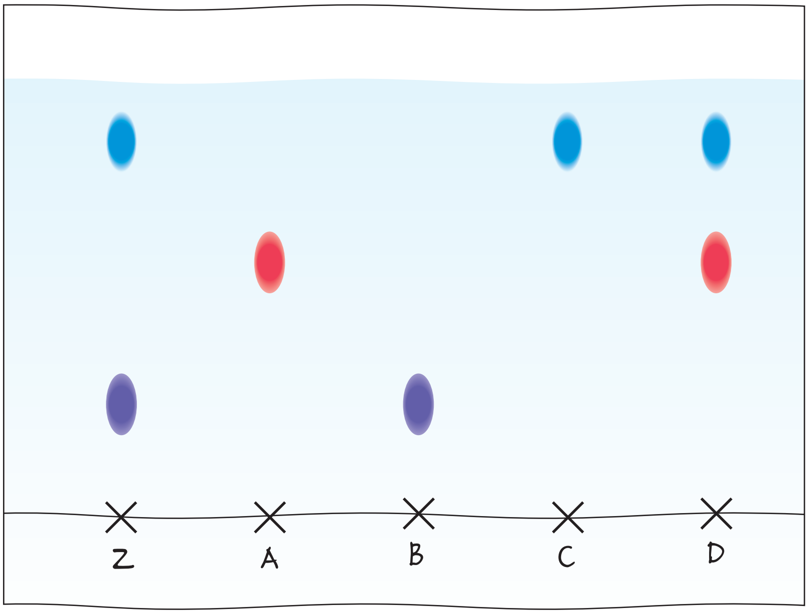
Black ink is a mixture of different coloured inks. The diagrams in Fig. 1.17 show how paper chromatography is used to find the colours that make up a black ink.



Δ Fig. 1.17 Paper chromatography separates a solution to find the colours in black ink. The left part of the diagram shows the paper before the inks have been separated, and the right part shows the paper after the inks have been separated.

A spot of ink is placed on the × mark and the paper is suspended in water. As the water rises up the paper, the different dyes travel different distances and so are separated on the **chromatogram**.

Paper chromatography can be used to identify what an unknown liquid is made of. This is called to interpreting a chromatogram.

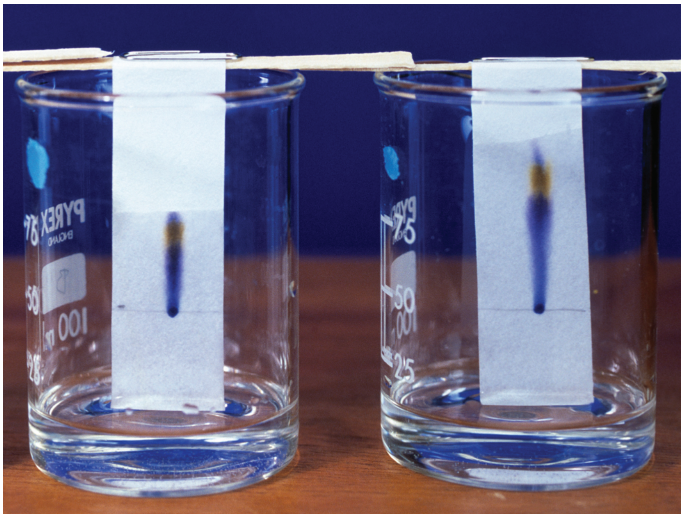


Δ Fig. 1.18 A chromatogram.

The unknown liquid Z is compared with known liquids – in this case A to D.

Z must be made of B and C because the pattern of their dots matches the pattern shown by Z.

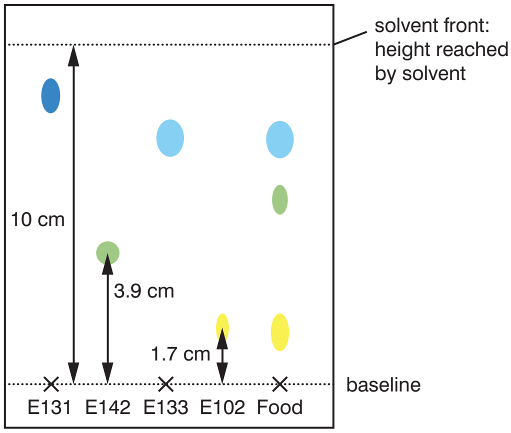
Δ Fig. 1.19 A piece of filter paper is marked with black ink and dipped into water in a beaker.



Δ Fig. 1.20 After a few minutes the chromatogram has been created by the action of the water on the ink.

EXTENDED

Even colourless liquids can be separated using this technique. The difference is that the chromatogram paper is still white/colourless at the end. To see the pattern of dots, the paper is viewed under ultraviolet light or sprayed with a liquid **locating agent** that shows up the dots as colours.



Δ Fig. 1.21 The R_f value for the food additive E102 is 0.17.

Retention factors

Substances can also be identified using chromatography by measuring their **retention factor** on the filter paper. The retention factor (R_f) for a particular substance compares the distance the substance has travelled up the filter paper with the distance travelled by the **solvent**. The retention factor can be calculated using the following formula:

$$R_f = \frac{\text{Distance moved by a substance from the baseline}}{\text{Distance moved by the solvent from the baseline}}$$

As the solvent will always travel further than the substance, R_f values will always be less than 1.

END OF EXTENDED

The purity of solids and liquids

It is very important that manufactured foods and drugs contain only the substances the manufacturers want in them – that is, they must not contain any contaminants.

The simplest way of checking the purity of solids and liquids is using heat to find the temperature at which they melt or boil.

An impure solid will have a lower melting point than the pure solid.

A liquid containing a dissolved solid (solute) will have a higher boiling point than the pure solvent.

The best examples to use to remember these facts are water and ice:

- Pure water boils at 100 °C – salted water for cooking vegetables boils at about 102 °C.
- Pure ice melts at 0 °C – ice with salt added to it melts at about –4 °C.

QUESTIONS

1. The start line, or baseline, in chromatography should be drawn in pencil. Explain why.
2. In a chromatography experiment, why must the solvent level in the beaker be below the baseline?

3. In a chromatography experiment to compare the dyes in two different inks, one of the inks does not move at all from the baseline. Suggest a reason for this.

4. A sample of water contains some dissolved impurities. What would you expect the boiling point of the sample to be?

5. EXTENDED Look at the diagram in Fig. 1.21. Explain why the retention factor for the food additive E102 is 0.17.

METHODS OF PURIFICATION

Techniques for purifying solids and liquids rely on finding different properties of the substances that make up the impure mixture.

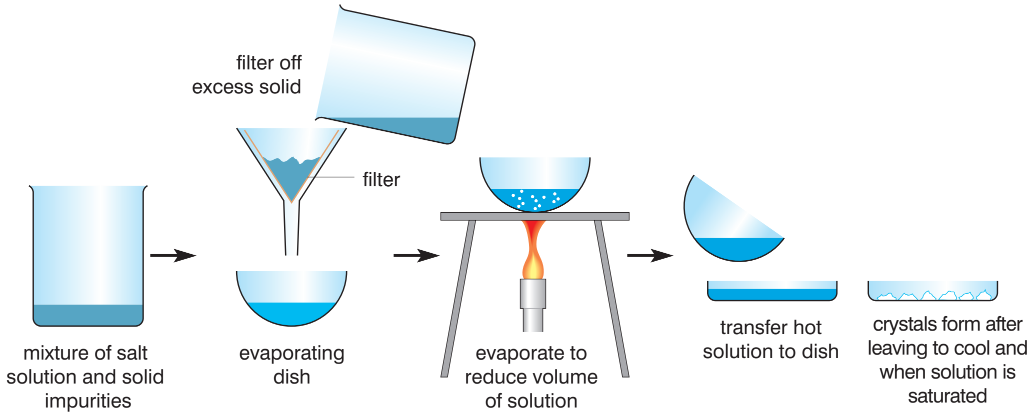
∆ Fig. 1.22 Filtration of copper (II) hydroxide.

Purifying impure solids

The method is:

1. Add a solvent that the required solid is **soluble** in, and dissolve it.
2. Filter the mixture to remove the insoluble impurity.
3. Heat the solution to remove some solvent and leave it to crystallise.
4. Filter off the crystals, wash with a small amount of cold solvent and dry them – this is the pure solid.

An example of using this technique would be separating salt from'rock salt' (the impure form of sodium chloride). Water is added to dissolve the salt but leave the other solids undissolved. Filter off the insoluble impurities, warm the salt solution and leave it to crystallise to form salt crystals.



∆ Fig. 1.23 Separating impurities in rock salt.

Purifying impure liquids

There are two methods:

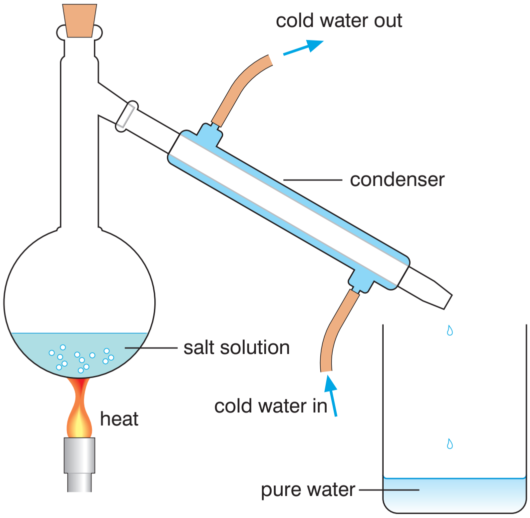
1. Liquids contaminated with soluble solids dissolved in them.

The method is **distillation**.

The solution is heated, the solvent boils and turns into a vapour. It is condensed back to the pure liquid and collected.

This is the technique used in **desalination** plants, which produce pure drinking water from sea water. The solids are left behind after boiling off the water.

Δ Fig.1.24 Distillation apparatus.



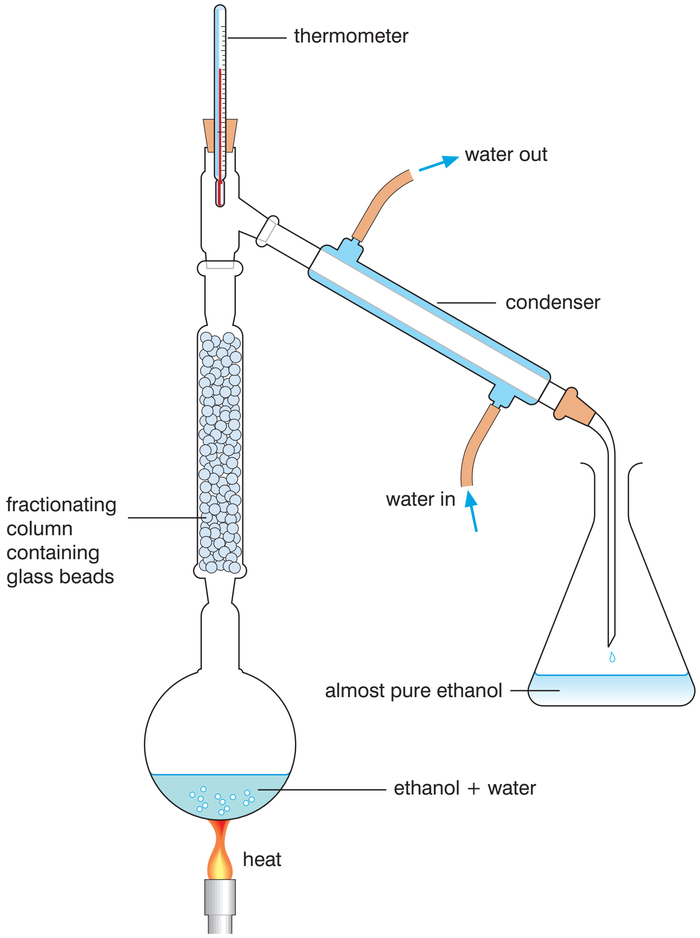
Δ Fig. 1.25 Distillation of salt water.

2. Liquids contaminated with other liquids.

In this case the technique is **fractional distillation**, which uses the difference in boiling points of the different liquids mixed together.

The mixture is boiled, and the liquid with the lowest boiling point turns to a vapour first, rises up the fractionating column and is condensed back to liquid in the condenser. The next lowest boiling point liquid comes off, and so on until all the liquids have been separated. You can identify the fraction you want to collect by the temperature reading on the thermometer. The fractionating column increases the purity of the distilled product by reducing the amount of other substances in the vapour when it condenses.

Fractional distillation is the method used in the separation of crude oil and collecting ethanol from the fermentation mixture.



Δ Fig. 1.26 Apparatus for fractional distillation of an alcohol/water mixture.

QUESTIONS

1. What is a *solvent*?
2. What does the term *soluble* mean?
3. What method would you use to separate a pure liquid from a solution of a solid and the liquid?
4. To separate two liquids by fractional distillation they must have different:
   a) melting points
   b) boiling points
   c) colours
   d) viscosities.

End of topic checklist

Key terms

chromatogram, desalination, distillation, fractional distillation, locating agent, paper chromatography, retention factor, soluble, solvent

During your study of this topic you should have learned:

❍ About the appropriate apparatus for the measurement of time, temperature, mass and volume, including burettes, pipettes and measuring cylinders.

❍ About the technique of paper chromatography.

❍ How to interpret simple chromatograms.

❍ How to identify substances and assess their purity from melting point and boiling point information.

❍ About the importance of purity in substances in everyday life.

❍ **EXTENDED** How to interpret simple chromatograms, including the use of R_f values.

❍ **EXTENDED** How to outline how chromatography techniques can be applied to colourless substances by exposing chromatograms to substances called locating agents.

❍ How to describe methods of purification by the use of:

- a suitable solvent – to separate a soluble solid from an insoluble solid
- filtration – to separate a solid from a liquid
- crystallisation – to separate a solid from its solution
- distillation – to separate a solid and a liquid from a solution
- fractional distillation – to separate liquids with different boiling points.

❍ How to suggest suitable purification techniques given information about the substances involved.

End of topic questions

Note: The marks awarded for these questions indicate the level of detail required in the answers. In the examination, the number of marks awarded to questions like these may be different.

1. You are provided with four samples of black water-soluble inks. Two of the ink samples are identical. Describe how you would use paper chromatography to identify which two ink samples are the same. **(3 marks)**

2. You are trying to separate the dyes in a sample of ink using paper chromatography. You set up the apparatus as shown in Fig. 1.19. After 20 minutes the black spot is unchanged and the water has risen nearly to the top of the filter paper.

a) Suggest a reason why the black spot has remained unchanged. **(1 mark)**

b) What could you change that might lead to a successful separation of the dyes? **(1 mark)**

3. Pure ethanol has a boiling point at normal pressure of 78 °C. What temperature might a sample of ethanol contaminated with sugar boil at? **(1 mark)**

4. What effect will impurities have on the melting point of ice? **(1 mark)**

5. In the fractional distillation of ethanol and water, why does the ethanol vapour condense in the condenser? **(1 mark)**

6. Describe how you would produce crystals of sodium chloride from a sodium chloride solution. **(2 marks)**

7. What process could be used to separate the following mixtures:

a) sand from a sand/water mixture? **(1 mark)**

b) petrol from a petrol/diesel mixture? **(1 mark)**

c) pure water from salt solution? **(1 mark)**

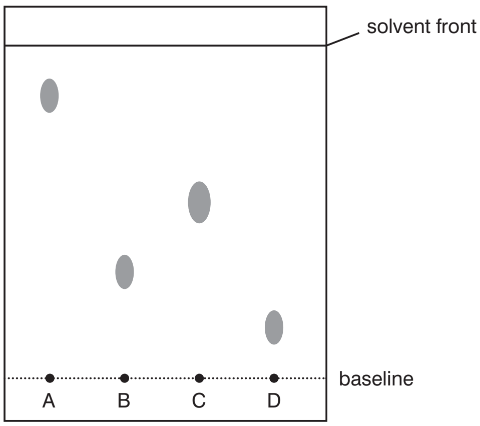
8. EXTENDED Look at the chromatogram produced when testing four food colouring compounds A, B, C and D.

a) Which compound has the largest retention factor (R_f)? **(1 mark)**

b) Which compound has the smallest R_f? **(1 mark)**

c) Estimate the R_f for compound C. Explain how you made the estimate. **(2 marks)**

d) Why are all R_f values less than 1.0? **(1 mark)**



Δ Fig. 1.27

End of topic questions continued

6. Magnesium burns in oxygen to form magnesium oxide:

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

(A_r: O = 16; Mg = 24)

Calculate:

a) the mass of magnesium required to make 8 g of magnesium oxide **(3 marks)**

b) the mass of oxygen required to make 8 g of magnesium oxide. **(1 mark)**

7. What mass of sodium hydroxide can be made by reacting 2.3 g of sodium with water? **(3 marks)**

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

(A_r: H = 1, O = 16, Na = 23)

8. EXTENDED How many moles are in the following?

a) 64 g of S_8 **(1 mark)**

b) 9.8 g of H_2SO_4 **(1 mark)**

c) 21 g of Li **(1 mark)**

(A_r: S = 32; H = 1; O = 16; Li = 7)

9. EXTENDED What is the mass of the following?

a) 2.5 moles of Sr **(1 mark)**

b) 0.25 moles of MgO **(1 mark)**

c) 0.1 moles of C_2H_5Br **(1 mark)**

(Sr = 88; Mg = 24; O = 16; C = 12; H = 1; Br = 80)

10. EXTENDED How many moles are in the following?

a) 24 000 cm^3 of hydrogen gas, measured at room temperature and pressure. **(1 mark)**

b) 1200 cm^3 of nitrogen gas measured at room temperature and pressure. **(1 mark)**

11. EXTENDED 0.64 g of copper when heated in air forms 0.80 g of copper oxide. What is the simplest formula of copper oxide? **(2 marks)**

(A_r: O = 16; Cu = 64)

12. **EXTENDED** Calculate the simplest formulae of the compounds formed in the following reactions:

a) 2.3 g of sodium reacting with 8.0 g of bromine **(2 marks)**

b) 0.6 g of carbon reacting with oxygen to make 2.2 g of a compound **(2 marks)**

c) 11.12 g of iron reacting with chlorine to make 32.20 g of a compound. **(2 marks)**

(A_r: C = 12; O = 16; Na = 23; Cl = 35.5; Fe = 56; Br = 80)

13. **EXTENDED** Titanium chloride contains 25% titanium and 75% chlorine by mass. Work out the simplest formula of titanium chloride. (A_r: Ti = 48, Cl = 35.5) **(3 marks)**

14. **EXTENDED** Ethene has an empirical formula of CH_2 and a relative formula mass of 28. What is the molecular formula of ethene? **(3 marks)**

15. **EXTENDED** A hydrocarbon contains 92.3% carbon and 7.7% hydrogen.

a) What is its empirical formula? **(2 marks)**

b) Its relative formula mass is 26. What is its molecular formula? **(2 marks)**

16. **EXTENDED** What mass of barium sulfate can be produced from 50 cm^3 of 0.2 mol/dm^3 barium chloride solution and excess sodium sulfate solution? **(3 marks)**

(A_r: O = 16, S = 32, Ba = 137)

$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

17. **EXTENDED** Iron(III) oxide is reduced to iron by carbon monoxide.

(A_r: C = 12, O = 16, Fe = 56)

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

a) Calculate the mass of iron that could be obtained by the reduction of 800 tonnes of iron(III) oxide. **(3 marks)**

b) What volume of carbon dioxide, measured at room temperature and pressure, would be obtained by the reduction of 320 g of iron(III) oxide? **(3 marks)**

18. **EXTENDED** In an experiment to make calcium oxide, the predicted yield was 2.8 g. The actual yield was 2.1 g. Calculate the percentage yield achieved. **(2 marks)**

19. **EXTENDED** An impure sample of solid X has a mass of 1.20 g. After purification the mass of pure X was 0.80 g. What was the percentage purity of the original sample? **(2 marks)**

20. **EXTENDED** Write ionic equations for the following reactions:

a) calcium ions and carbonate ions form calcium carbonate **(2 marks)**

b) iron(III) ions and hydroxide ions form iron(III) hydroxide **(2 marks)**

c) silver(I) ions and bromide ions form silver(I) bromide **(2 marks)**

Exam-style questions

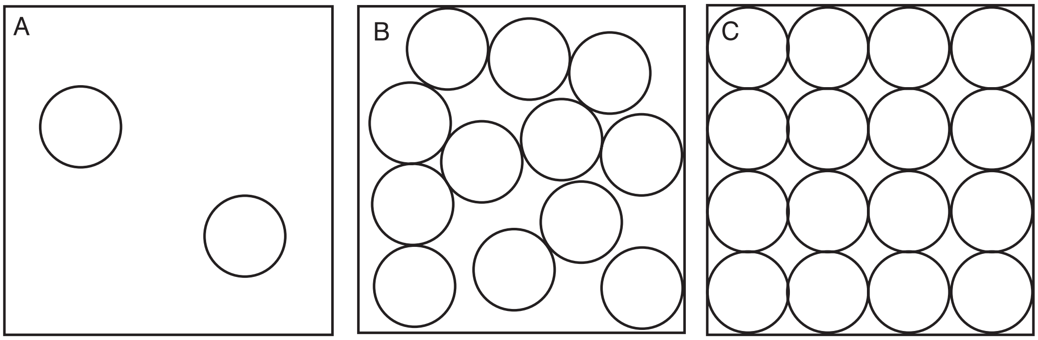
Sample student answer

Note: The questions, sample answers and marks in this section have been written by the authors as a guide only. The marks awarded for these questions indicate the level of detail required in the answers. In the examination, the number of marks awarded to questions like these may be different.

Question 1

a) The diagrams show the arrangement of particles in the three states of matter.

Each circle represents a particle.



Use the letters A, B, and C to give the starting and finishing states of matter for each of the changes in the table. For the mark, both the starting state and the finishing state need to be correct.

Change	Starting state	Finishing state	
i) The formation of water vapour from a puddle of water on a hot day	*B*	*A*	✓ 1
ii) The formation of solid iron from molten iron	*B*	*C*	✓ 1
iii) The manufacture of poly(ethene) from ethene	*B*	*A*	✗
iv) The reaction whose equation is ammonium hydrogen chloride(s) → ammonia(g) + hydrogen chloride(g)	*B*	*A*	✓ 1

(4)

TEACHER'S COMMENTS

a) It is important to identify the states of matter:

A = gas, B = liquid, C = solid.

i) Correct – evaporation process.

ii) Correct – solidifying.

iii) Incorrect – should be 'AC' order because ethene is a gas, poly(ethene) a solid.

iv) Correct – equation shows solid → gases (sublimation).

b) Answer is 'liquid'. In the Periodic Table, at room temperature the majority of elements are solids, a few are gases but only two are liquids – mercury and bromine.

c) **i)** Correct – sulfur.

ii) Incorrect – this is a 'mixture' of two elements.

iii) Correct – a mixture of an element (O_2) and a compound (H_2O).

iv) Correct – sulfuric acid.

The answers rely on using the state symbols for the equation and a thorough knowledge of the terms elements, mixtures and compounds.

b) Which state of matter is the *least* common for the elements of the Periodic Table at room temperature?

gases ✗ **(1)**

c) The manufacture of sulfuric acid can be summarised by the equation:

$$2S(s) + 3O_2(g) + 2H_2O(l) \rightarrow 2H_2SO_4(l)$$

Tick one box in each line to show whether the formulae in the table represents a compound, an element or a mixture.

	Compound	**Element**	**Mixture**	
i) $2S(s)$		✓		✓ 1
ii) $2S(s) + 3O_2(g)$		✓		✗
iii) $3O_2(g) + 2H_2O(l)$			✓	✓ 1
iv) $2H_2SO_4(l)$	✓			✓ 1

(4)

(Total 9 marks)

6/9

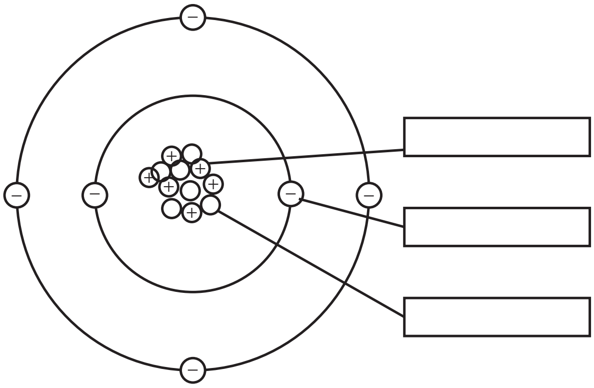
Exam-style questions continued

Question 2

This question is about atoms.

a) i) Choose words from the box to label the diagram of an atom. **(3)**

proton	neutron	electron	ion



ii) What is the proton number of this atom? **(1)**

iii) What is the nucleon number of this atom? **(1)**

b) Carbon has three isotopes. State one way in which the atoms of the three isotopes are:

i) the same **(1)**

ii) different. **(1)**

(Total 7 marks)

Question 3

a) Some elements combine together to form ionic compounds. Use words from the box to complete the sentences.

Each word may be used once, more than once or not at all.

gained	high	lost	low
medium	metals	non-metals	shared

Ionic compounds are formed between .... and ....

Electrons are .... by atoms of one element and .... by atoms of the other element.

The ionic compound formed has a .... melting point and a .... boiling point. **(6)**

b) Two elements react to form an ionic compound with the formula $MgCl_2$. (proton number of Mg = 12; proton number of Cl = 17)

i) Give the electronic configurations of the two elements in this compound *before* the reaction. **(2)**

ii) Give the electronic configurations of the two elements in this compound *after* the reaction. **(2)**

(Total 10 marks)

Question 4

9.12 g of iron(II) sulfate was heated. It decomposes to sulfur dioxide ($SO_2(g)$) and sulfur trioxide ($SO_3(g)$) and iron (III)oxide. Calculate the mass of iron(III) oxide formed and the volume of sulfur trioxide produced (measured at room temperature and pressure).

$2FeSO_4(s) \rightarrow Fe_2O_3(s) + SO_2(g) + SO_3(g)$

(A_r: O = 16, S = 32, Fe = 56; 1 mole of gas at room temperature and pressure occupies 24 000 cm^3)

(Total 6 marks)

Question 5

The following equation shows the reaction between potassium hydroxide solution and dilute sulfuric acid:

$$2KOH\ (aq) + H_2SO_4\ (aq) \rightarrow K_2SO_4\ (aq) + 2H_2O\ (l)$$

a) A 25.0 cm^3 sample of 0.15 mol/dm^3 potassium hydroxide solution was titrated with dilute sulfuric acid. It was found that 15.0 cm^3 of dilute sulfuric acid was needed to neutralise the potassium hydroxide solution.

b) Describe how you would carry out the titration experiment. You should include details of the apparatus you would use and how you would know when the potassium hydroxide had been neutralised. **(4)**

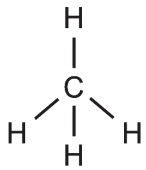
c) Use the equation and the experimental results to calculate the concentration of the sulfuric acid. **(4)**

(Total 8 marks)

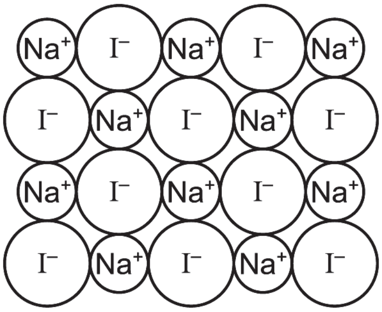
Exam-style questions continued

Question 6

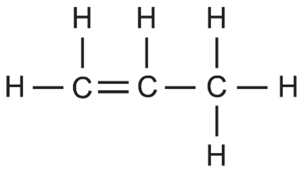
The structures of some substances are shown here:



A



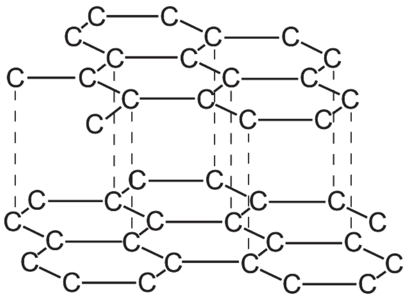
B



C

H — Br

D



E

a) Answer these questions using the letters **A**, **B**, **C**, **D** or **E**.

i) Which structure is methane? **(1)**

ii) Which two structures are giant structures? **(1)**

iii) Which two structures are hydrocarbons? **(1)**

iv) Which structure contains ions? **(1)**

v) Which two structures have very high melting points? **(1)**

b) Structure **E** is a form of carbon.

i) What is the name of this structure? Put a ring around the correct answer.

carbide graphite lead poly(hexene) **(1)**

ii) Name another form of carbon. **(1)**

c) Write the simplest formula for substance **B**. **(1)**

d) Is substance **D** an element or a compound? Explain your answer. **(3)**

(Total 11 marks)

Question 7

Strontium and sulfur chlorides both have a formula of the type XCl_2 but they have different properties.

Property	Strontium chloride	Sulfur chloride
Appearance	White crystalline solid	Red liquid
Melting point /°C	873	–80
Particles present	Ions	Molecules
Electrical conductivity of solid	Poor	Poor
Electrical conductivity of liquid	Good	Poor

a) The formulae of the chlorides are similar because both elements have a valency of 2.

Explain why Group II and Group VI elements both have a valency of 2. **(2)**

b) Draw a dot and cross diagram of one covalent molecule of sulfur chloride. Use x to represent an electron from a sulfur atom. Use o to represent an electron from a chlorine atom. **(3)**

c) Explain the difference in electrical conductivity between the following.

i) solid and liquid strontium chloride **(1)**

ii) liquid strontium chloride and liquid sulfur chloride. **(1)**

(Total 7 marks)