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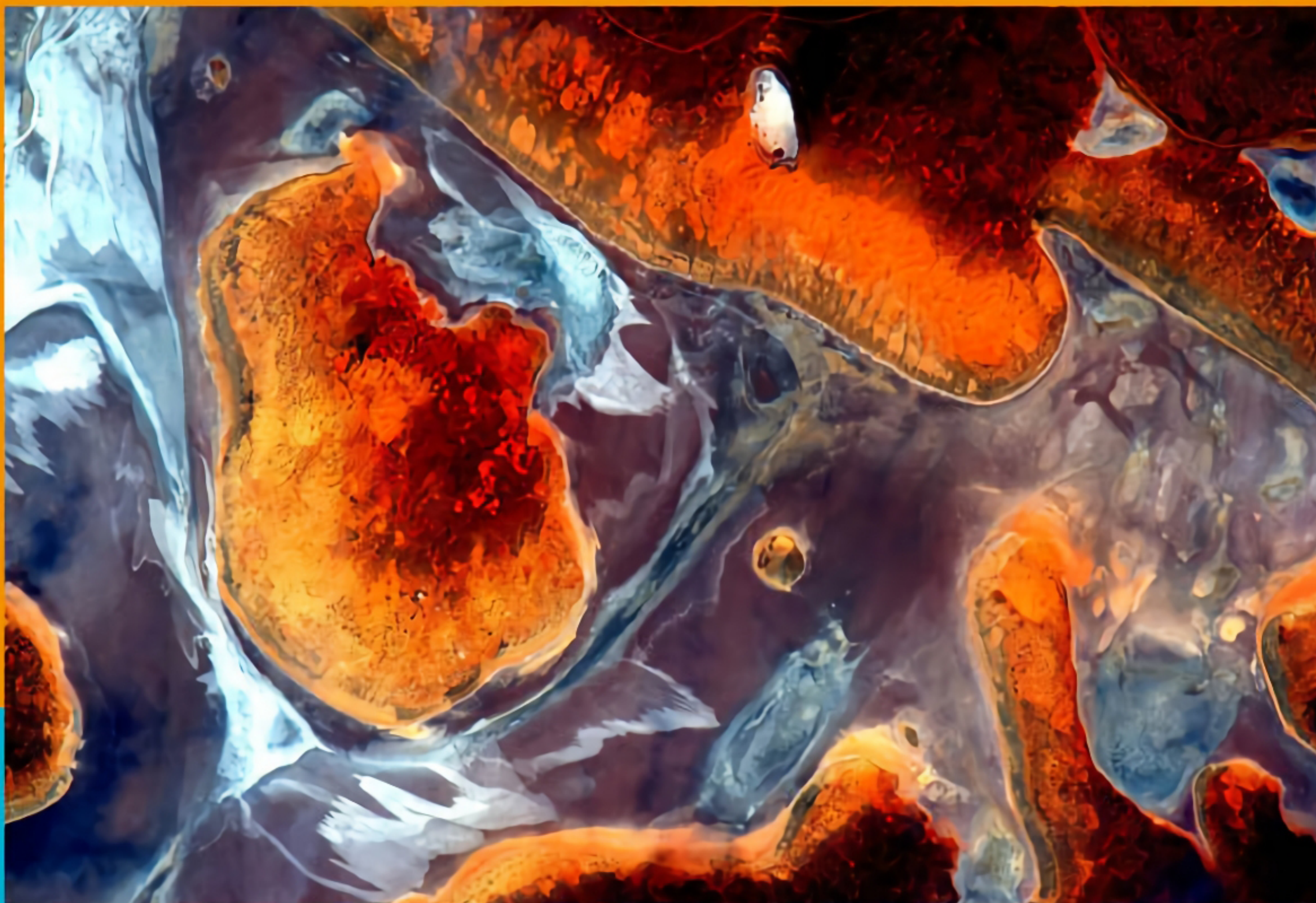

EDEXCEL INTERNATIONAL GCSE (9–1)

CHEMISTRY

Student Book

Jim Clark, Steve Owen, Rachel Yu

eBook included

PEARSON EDEXCEL INTERNATIONAL GCSE (9–1)

CHEMISTRY

Student Book

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Rachel Yu

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UNIT 1

PRINCIPLES OF CHEMISTRY

UNIT 2

INORGANIC CHEMISTRY

UNIT 3

PHYSICAL CHEMISTRY

UNIT 4

ORGANIC CHEMISTRY

ABOUT THIS BOOK

This book is written for students following the Pearson Edexcel International GCSE (9–1) Chemistry specification and the Edexcel International GCSE (9–1) Science Double Award specification. You will need to study all of the content in this book for your Chemistry examination. However, you will only need to study some of it if you are taking the Double Award specification. The book clearly indicates which content is in the Chemistry examination and not in the Double Award specification. To complete the Double Award course you will also need to study the Physics and Biology parts of the course.

In each unit of this book, there are concise explanations and worked examples, plus numerous exercises that will help you build up confidence. The book also describes the methods for carrying out all of the required practicals.

The language throughout this textbook is graded for speakers of English as an additional language (EAL), with advanced Chemistry-specific terminology highlighted and defined in the glossary at the back of the book. A list of command words, also at the back of the book, will help you to learn the language you will need in your examinations.

You will also find that questions in this book have Progression icons and Skills tags. The Progression icons refer to Pearson's Progression scale. This scale – from 1 to 12 – tells you what level you have reached in your learning and will help you to see what you need to do to progress to the next level. Furthermore, Edexcel have developed a Skills grid showing the skills you will practise throughout your time on the course. The skills in the grid have been matched to questions in this book to help you see which skills you are developing. Both Skills tags and Progression icons are not repeated where they are same in consecutive questions. You can find Pearson's Progression scale and Edexcel's Skills grid at www.pearsonglobalschools.com/igscienceprogression along with guidelines on how to use them.


160 INORGANIC CHEMISTRY EXTRACTION AND USES OF METALS

CHEMISTRY ONLY

15 EXTRACTION AND USES OF METALS

Metals are some of the most important materials that we use in everyday life. This chapter explores the principles behind the extraction of metals from their ores. We will also look at the properties of alloys and the uses of some metals.

LEARNING OBJECTIVES

- Know that most metals are extracted from ores found in the Earth's crust and that unreactive metals are often found as the uncombined element.
- Explain how the method of extraction of a metal is related to its position in the reactivity series, illustrated by carbon extraction for iron and electrolysis for aluminium.
- Be able to comment on a metal extraction process, given appropriate information *(detailed knowledge of the processes used in the extraction of a specific metal is not required).*
- Explain the uses of aluminium, copper, iron and steel in terms of their properties *(the types of steel will be limited to low-carbon (mild), high-carbon and stainless).*
- Know that an alloy is a mixture of a metal and one or more elements, usually other metals or carbon.
- Explain why alloys are harder than pure metals.

EXTRACTING METALS FROM THEIR ORES

MINERALS AND ORES

Most metals are found in the Earth's crust combined with other elements. The individual compounds are called minerals.

▲ Figure 15.1 Pyrite (iron pyrites), FeS_2

▲ Figure 15.2 Magnetite, Fe_3O_4

Figures 15.1 and 15.2 show samples of some iron-containing minerals; they are normally found mixed with other unwanted minerals in rocks. An ore is a sample of rock that contains enough of a mineral for it to be worthwhile to extract the metal. Most metals are extracted from ores found in the Earth's crust.

A few very unreactive metals, such as gold, are found native. That means that they exist naturally as the uncombined element. Silver and copper are also sometimes found native, although much more rarely.

DID YOU KNOW?
The price of a metal is affected by how common the ore is and how difficult it is to extract the metal from the ore.

HINT
You do not need to remember the names of these ores/minerals for the exam.

EXTRACTING THE METAL

Many ores contain either oxides or compounds that are easily converted to oxides. Sulfides such as sphalerite (zinc blende), ZnS, can be easily converted into an oxide by heating in air, a process known as roasting.

$$2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$$

Chemistry Only features show the content that is on the Chemistry specification only and not the Double Award specification. All other content in this book applies to Double Award students.

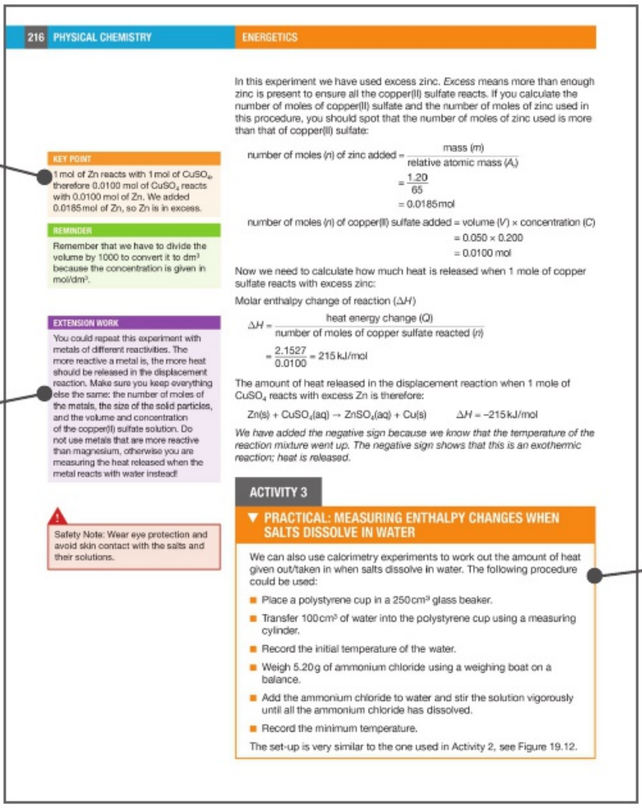
Learning Objectives show what you will learn in each chapter.

Did you know? Interesting facts to encourage wider thought and understanding around course texts.

Hint boxes give you tips on important points to remember in your examination.

Key Points boxes summarise the essentials.

Extension boxes include content which is not on the specification and which you do not have to learn for your examination. However, it will help to extend your understanding of the topic.

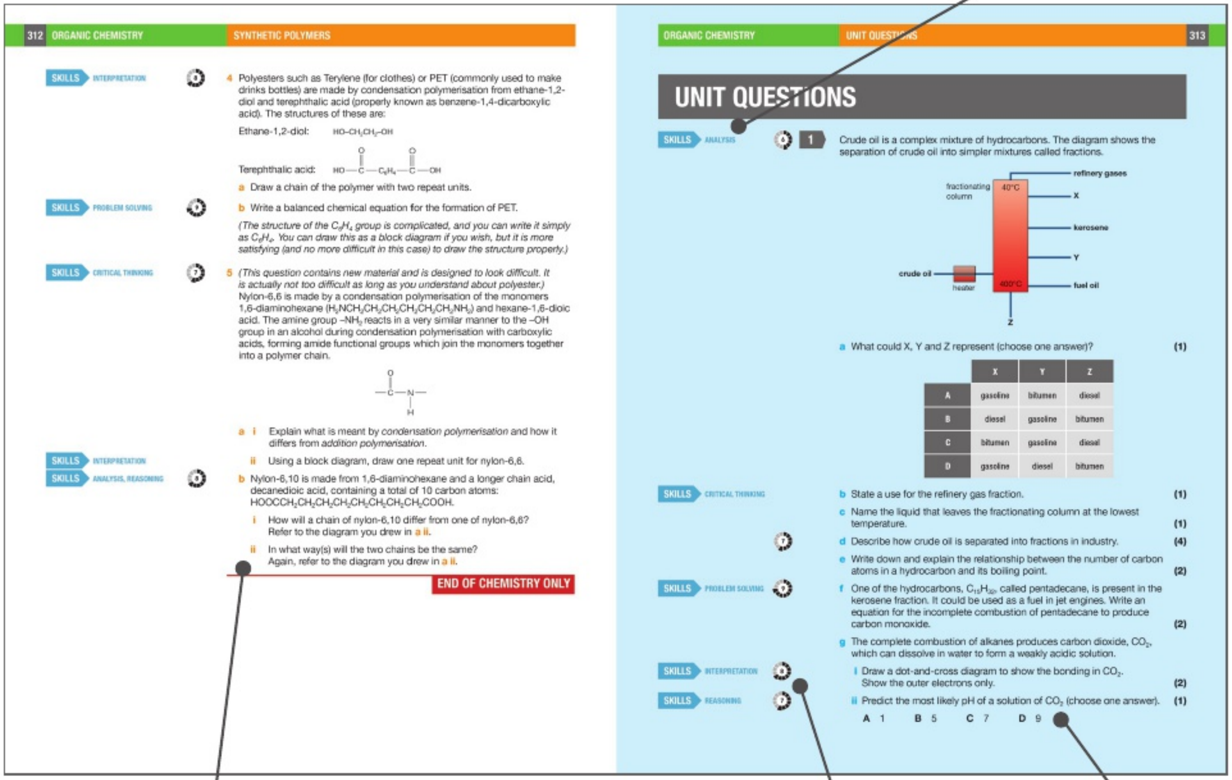


Looking Ahead tells you what you would learn if you continued your study of Chemistry to a higher level, such as International A Level.

Examples provide a clear, instructional framework.

Practicals describe the methods for carrying out all of the practicals you will need to know for your examination.

Transferable Skills are highlighted to show what skill you are using and where.



Chapter Questions test your knowledge of the topic in that chapter.

Progression icons show the level of difficulty according to the Pearson International GCSE Science Progression Scale.

Unit Questions test your knowledge of the whole unit and provide quick, effective feedback on your progress.

ASSESSMENT OVERVIEW

The following tables give an overview of the assessment for this course.

We recommend that you study this information closely to help ensure that you are fully prepared for this course and know exactly what to expect in the assessment.

PAPER 1	SPECIFICATION	PERCENTAGE	MARK	TIME	AVAILABILITY
Written examination paper Paper code 4CH1/1C and 4SD0/1C Externally set and assessed by Edexcel	Chemistry Science Double Award	61.1%	110	2 hours	January and June examination series First assessment June 2019
PAPER 2	**SPECIFICATION**	**PERCENTAGE**	**MARK**	**TIME**	**AVAILABILITY**
Written examination paper Paper code 4CH1/2C Externally set and assessed by Edexcel	Chemistry	38.9%	70	1 hour 15 mins	January and June examination series First assessment June 2019

If you are studying Chemistry then you will take both Papers 1 and 2. If you are studying Science Double Award then you will only need to take Paper 1 (along with Paper 1 for each of the Physics and Biology courses).

ASSESSMENT OBJECTIVES AND WEIGHTINGS

ASSESSMENT OBJECTIVE	DESCRIPTION	% IN INTERNATIONAL GCSE
AO1	Knowledge and understanding of chemistry	38%–42%
AO2	Application of knowledge and understanding, analysis and evaluation of chemistry	38%–42%
AO3	Experimental skills, analysis and evaluation of data and methods in chemistry	19%–21%

EXPERIMENTAL SKILLS

In the assessment of experimental skills, students may be tested on their ability to:

- solve problems set in a practical context
- apply scientific knowledge and understanding in questions with a practical context
- devise and plan investigations, using scientific knowledge and understanding when selecting appropriate techniques
- demonstrate or describe appropriate experimental and investigative methods, including safe and skilful practical techniques
- make observations and measurements with appropriate precision, record these methodically and present them in appropriate ways
- identify independent, dependent and control variables
- use scientific knowledge and understanding to analyse and interpret data to draw conclusions from experimental activities that are consistent with the evidence
- communicate the findings from experimental activities, using appropriate technical language, relevant calculations and graphs
- assess the reliability of an experimental activity
- evaluate data and methods, taking into account factors that affect accuracy and validity.

CALCULATORS

Students are permitted to take a suitable calculator into the examinations. Calculators with QWERTY keyboards or that can retrieve text or formulae will not be permitted.

UNIT 1 PRINCIPLES OF CHEMISTRY

The universe is made of three things!

Up to the present day scientists have discovered 118 elements. Most of these have been made naturally in stars but some are made artificially. As far as we know these are the only elements in the universe, so we basically have a model kit containing 118 different atoms. Chemistry can be described as the study of how these different atoms are joined together in various ways to make everything around us, from a tree, to a person, to the tallest skyscraper. Many of these elements are not very common so most of the things we see around us are made up of different combinations of only about a quarter of these elements. What makes this even more amazing is that each atom is made up of just three subatomic particles, which are called protons, neutrons and electrons. So, the world around us is made of only three things arranged in different ways.

▲ Figure 1.1 Southern view of the Milky Way

1 STATES OF MATTER

Everything around us is made of particles that we can't see because they are so small. This chapter looks at the arrangement of particles in solids, liquids and gases, and the ways in which the particles can move around. The nature of the different sorts of particles will be explored in Chapter 3.

▲ Figure 1.2 Everything you look at is a solid, a liquid or a gas . . .

▲ Figure 1.3 . . . metals, concrete, water, air, clouds – everything!

LEARNING OBJECTIVES

- Understand the three states of matter in terms of the arrangement, movement and energy of the particles
- Understand the interconversions between the three states of matter in terms of:
  - the names of the interconversions
  - how they are achieved
  - the changes in arrangement, movement and energy of the particles
- Understand how the results of experiments involving the dilution of coloured solutions and diffusion of gases can be explained
- Know what is meant by the terms:
  - solvent
  - solute
  - solution
  - saturated solution

CHEMISTRY ONLY

- Know what is meant by the term solubility in the units g per 100 g of solvent.
- Understand how to plot and interpret solubility curves.
- Practical: Investigate the solubility of a solid in water at a specific temperature.

STATES OF MATTER

Solids, liquids and gases are known as the three states of matter.

THE ARRANGEMENT OF THE PARTICLES

Think about these facts:

- You can't walk through a brick wall, but you can move (with some resistance – it pushes against you) through water. Moving through air is easy.
- When you melt most solids their volume increases slightly. Most liquids are less dense than the solid they come from.
- If you boil about 5 cm^3 of water, the steam will fill an average bucket.

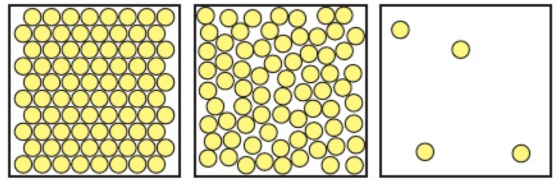
HINT

The packing in the solid might be completely different. What is important is that the particles are close together and, usually, regularly packed. When you draw a liquid, make sure the particles are mostly touching the particles next to them.

KEY POINTS

- You can't walk through a brick wall because of the strong forces of attraction between the particles – the particles can't move out of your way.
- You can swim through water because you can push the particles out of the way.
- It is easy to move through a gas because there are no forces between the particles.

The arrangement of the particles in solids, liquids and gases explains these facts.



▲ Figure 1.4 The arrangement of particles in different states of matter

In a solid, the particles are usually arranged regularly and packed closely together. The particles are only able to vibrate about fixed positions; they can't move around. The particles have strong forces of attraction between them, which keep them together.

In a liquid, the particles are still mostly touching, but some gaps have appeared. This is why liquids are usually less dense than solids. The forces between the particles are less effective, and the particles can move around each other. The particles in a liquid are arranged randomly.

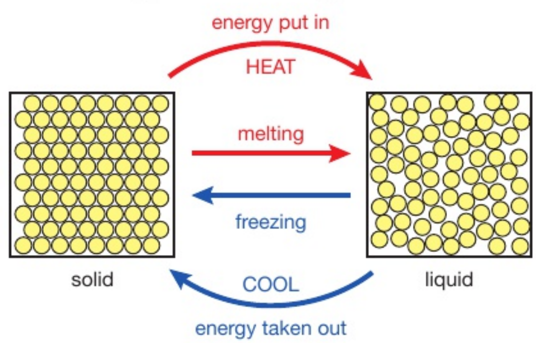
The particles in a gas are moving randomly at high speed in all directions. In a gas, the particles are much further apart and there are (almost) no forces of attraction between them.

The particles in a solid have less kinetic (movement) energy than the particles in a liquid, which have less kinetic energy than the particles in a gas.

INTERCONVERSIONS BETWEEN THE THREE STATES OF MATTER

CHANGING STATE BETWEEN SOLID AND LIQUID

If you heat a solid, the energy provided by the heat source makes the particles in the solid vibrate faster and faster. Eventually, they vibrate so fast that the forces of attraction between the particles are no longer strong enough to hold them together; the particles are then able to move around each other – the solid melts to form a liquid. The temperature at which the solid **melts** is called its **melting point**. The particles in the liquid have more kinetic energy than the particles in the solid so energy has to be supplied to convert a solid to a liquid.



▲ Figure 1.5 Melting to become a liquid – and freezing to become a solid.

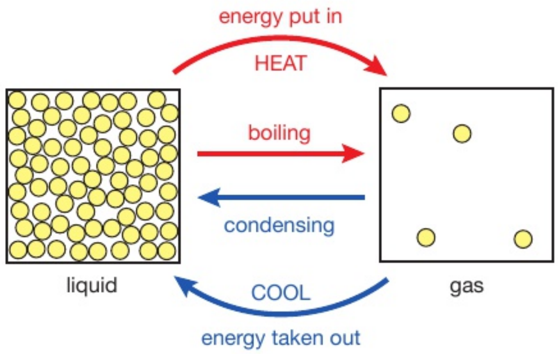
If the liquid is cooled again, the liquid particles will move around more and more slowly. Eventually, they are moving so slowly that the forces of attraction between them will hold them in a fixed position and the particles pack more closely together into a solid. The liquid **freezes**, forming a solid. The temperature at which this occurs is called the **freezing point**.

Although they are called different things depending which way you are going, the temperature of the melting point and that of the freezing point of a substance are exactly the same.

CHANGING STATE BETWEEN LIQUID AND GAS

There are two different ways this can happen, called **boiling** and **evaporation**.

BOILING



▲ Figure 1.6 Boiling to become a gas – and condensing to become a liquid.

Boiling occurs when a liquid is heated so strongly that the particles are moving fast enough to overcome all the forces of attraction between them. The stronger the forces of attraction between particles, the higher the boiling point of the liquid. This is because more energy is needed to overcome these forces of attraction.

If a gas is cooled, the particles eventually move slowly enough that forces of attraction between them start to form and hold them together as a liquid. The gas condenses.

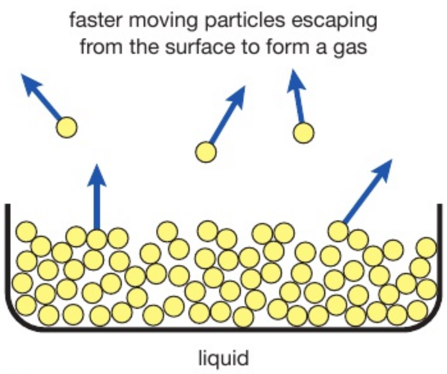
KEY POINT

Evaporation occurs at any temperature, but boiling only occurs at one temperature – the boiling point of the liquid. Puddles of water disappear quite quickly despite the outside temperature often being below 5 °C in the winter in the UK. The water in the puddles certainly does not boil at this temperature; the water evaporates. So water will evaporate at, for example, 5 °C but only boil at 100 °C.

EVAPORATION

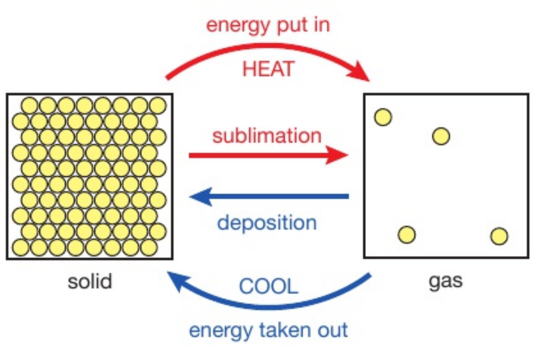
Evaporation is different. In any liquid or gas, the average speed of the particles varies with the temperature. But at each temperature, some particles will be moving faster than the average and others more slowly.

Some very fast particles at the surface of the liquid will have enough energy to overcome the forces of attraction between the particles – they will break away to form a gas. This is evaporation. You don't see any bubbling; the liquid just slowly disappears if it is open to the air. If the liquid is in a closed container, particles in the gas will also be colliding with particles at the surface of the liquid. If they are moving slowly enough they will be held by the attractive forces and become part of the liquid. In a closed container evaporation and condensation will both be occurring at the same time.



▲ Figure 1.7 Evaporation.

CHANGING STATE BETWEEN SOLID AND GAS: SUBLIMATION



▲ Figure 1.8 This change of state goes directly from a solid to a gas and from a gas to a solid.

A small number of substances can change directly from a solid to a gas, or from a gas to a solid, at normal pressure without involving any liquid in the process. The conversion of a solid into a gas is known as **sublimation** and the reverse process is usually called **deposition**.

KEY POINT

The process of a gas changing into a solid is given various names. Some people call it 'de-sublimation' or 'deposition' and others just use the word 'sublimation' again.

Figure 1.9 Dry ice subliming. Notice the white solid carbon dioxide in the beaker. The white cloud is because the carbon dioxide gas produced is so cold that it causes water vapour in the air to condense. Carbon dioxide gas itself is invisible.

An example of a substance that sublimes is carbon dioxide. At ordinary pressures, there is no such thing as liquid carbon dioxide – it turns directly from a solid to a gas at –78.5 °C. Solid carbon dioxide is known as dry ice.

WORKING OUT THE PHYSICAL STATE OF A SUBSTANCE AT A PARTICULAR TEMPERATURE

A substance is a solid at temperatures below its melting point, between its melting point and its boiling point it is a liquid, and above its boiling point it is a gas.

In science we can decide whether a substance is a solid, a liquid or a gas at room temperature by looking at where its melting and boiling points are in relation to room temperature.

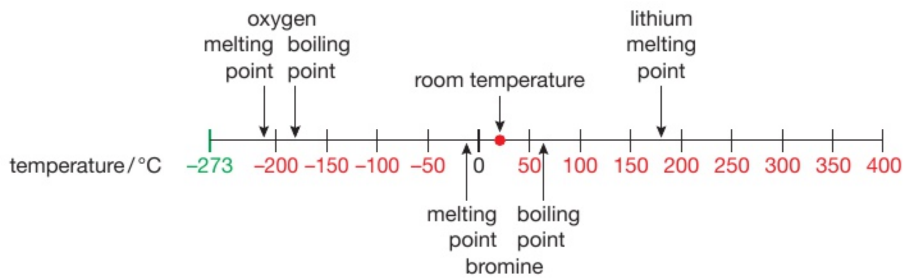


Figure 1.10 A temperature line can be used to work out whether substances are solids, liquids or gases.

KEY POINT

Room temperature is different in different places but in science it is usually taken to mean a temperature between 20 and 25 °C. Because there is not just one fixed value, for changes of state that occur near room temperature we must be careful when making comparisons and make clear what value is being used as room temperature.

If we look at the temperature line in Figure 1.10 we can see that room temperature is above the boiling point of oxygen; this means that oxygen is a gas at room temperature.

Let's look at what happens when we heat bromine from –100 °C to 100 °C. As –100 °C is below bromine's melting point, bromine is a solid at –100 °C. As it is heated to –7 °C (its melting point) it becomes a liquid and it remains as a liquid until its temperature reaches the boiling point at 59 °C. Room temperature is between the melting point and the boiling point, which means that bromine is a liquid at room temperature. Above 59 °C bromine is a gas.

Lithium's melting point is above room temperature and so it is a solid at room temperature.

DIFFUSION

DIFFUSION IN GASES

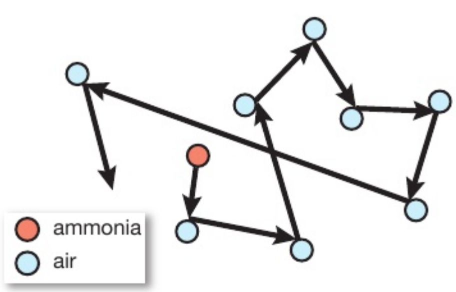


Figure 1.11 An ammonia particle bouncing off air particles.

Suppose someone accidentally releases some smelly gas in the lab, ammonia for example. Within a minute or so, everybody in the lab will be able to smell it. That isn't surprising – particles in the gas are free to move around. What does need explaining, though, is why it takes so long.

At room temperature, ammonia particles travel at speeds of about 600 metres per second so they should be able to travel from one end of a lab to the other in less than 1/100 s (0.01 s). This would be the case if they travelled in a straight line without bumping into anything else. However, each particle is bouncing off air particles on its way. In the time that it takes for the smell to reach all corners of the lab, each ammonia particle may have travelled 30 or more kilometres!

The spreading out of particles in a gas or liquid is known as **diffusion**. We say that ammonia particles *diffuse* through the air. A formal definition of diffusion is:

Diffusion is the spreading out of particles from where they are at a high concentration (there are lots of them in a certain volume) to where they are at a low concentration (there are fewer of them in a certain volume).

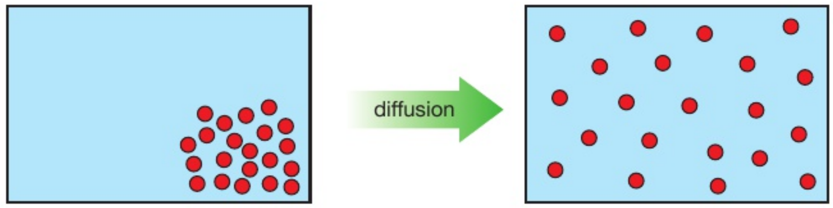


Figure 1.12 Diffusion involves the spreading out of particles.

You can show diffusion in gases very easily by using the apparatus in Figure 1.13. The lower gas jar contains bromine gas; the top one contains air. If the lids are removed, the brown colour of the bromine diffuses upwards until both gas jars are uniformly brown (the air particles also diffuse downwards). The bromine particles and air particles move around at random to give an even mixture – both gas jars contain air and bromine particles.

You can carry out the same experiment with hydrogen and air, but in this example you have to put a lighted splint in at the end to find out where the gases have gone. People often expect that the much less dense hydrogen will all go to the top gas jar. In fact, you will get identical explosions from both jars.

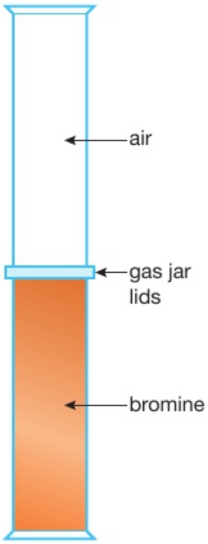


Figure 1.13 Demonstrating diffusion in gases

Safety Note: The teacher demonstration must be prepared in a working fume cupboard wearing eye protection and chemical-resistant gloves. Inhalation of bromine by anyone with breathing difficulties may produce a reaction, possibly delayed, requiring urgent medical attention.

SHOWING THAT PARTICLES OF DIFFERENT GASES TRAVEL AT DIFFERENT SPEEDS

This experiment relies on the reaction between ammonia (NH_3) and hydrogen chloride (HCl) gases to give white solid ammonium chloride (NH_4Cl):

$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

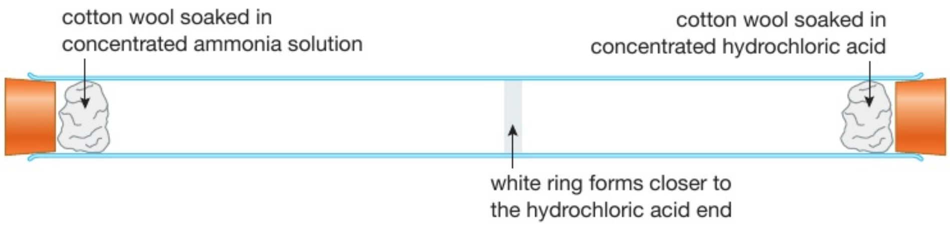


Figure 1.14 Demonstrating that particles in ammonia and hydrogen chloride travel at different speeds.

Pieces of cotton wool are soaked in concentrated ammonia solution (as a source of ammonia gas) and concentrated hydrochloric acid (as a source of hydrogen chloride gas). These are placed in the ends of a long glass tube with rubber bungs to stop the poisonous gases escaping.

Ammonia particles and hydrogen chloride particles diffuse along the tube. A white ring of solid ammonium chloride forms where they meet. The white ring of ammonium chloride takes time to form (as it takes some time for the particles of ammonia and hydrogen chloride to diffuse along the tube), and appears *closer to the hydrochloric acid end*. Ammonia particles are lighter than hydrogen chloride particles and therefore move faster. The ammonia particles travel further in the same amount of time, which means that the ring forms further away from the ammonia end.

HINT

Don't worry if you don't know how to write symbol equations. This one is included here so that you can refer to it again in later revision.

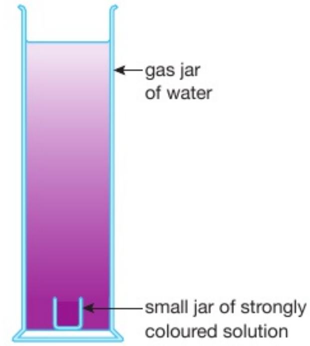
Safety Note: The teacher demonstration requires eye protection and the avoidance of skin contact and inhalation of any fumes. The apparatus has to be cleaned up in a working fume cupboard.

KEY POINT

You will learn about relative molecular mass in Chapter 5. The relative molecular mass of ammonia is 17 and that of hydrogen chloride is 36.5.

DIFFUSION IN LIQUIDS

Diffusion through a liquid is very slow if the liquid is completely still. For example, if a small jar of strongly coloured solution (such as potassium manganate(VII) solution) is placed in a gas jar of water, it can take days for the colour to diffuse throughout all the water. This is because *the particles in a liquid move more slowly than the particles in a gas*. The particles in a liquid are also much closer together than those in a gas and so there is less space for particles to move into without colliding with another one.



▲ Figure 1.15 Demonstrating diffusion in liquids

THE DILUTION OF COLOURED SOLUTIONS

Imagine you dissolve 0.01 g of potassium manganate(VII) in 1 cm^3 of water to make a deep purple solution. If we take the volume of 1 drop as 0.05 cm^3 we can work out that there are 20 drops in 1 cm^3 and each drop will contain 0.0005 g of potassium manganate(VII).

If you dilute this solution by adding water until the total volume is 10 000 cm^3, you should still just be able to see the purple colour.

There are now 200 000 drops in the solution. In order to see the colour each drop must contain at least one 'particle' of potassium manganate(VII), so there must be at least 200 000 'particles' in 0.01 g of potassium manganate(VII). This means that each 'particle' can't weigh more than 50 billionths of a gram (0.00000005 g).

This answer is not even close to the real answer. A potassium manganate(VII) 'particle' actually weighs about 0.00000000000000000000026 g and there are about 38 000 000 000 000 000 000 particles in 0.01 g! In reality, you need very large numbers of particles in each drop in order to see the colour.

REMINDER

Why the inverted commas around 'particle'? Potassium manganate(VII) is an *ionic compound* and contains more than one sort of particle. You will find out more about ionic compounds in Chapter 7.

THE SOLUBILITY OF SOLIDS

SOLUTES, SOLVENTS AND SOLUTIONS

When a solid dissolves in a liquid:

- the substance that dissolves is called the **solute**
- the liquid it dissolves in is called the **solvent**
- the liquid formed is a **solution**.

When you make a solution, the attractive forces between the particles in the solute (the solid) are being broken. At the same time, new attractive forces are being formed between the solvent particles and the solute particles. Whether a particular solid is soluble in any solvent depends on whether the new attractive forces are strong enough to overcome the old ones.

Only a certain amount of solute will dissolve in a fixed amount of solute at a particular temperature. When the maximum amount is dissolved a saturated solution is obtained. A **saturated solution** is a solution which contains as much dissolved solid as possible at a particular temperature. There must be some undissolved solute present.

MEASURING SOLUBILITY

CHEMISTRY ONLY

The **solubility** of a solid in a solvent at a particular temperature is usually defined as *'the mass of solute which must dissolve in 100 g of solvent at that temperature to form a saturated solution'*.

Figure 1.16 A saturated solution

EXTENSION WORK

It is possible to get supersaturated solutions with some solutes. These contain more dissolved solid than you would expect at a particular temperature. If you add even one tiny crystal of solid to these solutions, all the extra solute will crystallise out, and you are left with a normal saturated solution. You don't have to worry about this at International GCSE. Having undissolved solid present when you make a saturated solution prevents supersaturated solutions forming.

Safety Note: Wear eye protection and heat gently to avoid burns from hot solid 'spitting' out of the basin.

In other words, it is the maximum mass of solute that dissolves in 100 g of solvent at a particular temperature.

For example, the solubility of sodium chloride (common salt) in water at 25 °C is about 36 g per 100 g of water.

ACTIVITY 1

PRACTICAL: INVESTIGATING THE SOLUBILITY OF A SOLID IN WATER

A procedure we can use to measure the solubility of potassium nitrate in water at 40 °C is as follows:

1. Weigh an evaporating basin.
2. Heat a boiling tube of water to just above 40 °C.
3. Add potassium nitrate to the water in the boiling tube and stir rapidly until no more of it will dissolve and there is undissolved solid left over.
4. Allow the solution to cool to exactly 40 °C.
5. Pour off some of the solution into the evaporating basin (it is important that you only pour off solution and no solid). You do not have to pour off all the solution.
6. Weigh the evaporating basin and contents.
7. Heat the evaporating basin and contents gently to evaporate off all the water.
8. When it looks as if all the water has evaporated weigh the evaporating basin and contents.
9. Heat the evaporating basin and contents again and then re-weigh. This is to make sure that all the water has, indeed, evaporated and is called *heating to constant mass*.

This procedure is summarised in Figure 1.17.

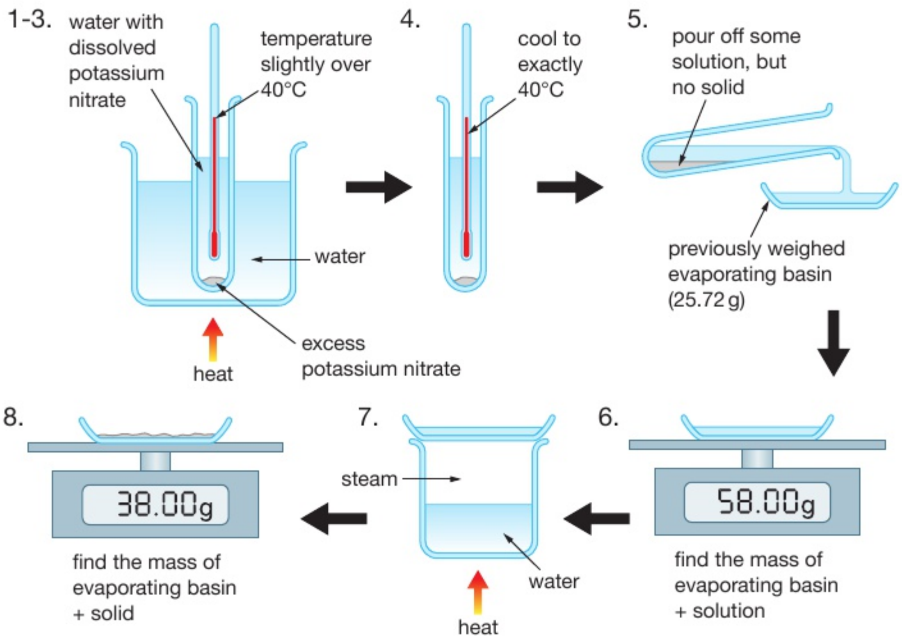


Figure 1.17 Finding the solubility of potassium nitrate in water at 40 °C.

We heat the solution gently to make sure that none spits out. If some did spit out we would record a lower mass of solid and the solubility would appear to be lower than the actual value.

The results for this experiment could be:

Mass of evaporating basin/g	25.72
Mass of evaporating basin + solution/g	58.00
Mass of evaporating basin + dry crystals/g	38.00

We need to calculate the mass of the solid and also the mass of water evaporated from the solution:

mass of crystals = 38.00 – 25.72 = 12.28 g

mass of water = 58.00 – 38.00 g = 20.00 g

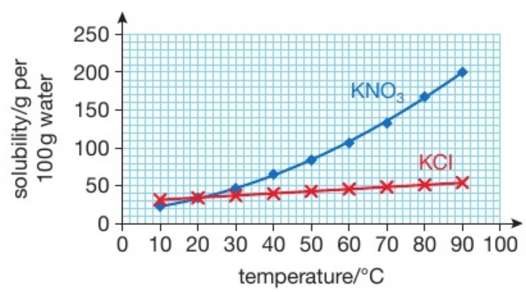
12.28 g of solid is the maximum mass that dissolves in 20.00 g of water, therefore 5 times as much would dissolve in 100 g of water. That works out at 61.4 g. The solubility of potassium nitrate at 40 °C is therefore 61.4 g per 100 g of water.

More generally, we can calculate the solubility of a substance in 100 g of solvent using the equation:

$$\text{solubility (g/100 g)} = \frac{\text{mass of solute}}{\text{mass of solvent}} \times 100$$

SOLUBILITY CURVES

The solubility of solids changes with temperature and you can plot this on a **solubility curve**. Most solids have solubility curves like those for the salts shown in Figure 1.18. Their solubility increases with temperature – either dramatically or just a little.



▲ Figure 1.18 Solubility curves for potassium nitrate and potassium chloride

HINT

Because solubility varies with temperature, you must always quote a temperature with a solubility value, e.g. the solubility of sodium chloride at 30 °C is…

You can use solubility curves to work out what mass of crystals you would get if you cooled a saturated solution.

Consider the solubility curve for potassium nitrate (KNO_3) in Figure 1.18. At 90 °C 200 g of potassium nitrate dissolves in 100 g water. At 30 °C only 50 g will dissolve. Therefore, if we have a solution containing 200 g of potassium nitrate dissolved in 100 g of water and let it cool down from 90 °C to 30 °C, 150 g of potassium nitrate must be released from the solution, which it does as crystals. We say that potassium nitrate *crystallises out of the solution* or *precipitates out of the solution*.

KEY POINT

The dashed lines marked on the graph come from answers to the next part of the question. In this case the solubility curve is virtually a straight line – this will not always be the case. If you are asked to draw a line of best fit this can be either a straight line or a curve.

EXAMPLE

The table shows the solubility of potassium chloride at various temperatures.

Temperature/°C	10	30	40	70	90
Solubility/g per 100 g of water	31.2	37.2	40.0	48.5	53.9

a Plot a solubility curve for potassium chloride.

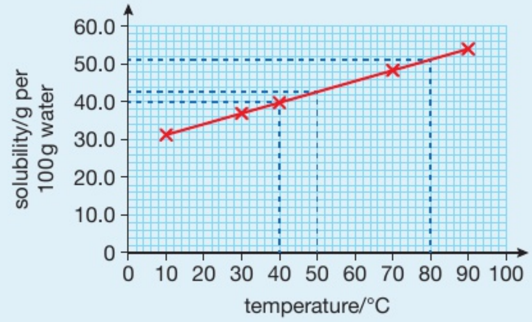


Figure 1.19 Solubility curve for potassium chloride

b Use the solubility curve to find:

- **i** the solubility of potassium chloride at 50 °C
- **ii** the maximum mass of potassium chloride that would dissolve in 50 g of water at 40 °C.
- **iii** the temperature at which crystals will first appear if you cooled a hot solution containing 51.0 g of potassium chloride in 100 g of water.

Use the graph to find the information you want.

Part i: The graph shows that at 50 °C the solubility is 42.5 g per 100 g of water.

Part ii: From the graph we can see that the solubility at 40 °C is 40 g per 100 g of water. From this we can deduce that half as much, that is 20 g, will dissolve in 50 g (half the mass) of water.

The numbers we have used here are quite simple, but if you were, for instance, asked to work out the maximum mass that would dissolve in 34.6 g of water at 40 °C you could use the equation:

$$\frac{\text{mass of water (g)}}{100\text{ (g)}} \times \text{solubility (g per 100 g)} = \text{maximum mass that dissolves (g)}$$

So, with 34.6 g of water we would get:

$$\frac{34.6}{100} \times 40 = 13.84\text{ g of potassium chloride dissolves}$$

Part iii: Solubility measures the maximum mass of potassium chloride which will dissolve in 100 g of water at a particular temperature. Crystals will start to appear as soon as the solution becomes saturated. From the graph, it can be seen that more than 51.0 g of solid will be soluble at 90 °C, but as the solution is cooled the solubility decreases. Drawing a line across at 51.0 g shows that this is the maximum mass that will dissolve at 80 °C and therefore crystals will first appear at temperatures below this.

c What mass of potassium chloride would crystallise from the solution in biii if the temperature fell to 10 °C?

You can use the data in the table to find the solubility at 10 °C. This value is 31.2 g per 100 g of water. That means that 31.2 g of potassium chloride will stay in solution at 10 °C. Since you started with 51.0 g, the rest of it must have formed crystals:

mass of crystals = 51.0 − 31.2 g = 19.8 g

19.8 g of potassium chloride will crystallise out.

HINT

In different circumstances, you might have to find the 10 °C figure from the graph as well.

END OF CHEMISTRY ONLY

CHAPTER QUESTIONS

SKILLS CRITICAL THINKING 2

1 What name is given to each of the following changes of state?
- a Solid to liquid
- b Liquid to solid
- c Solid to gas
- d Gas to solid

SKILLS INTERPRETATION 5

2 a Draw diagrams to show the arrangement of the particles in a solid, a liquid and a gas.
- b Describe the difference between the movement of particles in a solid and a liquid.
- c The change of state from a liquid to a gas can be either evaporation or boiling. Explain the difference between evaporation and boiling.

SKILLS ANALYSIS

3 The questions refer to the substances in the table.

	Melting point/°C	Boiling point/°C
A	−259	−253
B	0	100
C	3700 (sublimes)	
D	−116	34.5
E	801	1413

a Write down the physical states of each compound at
- i 30 °C.
- ii −100 °C
- iii 80 °C

SKILLS PROBLEM SOLVING 6

b Which substance has the greatest distance between its particles at 25 °C? Explain your answer.

SKILLS REASONING 5

c Why is no boiling point given for substance **C**?

6

d Which liquid substance would evaporate most quickly in the open air at 25 °C? Explain your answer.

SKILLS REASONING

6 **4** Refer to Figure 1.14 on page 7 showing the diffusion experiment.

a Explain why the ring takes a few minutes to form.

5 **b i** If you heat a gas, what effect will this have on the movement of the particles?

6 **ii** In the light of your answer to **i**, what difference would you find if you did this experiment outside on a day when the temperature was 2 °C instead of in a warm lab at 25 °C? Explain your answer.

c Explain why the ring was formed nearer the hydrochloric acid end of the tube.

d Suppose you replaced the concentrated hydrochloric acid with concentrated hydrobromic acid. This releases the gas hydrogen bromide (HBr). Hydrogen bromide also reacts with ammonia to form a white ring.

i Suggest a name for the white ring in this case.

7 **ii** Hydrogen bromide particles are about twice as heavy as hydrogen chloride particles. What effects do you think this would have on the experiment?

SKILLS CRITICAL THINKING

4 **5** Use the words given below to complete the following paragraph. Each word may be used once, more than once or not at all.

Sodium chloride dissolves in water to form a ______. The water is called the ______ and the sodium chloride is the ______. If the solution is heated to 50 °C some of the water ______ until the solution becomes ______ and sodium chloride crystals start to form.

boils solution solute saturated evaporates solvent condenses

CHEMISTRY ONLY

SKILLS INTERPRETATION

6 **6** The solubility of sodium chlorate in water was measured at a number of different temperatures.

Temperature/°C	0	20	40	60	80	100
Solubility/g per 100 g of water	3	8	14	23	38	55

a Use these figures to plot a solubility curve, with the temperature on the horizontal axis and the solubility on the vertical one.

SKILLS ANALYSIS

5 **b** Use your graph to find the solubility of sodium chlorate at 50 °C.

6 **c** Determine the maximum mass of sodium chlorate that will dissolve in 40 g of water at 30 °C.

SKILLS PROBLEM SOLVING

d 20 g of sodium chlorate was added to 100 g of water and the mixture heated to about 70 °C. It was then left to cool with the thermometer in the solution. Use your graph to answer the following questions.

i At what temperature would crystals first appear in the solution?

7 **ii** If the solution was cooled to 17 °C, calculate the total mass of crystals formed.

END OF CHEMISTRY ONLY

2 ELEMENTS, COMPOUNDS AND MIXTURES

Most of the substances that we are familiar with from everyday life are mixtures. For example, the air that we breathe is a mixture containing elements such as nitrogen and oxygen, and compounds such as carbon dioxide and nitrogen oxides. The food that we eat and the drinks that we drink are mixtures. This chapter looks at the properties of elements, compounds and mixtures, and also how to separate the components of a mixture. Separation of mixtures is very important in the analysis of substances, such as in forensics.

▲ Figure 2.1 Gold is an element, but a gold ring made from 18-carat gold only contains 75% gold. The metal is a mixture of gold and, usually, copper.

▲ Figure 2.2 Pure water is a compound, but the water we drink is a mixture of water and other dissolved substances.

LEARNING OBJECTIVES

- Understand how to classify a substance as an element, compound or mixture
- Understand that a pure substance has a fixed melting and boiling point, but that a mixture may melt or boil over a range of temperatures
- Describe these experimental techniques for the separation of mixtures:
  - simple distillation
  - fractional distillation
  - filtration
  - crystallisation
  - paper chromatography
- Understand how a chromatogram provides information about the composition of a mixture
- Understand how to use the calculation of R_f values to identify the components of a mixture
- Practical: Investigate paper chromatography using inks/food colourings

REMINDER

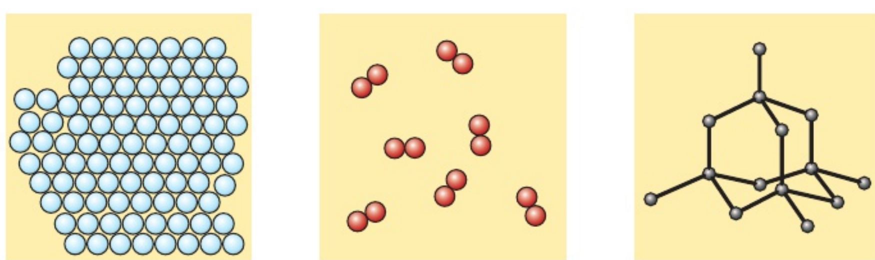
You might want to look at Chapter 3 if you do not already know the term 'atom'.

KEY POINT

It isn't completely true to say that elements consist of only one type of atom. A better way of saying it would be that *all the atoms in an element have the same atomic number*. Most elements consist of mixtures of isotopes, which have the same atomic number, but different mass numbers (due to different numbers of neutrons). When we draw diagrams or make models, we aren't usually interested in the differences between the isotopes. Isotopes will be discussed in Chapter 3.

ELEMENTS

Elements are *substances that can't be split into anything simpler by chemical means*. An element contains only one type of atom (but see the key point in the margin). In models or diagrams they are shown as atoms of a single colour or size.



a pure metal such as magnesium

oxygen gas

diamond (a form of carbon)

▲ Figure 2.3 Elements contain only one type of atom.

There are 118 elements and these are shown in the Periodic Table. Most of the elements occur naturally, such as hydrogen, helium and sulfur. Some others have to be made artificially, such as einsteinium.

COMPOUNDS

Compounds are formed when *two or more elements chemically combine*. The elements always combine in fixed proportions. For example, hydrogen and fluorine always combine to form hydrogen fluoride, with formula HF, whereas magnesium and fluorine always combine to form magnesium fluoride, with formula MgF_2 – the elements must combine in these ratios. Examples of other compounds are carbon dioxide (CO_2) and methane (CH_4). Diagrams of compounds show more than one type of atom bonded together.

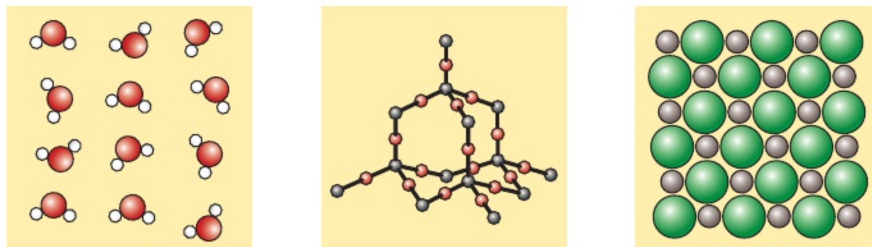


Figure 2.4 Some compounds

MIXTURES

In a **mixture**, the various substances are mixed together and no chemical reaction occurs. Mixtures can be made from elements and/or compounds. The various components can be in any proportion, for example you can put any amount of sugar into your cup of tea or coffee (until it becomes saturated).

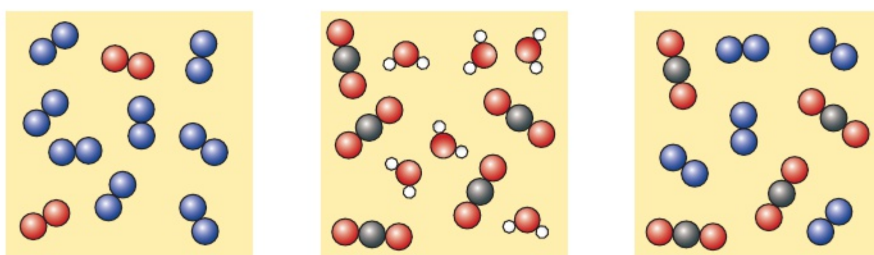


Figure 2.5 Some mixtures

SIMPLE DIFFERENCES BETWEEN MIXTURES AND COMPOUNDS

PROPORTIONS

In water (a compound), every single water molecule has two hydrogen atoms combined with one oxygen atom. It never varies. In a mixture of hydrogen and oxygen gases, the two could be mixed together in any proportion.

If you had some iron metal and some sulfur, you could mix them in any proportion you wanted to. In iron sulfide (FeS), a compound, the proportion of iron to sulfur is always exactly the same.

PROPERTIES

REMINDER

You can find out about the reactions of metals with dilute acids on pages 174–175. The reaction between iron sulfide and acids isn't needed for exam purposes at International GCSE.

In a mixture of elements, each element keeps its own properties, but the properties of the compound are quite different. For example, in a mixture of iron and sulfur, the iron is grey and the sulfur is yellow. The iron reacts with dilute acids such as hydrochloric acid to produce hydrogen; the sulfur doesn't react with the acid. However, the compound iron sulfide (FeS) reacts quite differently with acids to produce poisonous hydrogen sulfide gas, which smells of bad eggs.

A mixture of hydrogen and oxygen is a colourless gas which explodes when you put a flame to it. The compound, water, is a colourless liquid which just puts out a flame.

EASE OF SEPARATION

Mixtures can be separated by **physical means**. Physical means are things like changing the temperature or dissolving part of the mixture in a solvent such as water; in other words, methods that don't involve any chemical reactions.

For example, a mixture of iron and sulfur is quite easy to separate into the two elements using a magnet. The iron sticks to the magnet and the sulfur doesn't. The elements in a compound cannot be separated by physical means. To convert iron sulfide into separate samples of iron and sulfur requires chemical reactions.

You can cool a mixture of hydrogen and oxygen gases to separate it by a physical process. Oxygen condenses into a liquid at a much higher temperature than hydrogen (–183 °C as opposed to –253 °C). This would leave you with liquid oxygen and hydrogen gas, which are easy to separate. But to separate water into hydrogen and oxygen, you have to change it chemically using electrolysis. Electrolysis is explained in Chapter 10.

MELTING POINT AND BOILING POINT

Pure substances, such as elements and pure compounds, melt and boil at fixed temperatures. For example, the melting point of water is 0 °C and the boiling point 100 °C. However, mixtures usually melt or boil over a *range of temperatures*.

KEY POINT

A mixture is not a pure substance. If a sample contains only a small amount of an unwanted substance, the unwanted substance might be called an *impurity*.

The presence of impurities lowers the melting point of a substance and raises the boiling point. For instance, dissolving 10 g of common (table) salt (sodium chloride) in 1 litre of water lowers the melting point to about –0.6 °C and raises the boiling point to about 100.2 °C.

The melting point can be very useful in determining whether or not a substance is pure. If you continue to study chemistry you might carry out a practical experiment to make some aspirin. In order to determine whether your sample is pure or not you can measure the melting point. You would record the temperature at which your sample starts to melt, and then you would record the temperature at which it has fully melted to completely form a liquid. Aspirin is a white powder that melts at 138 °C. If the melting point of the sample you made is 128–134 °C you can see that it is quite impure because it melts over a wide range of temperature (below the melting point of pure aspirin).

SEPARATION OF MIXTURES

Separating mixtures is extremely important in chemistry. For example, we can see this in the processing of crude oil, in producing fresh water from salt water and in the enrichment of uranium. In forensic science, the components of a mixture usually have to be separated before they can be analysed.

FILTRATION

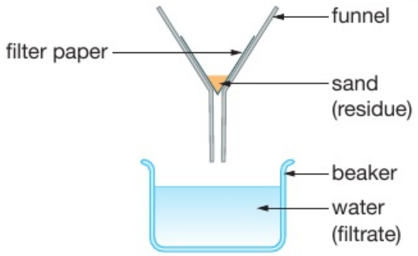


Figure 2.6 Filtration can be used to separate a mixture of sand and water.

Filtration can be used to separate a solid from a liquid.

For example, sand can be separated from water by filtration. The apparatus for filtration is shown in Figure 2.6.

The substance left in the filter paper is called the **residue** and the liquid that comes through is called the **filtrate**.

Filtration can also be used to separate two solids from each other if only one of them is soluble in water (see below – rock salt).

CRYSTALLISATION

Crystallisation can be used to separate a solute from a solution. For example, it could be used to separate sodium chloride from a sodium chloride solution. The solution is heated in an evaporating basin to boil off some of the water until an almost saturated solution is formed. This can be tested by dipping a glass rod into the solution and seeing if crystals form quickly on its surface when it is removed. The Bunsen burner is then turned off and the crystals allowed to form as more water evaporates and the solution cools. The crystals can now be removed from the mixture by filtration.

The apparatus for crystallisation is shown in Figure 2.7.

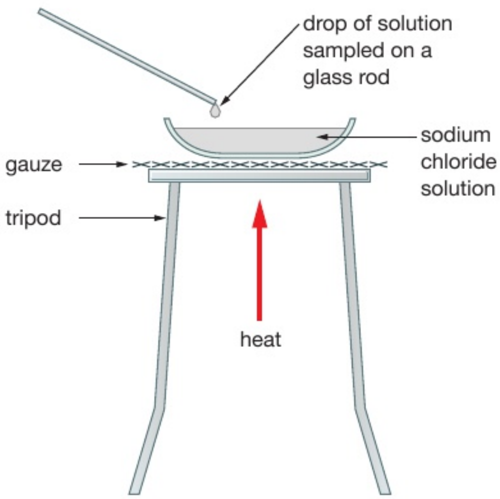


Figure 2.7 Crystallisation can be used to separate a solute from a solution.

MAKING PURE SALT FROM ROCK SALT

Figure 2.8 Rock salt

We can use filtration and crystallisation to obtain pure salt from rock salt.

Rock salt consists of salt contaminated by various earthy or rocky impurities. These impurities aren't soluble in water.

If you crush the rock salt and mix it with hot water, the salt dissolves, but the impurities don't. The impurities can be filtered off, and remain on the filter paper. The filtrate is then a salt solution. The solid salt can be obtained from the solution by crystallisation.

This is typical of the way you can separate any mixture of two solids, one of which is soluble in water and one of which isn't.

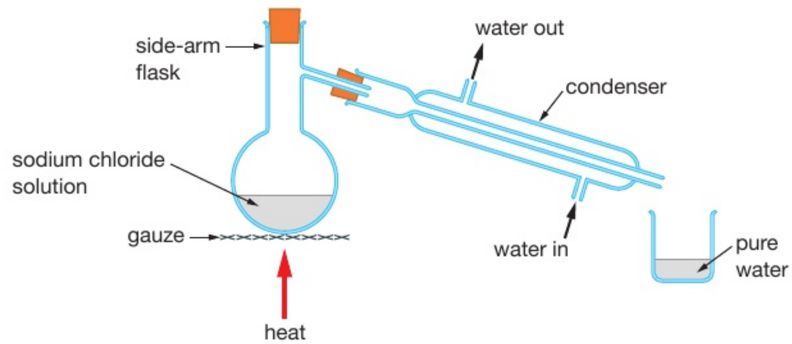
SIMPLE DISTILLATION

Simple distillation can be used to separate the components of a solution. Although we can use crystallisation to separate sodium chloride from a sodium chloride solution, we can also collect the water if we use simple distillation.

The water boils and is condensed back to a liquid by the condenser. The salt remains in the flask.

KEY POINT

Notice that water is always fed into the condenser at the lower end. That way it fills the condenser jacket better and if the flow of water stops for any reason the condenser jacket remains full of water.



▲ Figure 2.9 Distilling pure water from sodium chloride solution.

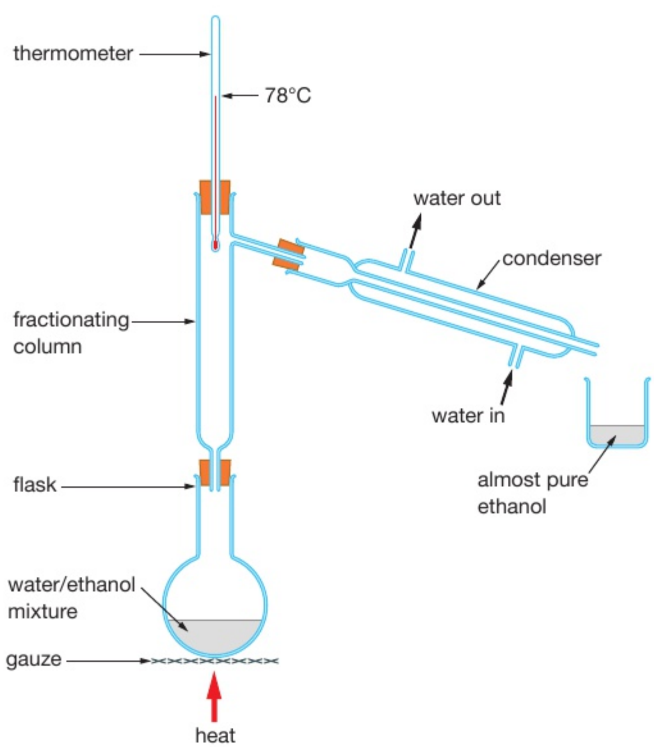
You could, of course, collect the salt from the solution as well as collecting pure water. The sodium chloride solution eventually becomes so concentrated that the salt will crystallise out.

FRACTIONAL DISTILLATION

Fractional distillation is used to separate a mixture of liquids such as ethanol (alcohol) and water. Ethanol and water are completely miscible with each other. That means you can mix them together in any proportion and they will form a single liquid layer. You can separate them by taking advantage of their different boiling points: water boils at 100 °C, ethanol at 78 °C.

EXTENSION WORK

The fractionating column is often packed with glass beads or something similar, although the separation of ethanol and water in the lab works perfectly well just with an empty column. For reasons that are beyond International GCSE, a high surface area in the column helps separation of the two vapours. The ethanol produced by this experiment is about 96% pure. For complicated reasons, again beyond International GCSE, it is impossible to remove the last 4% of water by distillation.



▲ Figure 2.10 Fractional distillation

Both liquids boil, but by careful heating you can control the temperature of the column so that all the water condenses in the column and trickles back into the flask. Only the ethanol remains as a vapour all the way to the top of the fractionating column and out into the condenser.

PAPER CHROMATOGRAPHY

Paper chromatography can be used to separate a variety of mixtures. However, at International GCSE level we will usually use it to separate mixtures of coloured inks or food colourings. Most inks and food colourings are not just made up of one colour but contain a mixture of dyes.

Paper chromatography can also be used to separate a mixture of colourless substances such as sugars, but then some method must be used to make the spots visible on the paper.

Safety Note: Avoid skin contact with the solvents and dyes, especially if you have sensitive skin.

ACTIVITY 2

PRACTICAL: INVESTIGATING THE COMPOSITION OF DYE WITH PAPER CHROMATOGRAPHY

We can investigate the composition of a mixture of coloured dyes using paper chromatography. To do this we carry out the following steps.

1. Draw a line with a pencil across a piece of chromatography paper; this line should be about 1 cm from the bottom of the paper. Do not use a pen as the colours in the ink may move up the chromatography paper with the solvent.
2. Put a spot (use a teat pipette or a capillary tube) of the mixture of dyes on the pencil line and allow it to dry.
3. Suspend the chromatography paper in a beaker that contains a small amount of solvent so that the bottom of the paper goes into the solvent. It is important that the solvent is below the pencil line so that the inks/colourings don't just dissolve in the solvent.
4. Put a lid (such as a watch glass) on the beaker so that the atmosphere becomes saturated with the solvent. This is to stop evaporation of the solvent from the surface of the paper.
5. When the solvent has moved up the paper to about 1 cm from the top, remove the paper from the beaker and draw a pencil line to show where the solvent got to. The highest level of the solvent on the paper at any time is called the *solvent front*.
6. Leave the paper to dry so that all the solvent evaporates.

For the solvent you can use water or a non-aqueous solvent (a solvent other than water). Which solvent you use depends on what substances are present in the mixture. A suitable solvent is usually found by experimenting with different ones.

The dyes that make up the mixture will be different in two important ways:

- the affinity they have for the paper (how well they 'stick' to the paper)
- how soluble they are in the solvent which moves up the paper.

In Figure 2.11 spot C has hardly moved. Either it was not very soluble in the solvent or it has a very high affinity for the paper (or both). On the other hand, spot A has moved almost as far as the solvent. It must be very soluble in the solvent and not have much affinity for the paper. The pattern you get is called a **chromatogram**.

KEY POINT

If the dye does not move from the pencil line during an experiment, then the dye is not at all soluble in the solvent you are using. In this case, you need to find a different solvent. If the dye moves up the paper with the solvent front, the dye is too soluble in that solvent and, again, you have to try a different solvent.

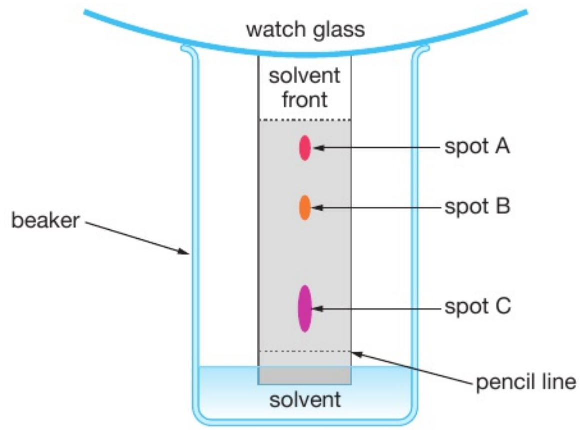


Figure 2.11 Paper chromatography

In this example, the mixture must have contained a minimum of three different dyes. We say a *minimum* of three dyes because there could be more – it is possible that one of the spots is made up of two coloured dyes that by concidence moved the same distance. You could only confirm this by doing the experiment again with a different solvent.

Figure 2.12 A paper chromatography experiment

USING PAPER CHROMATOGRAPHY IN ANALYSIS

You can use paper chromatography to identify the particular dyes in a mixture. If you think that your mixture (m) could contain dyes d1, d2, d3 and d4, you can carry out an experiment to determine this.

A pencil line is drawn on a larger sheet of paper and pencil marks are drawn along the line to show the original positions of the various dyes placed on the line (see Figure 2.13). One spot is your unknown mixture; the others are single, known dyes. The chromatogram is then allowed to develop as before.

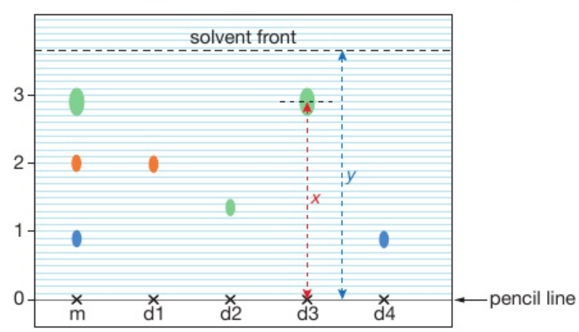


Figure 2.13 Paper chromatography can be used to analyse a mixture. Lines will not be present on your paper, but they have been added here to help you measure the distances.

The mixture (m) has spots corresponding to dyes d1, d3 and d4. They have the same colour as spots in the mixture, and have travelled the same distance on the paper. Although dye d2 is the same colour as one of the spots in the mixture, it has travelled a different distance and so must be a different compound.

Instead of just saying the spots move different distances we can use the R_f value to describe how far the spots move. R_f stands for **retardation factor**. Each time we do a chromatography experiment the solvent (and therefore the spots) will move different distances along the paper. This means we can't just report the distance moved by a particular spot so we have to work out a ratio instead.

$$R_f = \frac{\text{distance moved by a spot (from the pencil line)}}{\text{distance moved by the solvent front (from the pencil line)}}$$

In Figure 2.13 $R_f = \frac{x}{y}$.

So in Figure 2.13 the R_f value for dye d3 is:

$$R_f = \frac{2.9\,\text{cm}}{3.6\,\text{cm}} = 0.81$$

HINT

Measure to the centre of the spot.

The R_f values of the dyes in mixture m are:

blue spot: $R_f = \frac{0.9}{3.6} = 0.25$

orange spot: $R_f = \frac{2.0}{3.6} = 0.56$

green spot: $R_f = \frac{2.9}{3.6} = 0.81$

The R_f values of dyes d1 to d4 are:

d1: $R_f = 0.56$

d2: $R_f = 0.36$

d3: $R_f = 0.81$

d4: $R_f = 0.25$

Because the spots in mixture m have the same R_f values as d1, d3 and d4, we can conclude that the mixture contains these dyes.

An R_f value must be between 0 and 1. If you get a number bigger than 1 you have probably divided the numbers the wrong way round. An R_f value has no units.

You have to be careful when using R_f values as they depend on the solvent used and on the type of paper. There was no problem in the experiment described above because the mixture and the individual dyes were all put on the same piece of paper. However, if the mixture was put on one piece of chromatography paper and the individual dyes on a separate piece, you can still compare R_f values as long as you use the same type of paper and the same solvent.

CHAPTER QUESTIONS

SKILLS CRITICAL THINKING

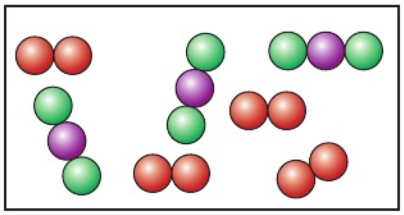
1 Classify each of the following substances as an element, compound or mixture:

sea water	hydrogen	honey
magnesium oxide	copper(II) sulfate	blood
calcium	mud	potassium iodide solution

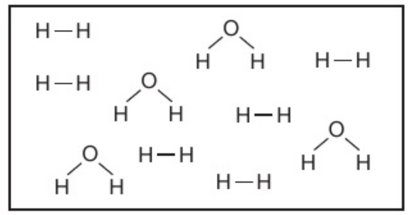
SKILLS ANALYSIS

2 Look at the diagrams below and classify each one as an element, compound or mixture.

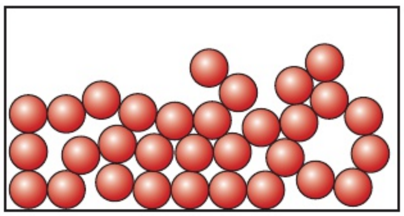
a



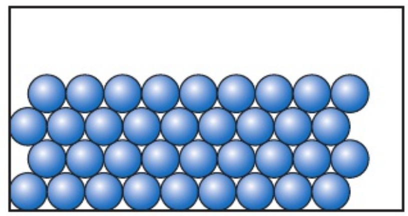
b



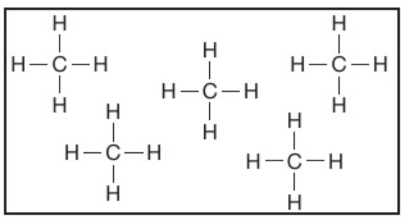
c



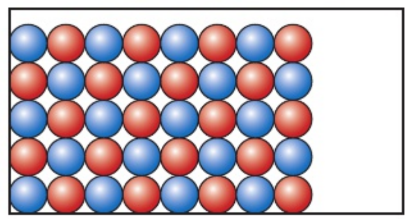
d



e



f



SKILLS REASONING, PROBLEM SOLVING

3 A teacher has found two white powders on a desk in the chemistry laboratory. She wants to test to see if they are pure substances, so she measures the melting points. Substance X melts at 122 °C and substance Y melts between 87 and 93 °C. Explain which one is the pure substance.

SKILLS DECISION MAKING

4 State which separation method you would use to carry out the following separations:

a Potassium iodide from a potassium iodide solution.

b Water from a potassium iodide solution.

c Ethanol from a mixture of ethanol and water.

d Red dye from a mixture of red and blue dyes.

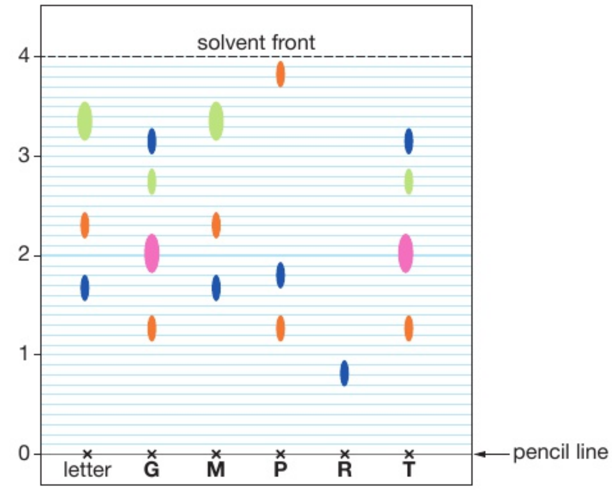
e Calcium carbonate (insoluble in water) from a mixture of calcium carbonate and water.

SKILLS CREATIVITY, DECISION MAKING

5 Suppose you had a valuable collection of small diamonds, which you kept safe from thieves by mixing them with white sugar crystals. You store the mixture in a jar labelled ‘sugar’. Now you want to sell the diamonds. Describe how you would separate all the diamonds from the sugar.

SKILLS ANALYSIS 5

6 In order to identify the writer of an anonymous letter, a sample of ink from the letter was dissolved in a solvent and then placed on some chromatography paper. Spots of ink from the pens of five possible writers, **G**, **M**, **P**, **R** and **T**, were placed next to the sample on the chromatography paper. The final chromatogram looked like this:



a Which of the five writers is using ink that matches the sample from the letter?

b Which of the writers is using a pen that contains ink made from a single dye?

SKILLS PROBLEM SOLVING 6

c What is the R_f value of the blue dye in suspect **P**'s pen?

SKILLS ANALYSIS 5

d Which two of the five writers are using pens containing the same ink?

6

e Whose pen contained the dye that was most soluble in the solvent?

3 ATOMIC STRUCTURE

This chapter explores the nature of atoms and how they differ from element to element. The 118 elements are the building blocks from which everything is made, from a simple substance, such as carbon, to a more complex one, such as DNA.

LEARNING OBJECTIVES

- Know what is meant by the terms atom and molecule
- Know the structure of an atom in terms of the positions, relative masses and relative charges of sub-atomic particles
- Know what is meant by the terms atomic number, mass number, isotopes and relative atomic mass (A_r)
- Be able to calculate the relative atomic mass of an element (A_r) from isotopic abundances

Copper is an element. If you tried to cut it up into smaller and smaller pieces, the final result would be the smallest possible piece of copper. At that stage, you would have an individual copper atom. You can, of course, split that atom into smaller pieces (protons, neutrons and electrons), but you would no longer have copper. Therefore, an **atom** is the smallest piece of an element that can still be recognised as that element.

▲ Figure 3.1 New atoms are produced in stars . . .

▲ Figure 3.2 . . . or in nuclear processes such as nuclear bombs, nuclear reactors or radioactive decay.

ATOMS AND MOLECULES

Atoms can be joined together to make molecules. A **molecule** consists of two or more atoms chemically bonded (by covalent bonds). The atoms that make up a molecule can be from the same elements or different elements. A hydrogen (H_2) molecule (Figure 3.3a) consists of 2 hydrogen atoms chemically bonded together. A water (H_2O) molecule (Figure 3.3b) consists of 2 hydrogen atoms and an oxygen atom chemically bonded.

▲ Figure 3.3 (a) A H_2 molecule and (b) a H_2O molecule. The lines between the atoms represent chemical bonds.

THE STRUCTURE OF THE ATOM

Atoms are made of protons, neutrons and electrons. These particles are sometimes called *sub-atomic particles* because they are smaller than an atom.

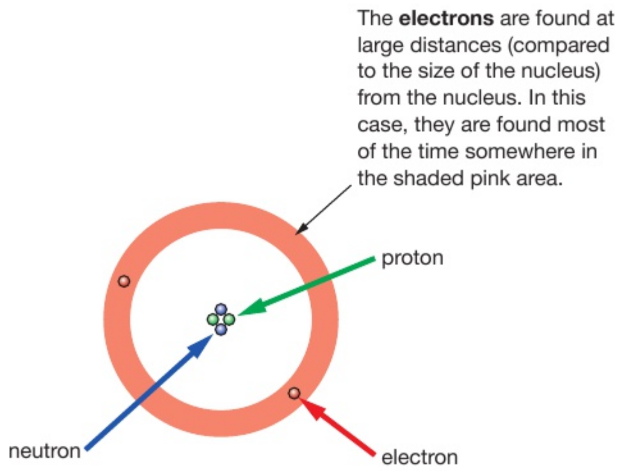


Figure 3.4 The structure of a helium atom

The nucleus of the atom contains protons and neutrons, and is shown highly magnified in Figure 3.4. In reality, if you scale up a helium atom to the size of a sports hall the nucleus would be no more than the size of a grain of sand.

The relative masses and charges of protons, neutrons and electrons are shown in Table 3.1.

Table 3.1 The properties of protons, neutrons and electrons

Particle	Relative mass	Relative charge
proton	1	+1
neutron	1	0
electron	1/1836	–1

Virtually all the mass of the atom is concentrated in the nucleus because electrons have a much smaller mass than protons and neutrons.

The masses and charges are measured relative to each other because the actual values are incredibly small. For example, it would take about 600 000 000 000 000 000 000 000 (6×10^{23}) protons to weigh 1 g.

ATOMIC NUMBER AND MASS NUMBER

The number of protons in an atom's nucleus is called its **atomic number** or **proton number**. Each of the 118 different elements has a different number of protons. For example, if an atom has 8 protons it must be an oxygen atom:

atomic number = number of protons

The **mass number** (sometimes known as the **nucleon number**) counts the total number of protons and neutrons in the nucleus of the atom:

mass number = number of protons + number of neutrons

KEY POINT

You may have come across diagrams of the atom in which the electrons are drawn orbiting the nucleus rather like planets around the sun. This can be misleading.
Electrons are constantly moving in the atom and it is impossible to know exactly where they are at any moment in time. You can only identify that they have a particular energy and that they are likely to be found in a certain region of space at some particular distance from the nucleus. Electrons with different energies are found at different distances from the nucleus.

HINT

$\frac{1}{1836}$ is approximately 0.0005.

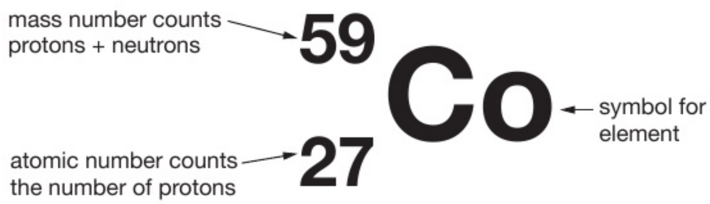
KEY POINT

The atomic number defines an element and is unique to that element. We can identify an element by its atomic number instead of its name. We could talk about a wristwatch made from the element with atomic number 79 instead of talking about 'a gold wristwatch', or say that the element with atomic number 17 is poisonous instead of saying 'chlorine is poisonous'. However, these are more complicated ways of describing things!

HINT

Be careful! When you are writing symbols with two letters, the first is a capital letter and the second must be lower case. If you write CO you are talking about carbon monoxide, not cobalt.

For any particular atom, this information can be shown as, for example:

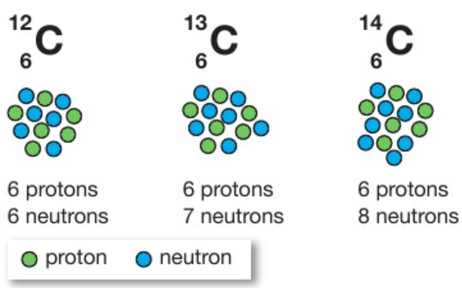


This particular atom of cobalt contains 27 protons. To make the total number of protons and neutrons up to 59, there must also be 32 neutrons.

You can see from this that:

number of neutrons = mass number – atomic number

ISOTOPES



▲ Figure 3.5 The nuclei of the three isotopes of carbon

The number of neutrons in an atom can vary slightly. For example, there are three kinds of carbon atom called carbon-12, carbon-13 and carbon-14. They all have the same number of protons (because all carbon atoms have 6 protons, its atomic number), but the number of neutrons varies. These different atoms of carbon are called **isotopes**.

Isotopes are atoms (of the same element) which have the same atomic number but different mass numbers. They have the same number of protons but different numbers of neutrons.

The fact that they have varying numbers of neutrons makes no difference to their chemical reactions. The chemical properties (how something reacts) are controlled by the number and arrangement of the electrons, and that is identical for all three isotopes.

RELATIVE ATOMIC MASS

You might have seen the following in a Periodic Table:

35.5
Cl
chlorine
17

Chlorine appears to have a mass number of 35.5. If you calculate the number of neutrons for chlorine you obtain:

number of neutrons = 35.5 – 17 = 18.5

It is not possible to have half a neutron and so there must be something wrong with this. The number 35.5 is not actually the mass number for chlorine but rather the **relative atomic mass** (A_r). Chlorine consists of two isotopes, ^{35}Cl and ^{37}Cl, and a naturally occurring sample contains a mixture of these.

KEY POINT

The number above each symbol in the International GCSE Periodic Table papers is a relative atomic mass and not a mass number. However, in most cases the relative atomic mass stated is the same as the mass number of the most common isotope. The only exceptions to this are chlorine (35.5) and copper (63.5).

KEY POINT

This type of average is called a **weighted average** or weighted mean.

Relative atomic mass is the average mass of an atom, taking into account the amount of each isotope present in a naturally occurring sample of an element. It is explained in more detail in Chapter 5.

You can probably see that a naturally occurring sample of chlorine must contain more of the ^{35}Cl isotope than the ^{37}Cl isotope. This is because the relative atomic mass is closer to 35 than to 37.

We can calculate the relative atomic mass of an element by knowing how much of each isotope is present in a sample (the isotopic abundances) of that element, and then working out the average mass of an atom. This is done in exactly the same way as you would calculate a weighted average in maths. It can be understood more easily by looking at a worked example.

EXAMPLE 1

A naturally occurring sample of the element boron contains 20% ^{10}B and 80% ^{11}B. Calculate the relative atomic mass.

If we imagine there are 100 atoms we can work out that 20% of them, that is 20, will have mass 10 and 80 will have mass 11.

The total mass of the 20 atoms with mass 10 is 20×10.

The total mass of the 80 atoms with mass 11 is 80×11.

The total mass of all the atoms in the sample is $20 \times 10 + 80 \times 11$.

There are 100 atoms so we can work out the average by dividing the total mass by the total number of atoms (100):

$$\text{relative atomic mass} = \frac{20 \times 10 + 80 \times 11}{100} = 10.8$$

Therefore, the relative atomic mass of boron is 10.8.

Even if there are three or four different isotopes, you still do the calculation in the same way: calculate the total mass of 100 atoms, then divide the answer by 100.

THE ELECTRONS

COUNTING THE NUMBER OF ELECTRONS IN AN ATOM

Atoms are electrically neutral (they have no overall charge). The charge on a proton (+1) is equal but opposite to the charge on an electron (–1), and therefore in an atom:

number of electrons = number of protons

HINT

Remember that the number of protons is the same as the atomic number of the element.

So, if an oxygen atom (atomic number = 8) has 8 protons, it must also have 8 electrons; if a chlorine atom (atomic number = 17) has 17 protons, it must also have 17 electrons.

You will see that the key feature in this is knowing the atomic number. You can find the atomic number from the Periodic Table.

The number of protons in an atom is equal to the number of electrons. However, the atomic number is defined in terms of the number of protons because the number of electrons can change in chemical reactions, for example when atoms form ions (see Chapter 7).

THE PERIODIC TABLE

Chapter 4 deals in detail with what you need to know about the Periodic Table for International GCSE purposes.

Atoms are arranged in the Periodic Table in order of increasing atomic number. You will find a full version of the Periodic Table in Appendix A on page 320. Most Periodic Tables have two numbers against each symbol; be careful to choose the right one. The atomic number will always be the smaller number. The other number will either be the mass number of the most common isotope of the element or the relative atomic mass of the element. The Periodic Table will clarify this.

You can use a Periodic Table to find out the number of protons, neutrons and electrons in an atom. Remember:

- the number of protons in an atom is equal to the atomic number
- the number of electrons in an atom is equal to the number of protons
- the number of neutrons in an atom = mass number – atomic number.

EXAMPLE 2

The symbol for uranium is given in a Periodic Table as:

238
U
Uranium
92

Calculate the number of protons, neutrons and electrons in an atom of uranium.

The atomic number is the smaller number, so the atomic number of uranium is 92. The atomic number tells us the number of protons, therefore a uranium atom contains 92 protons.

The number of protons is equal to the number of electrons, therefore a uranium atom contains 92 electrons.

The number of neutrons = mass number – atomic number.

The number of neutrons = 238 – 92 = 146.

CHAPTER QUESTIONS

You will need to use the Periodic Table in Appendix A on page 320.

SKILLS CRITICAL THINKING 7

1 Atoms contain three types of particle: proton, neutron and electron.

a State where the protons and neutrons are in an atom.

b State which type of particle in the atom orbits the nucleus.

c State which one of the particles has a positive charge.

d State which two particles have approximately the same mass.

SKILLS CRITICAL THINKING

7

2 Fluorine atoms have a mass number of 19.

a Use the Periodic Table to find the atomic number of fluorine.

b Explain what *mass number* means.

c State the number of protons, neutrons and electrons in a fluorine atom.

SKILLS REASONING

8

d Explain why the number of protons in an atom must always equal the number of electrons.

SKILLS PROBLEM SOLVING

9

3 Work out the numbers of protons, neutrons and electrons in each of the following atoms:

a $^{56}_{26}Fe$ b $^{93}_{41}Nb$ c $^{235}_{92}U$

SKILLS CRITICAL THINKING

7

4 Chlorine has two isotopes, chlorine-35 and chlorine-37.

a Explain what *isotopes* are.

SKILLS PROBLEM SOLVING

b State the numbers of protons, neutrons and electrons in the two isotopes.

10

5 Lithium has two naturally occurring isotopes, ^{6}Li (abundance 7%) and ^{7}Li (abundance 93%). Calculate the relative atomic mass of lithium, giving your answer to 2 decimal places.

6 Magnesium has three naturally occurring stable isotopes, ^{24}Mg (abundance 78.99%), ^{25}Mg (abundance 10.00%) and ^{26}Mg (abundance 11.01%). Calculate the relative atomic mass of magnesium, giving your answer to 2 decimal places.

7 Lead has four naturally occurring stable isotopes. Calculate the relative atomic mass of lead given the data in the table.

Mass number	Natural abundance/%
204	1.4
206	24.1
207	22.1
208	52.4

7

8 Iridium has two naturally occurring isotopes, ^{191}Ir and ^{193}Ir.

a State the number of protons, neutrons and electrons in an ^{191}Ir atom.

SKILLS REASONING

8

b Explain the difference between the two isotopes.

9

c The relative atomic mass of iridium is 192.22. Explain whether a naturally occurring sample of iridium contains more ^{191}Ir or ^{193}Ir.

10

9 Use the Periodic Table to explain whether the following statement is true or false.

Considering only the most common isotope of each element, there is only one element that has more protons than neutrons.