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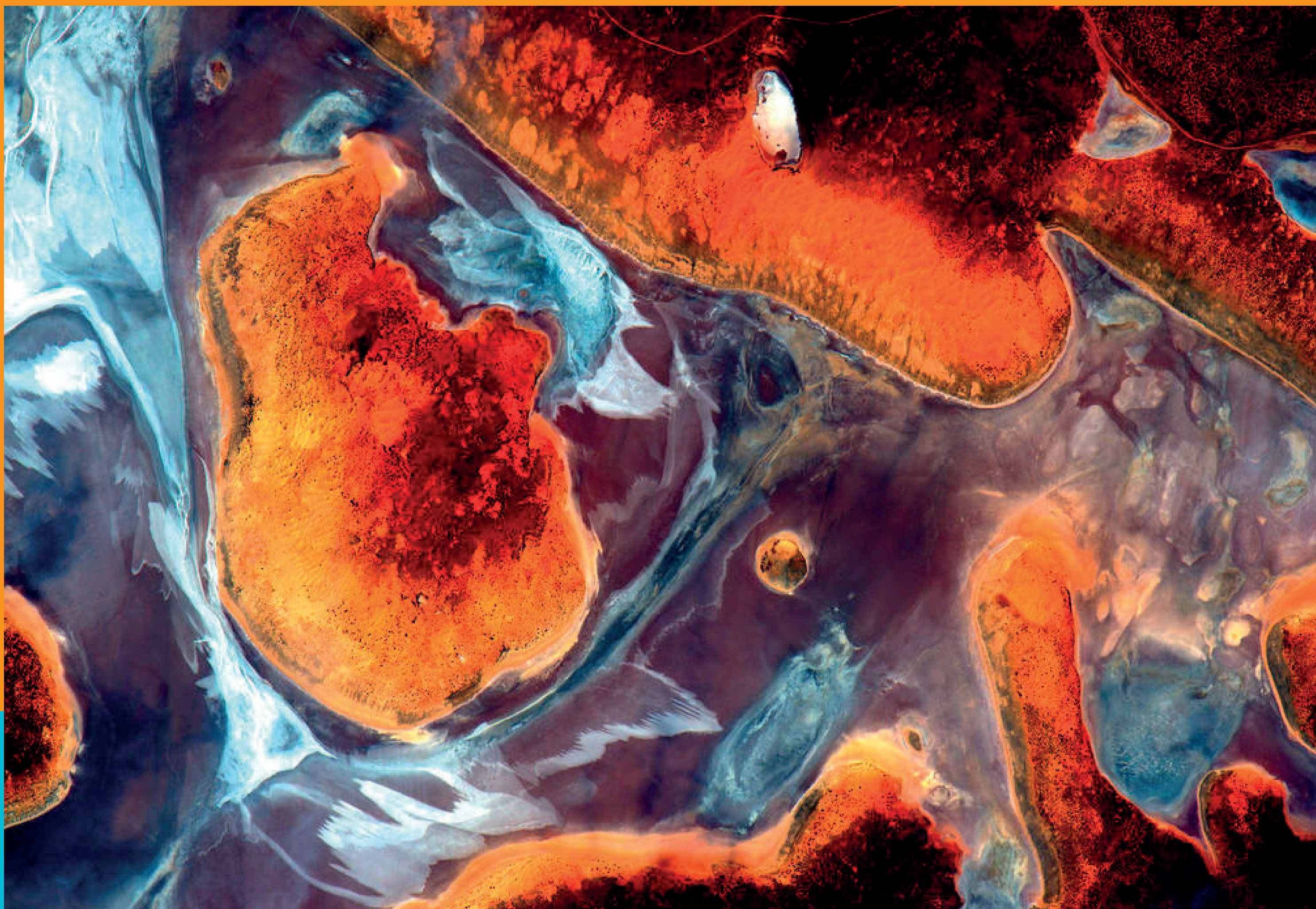

EDEXCEL INTERNATIONAL GCSE (9–1)

CHEMISTRY

Student Book

Jim Clark, Steve Owen, Rachel Yu

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CHEMISTRY

Student Book

Jim Clark
Steve Owen
Rachel Yu

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UNIT 1

PRINCIPLES OF CHEMISTRY

UNIT 2

INORGANIC CHEMISTRY

UNIT 3

PHYSICAL CHEMISTRY

UNIT 4

ORGANIC CHEMISTRY

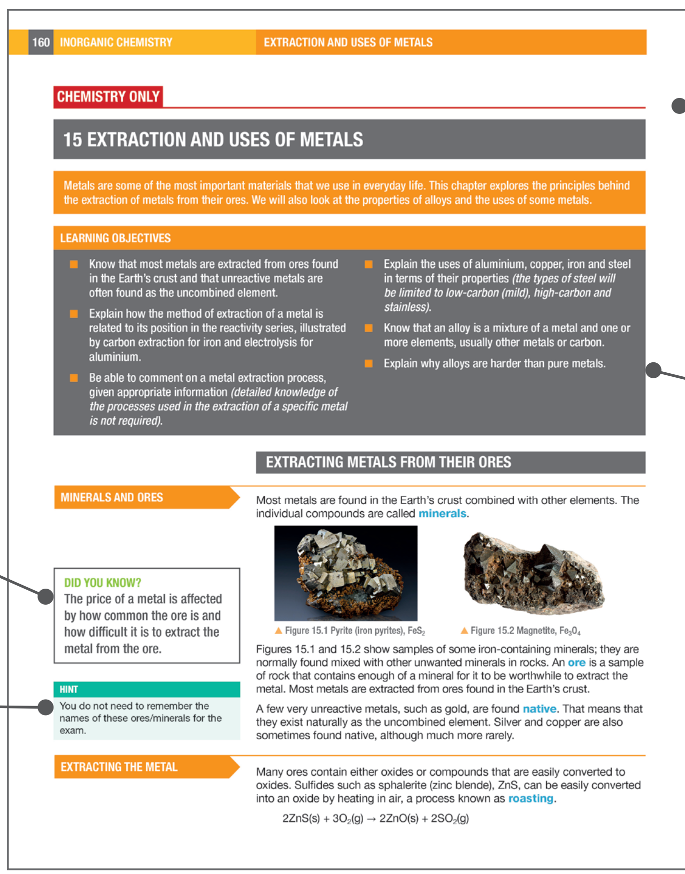
ABOUT THIS BOOK

This book is written for students following the 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.

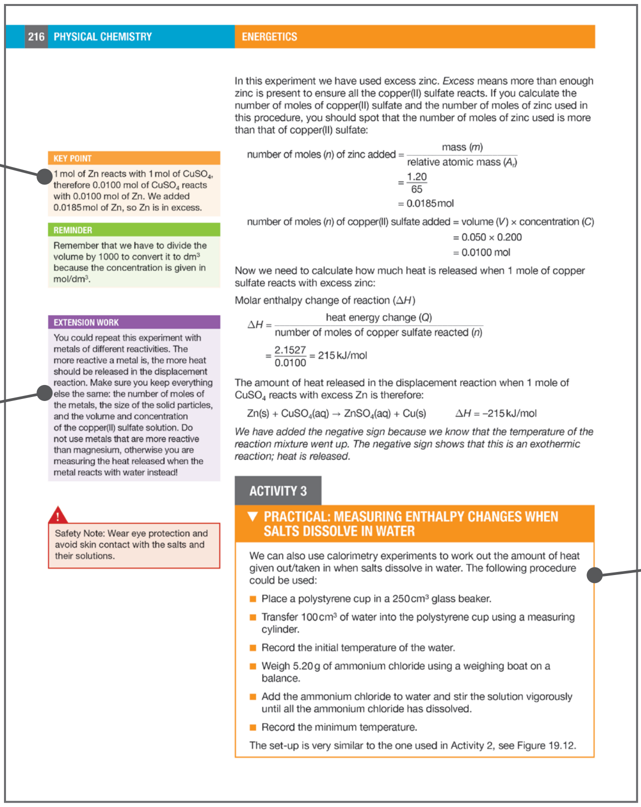
Learning Objectives show what you will learn in each chapter.

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.


216 PHYSICAL CHEMISTRY ENERGETICS

In this experiment we have used excess zinc. *Excess* means more than enough zinc is present to ensure all the copper(II) sulfate reacts. If you calculate the number of moles of copper(II) sulfate and the number of moles of zinc used in this procedure, you should spot that the number of moles of zinc used is more than that of copper(II) sulfate:

$$\text{number of moles } (n) \text{ of zinc added} = \frac{\text{mass } (m)}{\text{relative atomic mass } (A_r)} = \frac{1.20}{65} = 0.0185\,\text{mol}$$

$$\text{number of moles } (n) \text{ of copper(II) sulfate added} = \text{volume } (V) \times \text{concentration } (C) = 0.050 \times 0.200 = 0.0100\,\text{mol}$$

Now we need to calculate how much heat is released when 1 mole of copper sulfate reacts with excess zinc:

Molar enthalpy change of reaction (ΔH)

$$\Delta H = \frac{\text{heat energy change } (Q)}{\text{number of moles of copper sulfate reacted } (n)} = \frac{2.1527}{0.0100} = 215\,\text{kJ/mol}$$

The amount of heat released in the displacement reaction when 1 mole of $CuSO_4$ reacts with excess Zn is therefore:

$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$ $\Delta H = -215\,\text{kJ/mol}$

We have added the negative sign because we know that the temperature of the reaction mixture went up. The negative sign shows that this is an exothermic reaction; heat is released.

KEY POINT
1 mol of Zn reacts with 1 mol of $CuSO_4$, therefore 0.0100 mol of $CuSO_4$ reacts with 0.0100 mol of Zn. We added 0.0185 mol of Zn, so Zn is in excess.

REMINDER
Remember that we have to divide the volume by 1000 to convert it to dm^3 because the concentration is given in mol/dm^3.

EXTENSION WORK
You could repeat this experiment with metals of different reactivities. The more reactive a metal is, the more heat should be released in the displacement reaction. Make sure you keep everything else the same: the number of moles of the metals, the size of the solid particles, and the volume and concentration of the copper(II) sulfate solution. Do not use metals that are more reactive than magnesium, otherwise you are measuring the heat released when the metal reacts with water instead!

Safety Note: Wear eye protection and avoid skin contact with the salts and their solutions.

ACTIVITY 3
PRACTICAL: MEASURING ENTHALPY CHANGES WHEN SALTS DISSOLVE IN WATER

We can also use calorimetry experiments to work out the amount of heat given out/taken in when salts dissolve in water. The following procedure could be used:

- Place a polystyrene cup in a 250 cm^3 glass beaker.
- Transfer 100 cm^3 of water into the polystyrene cup using a measuring cylinder.
- Record the initial temperature of the water.
- Weigh 5.20 g of ammonium chloride using a weighing boat on a balance.
- Add the ammonium chloride to water and stir the solution vigorously until all the ammonium chloride has dissolved.
- Record the minimum temperature.

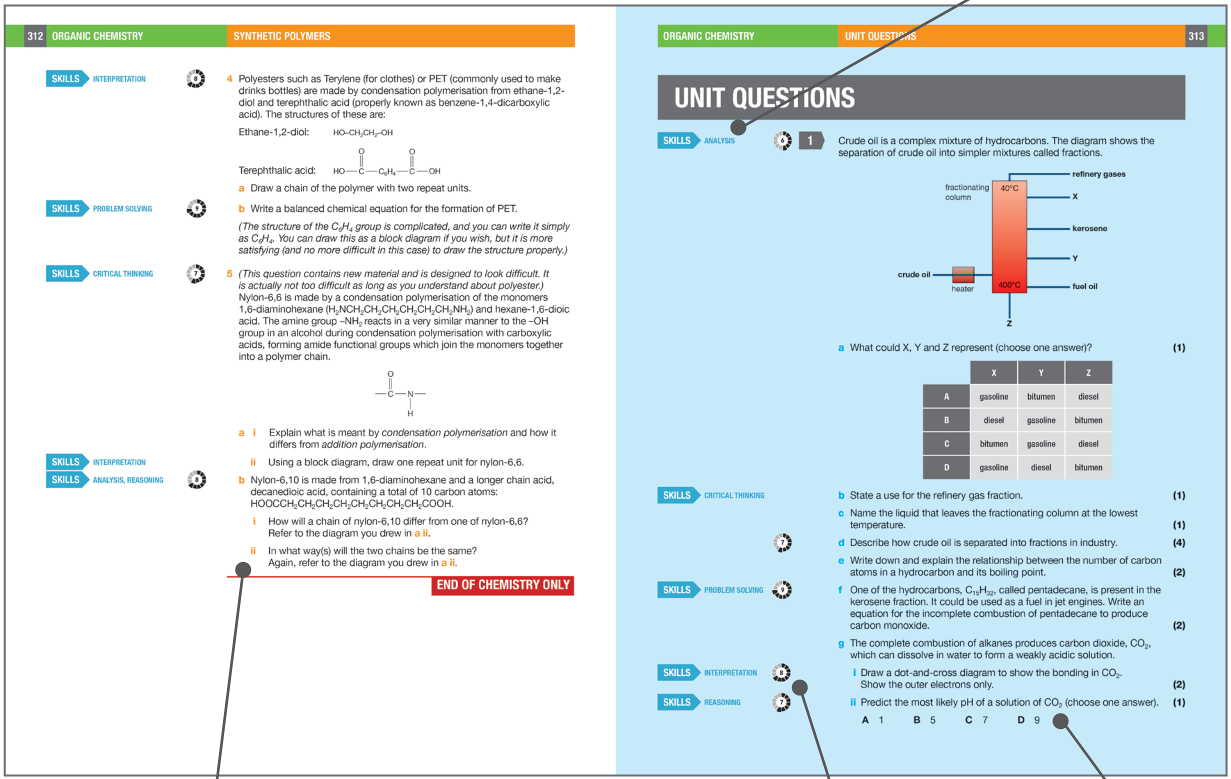
The set-up is very similar to the one used in Activity 2, see Figure 19.12.

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.


312 ORGANIC CHEMISTRY SYNTHETIC POLYMERS

SKILLS INTERPRETATION

4 Polyesters such as Terylene (for clothes) or PET (commonly used to make drinks bottles) are made by condensation polymerisation from ethane-1,2-diol and terephthalic acid (properly known as benzene-1,4-dicarboxylic acid). The structures of these are:

Ethane-1,2-diol: $HO{-}CH_2CH_2{-}OH$

Terephthalic acid: $HO{-}C({=}O){-}C_6H_4{-}C({=}O){-}OH$

a Draw a chain of the polymer with two repeat units.

SKILLS PROBLEM SOLVING

b Write a balanced chemical equation for the formation of PET.

(The structure of the C_6H_4 group is complicated, and you can write it simply as C_6H_4. You can draw this as a block diagram if you wish, but it is more satisfying (and no more difficult in this case) to draw the structure properly.)

SKILLS CRITICAL THINKING

5 *(This question contains new material and is designed to look difficult. It is actually not too difficult as long as you understand about polyester.)* Nylon-6,6 is made by a condensation polymerisation of the monomers 1,6-diaminohexane ($H_2NCH_2CH_2CH_2CH_2CH_2CH_2NH_2$) and hexane-1,6-dioic acid. The amine group $-NH_2$ reacts in a very similar manner to the $-OH$ group in an alcohol during condensation polymerisation with carboxylic acids, forming amide functional groups which join the monomers together into a polymer chain.

$-C(=O)-N(H)-$

a i Explain what is meant by *condensation polymerisation* and how it differs from *addition polymerisation*.

SKILLS INTERPRETATION

ii Using a block diagram, draw one repeat unit for nylon-6,6.

SKILLS ANALYSIS, REASONING

b Nylon-6,10 is made from 1,6-diaminohexane and a longer chain acid, decanedioic acid, containing a total of 10 carbon atoms: $HOOCCH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2COOH$.

i How will a chain of nylon-6,10 differ from one of nylon-6,6? Refer to the diagram you drew in a ii.

ii In what way(s) will the two chains be the same? Again, refer to the diagram you drew in a ii.

END OF CHEMISTRY ONLY

ORGANIC CHEMISTRY UNIT QUESTIONS 313

UNIT QUESTIONS

SKILLS ANALYSIS

1 Crude oil is a complex mixture of hydrocarbons. The diagram shows the separation of crude oil into simpler mixtures called fractions.

fractionating column; 40°C; refinery gases; X; kerosene; Y; crude oil; heater; 400°C; fuel oil; Z

a What could X, Y and Z represent (choose one answer)? (1)

	X	Y	Z
A	gasoline	bitumen	diesel
B	diesel	gasoline	bitumen
C	bitumen	gasoline	diesel
D	gasoline	diesel	bitumen

SKILLS CRITICAL THINKING

b State a use for the refinery gas fraction. (1)

c Name the liquid that leaves the fractionating column at the lowest temperature. (1)

d Describe how crude oil is separated into fractions in industry. (4)

e Write down and explain the relationship between the number of carbon atoms in a hydrocarbon and its boiling point. (2)

SKILLS PROBLEM SOLVING

f One of the hydrocarbons, $C_{15}H_{32}$, called pentadecane, is present in the kerosene fraction. It could be used as a fuel in jet engines. Write an equation for the incomplete combustion of pentadecane to produce carbon monoxide. (2)

g The complete combustion of alkanes produces carbon dioxide, CO_2, which can dissolve in water to form a weakly acidic solution.

SKILLS INTERPRETATION

i Draw a dot-and-cross diagram to show the bonding in CO_2. Show the outer electrons only. (2)

SKILLS REASONING

ii Predict the most likely pH of a solution of CO_2 (choose one answer). (1)

A 1 B 5 C 7 D 9

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.

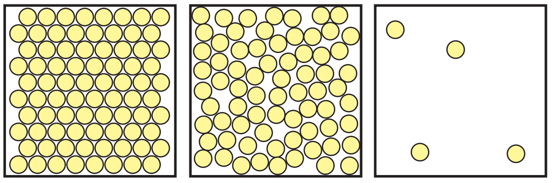


Figure 1.4 The arrangement of particles in different states of matter

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.

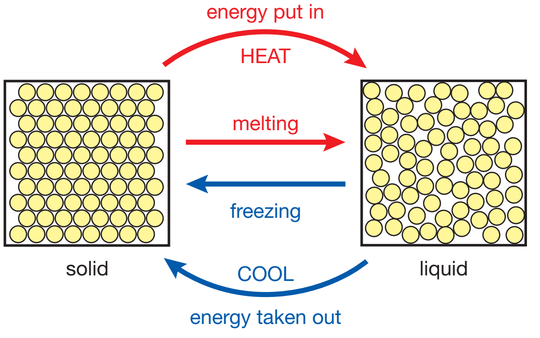


Figure 1.5 Melting to become a liquid – and freezing to become a solid.

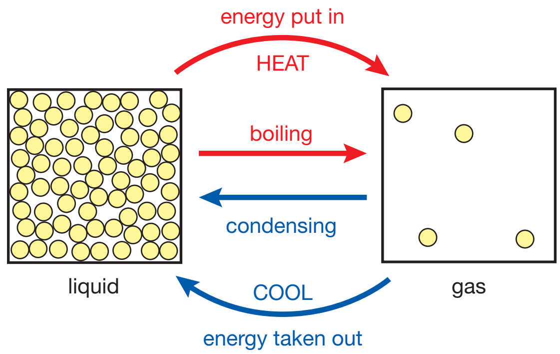
If the liquid is cooled again, the liquid particles will move around more and more slowly. Eventually, they are moving so slowly that the forces of attraction between them will hold them in a fixed position and the particles pack more closely together into a solid. The liquid **freezes**, forming a solid.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

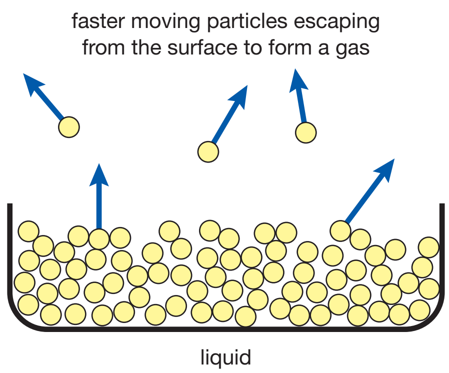
KEY POINT

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

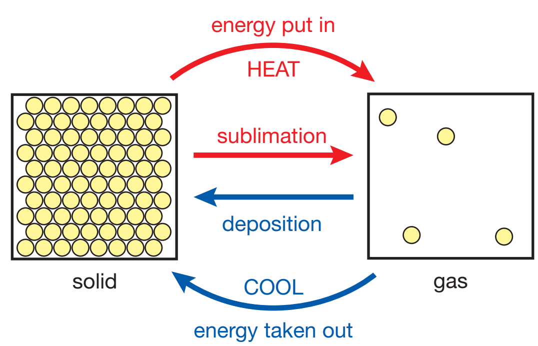
Evaporation is different. In any liquid or gas, the average speed of the particles varies with the temperature. But at each temperature, some particles will be moving faster than the average and others more slowly.

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.

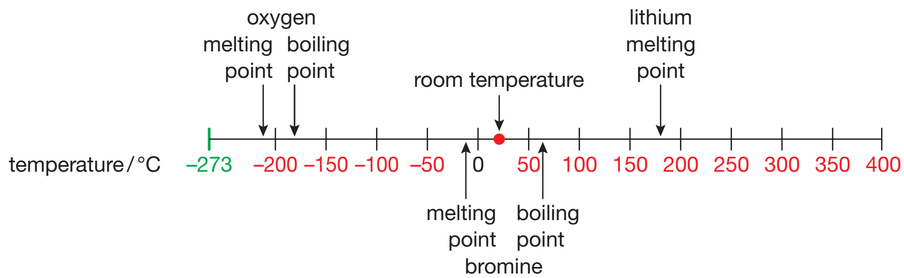
▲ Figure 1.9 Dry ice subliming. Notice the white solid carbon dioxide in the beaker. The white cloud is because the carbon dioxide gas produced is so cold that it causes water vapour in the air to condense. Carbon dioxide gas itself is invisible.

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.

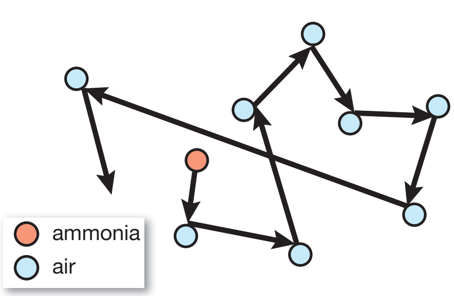
If we look at the temperature line in Figure 1.10 we can see that room temperature is above the boiling point of oxygen; this means that oxygen is a gas at room temperature.

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:

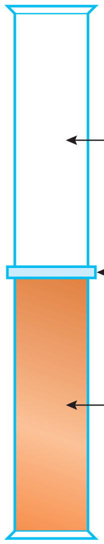


Figure 1.13 Demonstrating diffusion in gases

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.

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).

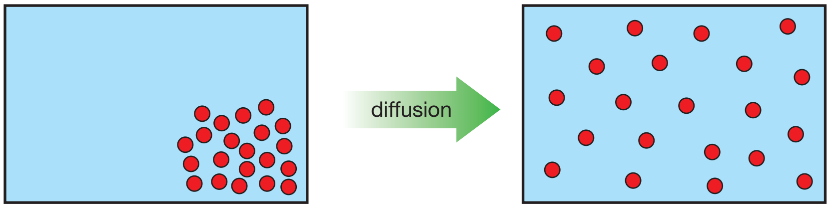


Figure 1.12 Diffusion involves the spreading out of particles.

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.

SHOWING THAT PARTICLES OF DIFFERENT GASES TRAVEL AT DIFFERENT SPEEDS

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.

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

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

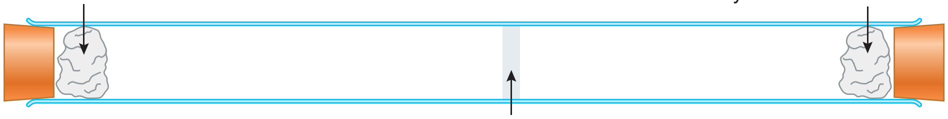


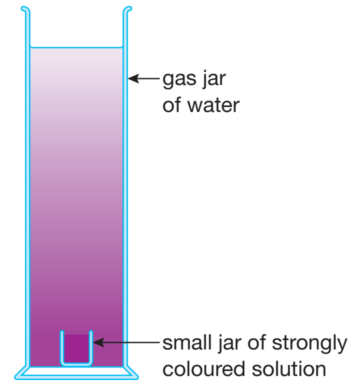
Figure 1.14 Demonstrating that particles in ammonia and hydrogen chloride travel at different speeds.

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.

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.

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

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.

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

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.

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.

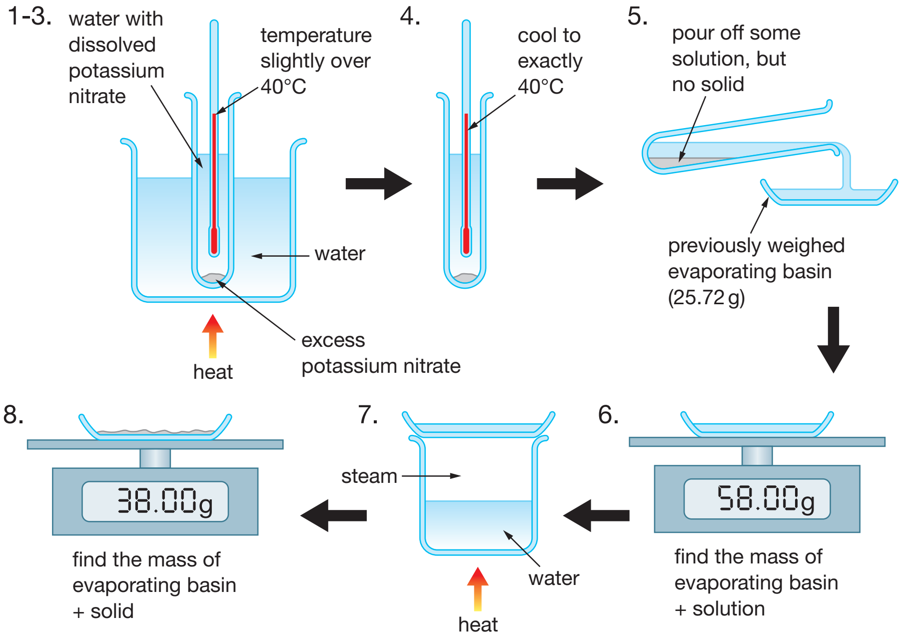


Figure 1.17 Finding the solubility of potassium nitrate in water at 40 °C.

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.

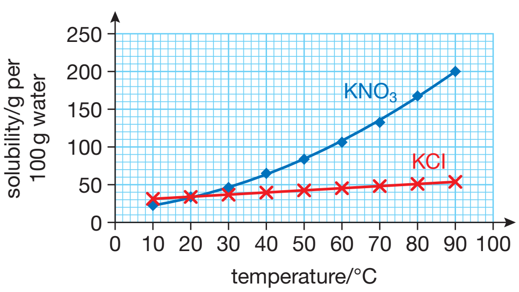


Figure 1.18 Solubility curves for potassium nitrate and potassium chloride

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*.

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…

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.

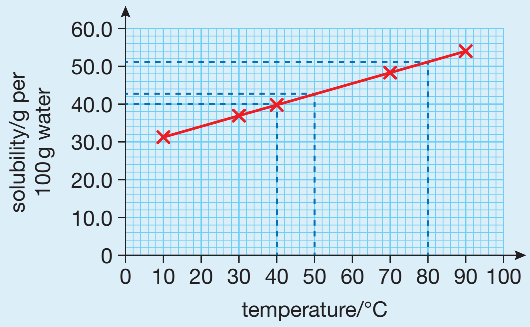


Figure 1.19 Solubility curve for potassium chloride

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

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

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

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

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.

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

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 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

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

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?

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

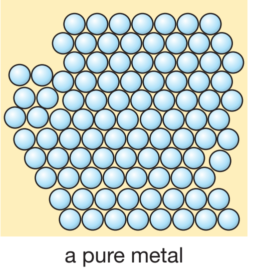
You might want to look at Chapter 3 if you do not already know the term 'atom'.

KEY POINT

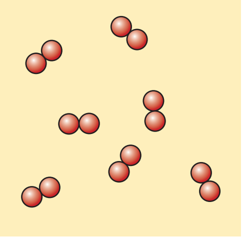
It isn't completely true to say that elements consist of only one type of atom. A better way of saying it would be that *all the atoms in an element have the same atomic number*. Most elements consist of mixtures of isotopes, which have the same atomic number, but different mass numbers (due to different 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

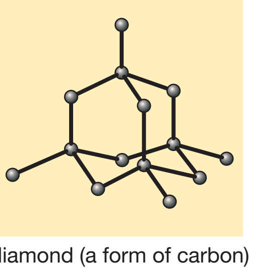
Elements are *substances that can't be split into anything simpler by chemical means*. An element contains only one type of atom (but see the key point in the margin). In models or diagrams they are shown as atoms of a single colour or size.



a pure metal such as magnesium



oxygen gas



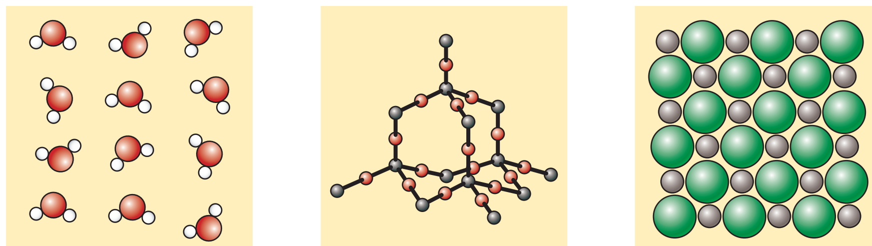
diamond (a form of carbon)

▲ 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.

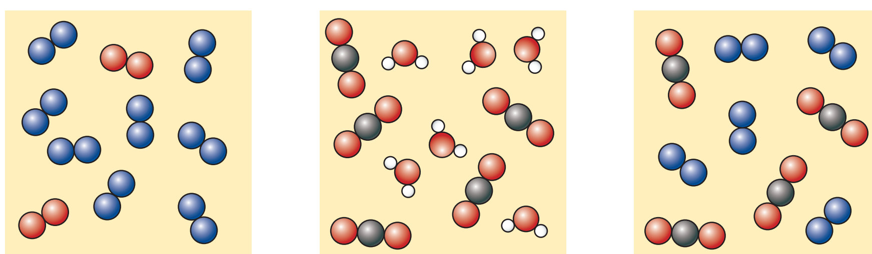


water | silicon dioxide | sodium chloride

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).



mixture of elements – nitrogen and oxygen | mixture of compounds – carbon dioxide and water (vapour) | mixture of an element with a compound – carbon dioxide and nitrogen

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

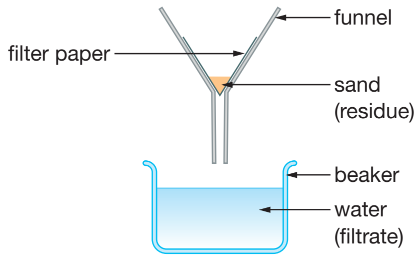


Figure 2.6 Filtration can be used to separate a mixture of sand and water.

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.

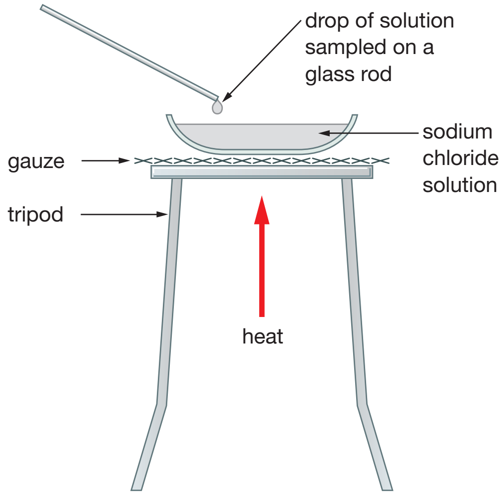


Figure 2.7 Crystallisation can be used to separate a solute from a solution.

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.

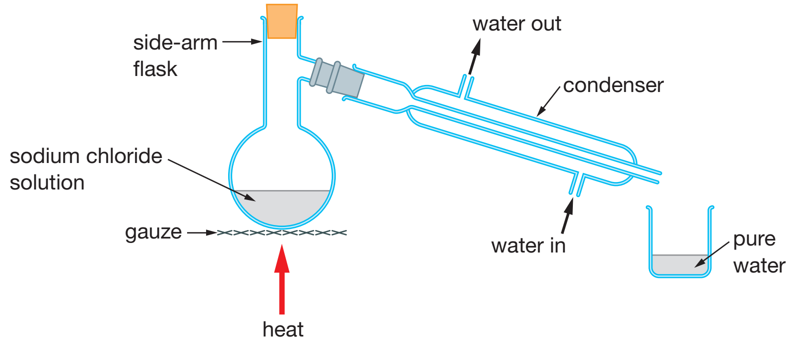
SIMPLE DISTILLATION

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.

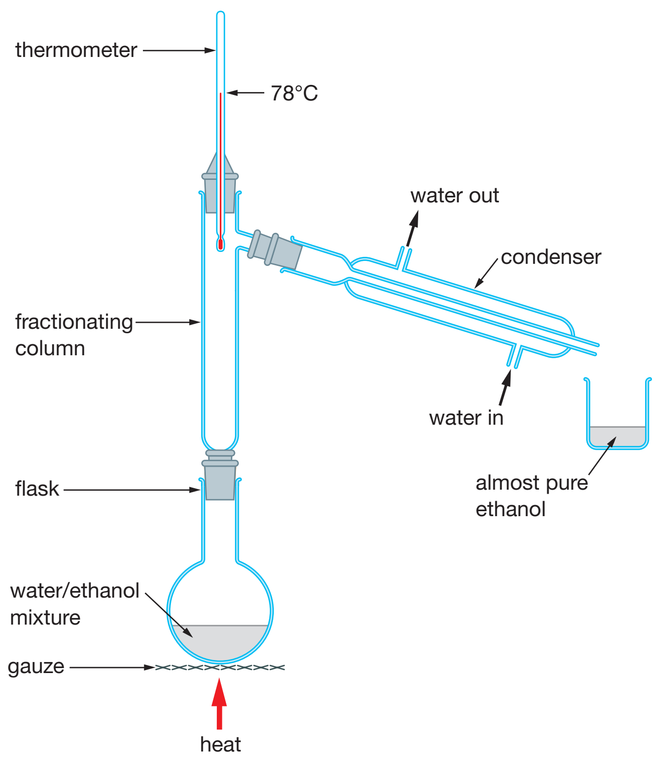
You could, of course, collect the salt from the solution as well as collecting pure water. The sodium chloride solution eventually becomes so concentrated that the salt will crystallise out.

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.

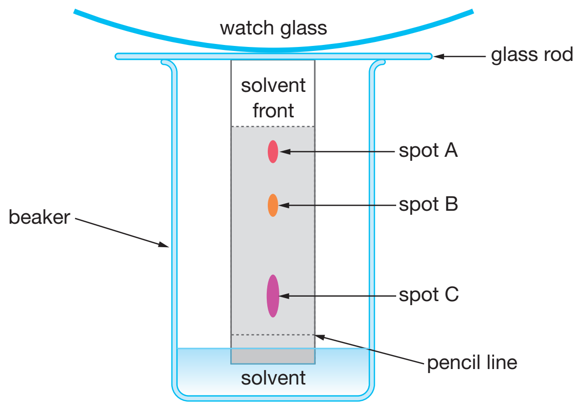


Figure 2.11 Paper chromatography

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.

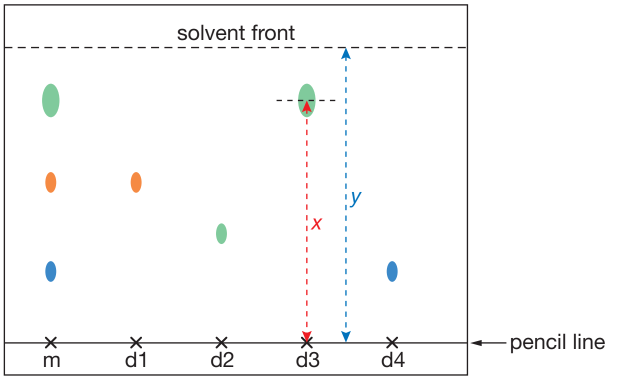


Figure 2.13 Paper chromatography can be used to analyse a mixture.

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

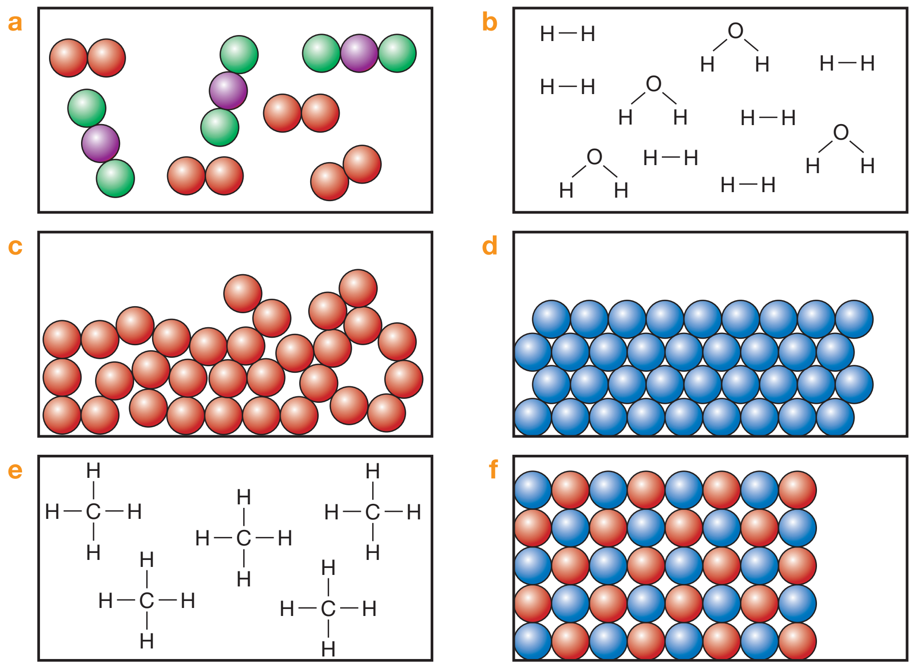
SKILLS CRITICAL THINKING

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.



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.

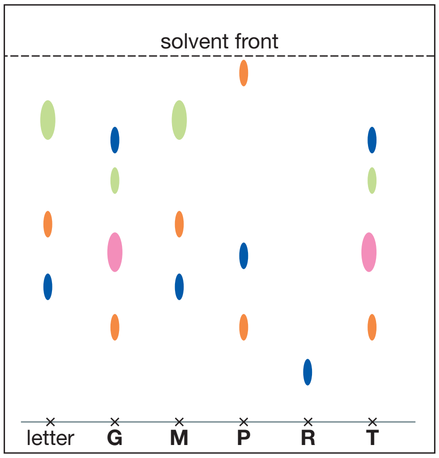
e Calcium carbonate (insoluble in water) from a mixture of calcium carbonate and water.

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.

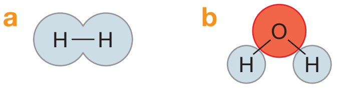
Copper is an element. If you tried to cut it up into smaller and smaller pieces, the final result would be the smallest possible piece of copper. At that stage, you would have an individual copper atom. You can, of course, split that atom into smaller pieces (protons, neutrons and electrons), but you 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.

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.

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.

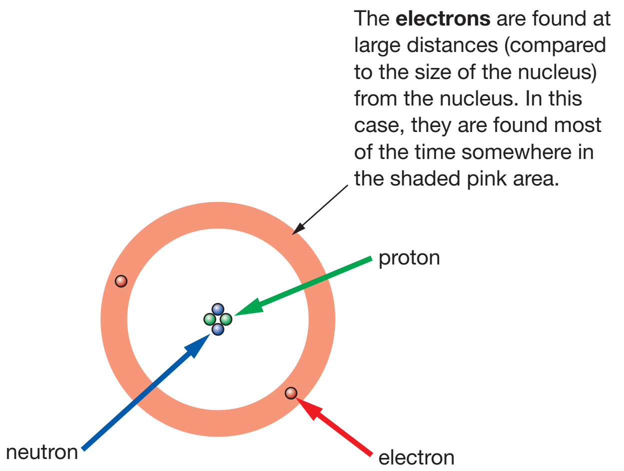


Figure 3.4 The structure of a helium atom

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.

HINT

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

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.

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!

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

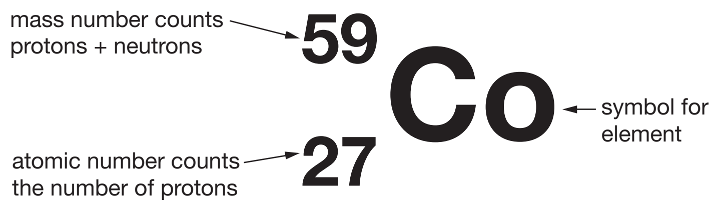
The **mass number** (sometimes known as the **nucleon number**) counts the total number of protons and neutrons in the nucleus of the atom:

mass number = number of protons + number of neutrons

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

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.

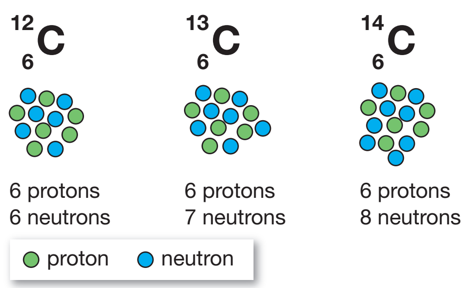


Figure 3.5 The nuclei of the three isotopes of carbon

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 × 10 + 80 × 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

SKILLS CRITICAL THINKING 7

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

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

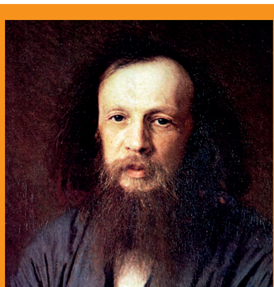
c The relative atomic mass of iridium is 192.22. Explain whether a naturally occurring sample of iridium contains more ^{191}Ir or ^{193}Ir.

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.

4 THE PERIODIC TABLE



The Periodic Table shows all the elements in the universe and is one of the most important tools that a chemist has. The arrangement of the elements allows us to understand trends in properties and make predictions. The modern Periodic Table was first presented in 1869 by a famous Russian chemist, Dmitri Mendeleev (left). This chapter explores some of the features of the Periodic Table.

LEARNING OBJECTIVES

- Understand how elements are arranged in the Periodic Table:
  - in order of atomic number
  - in groups and periods.
- Understand how to deduce the electronic configurations of the first 20 elements from their positions in the Periodic Table.
- Understand how to use electrical conductivity and the acid–base character of oxides to classify elements as metals or non-metals.
- Identify an element as a metal or a non-metal according to its position in the Periodic Table.
- Understand how the electronic configuration of a main group element is related to its position in the Periodic Table.
- Understand why elements in the same group of the Periodic Table have similar chemical properties.
- Understand why the noble gases (Group 0) do not readily react.

THE PERIODIC TABLE

The search for patterns in chemistry during the 19th century resulted in the modern Periodic Table. *The elements are arranged in order of atomic number* – the number of protons in the nuclei of the atoms.

The vertical columns are called **groups**

The horizontal rows are called **periods**

	1	2											3	4	5	6	7	0
1	H																	He
2	Li	Be											B	C	N	O	F	Ne
3	Na	Mg					transition metals						Al	Si	P	S	Cl	Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba	•	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
7	Fr	Ra	:	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og

lanthanoids

•	La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
:	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

actinoids

▲ Figure 4.1 The Periodic Table

DID YOU KNOW?

You will sometimes see Group 0 labelled as Group 8. In more advanced Periodic Tables the groups are numbered from 1 to 18, including the transition metals.

REMINDER

For a larger version of the Periodic Table, including atomic numbers and other information, see Appendix A.

DID YOU KNOW?

Most International GCSE Periodic Tables stop at the end of the actinoid series, even though more elements continue to be discovered. Some tables completely omit the lanthanoids and actinoids because the emphasis in International GCSE is placed on a general understanding, which can easily be achieved with the main group and transition elements. Although sometimes called the rare earth metals, the lanthanoids are actually not that rare and are quite important in modern technology. Most of the actinoids are unstable and undergo radioactive decay.

EXTENSION WORK

The third shell can actually hold up to 18 electrons. To explain why 2 electrons go into the fourth shell before the third shell is completely filled requires a more advanced understanding of electronic configurations. Each shell is actually made up of a series of subshells and some of these subshells overlap from one shell to another. The 4s subshell (which can hold 2 electrons) is lower in energy than the 3d subshell (which can hold 10 electrons) and so 2 electrons go into the fourth shell before the next 10 go into the third shell.

The vertical columns in the Periodic Table are called **groups**. The first seven groups are numbered from 1 to 7 and the final group is numbered 0. The elements in orange in Figure 4.1 are called the transition metals or transition elements. At this level, they are not usually included in the numbering of the groups. Some of the groups have names, e.g. Group 1 is the *alkali metals*, Group 7 is the *halogens* and Group 0 is the *noble gases*.

The horizontal rows in the Periodic Table are called **periods**. It is important to remember that hydrogen and helium make up Period 1.

The lanthanoids and actinoids are usually dropped out of their proper places and written separately at the bottom of the Periodic Table. There is a good reason for this. If you put them where they should be (as in Figure 4.2), everything has to be drawn slightly smaller to fit on the page. That makes it more difficult to read.

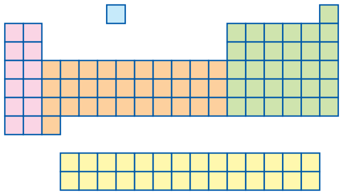


Figure 4.2 The real shape of the Periodic Table

THE PERIODIC TABLE AND THE NUMBER OF PROTONS, NEUTRONS AND ELECTRONS

Most Periodic Tables have two numbers against each symbol. 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 tell you which.

You can use a Periodic Table to work out the number of protons, neutrons and electrons there are in atoms. 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.

THE ARRANGEMENT OF THE ELECTRONS IN AN ATOM

The electrons move around the nucleus in a series of levels called **energy levels** or **shells**. Each energy level (shell) can only hold a certain number of electrons. Lower energy levels are always filled before higher ones. The lowest energy level is the closest one to the nucleus.

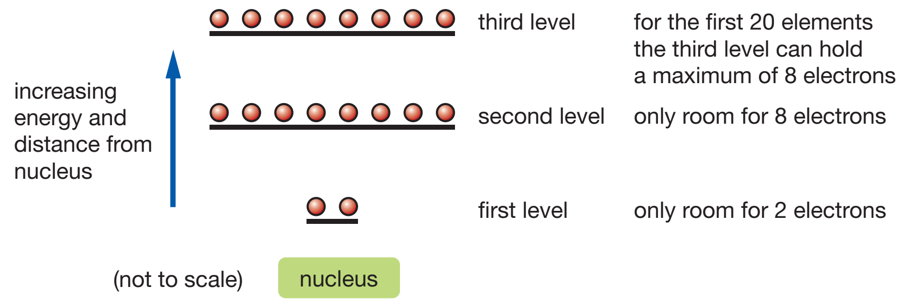


Figure 4.3 The different energy levels for electrons in an atom showing the maximum number of electrons that each energy level (shell) can hold.

HOW TO WORK OUT THE ARRANGEMENT OF ELECTRONS IN AN ATOM

The arrangement of electrons in an atom is called its **electronic configuration**.

We will use chlorine as an example.

Look up the atomic number in the Periodic Table. (Make sure that you don't use the wrong number if you have a choice. The atomic number will always be the smaller one.)

The Periodic Table tells you that chlorine's atomic number is 17.

This tells you the number of protons. The number of electrons is equal to the number of protons.

There are 17 protons and so 17 electrons in a neutral chlorine atom.

Arrange the electrons in shells (energy levels), always completing an inner shell (lower energy level) before you go to an outer one. Remember that the first shell (lowest energy level) can take up to 2 electrons, the second one can take up to 8, and the third one also takes up to 8.

For chlorine the electrons will be arranged as follows: 2 in the first shell, 8 in the second shell and 7 in the third shell. This is written as 2, 8, 7. When you have finished, always check to make sure that the electrons add up to the right number, in this case 17.

DRAWING DIAGRAMS OF ELECTRONIC CONFIGURATIONS

When we draw a diagram of an atom we usually draw circles to represent the shells (energy levels); dots or crosses are then drawn on the circles to represent the electrons. You can choose to draw dots or crosses.

Hydrogen has 1 electron and helium has 2 in the first shell (lowest energy level).

Figure 4.4 Electronic arrangements of hydrogen and helium

The helium electrons are sometimes shown as a pair (as here), and sometimes as two separate electrons on opposite sides of the circle. Either form is acceptable.

The next 8 atoms are drawn like this:

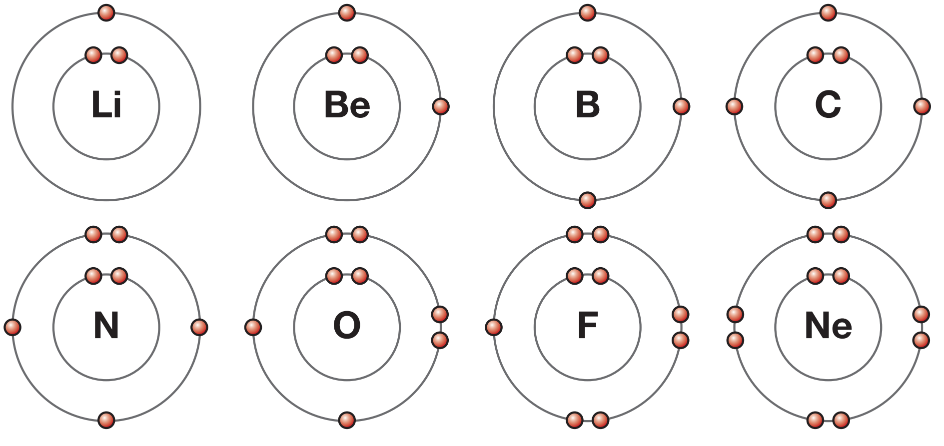


Figure 4.5 The electronic arrangements of the elements in Period 2

It does not matter at this level whether you draw the electrons singly or in pairs.

KEY POINT

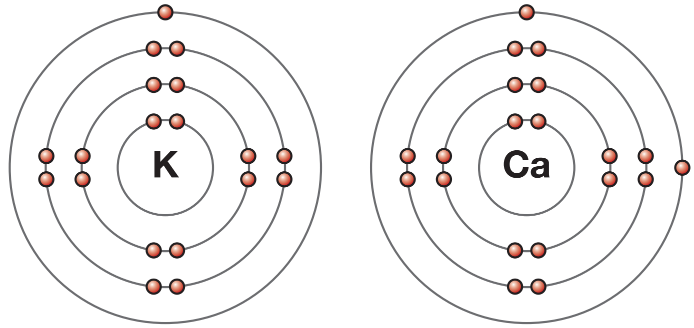
Drawing circles like this does not imply that the electrons are orbiting the nucleus along the circles.
The circles represent energy levels. The further away the level is from the nucleus, the higher its energy.
It is impossible to work out exactly how an electron is moving in that energy level.

HINT

The terms 'shell' and 'energy level' are used quite interchangeably by chemists. To make things easier, we will just use one term, shell, from now on. If, however, you prefer the term energy level, that is absolutely fine and both terms will be accepted in the examination.

The atoms in the Periodic Table from sodium to argon fill the third shell in exactly the same way, and potassium and calcium start to fill the fourth shell.

Potassium and calcium will look like this:



▲ Figure 4.6 Electronic arrangements of potassium and calcium

ELECTRONIC CONFIGURATIONS AND THE PERIODIC TABLE

The electronic configurations of the first 20 elements in the Periodic Table are shown in Figure 4.7.

Group 1	Group 2		Group 3	Group 4	Group 5	Group 6	Group 7	Group 0
		H 1						He 2
Li 2,1	Be 2,2		B 2,3	C 2,4	N 2,5	O 2,6	F 2,7	Ne 2,8
Na 2,8,1	Mg 2,8,2		Al 2,8,3	Si 2,8,4	P 2,8,5	S 2,8,6	Cl 2,8,7	Ar 2,8,8
K 2,8,8,1	Ca 2,8,8,2	10 more elements						

▲ Figure 4.7 The electronic configurations of the first 20 elements in the Periodic Table

Two important facts:

- ***Elements in the same group in the Periodic Table have the same number of electrons in their outer shell.***
- ***The number of electrons in the outer shell is the same as the group number for Groups 1 to 7.***

So if you know that barium is in Group 2, you know it has 2 electrons in its outer shell, the same as all the other elements in Group 2. Iodine (Group 7) has 7 electrons in its outer shell. Lead (Group 4) has 4 electrons in its outer shell. Working out what is in their inner shells is much more difficult – the simple patterns we have described don't work beyond calcium.

The period number gives the number of occupied shells or the highest occupied shell.

So, calcium (electronic configuration = 2, 8, 8, 2) is in Period 4 and has four shells occupied; the outermost electron is in the fourth shell. Iodine is in Period 5 and has five occupied shells, and because it is also in Group 7 we can deduce that it has 7 electrons in the fifth shell.

ELEMENTS IN THE SAME GROUP IN THE PERIODIC TABLE HAVE SIMILAR CHEMICAL PROPERTIES

Groups in the Periodic Table contain elements with similar chemical properties – they react in the same way.

For example:

- all the elements in Group 1 react vigorously with water to form hydrogen and hydroxides with similar formulae: LiOH (lithium hydroxide), NaOH (sodium hydroxide), KOH (potassium hydroxide)
- all the elements in Group 7 react with hydrogen to form compounds with similar formulae: HF (hydrogen fluoride), HCl (hydrogen chloride), HBr (hydrogen bromide)
- all the elements in Group 2 form chlorides with similar formulae: $MgCl_2$, $CaCl_2$.

The reactions of atoms depend on how many electrons there are in their outer shell. These are the electrons which normally get involved when elements bond to other elements. Elements in the same group (apart from helium in Group 0) have the same number of electrons in their outer shell, therefore they react in similar ways.

THE NOBLE GASES

KEY POINT

The reason the noble gas group is usually called Group 0 and not Group 8 (which would seem more sensible because most of the elements have 8 electrons in their outer shell) is because when they were first discovered it was believed that noble gases did not combine with anything, they had zero combining power (valency). This is the only group where the group number does not tell you the number of electrons in the outer shell! The key point is that the noble gases (except helium) have 8 electrons in their outer shell. You will see in Chapters 7 and 8 that atoms tend to form compounds by losing/gaining or sharing electrons so that they have 8 electrons in their outer shell. The noble gases already have 8 electrons in their outer shell so they do not do that.

The Group 0 elements are known as the **noble gases** because they are almost completely unreactive, in fact the two at the top of the group, helium and neon, don't react with anything. The elements in Group 0 have 8 electrons in their outer shell (apart from helium, which has 2).

The lack of reactivity of the elements in Group 0 is associated with their electronic configurations. The noble gases are unreactive because *the outer shell is full, and so there is no tendency to lose, gain or share electrons in a chemical reaction*. You will learn more about how elements form compounds in Chapters 7 and 8.

He
Ne
Ar
Kr
Xe
Rn

▲ Figure 4.8 The noble gases

EXTENSION WORK

In reality, the outer shells of argon, krypton, xenon and radon are not full; in fact only helium and neon have full outer shells. The third shell can actually hold up to 18 electrons and the fourth up to 32 electrons.

METALS AND NON-METALS IN THE PERIODIC TABLE

The metals are on the left-hand side of the Periodic Table and the non-metals are on the right-hand side. Although the division into metals and non-metals is shown clearly in Figure 4.9, in practice there is a lot of uncertainty on the dividing line. For example, arsenic (As) has properties of both metals and non-metals.

KEY POINT

You may have noticed by now that hydrogen does not really fit into the Periodic Table properly. Although it has 1 electron in its outer shell, it is not in Group 1 because it does not have similar properties to the other elements in Group 1: it is not a metal and does not have similar chemical properties to the alkali metals.

EXTENSION WORK

If you are interested you could try an internet search for ‘metallic hydrogen’.

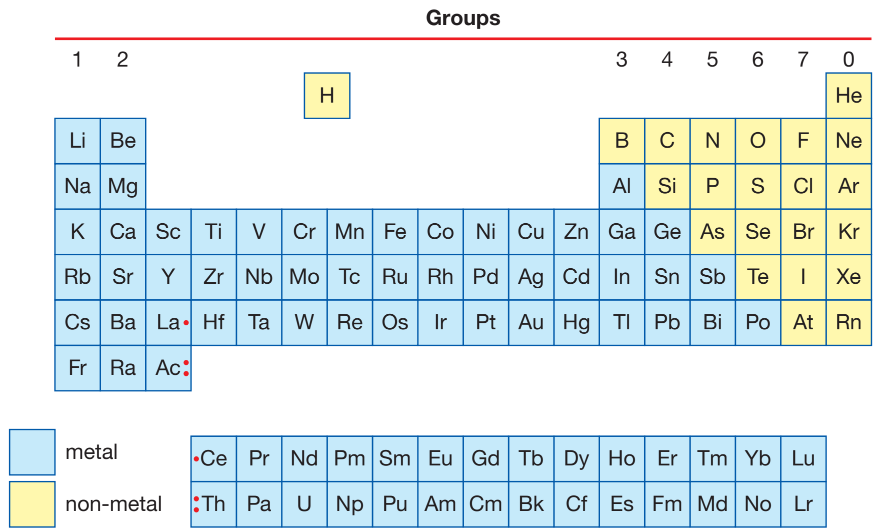


Figure 4.9 Metals and non-metals

DIFFERENCES BETWEEN THE PROPERTIES OF METALS AND NON-METALS

There are many differences between the properties of metals and non-metals. Here, we will use two main ones to classify them: electrical conductivity and the acid–base character of their oxides.

Metals conduct electricity and non-metals do not generally conduct electricity. The reason that metals conduct electricity will be explained in Chapter 9 (it is due to the presence of delocalised electrons that are free to move). You may be familiar with the use of metals such as copper in electrical wires.

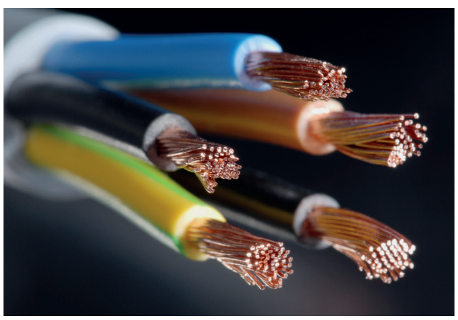


Figure 4.10 Copper, like all other metals, conducts electricity.

Non-metals do not conduct electricity (because there are no electrons that are free to move or mobile ions) but there are a few exceptions, such as graphite (a form of carbon) and silicon.

Metals generally form basic oxides. A basic oxide is one which reacts with acids to form salts. For example, copper forms copper(II) oxide (CuO). This reacts, for example, with sulfuric acid to form the salt copper(II) sulfate:

$$\underset{\text{copper(II) oxide}}{CuO} + \underset{\text{sulfuric acid}}{H_2SO_4} \rightarrow \underset{\text{copper(II) sulfate}}{CuSO_4} + \underset{\text{water}}{H_2O}$$

The reactions between bases and acids will be discussed in more detail in Chapter 17.

KEY POINT

If a basic oxide is soluble in water it will dissolve to form an alkali. For example, sodium oxide reacts with water to form sodium hydroxide solution, an alkali. If an acidic oxide is soluble in water it will dissolve to form an acidic solution. For example, sulfur(IV) oxide reacts with water to form sulfurous acid.

Non-metals generally form acidic oxides. You may be familiar, for example, with sulfur(IV) oxide (SO_2) being one of the gases responsible for acid rain. Acidic oxides react with bases/alkalis to form salts. For example, carbon dioxide, an acidic oxide, reacts with sodium hydroxide, an alkali:

$$\underset{\text{carbon dioxide}}{CO_2} + \underset{\text{sodium hydroxide}}{2NaOH} \rightarrow \underset{\text{sodium carbonate}}{Na_2CO_3} + \underset{\text{water}}{H_2O}$$

Some other properties you might also associate with metals and non-metals are included in the list on the next page.

KEY POINT

There are some exceptions to the rules. For example, some metals form amphoteric oxides (e.g. Al_2O_3), which react with acids and bases, and some non-metal oxides (e.g. CO) are neutral.

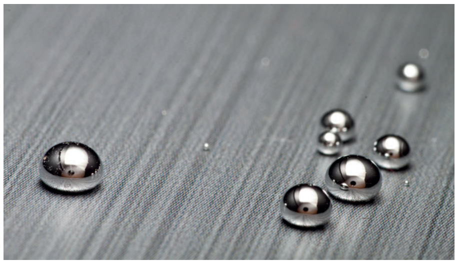


Figure 4.11 Mercury has most of the properties of a metal (high density, shiny, conducts electricity, forms positive ions); the exception is that it is a liquid at room temperature.

Figure 4.12 Sulfur crystals are shiny, but you wouldn't mistake them for a metal.

Metals:

- tend to be solids with high melting and boiling points, and with relatively high densities (but as with several of the properties in this list, there are exceptions, for example mercury is a liquid)
- are shiny (have a metallic lustre) when they are polished or freshly cut
- are **malleable** (can be hammered into shape)
- are **ductile** (can be drawn into wires)
- are good conductors of electricity and heat
- form ionic compounds (see Chapter 7)
- form positive ions in their compounds (see Chapter 7).

Non-metals:

- tend to have low melting and boiling points (there are some exceptions, e.g. carbon and silicon)
- tend to be brittle when they are solids
- don't have the same type of shine as metals
- don't usually conduct electricity; carbon (in the form of graphite) and silicon are again exceptions
- are poor conductors of heat (diamond is an exception; it is the best conductor of heat of all the elements)
- form both ionic and covalent compounds (see Chapters 7 and 8)
- tend to form negative ions in ionic compounds (see Chapter 7).

CHAPTER QUESTIONS

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

SKILLS CRITICAL THINKING 6

1 Answer the questions that follow using only the elements in this list:

caesium, chlorine, molybdenum, neon, nickel, nitrogen, strontium, tin.

a State the name of an element which is:

i in Group 2
ii in the same period as silicon
iii in the same group as phosphorus
iv in Period 6
v a noble gas

b Divide the list of elements at the beginning of the question into two groups, metals and non-metals.

SKILLS INTERPRETATION 8

2 Draw diagrams to show the arrangement of the electrons in:

a sodium
b silicon
c sulfur

SKILLS CRITICAL THINKING

8

3 State the electronic configurations of the following atoms:

a fluorine

b aluminium

c calcium

SKILLS PROBLEM SOLVING

9

4 Find each of the following elements in the Periodic Table and state the number of electrons in its outer shell:

a arsenic, As

b bromine, Br

c tin, Sn

d xenon, Xe

SKILLS REASONING

8

5 The questions refer to the electronic configurations below. Don't worry if some of these are unfamiliar to you. All of these are the electronic configurations of neutral atoms.

A 2, 4

B 2, 8, 8

C 2, 8, 18, 18, 7

D 2, 8, 18, 18, 8

E 2, 8, 8, 2

F 2, 8, 18, 32, 18, 4

a Explain which of these atoms are in Group 4 of the Periodic Table.

b State which of these electronic configurations represents carbon.

c Explain which atoms are in Period 5 of the Periodic Table.

d Explain which of these electronic configurations represents an element in Group 7 of the Periodic Table.

e State which of these electronic configurations represent noble gases.

SKILLS PROBLEM SOLVING

f State the name of element **E** and explain how you arrived at your answer.

g State how many protons are present in an atom of element **F**. State the name of the element.

SKILLS INTERPRETATION

h Element **G** has one more electron than element **B**. Draw a diagram to show how the electrons are arranged in an atom of **G**.

SKILLS REASONING

6 Predict two properties of the element palladium, Pd (atomic number 46), or its compounds. The properties can be either physical or chemical.

7 Helium and neon do not form any compounds. Explain why the noble gases are unreactive.

9

8 The elements in the Periodic Table are arranged in order of atomic number. If they were arranged in order of mass number give the names of two elements that would be in different positions. Explain why this would cause a problem.

5 CHEMICAL FORMULAE, EQUATIONS AND CALCULATIONS: PART 1

It is often important in chemistry to know how much of one substance reacts with a certain amount of another. To do this, we will learn about moles in this chapter.

▲ Figure 5.1 Each gold bar contains almost 4×10^{25} gold atoms. That's 4 followed by 25 noughts! Learning about the mole allows us to work out how many atoms are in something.

LEARNING OBJECTIVES

- Write word equations and balanced chemical equations (including state symbols):
  - for reactions studied in this course
  - for unfamiliar reactions where suitable information is provided.
- Calculate relative formula masses (including relative molecular masses) (M_r) from relative atomic masses (A_r).
- Know that the mole (mol) is the unit for the amount of a substance.
- Understand how to carry out calculations involving amount of substance, relative atomic mass (A_r) and relative formula mass (M_r).
- Calculate reacting masses using experimental data and chemical equations.
- Calculate percentage yield.
- Understand how the formulae of simple compounds can be obtained experimentally, including metal oxides, water and salts containing water of crystallisation.
- Know what is meant by the terms empirical formula and molecular formula.
- Calculate empirical and molecular formulae from experimental data.
- Practical: Know how to determine the formula of a metal oxide by combustion (e.g. magnesium oxide) or by reduction (e.g. copper(II) oxide).

WRITING EQUATIONS

There are two types of chemical equation that you could be asked to write: word equations and symbol equations. Symbol equations are usually called *chemical equations* and you should only write a word equation if you are specifically asked to. All chemical equations must be balanced.

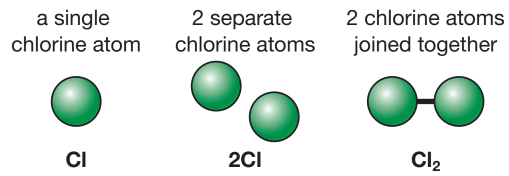
WHAT ALL THE NUMBERS MEAN

An example of a balanced chemical equation is:

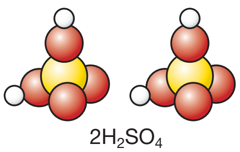
$$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$$

When you write equations, it is important to be able to count how many of each type of atom you have. In particular, you must understand the difference between big numbers written in front of formulae (sometimes called **coefficients**), such as the **2** in **2**HCl, and the smaller, subscript (slightly lower on the line) numbers, such as the **3** in $CaCO_3$.

Another chemical equation is $2Cl \rightarrow Cl_2$. Here, **2Cl** represents 2 separate Cl atoms and **Cl_2** means that the atoms are joined together in a molecule.



▲ Figure 5.2 The difference between 2Cl and Cl_2



▲ Figure 5.3 $2H_2SO_4$. The sulfur atoms are shown in yellow, the oxygens in red and the hydrogens in white.

Another balanced chemical equation is $H_2S_2O_7 + H_2O \rightarrow 2H_2SO_4$.

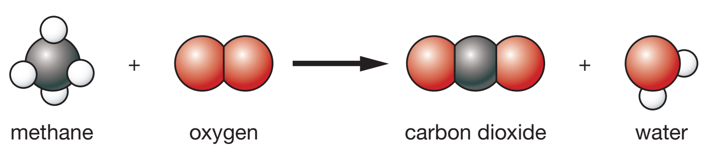
Look at the way the numbers work in $2H_2SO_4$. The big number in front tells you that you have 2 sulfuric acid (H_2SO_4) molecules. The subscript 4 tells you that you have 4 oxygen atoms in each molecule. A small, subscript number in a formula applies only to the atom immediately before it in the formula. If you count the atoms in $2H_2SO_4$ you will find 4 hydrogens, 2 sulfurs and 8 oxygens.

If you have brackets in a formula, the small number refers to everything inside the brackets. For example, in the formula $Ca(OH)_2$, the 2 applies to both the oxygen and the hydrogen. The formula shows 1 calcium, 2 oxygens and 2 hydrogens. You will learn more about how to work out the formulae of compounds in Chapter 7.

BALANCING EQUATIONS

Chemical reactions involve taking elements or compounds and moving their atoms around into new combinations. It follows that you must always end up with the same number of atoms that you started with.

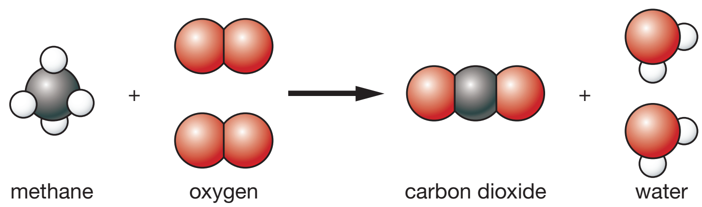
Imagine you had to write an equation for the reaction between methane, CH_4, and oxygen, O_2. Methane burns in oxygen to form carbon dioxide and water. Think of this in terms of rearranging the atoms in some models.



▲ Figure 5.4 Rearranging the atoms in methane and oxygen

If you count the atoms you had at the beginning (on the left-hand side of the arrow) and the atoms you have at the end (on the right-hand side of the arrow), you can see that this can't be right! During the rearrangement, we seem to have gained an oxygen atom and lost two hydrogens. The reaction must be more complicated than this. Since the substances are all correct, the proportions must be wrong.

Try again:



▲ Figure 5.5 A balanced equation for the reaction between methane and oxygen.

There are now the same number of each type of atom before and after. This is called **balancing the equation**.

In symbols, this equation would be:

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

Think of each symbol (C or H or O) as representing one atom of that element. Count them up in the equation, and check that there is the same number of atoms on both sides.

HOW TO BALANCE EQUATIONS

In order to balance equations you should adopt a systematic approach.

- Work across the equation from left to right, checking one element after another, except if an element appears in several places in the equation. In that case, leave the element until the end and you will often find that it has sorted itself out.
- If you have a group of atoms (like a sulfate group (SO_4), for example), which is unchanged from one side of the equation to the other, there is no reason why you can't just count that up as a whole group, rather than counting individual sulfurs and oxygens. It saves time.
- Check everything at the end to make sure you haven't changed something that you have already counted.

HINT

This is really important! You must never, never change a formula when balancing an equation. All you are allowed to do is to write big numbers in front of the formula.

EXAMPLE 1

Balance the equation for the reaction between zinc and hydrochloric acid:

$$Zn + HCl \rightarrow ZnCl_2 + H_2$$

Work from left to right. Count the zinc atoms: 1 on each side; no problem!

Count the hydrogen atoms: 1 on the left, 2 on the right. If you have 2 at the end, you must have started with 2. The only way of achieving this is to have 2HCl. (You must not change the formula to H_2Cl because this substance does not exist.)

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

Now count the chlorines: there are 2 on each side. Good! Finally check everything again to make sure and you've finished.

	$Zn + 2HCl$		$\rightarrow$	$ZnCl_2 + H_2$	
numbers of atoms	Zn	1		Zn	1
	H	2		H	2
	Cl	2		Cl	2

EXAMPLE 2

Balance the equation for the combustion of ethane:

$$C_2H_6 + O_2 \rightarrow CO_2 + H_2O$$

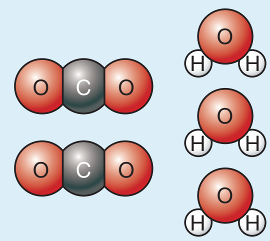
Starting from the left, balance the carbons:

$$C_2H_6 + O_2 \rightarrow 2CO_2 + H_2O$$

Now the hydrogens:

$$C_2H_6 + O_2 \rightarrow 2CO_2 + 3H_2O$$

Finally the oxygens: there are 7 oxygens ((2 × 2) + 3) on the right-hand side, but only 2 on the left. The problem is that the oxygens have to go around in pairs. So how can you obtain an odd number (7) of oxygens on the left-hand side?



$2CO_2$ $3H_2O$

▲ Figure 5.6 There are 7 O atoms in $2CO_2 + 3H_2O$

The trick with this is to allow yourself to have halves in your equation. 7 oxygen atoms, O, is the same as $3\frac{1}{2}$ oxygen molecules, O_2.

$$C_2H_6 + 3\tfrac{1}{2}O_2 \rightarrow 2CO_2 + 3H_2O$$

Now double everything to get rid of the half:

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

KEY POINT

In fact, it is acceptable to have halves in equations, but you don't usually come across them at Internaional GCSE.
You might reasonably argue that you can't have half an oxygen molecule, but to remove that problem all you have to do is double everything.

STATE SYMBOLS

State symbols are often, but not always, written after the formulae of the various substances in an equation to show what physical state everything is in. You need to know four different state symbols:

(s) solid (l) liquid (g) gas (aq) in **aqueous** solution (dissolved in water)

So an equation might look like this:

$$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$$

This shows that *solid* potassium reacts with *liquid* water to make a *solution* of potassium hydroxide in water and hydrogen *gas*.

HINT

Don't worry if this chemistry is new to you, or if at this stage you don't know what the state symbols should be. That is not important at the moment.

HINT

Remember that water is a liquid (l), not an aqueous solution (aq). An aqueous solution is formed when something is *dissolved in water*.

HOW MUCH OF EACH SUBSTANCE REACTS IN A CHEMICAL REACTION?

You can make iron(II) sulfide by heating a mixture of iron and sulfur:

$$Fe + S \rightarrow FeS$$

How do you know what proportions to mix them in? You can't just mix equal masses of them because iron and sulfur atoms don't weigh the same. Iron atoms contain more protons and neutrons than sulfur atoms, so an iron atom is one and three-quarter times heavier than a sulfur atom. In this or any other reaction, you can get the right proportions only if you know about the masses of the individual atoms that take part in the reaction.

RELATIVE ATOMIC MASS (A_r)

We have already looked at how to calculate relative atomic masses from the isotopic abundances in Chapter 3. Here we will look a little more closely at what exactly the relative atomic mass is.

REMINDER

Remember that isotopes are atoms of the same element, but with different masses. Isotopes are explained in Chapter 3 (page 26).

EXTENSION WORK

This is a slight approximation. To be accurate, each of these hydrogen atoms has a mass of 1.008 on the carbon-12 scale. For International GCSE purposes, we take it as being exactly 1.

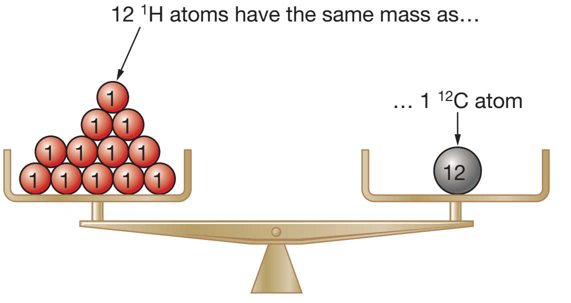
KEY POINT

Remember, to work out a weighted average we have to know how much of each isotope is present in a sample.

HINT

Avoid the term 'relative molecular mass' because it can only properly be applied to substances which are actually molecules, that is, to covalent substances. You shouldn't use it for things like magnesium oxide or sodium chloride, which are ionic. Relative formula mass covers everything.

Atoms are amazingly small. The mass of a hydrogen atom is about 1.67×10^{-24} g (0.00000000000000000000000167 g). It is really difficult to use numbers such as this and so we use a scale of *relative* masses instead. The masses of atoms (and molecules) are compared with the mass of an atom of the carbon-12 isotope. We call this the carbon-12 scale. On this scale, one atom of the carbon-12 isotope weighs *exactly* 12 units and the mass of the most common hydrogen isotope is 1, which is a much simpler number to use!



▲ Figure 5.7 The most common hydrogen atom weighs one-twelfth as much as a ^{12}C atom.

The basic unit on the carbon-12 scale is $\frac{1}{12}$ of the mass of a ^{12}C atom, which is approximately the mass of the most common hydrogen atom. A fluorine-19 atom has a relative mass of 19 because its atoms have a mass 19 times that of $\frac{1}{12}$ of a ^{12}C atom. An atom of the most common isotope of magnesium weighs 24 times as much as $\frac{1}{12}$ of a ^{12}C atom, and is therefore said to have a relative mass of 24.

The masses we are talking about here are the masses of individual *isotopes*, but samples of an actual element contain different isotopes and so we need a measure of the average mass of an *atom* taking into account the different isotopes. This is the relative atomic mass.

The relative atomic mass of an element (as opposed to one of its isotopes) is given the symbol $\boldsymbol{A_r}$ and it is defined like this:

The relative atomic mass of an element is the weighted average mass of the isotopes of the element. It is measured on a scale on which a carbon-12 (^{12}C) atom has a mass of exactly 12.

Because we are talking here about *relative* masses they have no units.

On this scale, the relative atomic mass of chlorine is 35.45, that of lithium is 6.94 and that of sodium is 22.99 because we are taking into account the different isotopes, and are quoting an average mass for an atom. Although all elements consist of a mixture of isotopes, at International GCSE we only use relative atomic masses, including decimal places for Cl (35.5) and Cu (63.5), therefore we will take the relative atomic mass of lithium as 7 and that of sodium as 23.

RELATIVE FORMULA MASS (M_r)

You can measure the masses of compounds on the same carbon-12 scale. For example, a water molecule, H_2O, has a mass of 18 on the carbon-12 scale. This means that a water molecule has 18 times the mass of $\frac{1}{12}$ of a ^{12}C atom.

If you are talking about compounds, you use the term *relative formula mass*. Relative formula mass is sometimes called relative molecular mass.

Relative formula mass is given the symbol $\boldsymbol{M_r}$.

CALCULATING SOME RELATIVE FORMULA MASSES

To find the relative formula mass (M_r) of magnesium carbonate, $MgCO_3$

Relative atomic masses: C = 12, O = 16, Mg = 24.

Add up the relative atomic masses to give the relative formula mass of the whole compound. In this case, you need to add up the masses of 1 × Mg, 1 × C and 3 × O.

$$M_r = \underset{Mg}{24} + \underset{C}{12} + \underset{(3 \times O)}{(3 \times 16)}$$

$$M_r = 84$$

To find the relative formula mass of calcium hydroxide, $Ca(OH)_2$

Relative atomic masses: H = 1, O = 16, Ca = 40.

$$M_r = \underset{Ca}{40} + (\underset{O}{16} + \underset{H}{1}) \times 2$$

$$M_r = 74$$

To find the M_r of copper(II) sulfate crystals, $CuSO_4{\cdot}5H_2O$

This formula looks very different and a lot more complicated than the ones we have looked at above. When some substances crystallise from solution, water becomes chemically bound up with the salt. This is called **water of crystallisation**. The $5H_2O$ is water of crystallisation. The water is necessary to form crystals of copper(II) sulfate (and some other substances). There are always 5 water molecules associated with 1 $CuSO_4$ unit and they are part of the formula. Salts containing water of crystallisation are said to be **hydrated**. Other examples include $Na_2CO_3{\cdot}10H_2O$ and $MgCl_2{\cdot}6H_2O$.

Relative atomic masses: H = 1, O = 16, S = 32, Cu = 63.5.

It is easiest to work out the relative formula mass of water first:

$$H_2O: \quad M_r = 2 \times 1 + 16 = 18$$

Now add the correct number of waters on to the rest of the formula:

$$CuSO_4{\cdot}5H_2O: \quad M_r = \underset{Cu}{63.5} + \underset{S}{32} + \underset{(4 \times O)}{(4 \times 16)} + \underset{(5 \times H_2O)}{(5 \times 18)} = 249.5$$

HINT

Remember, you do not have to calculate the relative atomic masses, they will always be given to you in an exam, either in the question or on the Periodic Table. **If you use a Periodic Table, be sure to use the right number!** It will always be the larger of the two numbers given.

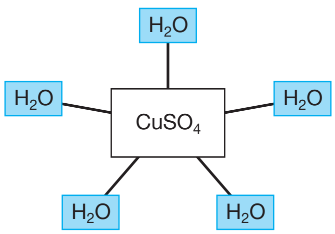


Figure 5.8 A schematic diagram showing the formula of copper(II) sulfate crystals: there are five water molecules bonded to each copper(II) sulfate unit.

HINT

It is dangerous to calculate the number of hydrogens and oxygens separately. The common mistake is to work out 10 hydrogens (quite correctly!), but then only count 1 oxygen rather than 5.

USING RELATIVE FORMULA MASS TO FIND PERCENTAGE COMPOSITION

Having found the relative formula mass of a compound, we can work out the percentage by mass of any part of it. Examples make this clear.

To find the percentage by mass of copper in copper(II) oxide, CuO

Relative atomic masses: O = 16, Cu = 63.5.

$$M_r \text{ of CuO} = 63.5 + 16 = 79.5$$

Of this, 63.5 is copper.

$$\text{Percentage of copper} = \frac{63.5}{79.5} \times 100$$

$$= 79.9\%$$

To find the percentage by mass of oxygen in sodium carbonate, Na_2CO_3

Relative atomic masses: Na = 23, C = 12, O = 16.

$$M_r = (2 \times 23) + 12 + (3 \times 16)$$

$$= 106$$

48 of this 106 is due to the oxygen (remember that there are 3 oxygens in the formula, so the total mass of oxygen is 3 × 16 = 48).

$$\text{Percentage of oxygen} = \frac{3 \times 16}{106} \times 100$$
$$= 45.3\%$$

Figure 5.9 1-mole quantities (clockwise from upper left) of carbon (12 g), sulfur (32 g), iron (56 g), copper (63.5 g) and magnesium (24 g).

THE MOLE

In chemistry, the **mole** is a unit of the amount of substance. We can talk about an amount of substance in grams or an amount of substance in moles. The difference between expressing the amount of substance in moles or in grams is that 1 mole of any substance has its own particular mass. For example, 1 mole of water has a mass of 18 g, 1 mole of sulfur has a mass of 32 g and 1 mole of magnesium oxide has a mass of 40 g. The abbreviation for mole is **mol**.

We can therefore talk, for example, about an amount of water as 36 g or 2 mol, since 1 mol of water has a mass of 18 g. 36 g of sulfur, however, would only be 1.125 mol and 36 g of magnesium would be 1.5 mol because the mass of 1 mol of sulfur is 32 g and that of 1 mol of magnesium is 24 g.

Some people talk about the *amount in moles* and others talk about the *number of moles*. This is like talking about the *mass of a substance* or the *number of grams*. For example, we could say that the amount in grams of water is 36 g or the number of grams of water is 36 g. Similarly, we can say that the amount in moles of water is 2 mol or the number of moles of water is 2 mol.

CALCULATING THE MASSES OF A MOLE OF SUBSTANCE

You find the mass of 1 mole of a substance by calculating the relative formula mass (M_r) and attaching the units, grams.

1 mole of oxygen gas, O_2

Relative atomic mass: O = 16.

$$M_r \text{ of } O_2 = 2 \times 16$$
$$= 32$$

1 mole of oxygen, O_2, has a mass of 32 g.

1 mole of calcium chloride, $CaCl_2$

Relative atomic masses: Cl = 35.5, Ca = 40.

$$M_r \text{ of } CaCl_2 = 40 + (2 \times 35.5)$$
$$= 111$$

1 mole of calcium chloride has a mass of 111 g.

1 mole of iron(II) sulfate crystals, $FeSO_4{\cdot}7H_2O$

Relative atomic masses: H = 1, O = 16, S = 32, Fe = 56.

$$M_r \text{ of crystals} = 56 + 32 + (4 \times 16) + \{7 \times [(2 \times 1) + 16]\}$$

1 mole of iron(II) sulfate crystals has a mass of 278 g.

THE IMPORTANCE OF QUOTING THE FORMULA

Whenever you talk about a mole of something, you *must* quote its formula, otherwise there is a risk of confusion.

For example, if you talk about 1 mole of oxygen, this could mean:

1 mole of oxygen atoms, O, with a mass of 16 g

1 mole of oxygen molecules, O_2, with a mass of 32 g.

HINT

We often use the term ‘molar mass’ instead of ‘mass of 1 mole’.

EXTENSION WORK

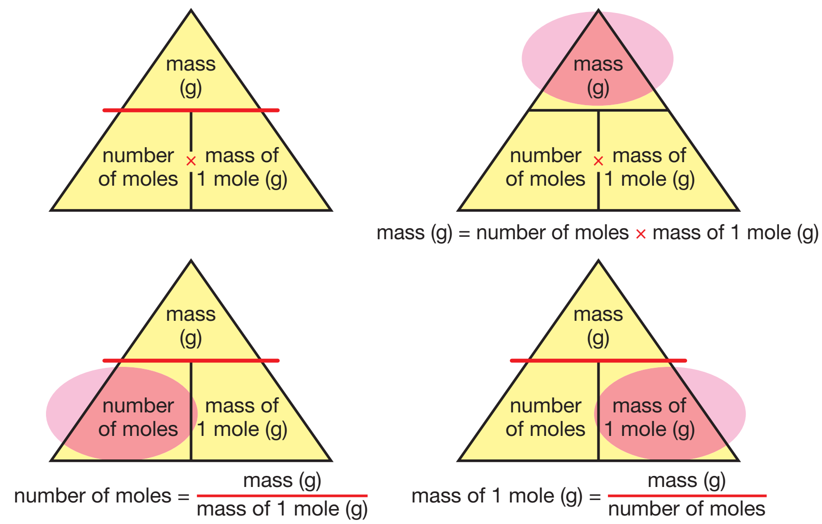
To be 100% correct 'molar mass' and 'mass of 1 mole' are not exactly the same thing and have different units: molar mass has units of g/mol and the mass of 1 mole has units of g. You are unlikely to come across the distinction between these at International GCSE.

SIMPLE CALCULATIONS WITH MOLES

You need to be able to interconvert between a mass in grams and an amount in moles for a given substance. There is a simple formula that you can learn:

$$\text{number of moles} = \frac{\text{mass (g)}}{\text{mass of 1 mole (g)}}$$

You can rearrange this formula to find whatever you want. If rearranging this expression causes you problems, you can learn a simple triangular arrangement to help you.



▲ Figure 5.10 Converting between mass in grams and number of moles.

Look at this carefully and make sure that you understand how you can use it to deduce the three equations that you might need.

Finding the mass of 0.2 mol of calcium carbonate, $CaCO_3$

Relative atomic masses: C = 12, O = 16, Ca = 40.

First find the relative formula mass of calcium carbonate.

$$\begin{aligned} M_r \text{ of } CaCO_3 &= 40 + 12 + (3 \times 16) \\ &= 100 \end{aligned}$$

1 mol of $CaCO_3$ has a mass of 100 g.

$$\begin{aligned} \text{Mass (g)} &= \text{number of moles} \times \text{mass of 1 mole (g)} \\ &= 0.2 \times 100 \\ &= 20\,\text{g} \end{aligned}$$

0.2 mol of $CaCO_3$ has a mass of 20 g.

HINT

Don’t get too reliant on learning formulae for doing simple calculations. If 1 mole weighs 100 g, it should be fairly obvious to you that 2 moles will weigh twice as much, 10 moles will weigh 10 times as much and 0.2 moles will weigh 0.2 times as much.

Finding the number of moles in 54 g of water, H_2O

Relative atomic masses: H = 1, O = 16.

1 mol of H_2O has a mass of 18 g.

$$\begin{aligned} \text{Number of moles} &= \frac{\text{mass (g)}}{\text{mass of 1 mole (g)}} \\ &= \frac{54}{18} \\ &= 3\,\text{mol} \end{aligned}$$

54 g of water is 3 mol.

HINT

If 1 mol weighs 18 g, then you need to find out how many times 18 goes into 54 to find out how many moles you have. On the other hand, if you feel safer using a formula, use it.

EXTENSION WORK

Moles and the Avogadro constant

Imagine you had 1 mole of ^{12}C. It would have a mass of 12 g and contain a very large number of carbon atoms, in fact about 6×10^{23} carbon atoms, which is 6 followed by 23 noughts. This number of atoms in 12 g of ^{12}C is called the **Avogadro constant**.

1 mole of anything else contains this same number of particles. For example:

1 mole of magnesium contains 6×10^{23} magnesium atoms, Mg, and has a mass of 24 g.

1 mole of water contains 6×10^{23} water molecules, H_2O, and has a mass of 18 g.

This is the reason that we use moles: if we know the number of moles we also know how many particles are present. If we know how many particles are present, we can work out how much of one substance reacts with a certain amount of another.

▲ Figure 5.11 Approximately 1×10^{32} water molecules go over Niagara Falls every second during the summer. That's 100 million million million million million water molecules per second.

FORMULAE

The formula of sulfur dioxide is SO_2. We can see that there are 2 O atoms for each S atom. If we had 1 mol SO_2 we would still have twice as many O atoms as S atoms; there would be 1 mol S atoms and 2 mol O atoms. If we know that in a certain sample of sulfur dioxide we have 0.1 mol of S atoms we also know that we must have 0.2 mol O atoms. 0.2 mol contains twice as many atoms as 0.1 mol.

If we did an experiment and found that a compound contained 0.2 mol Ca and 0.2 mol O, then we could work out that the formula must be CaO because there are the same number of Ca atoms as O atoms.

If a compound contains 0.4 mol Mn and 0.8 mol O then the formula must be MnO_2 because there are twice as many O atoms as Mn atoms.

The formulae that we have found here are called *empirical formulae*.

The* empirical formula *shows the simplest whole number ratio of the atoms present in a compound.

There is another type of formula that we will come across, called the *molecular formula*.

The* molecular formula *shows the actual number of atoms of each element present in a molecule (covalent compound) or formula unit (ionic compound) of a compound.

The molecular formula can be the same as the empirical formula or a multiple of the empirical formula. The empirical formula of calcium oxide is CaO and the molecular formula is also CaO. The empirical formula of hydrogen peroxide is HO and the molecular formula is H_2O_2.

In order to work out the molecular formula from the empirical formula we need more information – the M_r of the compound. We will look at this again below.

WORKING OUT EMPIRICAL FORMULAE

In order to find out the empirical formula of a compound such as copper oxide, we need to know how many atoms of copper combine with how many atoms of oxygen. We can work out the number of atoms from the number of moles. If we know the ratio between the number of copper atoms and oxygen atoms in the compound we know the formula.

KEY POINT

0.02 mol of atoms is twice as many atoms as 0.01 mol of atoms.

HINT

More significant figures should have been used for the number of moles in the examples on this page but these have been omitted for simplicity.

EXAMPLE 3

A sample of a compound contains 1.27 g of Cu and 0.16 g of O. Calculate the empirical formula. (A_r of Cu = 63.5, A_r of O = 16)

It is easiest to do this in columns using a table:

	Cu	O
masses/g	1.27	0.16
find the number of moles of atoms by dividing the mass by the mass of 1 mole of atoms	1.27/63.5	0.16/16
number of moles of atoms	0.02	0.01
divide by the smaller number to find the ratio	0.02/0.01	0.01/0.01
ratio of moles	2	1
empirical formula	Cu_2O	

From calculating the number of moles of Cu and O we can see that there are twice as many moles of Cu and therefore there must be twice as many Cu atoms as O atoms in the compound.

It is important to remember in this calculation that we are working out how many *atoms* of copper combine with how many *atoms* of oxygen so we divide the mass of oxygen by 16 (the mass of 1 mole of oxygen atoms) rather than by 32 (the mass of 1 mole of oxygen molecules).

EXAMPLE 4

A sample of a compound contains 0.78 g of K, 1.10 g of Mn and 1.28 g of O. (A_r of K = 39, A_r of Mn = 55, A_r of O = 16)

Again, we use a table:

	K	Mn	O
masses/g	0.78	1.10	1.28
find the number of moles of atoms by dividing the mass by the mass of 1 mole of atoms	0.78/39	1.10/55	1.28/16
number of moles of atoms	0.02	0.02	0.08
divide by the smallest number to find the ratio	0.02/0.02	0.02/0.02	0.08/0.02
ratio of moles	1	1	4
empirical formula	$KMnO_4$		

The empirical formula of this compound is $KMnO_4$.

Safety Note: Wear eye protection and take care not to get burnt when raising the crucible lid with tongs. The crucible will stay very hot for some time.

KEY POINT

How do we know when the reaction has finished? We can generally see when there is no more reaction because nothing is happening in the crucible. We can also check by weighing the crucible and contents, then heating it and weighing again. If the mass is higher when we weigh it again then this indicates that a reaction has occurred. We keep heating and weighing until the mass stays constant.

ACTIVITY 3

PRACTICAL: INVESTIGATING THE FORMULA OF A METAL OXIDE BY COMBUSTION

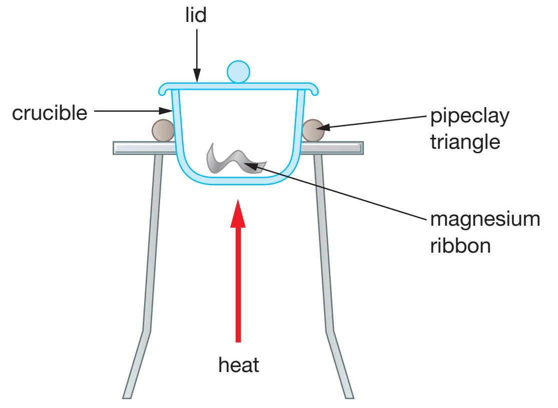


Figure 5.12 Apparatus for determining the formula of magnesium oxide

We can find the formula of magnesium oxide by burning magnesium in oxygen and looking at how the mass changes. The following procedure is usually used:

- Weigh a crucible with a lid.
- Place a piece of magnesium ribbon about 10 cm long in the crucible and weigh the crucible and contents.
- Set up the apparatus as shown in Figure 5.12.
- Heat the crucible strongly (a roaring flame).
- Lift the lid every few seconds.
- When the reaction is finished, allow the crucible and contents to cool.
- Weigh the crucible and contents.

When the magnesium burns it does so with a bright white flame. Magnesium oxide, a white powder, is produced in the reaction.

A lid is placed on the crucible to prevent the white powder (magnesium oxide) escaping but the lid must be lifted every few seconds to allow oxygen into the crucible to react with the magnesium.

A set of results for this practical could be:

mass of empty crucible/g	32.46
mass of crucible + magnesium/g	32.70
mass of crucible + contents at end of experiment/g	32.86

We can work out the mass of magnesium by subtracting the mass of the crucible from the mass of the crucible + magnesium:

mass of magnesium = 32.70 – 32.46 = 0.24 g

The reason that the mass increases is because the magnesium combines with the oxygen in the air. The mass of magnesium oxide is greater than the mass of just magnesium due to the extra oxygen.

KEY POINT

Remember: Number of moles is mass in grams divided by the mass of 1 mole in grams.

The overall reaction that occurred was:

$2Mg + O_2 \rightarrow 2MgO$

magnesium + oxygen → magnesium oxide

The type of reaction that occurred was combustion.

When we do this experiment in practice, it is not very often that the ratio comes out as 1:1. There are things that may go wrong with this experiment which could affect the results, such as:

- some of the magnesium oxide powder might escape when the lid is lifted
- not all the magnesium might have reacted
- the magnesium can also react with nitrogen in the air.

HINT

The original masses were measured to 2 significant figures so the number of moles should also be given to 2 significant figures (0.010). If you write 0.01 for the number of moles of atoms in the exam, this will be fine and still give you the correct answer.

!

Safety Note: The teacher demonstrating needs to wear a face shield and use safety screens. The pupils require eye protection and should be no closer than 2 metres. If a drying agent is needed anhydrous calcium chloride should be used NOT concentrated sulfuric acid.

KEY POINT

We must always make sure that the tube is filled with hydrogen before lighting the stream of hydrogen because a mixture of hydrogen and oxygen is explosive! It is therefore important to let the stream of hydrogen gas flow through the tube for a little while (to flush out all the oxygen) before lighting it or lighting the Bunsen burner.

We can work out the mass of oxygen that combines with the magnesium:

mass of oxygen = 32.86 – 32.70 = 0.16 g

In order to find the formula of magnesium oxide we need to work out how many moles of magnesium atoms combine with how many moles of oxygen atoms.

	Mg	O
masses/g	0.24	0.16
find the number of moles of atoms by dividing the mass by the mass of 1 mole	0.24/24	0.16/16
number of moles of atoms	0.010	0.010
divide by the smaller number to find the ratio	0.010/0.010	0.010/0.010
ratio of moles	1	1
empirical formula	MgO	

The empirical formula of magnesium oxide is MgO.

THE FORMULA FOR COPPER OXIDE

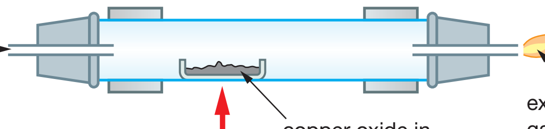
ACTIVITY 4

▼ PRACTICAL: INVESTIGATING THE FORMULA OF A METAL OXIDE BY REDUCTION

We can also find the formula of an oxide by removing the oxygen from it and looking at how the mass changes.

The following procedure can be used to find the formula of copper oxide:

- Weigh a ceramic dish.
- Put about 3 g of copper oxide in the ceramic dish and weigh the dish again.
- Place the ceramic dish in a tube as shown in Figure 5.13.
- Pass hydrogen gas over the copper oxide.
- Ignite the excess hydrogen, which comes out of the small hole in the boiling tube.
- Heat the copper oxide strongly until the reaction is finished (pink-brown copper metal will be seen).



▲ Figure 5.13 The experimental set-up for finding the formula of copper oxide.

A set of results for this practical could be:

mass of empty dish/g	23.78
mass of dish + copper oxide/g	26.96
mass of dish + contents at end of experiment/g	26.32

The reason that the mass decreases is because the hydrogen combines with the oxygen from the copper oxide to form water. The oxygen is removed from the copper oxide and we are left with only copper in the dish at the end of the experiment. Because oxygen is removed from the copper oxide we say that the copper oxide has been reduced. Reduction is explained in Chapter 14.

KEY POINT

This is a displacement (or competition) reaction. Hydrogen is more reactive than copper and displaces it from copper oxide. Displacement reactions are discussed in Chapter 14.

From this data we can calculate the mass of copper oxide at the beginning by subtracting the mass of the dish from the mass of the dish + copper oxide:

mass of copper oxide = 26.96 – 23.78 = 3.18 g

We can work out the mass of copper remaining at the end by subtracting the mass of the dish from the mass of the dish + copper at the end.

mass of copper = 26.32 – 23.78 = 2.54 g

The mass has decreased because the oxygen has been removed from the copper oxide and we can calculate the mass of oxygen by subtracting the mass of copper at the end from the mass of copper oxide:

mass of oxygen = 3.18 – 2.54 = 0.64 g

We now know that 2.54 g of copper combines with 0.64 g of oxygen in copper oxide and can deduce the empirical formula:

KEY POINT

The equation for the reaction that occurs in this experiment is:

$$CuO + H_2 \rightarrow Cu + H_2O$$

copper(II) oxide + hydrogen → copper + water

As well as a displacement reaction, this type of reaction can also be called a redox reaction. Redox reactions are discussed in Chapter 14.

	Cu	O
masses/g	2.54	0.64
find the number of moles of atoms by dividing the mass by the mass of 1 mole	2.54/63.5	0.64/16
number of moles of atoms	0.0400	0.040
divide by the smaller number to find the ratio	0.0400/0.040	0.040/0.040
ratio of moles	1	1
empirical formula	CuO	

Safety Note: The teacher demonstrating needs to wear a face shield and use safety screens. The pupils require eye protection and should be no closer than 2 metres. If a drying agent is needed anhydrous calcium chloride should be used NOT concentrated sulfuric acid.

DETERMINING THE FORMULA OF WATER

We can modify the apparatus in Figure 5.13 to allow us to determine the formula of water.

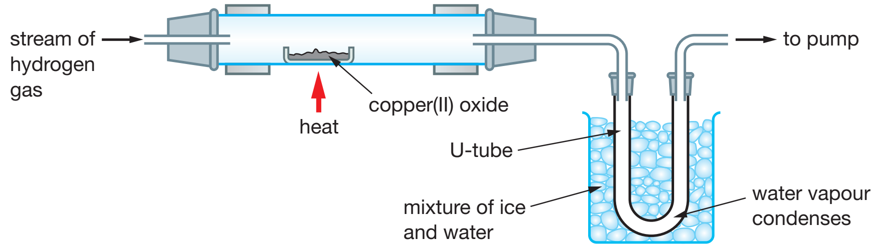


Figure 5.14 Apparatus that can be used to determine the formula of water.

The experiment is conducted in the same way except that this time the water vapour that is produced from the reaction between copper(II) oxide and hydrogen is condensed. The contents of the dish at the beginning and the end of the experiment are again weighed but this time the mass of water that collects in the U-tube must also be measured.

mass of empty dish/g	23.78
mass of dish + copper oxide/g	26.96
mass of dish + contents at end of experiment/g	26.32
mass of water/g	0.72

Here we are using the same results as above. The mass of oxygen lost from the copper(II) oxide is 0.64g.

All the oxygen lost from the copper(II) oxide combines with hydrogen to form water. This means that the water contains 0.64g of oxygen. 0.72g of water was collected so the mass of hydrogen in the water must be 0.72 – 0.64 = 0.08g. We can now determine the empirical formula of water:

	H	O
masses/g	0.08	0.64
number of moles of atoms	0.08/1	0.64/16
number of moles of atoms	0.08	0.040
divide by the smaller number to find the ratio	0.08/0.040	0.040/0.040
ratio of moles	2	1
empirical formula	H_2O	

WORKING OUT FORMULAE USING PERCENTAGE COMPOSITION FIGURES

In the worked examples and practical examples above, we have determined the empirical formulae of compounds using masses. However, we are often given percentages by mass instead of just masses.

EXAMPLE 5

Find the empirical formula of a compound containing 82.7% C and 17.3% H by mass (A_r of H = 1, A_r of C = 12).

The percentage figures apply to any amount of substance you choose, so let's choose 100 g. In this case the percentages convert simply into masses: 82.7% of 100 g is 82.7 g (Table 5.1).

Table 5.1 Finding the ratio from percentage by mass

	C	H
percentages/%	82.7	17.3
masses in 100 g/g	82.7	17.3
number of moles of atoms	82.7/12	17.3/1
number of moles	6.89	17.3
divide by smallest to get ratio	6.89/6.89	17.3/6.89
ratio of moles	1	2.5

From this we can see that there are 2.5 mol of H atoms for every mole of C atoms. The empirical formula, however, is the *whole number ratio* of the elements present in a compound. To obtain a whole number here we multiply both numbers by 2 to get C_2H_5, which is the empirical formula.

HINT

When you calculate 17.3/6.89 you actually obtain 2.51. When doing these calculations it is fine to round 2.51 to 2.5 or 1.01 to 1 but what you must not do is round 2.51 to 3!
If you got a ratio of C:H of 1:1.33 you would not round this down to 1:1 but rather multiply by 3 to obtain C_3H_4.

CONVERTING EMPIRICAL FORMULAE INTO MOLECULAR FORMULAE

When you have learnt a bit more organic chemistry in Unit 4 you will realise that, in the example we have just looked at, C_2H_5 can't possibly be the real formula of the compound. The molecular formula (the actual number of each atom present in a molecule) has to be a multiple of C_2H_5, like C_4H_{10}.

We can find the molecular formula if we know the relative formula mass of the compound (or the mass of 1 mole, which is just the M_r in grams).

In the previous example, suppose you were told that the relative formula mass of the compound was 58.

C_2H_5 has a relative formula mass of 29 (A_r of H = 1, A_r of C = 12).

All you need to find out is how many times 29 goes into 58.

$$58/29 = 2$$

So you need 2 lots of C_2H_5, in other words, C_4H_{10}.

The molecular formula is C_4H_{10}.

EXAMPLE 6

A compound has the empirical formula CH_2. If the relative formula mass is 56, work out the molecular formula.

The relative formula mass of CH_2 is $12 + (2 \times 1) = 14$.

$$56/14 = 4$$

Therefore there must be 4 lots of CH_2 in the actual molecule and the molecular formula is C_4H_8.

Remember, the ratio of elements in the molecular formula must be the same as in the empirical formula: C_4H_8 cancels down to CH_2.

EMPIRICAL FORMULA CALCULATIONS INVOLVING WATER OF CRYSTALLISATION

FINDING THE n IN $BaCl_2 \cdot nH_2O$

REMINDER

Remember that when some substances crystallise from solution, water becomes chemically bound up with the salt. This is called **water of crystallisation**. The salt is said to be **hydrated**.

REMINDER

Anhydrous means without water.

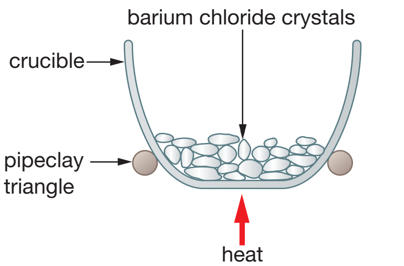


Figure 5.15 Heating barium chloride crystals in a crucible

Safety Note: Barium chloride is toxic but a Bunsen flame is not hot enough to release it into the atmosphere.

HINT

If you are doing this experiment, how can you be sure that all the water has been driven off? The best way to do this is to heat it to constant mass. You heat the crucible and weigh it. Keep doing this until the mass remains constant. If the mass has gone down after heating there may still be water present, but when the mass remains constant after heating you know that all the water has been driven off.

Usually, when you heat a salt that contains water of crystallisation, the water is driven off, leaving the **anhydrous** salt behind. Hydrated barium chloride is a commonly used example because the barium chloride itself doesn't decompose even on quite strong heating.

If you heated barium chloride crystals in a crucible you might obtain these results:

Mass of crucible	= 30.00 g
Mass of crucible + barium chloride crystals, $BaCl_2 \cdot nH_2O$	= 32.44 g
Mass of crucible + anhydrous barium chloride, $BaCl_2$	= 32.08 g

To find n, you need to find the ratio of the number of moles of $BaCl_2$ to the number of moles of water. It's just another empirical formula calculation (A_r of H = 1, A_r of O = 16, A_r of Cl = 35.5, A_r of Ba = 137).

Mass of $BaCl_2$ = 32.08 − 30.00 = 2.08 g
Mass of water = 32.44 − 32.08 = 0.36 g

Table 5.2 Finding the n in $BaCl_2 \cdot nH_2O$

	$BaCl_2$	H_2O
masses/g	2.08	0.36
divide by M_r to find the number of moles	2.08/208	0.36/18
number of moles	0.0100	0.020
ratio of moles	1	2
empirical formula	$BaCl_2 \cdot 2H_2O$	

CALCULATIONS USING MOLES, CHEMICAL EQUATIONS AND MASSES OF SUBSTANCES

INTERPRETING SYMBOLS IN EQUATIONS IN TERMS OF MOLES

KEY POINT

The big numbers in front of the formulae tell us the number of moles of each substance that react. Although there does not appear to be a big number in front of the CH_4 and CO_2, there should actually be a '1' in front of each but we just don't write it.

If we write a chemical equation:

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

we can say that 1 molecule of CH_4 combines with 2 molecules of O_2 to form 1 molecule of CO_2 and 2 molecules of water. When doing calculations, however, it is often more useful to think about this in terms of moles and take symbols as meaning moles of each substance. So here we can say:

1 mol CH_4 reacts with 2 mol O_2 to form 1 mol CO_2 and 2 mol H_2O

We have not changed the meaning of the equation as the ratios are still the same.

KEY POINT

Remember that 2 mol of O_2 contains twice as many molecules as 1 mol of CH_4.

We can therefore tell from this equation that 16 g of CH_4 (1 mol) reacts exactly with 64 g of O_2 (2 mol) to form 44 g of CO_2 (1 mol) and 36 g of H_2O (2 mol).

CALCULATIONS INVOLVING ONLY MASSES

Typical calculations will give you a mass of starting material and ask you to calculate how much product you are likely to obtain. You will also find examples done in reverse, where you are told the mass of the product and are asked to find out how much of the starting material you would need. In almost all the cases you will see at International GCSE you will be given the equation for the reaction.

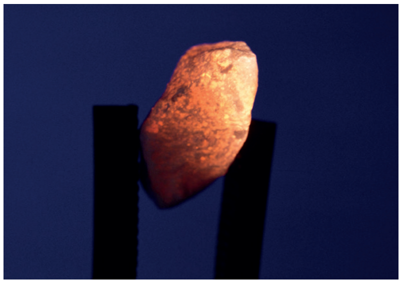


Figure 5.16 A glowing piece of limestone. Limestone is impure calcium carbonate – it decomposes when heated to form calcium oxide (quicklime) and carbon dioxide. This is a thermal decomposition reaction.

A PROBLEM INVOLVING HEATING CALCIUM CARBONATE

When calcium carbonate, $CaCO_3$, is heated, calcium oxide is formed. Imagine you wanted to calculate the mass of calcium oxide produced by heating 25 g of calcium carbonate (A_r: C = 12, O = 16, Ca = 40).

$$CaCO_3 \rightarrow CaO + CO_2$$

THE CALCULATION

The method we will use has three steps:

Step 1: Calculate the number of moles using the mass that you have been given.

Step 2: Use the chemical equation to work out the number of moles of the substance you are interested in.

Step 3: Convert the number of moles to a mass.

KEY POINT

The M_r of $CaCO_3$ is 100 and the M_r of CaO is 56.

In this reaction we have 25 g of $CaCO_3$ and the number of moles can be calculated as:

$$\text{number of moles} = \frac{\text{mass}}{\text{mass of 1 mole}} = \frac{25}{100} = 0.25\,\text{mol}$$

The chemical equation for this reaction is

$$CaCO_3 \rightarrow CaO + CO_2$$

We can see from the equation that 1 mole of $CaCO_3$ will decompose to produce 1 mole of CaO. In other words, the equation is telling us that if we start with a certain number of moles of $CaCO_3$ we will obtain the same number of moles of CaO at the end.

Therefore, we can deduce that 0.25 mol $CaCO_3$ will decompose to produce 0.25 mol CaO. Now that we know the number of moles of CaO (M_r = 56) we can convert this to a mass:

$$\text{mass} = \text{number of moles} \times \text{mass of 1 mole}$$

$$\text{mass} = 0.25 \times 56 = 14\,\text{g}$$

Therefore, the reaction will produce 14 g of calcium oxide.

ALTERNATIVE METHOD

The calculation can also be done in a different way using ratios:

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

Interpret the equation in terms of moles:

1 mol $CaCO_3$ produces 1 mol CaO (and 1 mol CO_2)

Substitute masses where relevant:

100 g (1 mol) $CaCO_3$ produces 56 g (1 mol) CaO

Do a proportion calculation:

If 100 g of calcium carbonate gives 56 g of calcium oxide

1 g of calcium carbonate gives $\frac{56}{100}$ g of calcium oxide = 0.56 g

25 g of calcium carbonate gives 25 × 0.56 g of calcium oxide = 14 g of calcium oxide

HINT

You do not have to learn both methods, just learn whichever method you are happier with.
Your maths may be so good that you don't need to take all these steps to find the answer. If you can do it more quickly, that's fine. However, you must still show all your calculations.
When you have finished a chemistry calculation, the impression should nearly always be that there are a lot of words with a few numbers between them, not vice versa.

A PROBLEM ABOUT EXTRACTING IRON

The equation below is for a reaction that occurs in the extraction of iron:

$Fe_2O_3 + 3C \rightarrow 2Fe + 3CO$

Calculate the mass of iron which can be formed from 1000 g of iron oxide.

We are given the mass of Fe_2O_3 in the question and so we can calculate the number of moles of Fe_2O_3:

$$\text{number of moles} = \frac{\text{mass}}{\text{mass of 1 mole}}$$

$$\text{number of moles} = \frac{1000}{160} = 6.25\,\text{mol}$$

From the chemical equation we can see that 1 mol Fe_2O_3 produces 2 mol Fe. We know this because of the big numbers in front of the formulae.

So we know that, if we start with a certain number of moles of Fe_2O_3, we will obtain twice as many moles of Fe.

So 6.25 mol Fe_2O_3 produces 2 × 6.25 = 12.5 mol Fe:

mass = number of moles × mass of 1 mole

The mass of 12.5 mol Fe is 12.5 × 56 = 700 g.

KEY POINT

The A_r of Fe is 56.
The M_r of Fe_2O_3 = 2 × 56 + 3 × 18 = 160.

HINT

To find the mass of Fe we multiply the number of moles of Fe by the mass of 1 mole of Fe (56 g). A common mistake is to multiply the number of moles by the mass of 2Fe (112 g), but we have already used the '2' when we worked out the number of moles of Fe.

ALTERNATIVE METHOD IN TERMS OF RATIOS

First interpret the equation in terms of moles:

1 mol Fe_2O_3 reacts with C to form 2 mol Fe (and CO_2).

We are only looking at how much iron is produced, so let's introduce just these masses.

160 g (1 mol) of Fe_2O_3 produces 2 × 56 g (2 mol) of Fe.

That is, 160 g of Fe_2O_3 produces 112 g of Fe.

From this we can calculate that 1 g of Fe_2O_3 will produce 112/160 g of Fe, and therefore 1000 g of Fe will produce 1000 × 112/160 = 700 g.

A PROBLEM INVOLVING THE EXTRACTION OF LEAD

Lead is extracted from galena, PbS. The ore is roasted in air to produce lead(II) oxide, PbO:

$$2PbS(s) + 3O_2(g) \rightarrow 2PbO(s) + 2SO_2(g)$$

The lead(II) oxide is then converted to lead by heating it with carbon in a blast furnace:

$$PbO(s) + C(s) \rightarrow Pb(l) + CO(g)$$

The molten lead is tapped from the bottom of the furnace.

Calculate the mass of lead that would be produced from 1 tonne of galena.

(A_r: O = 16, S = 32, Pb = 207)

Figure 5.17 Molten lead tapped from the bottom of a furnace.

HINT

You should be told how many grams there are in a tonne if you need this information in the exam.

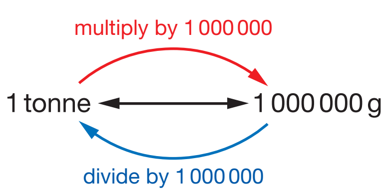


Figure 5.18 How to convert between grams and tonnes

HINT

1 000 000 might be given to you as 10^6 or 1×10^6. These are both the same.

We have been given the mass of PbS so we can calculate the number of moles. However, we have to be careful because the mass is in tonnes but to calculate a number of moles we need to know the mass in grams, so we must convert tonnes to grams. The conversion for this is shown in Figure 5.18. Just write down this relationship: *1* tonne is equivalent to *1 000 000* g so, to change from tonnes to grams, the number has to *increase* (from 1 to 1 000 000), so we *multiply* by 1 000 000.

Therefore we have 1 000 000 g PbS.

$$\text{Number of moles} = \frac{\text{mass}}{\text{mass of 1 mole}}$$

$$\text{The number of moles of PbS} = \frac{1\,000\,000}{239} = 4184\,\text{mol.}$$

From the first equation:

$$2PbS(s) + 3O_2(g) \rightarrow 2PbO(s) + 2SO_2(g)$$

we can see that 2 mol PbS produce 2 mol PbO, therefore we can work out that 4184 mol PbS will produce 4184 mol PbO.

From the second equation:

$$PbO(s) + C(s) \rightarrow Pb(l) + CO(g)$$

we can see that 1 mol PbO produces 1 mol Pb. But we are starting this reaction with 4184 mol PbO and so we can work out that we will produce 4184 mol Pb.

The mass of 4184 mol Pb is 4184 × 207 = 866 000 g.

We can convert this to tonnes by dividing by 1 000 000:

866 000/1 000 000 = 0.866 tonne

ALTERNATIVE METHOD

$$2PbS(s) + 3O_2(g) \rightarrow 2PbO(s) + 2SO_2(g)$$

$$PbO(s) + C(s) \rightarrow Pb(l) + CO(g)$$

Interpret the equation in terms of moles and trace the lead through the equations:

2 mol PbS produces 2 mol PbO

If we double the second equation:

$$2PbO(s) + 2C(s) \rightarrow 2Pb(l) + 2CO(g)$$

we can see that 2 mol PbO produces 2 mol Pb.

In other words, every 2 mol of PbS produces 2 mol of Pb.

Substitute masses where relevant. In this case, the relevant masses are only the PbS and the Pb:

2 × 239 g PbS produces 2 × 207 g Pb

478 g PbS produces 414 g Pb

Now there seems to be a problem. The question is asking about tonnes and not grams. You could calculate how many grams there are in a tonne. However, it's much easier to think a bit, and realise that the ratio is always going to be the same, whatever the units, which means that:

478 tonnes PbS produces 414 tonnes Pb

Do the proportion calculation:

If 478 tonnes PbS produces 414 tonnes Pb

then 1 tonne PbS gives $\frac{417}{478}$ tonne Pb = 0.866 tonne

0.866 tonne of lead is produced from 1 tonne of galena.

KEY POINT

We've doubled the second equation so that we can trace what happens to all the 2PbO from the first one.
This could also be simplified to 1 mol PbS produces 1 mol Pb. This would save you doing some arithmetic as you would not have to multiply everything by 2. However, in the end it does not make any difference to the answer, so you do not have to simplify it if you don't want to.

HINT

There are advantages and disadvantages of both methods. For the first method, we can use the same three steps in all the calculations we do below. However, a bit more maths will be involved: converting tonnes to grams and back again. The alternative method has fewer steps but we cannot use it in the same way when doing some of the calculations involving solutions in the next chapter.

CALCULATING PERCENTAGE YIELDS

WHAT IS A PERCENTAGE YIELD?

If you calculate how much product a reaction might produce, in real life you rarely obtain as much as you expected. If you expect to get 100 g, but only get 80 g, your percentage yield is 80%. The rest of it has been lost in some way. This could be due to spillages, or losses when you transfer a liquid from one container to another. Or it may be that there are all sorts of side reactions going on, so that some of your starting materials are changed into unwanted products. That happens a lot during reactions in organic chemistry.

Figure 5.19 Experiments done by students to make and purify organic compounds rarely give a high percentage yield.

CALCULATING THE PERCENTAGE YIELD

HINT

The percentage yield should always come out as less than 100%. If it comes out as more than 100% either you have made a mistake in the calculation or the question will continue by asking about what might have gone wrong in the experiment to give a value greater than 100%. If you are making a salt (see Chapter 17), for instance, a possible reason you could get a percentage yield greater than 100% is that it was not completely dry.

Suppose you do a calculation and work out that you should get 12.5 g of copper sulfate crystals, but you only get 11.2 g when you actually do the experiment. 12.5 g is the *theoretical yield* and can be obtained by carrying out a moles calculation. 11.2 g is the *actual yield* – this is what is obtained in the actual experiment and will be given to you in the question.

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

The percentage yield is $\frac{11.2}{12.5} \times 100 = 89.6\%$.

EXAMPLE 7

A student reacted 2.40 g of copper(II) oxide (CuO) with hot sulfuric acid. She made 5.21 g of copper(II) sulfate crystals ($CuSO_4{\cdot}5H_2O$).

The equations for the reactions are:

$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$

$$CuSO_4(aq) + 5H_2O(l) \rightarrow CuSO_4{\cdot}5H_2O(l)$$

First we need to calculate the theoretical yield. This is done by carrying out a moles calculation. We have enough information to calculate the number of moles of copper(II) oxide:

$$\text{number of moles of CuO} = \frac{2.40}{79.5} = 0.0302\text{ mol}$$

From the first equation we can deduce that 0.0302 mol CuO will produce 0.0302 mol $CuSO_4$.

From the second equation we can deduce that 0.0302 mol $CuSO_4$ will produce 0.0302 mol $CuSO_4{\cdot}5H_2O$.

Now we need to calculate the mass of $CuSO_4{\cdot}5H_2O$ by multiplying the number of moles by the mass of 1 mole:

$$\text{mass} = 0.0302 \times 249.5 = 7.53\text{ g}$$

This is the theoretical yield of copper(II) sulfate crystals.

The actual yield is what was obtained in the experiment. This was 5.21 g.

The percentage yield = (actual yield/theoretical yield) × 100

The percentage yield is 5.21/7.53 × 100 = 69.2%.

CALCULATIONS IN WHICH YOU HAVE TO CALCULATE WHICH SUBSTANCE IS IN EXCESS

When you make something in the lab, you rarely mix things together in exactly the right proportions. Usually, something is in excess, and you remove the excess (by filtering, for example) when the reaction is complete. If something is in excess there is more than enough of it to react with the other substance and so we don't need to worry about it.

In all the examples we have done above, we have only ever been given enough information to calculate the number of moles of one thing at the beginning of the problem. We have used this number of moles throughout the calculations. But what would happen if we were given the following question?

Magnesium reacts with hydrochloric acid according to the following equation:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

0.2 mol Mg is reacted with 0.2 mol HCl. Calculate the mass of hydrogen gas produced.