Chemistry

GRADE 8

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Contents

Introduction

For the student

Welcome to your new Chemistry textbook, *Grade 8 Chemistry*. Your textbook comes with a **Grade 8 Chemistry Student's Portfolio** and a range of *digital resources*. This book will build on your previous learning of Chemistry by helping you to understand the world around you. It aims to develop your learning skills in science. You will develop these skills yourself while also learning from your teacher and your fellow students.

Glossary

A comprehensive Glossary for the Textbook and Student's Portfolio book is included at the back of this book.

For the teacher

Written for the new Grade 8 Chemistry subject programme in Kazakhstan, *Grade 8 Chemistry* aims to give students a sense of enjoyment and an interest in the learning of science. The book is based on the Grade 8 Learning Objectives in the Grade 7- 9 Chemistry subject programme document. It develops students' knowledge of and about science through the four content and skills strands described in the Chemistry subject programme and highlighted throughout the text using four different logos (understanding science, researching and experimenting in science, communicating in science, and science and society).

Understanding U1 Research R4 Communicating C3 Society S2

- **Learning outcomes** are stated at the beginning of each module in student-friendly language.
- **Keywords** are listed at the start of each module to allow students to become familiar with important new terms.
- **Activities** allow students to build on their knowledge by completing research.
- **Diagrams** have been fully labelled and are drawn in a simple style so that students can replicate them easily.
- **Questions** are interspersed within the text to offer teachers the opportunity to use different teaching strategies. In particular, there are chances for group work and pair work.
- **Did you know?** boxes feature interesting facts to stimulate students' interest in science.

- The **language** used is clear and simple to allow for use by students of varying reading levels.
- Simple and helpful **logos** are used throughout to enhance student understanding.

Activity

Corresponding page in Student's Portfolio

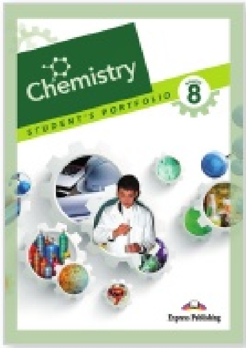
Key fact

Question

Group work

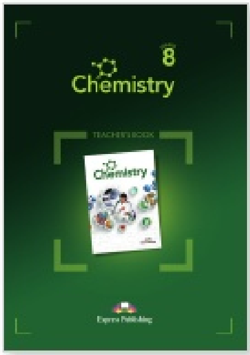
Research

Student's Portfolio



The Student's Portfolio provides additional material, activities and tasks. The portfolio book enables students to maintain a record of specific activities and reflect on their learning, as well as focusing on key words and key facts, through mind maps and comprehension and recall activities. It also contains templates for self-assessment and peer assessment. This book works in conjunction with the Textbook.

Teacher's Resource Book



The Teacher's Resource book works in conjunction with the Textbook and the Student's Portfolio book by providing:

- An outline of the Grade 8 content and skills priorities in the subject programme
- Learning outcomes for each module with explanations of how they are incorporated into lessons
- Information on topics, questions and research ideas that can be used to enhance the students' learning
- Answers to all student questions in the Textbook and Student's Portfolio book
- Outlines of digital resources for each module and suggestions for integrating them into classroom work
- Suggestions of ways to assess student activities with assessment templates
- A range of other information and suggestions to support teachers in the delivery of the new course
- Key skills, literacy and numeracy linked to relevant modules
- Guidance for the teacher through the module
- Additional activities and research activities

Digital resources

The *Grade 8 Chemistry* **digital resources** will further enhance classroom learning. These resources have been designed to integrate with the Textbook and to complement lessons suggested in the Teacher's Resource Book. Following the principles of the new national Chemistry subject programme, material is provided to suit a range of learner types and to encourage participation and engagement on the part of the student.

A series of **videos** allows students to observe science in action across all modules. These videos reinforce the topic at hand and allow for other perspectives, which may be discussed in class. Similarly, a series of **videos** about **scientist biographies** presents a lively gateway to develop students' interest in science and initiate student-led research.

Further classroom discussion and participation is opened up through **PowerPoint presentations**, including a thematic presentation of information from the Textbook. **Experiment videos** allow for a visual review of activities carried out in the classroom. **Extra assessment material** is provided to support teachers in carrying out a range of oral and written formative and summative assessments.

Guidance for integration of digital resources in the classroom is provided by the **digital resource symbol** used throughout the Textbook, as well as the provision of detailed notes and suggestions in the Teacher's Resource Book.

Laboratory equipment

Laboratory safety rules for pupils

The following rules are enforced to keep you and your classmates safe while in a school laboratory.

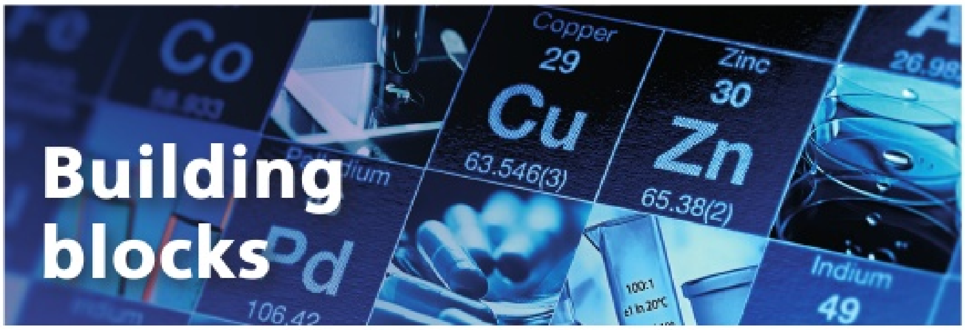
1. Do not enter the laboratory without permission.
2. Do not use any equipment unless permitted to do so by your teacher.
3. Make sure you know exactly what you are supposed to do. If in doubt, ask your teacher.
4. Make sure you know the position of all safety equipment in the laboratory, e.g. the fire extinguishers, first aid equipment etc.
5. Always wear eye protection or gloves when instructed to do so.
6. Long hair must be tied back during practical classes.
7. Place your bag and other personal items safely out of the way.
8. Never handle any chemicals with bare hands.
9. Nothing must be eaten, tasted or drunk in the laboratory.
10. Any cut, burn or other accident must be reported at once to your teacher.
11. Always check that the label on the bottle is exactly the same as the material you require. If in doubt, ask your teacher.
12. Any chemical spilled on the skin or clothing must be washed at once with plenty of water and reported to your teacher.
13. Test tubes should never be overfilled. When heating a test tube ensure that the mouth of the test tube is pointed away from you and everyone else.
14. All equipment should be cleaned and put back in its correct place after use.
15. Always wash your hands after practical work.
16. Students should behave in a responsible manner at all times in the laboratory.

Safety labels

The following labels appear on bottles in the laboratory. They also appear on many everyday chemicals such as cleaning products and solvents. These labels indicate chemicals that could be dangerous if not used or handled properly. We use these warning symbols on activities in this book.

Toxic		Substances which can cause death if they are swallowed, breathed in or absorbed through the skin. Example: weedkiller.
Harmful or irritant		Substances which should not be eaten, breathed in or handled without gloves. Though not as dangerous as toxic substances they may cause a rash, sickness or an allergic reaction.
Oxidising		Substances which provide oxygen, allowing other materials to burn more intensely. Example: hair bleach.
Highly flammable		Substances which easily catch fire. Example: petrol.
Corrosive		Substances which attack and destroy living tissue, including skin and eyes. Example: oven cleaner.
Warning sign		This sign is used to draw attention to a warning of danger, hazards and the unexpected.
Safety glasses		Wear safety glasses to protect your eyes.

MODULE 1



Building blocks

Learning outcomes

At the end of this module you will be able to:

- Describe an atom
- Distinguish between elements, compounds and molecules
- Describe electron distribution in atoms 8.1.3.1
- Recognise that the number of electrons in each electron shell has a maximum value 8.1.3.2
- Identify the electron configuration of the first 20 elements 8.1.3.4
- Explain ionic bonding 8.1.3.5
- Make formulas for ionic compounds 8.1.3.6

Keywords

✔ atom ✔ element ✔ compound ✔ molecule ✔ proton ✔ neutron ✔ electron ✔ symbol ✔ formula

What are atoms?

Atoms are the basic building blocks of all materials. The word 'atom' comes from a Greek word describing something that cannot be divided. When atoms were discovered, scientists thought of them as solid balls, like snooker balls, which could not be split into anything simpler.

Figure 1.1 *When atoms were discovered, they were thought to be solid, like this snooker ball*

Element

An element is a substance made up of only one type of atom.

If you burn a piece of wood, you will get a black solid. That solid is called carbon. Carbon is an element because it is made up of only carbon atoms and cannot be broken down into anything simpler.

Did you know?

Hydrogen is the most abundant element in the universe.

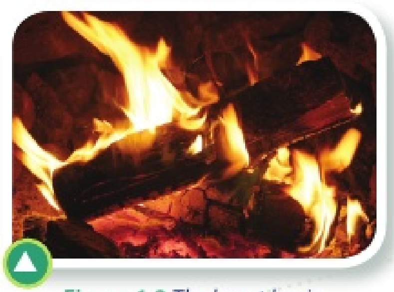


Figure 1.2 *The burnt log is carbon, and only carbon*

Symbol

Each element has a symbol. H is the symbol for hydrogen. Fe is the symbol for iron, which comes from the Latin name *ferrum*.

Communicating C1 Society S1

1.1 What is an element?

1.2 Give three examples of elements and their symbols.

1.3 Who was the scientist who first called the tiny particles in matter 'atoms'?

1.4 Where would you find a list of all the elements, with their symbols?

1.5 List the names of all the elements whose symbols start with the letter C.

Communicating C1 Communicating C2 Society S2

1.6 Divide the class into five different groups. Your teacher will give each group four different elements to research.

For each element, find out:

(a) Its symbol

(b) Its state of matter at room temperature

(c) Whether it is a metal, a non-metal or a semi-metal

(d) Its appearance

(e) Where it may be used.

Did you know?

The element helium was discovered on the Sun before it was found on Earth.

Communicating C1 Society S2

1.7 Which of the elements in Figure 1.3 are used in:

(a) Jewellery?

(b) Thermometers?

(c) Plumbing?

Figure 1.3 *Elements*

Molecules

A molecule is formed when two or more atoms join (**bond**) together chemically.

Compounds

When a molecule is made up of two or more different types of atoms, we call it a compound. All compounds are molecules but not all molecules are compounds.

Did you know?

There are about ten million known compounds that can be made with carbon.

Hydrogen gas (H_2) and oxygen gas (O_2) are not compounds because each is composed of one element. Water (H_2O) and carbon dioxide (CO_2) are compounds because each is made up of more than one type of element. All compounds are **non-elements**.

Q

1.8 State two differences between a mixture and a compound.

Q Understanding U1

1.9 Look at the following and state whether each is an element or a compound.

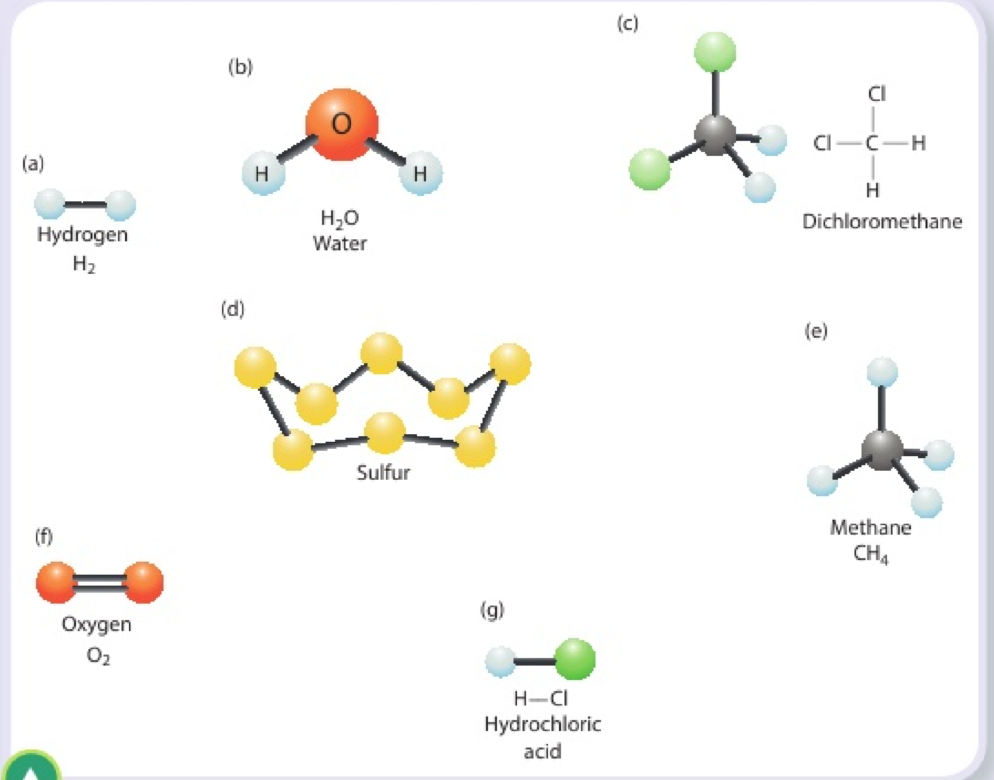


Figure 1.4 *Element or compound?*

Making models

You can use models to show how atoms bond (join) to each other.

Model kits have different coloured balls for each element and sticks to join one ball to another.

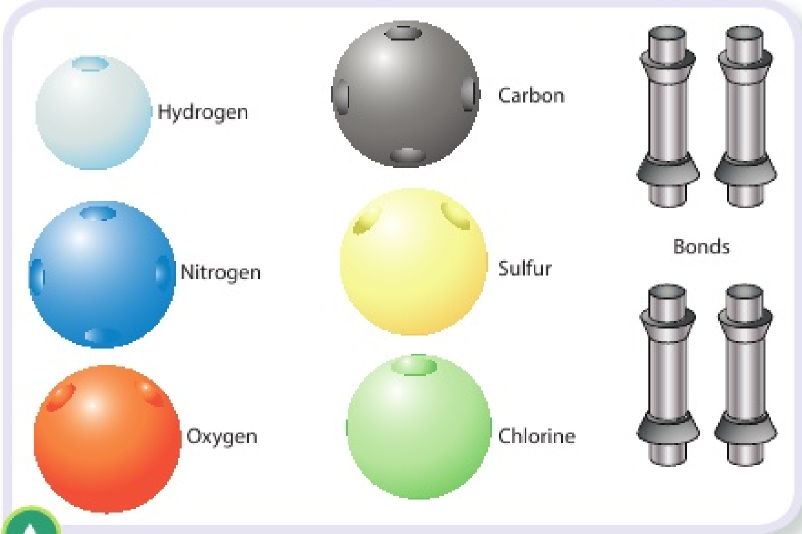


Figure 1.5 *Models are used to show how atoms bond together*

Q U1 C1 C2

1.10 Make your own model

(a) Using a model kit, design your own model of an element or compound. Remember to use the colours for each element as shown in **Figure 1.5**.

(b) Compare your model with others in your group. Do the other students know the name of the atoms in your model of an element or compound?

(c) Was the model you made an element or a compound?

(d) Draw diagrams of the molecules everyone in your group made.

Structure of the atom

The atom is made up of smaller particles called **sub-atomic particles**.

These are:

- Protons
- Neutrons
- Electrons.

The three particles in the atom are quite different from each other. **Table 1.1** summarises the properties of each.

Figure 1.6 *Computer image of the atomic structure*

Table 1.1 **The properties of protons, neutrons and electrons**

Particle	Charge	Mass	Location
Proton	+1	1	Nucleus
Neutron	0	1	Nucleus
Electron	−1	Negligible	Shells

All the particles are extremely small. Electrons are so small it would take almost 2000 of them to have the same mass as a single proton or neutron.

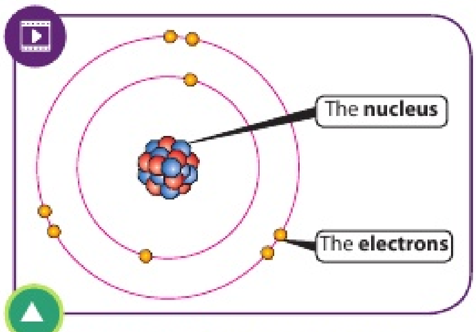


Figure 1.7 *Structure of the atom*

Atomic and mass numbers

Atoms of different elements differ from each other by the number of protons, neutrons and electrons they have.

Every element has its own **atomic number**. The atomic number tells you how many protons (which is the same as the number of electrons) there are in one atom of the element.

There is also the **mass number**. The mass number tells you how many protons and neutrons there are in the atom. Examples are given in Figure 1.9 and Figure 1.10.

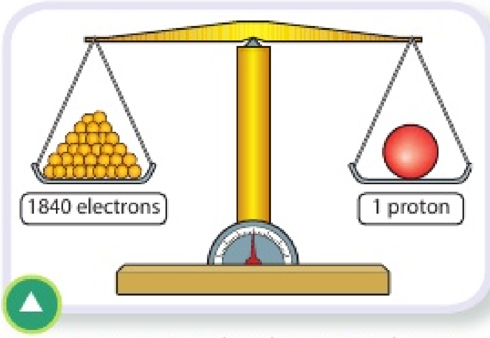


Figure 1.8 *It takes almost 2000 electrons to have the same mass as a single proton (or neutron)*

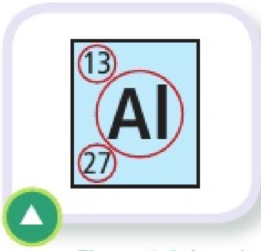


Figure 1.9 *Atomic and mass numbers of aluminium*

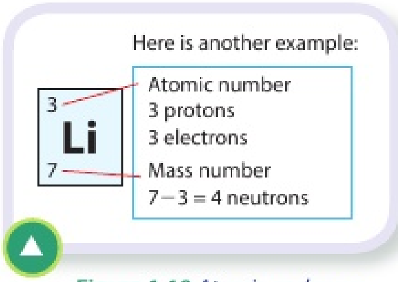


Figure 1.10 *Atomic and mass numbers of lithium*

How electrons are arranged

Electrons whizz around the nucleus in energy levels called **shells** or **orbits**. The first shell can hold only two electrons; the second shell can hold up to eight electrons and the third shell can have as many as eighteen electrons. Electrons fill up the shells one by one, starting with the first shell. When a shell is full, they start a new one.

All atoms would like to have full electron shells, but in most atoms the outer shell is not full and this makes the atom want to react to fill it.

1.11 Do you think an atom with full electron shells would be more or less likely to react chemically with another material?

Bohr model

A Danish scientist called Niels Bohr was the first person to suggest the idea of electron shells containing electrons orbiting the nucleus. The way these electrons are arranged is called the **electron configuration**. The way they are explained is called the **Bohr model**. **Figures 1.12**, **1.13** and **1.14** show some examples of atoms of elements using Bohr models.

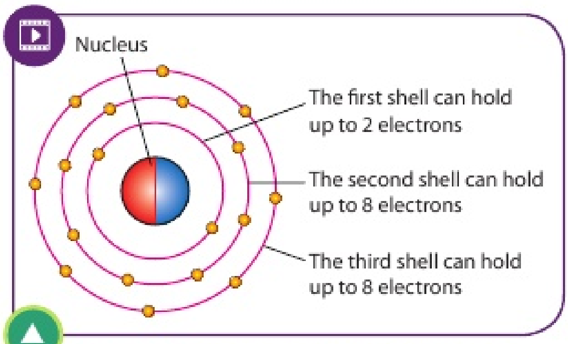


Figure 1.11 *The arrangement of electrons in an atom*

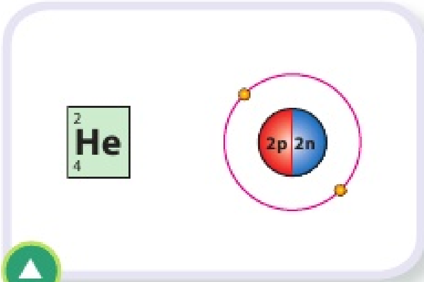


Figure 1.12 *Helium*

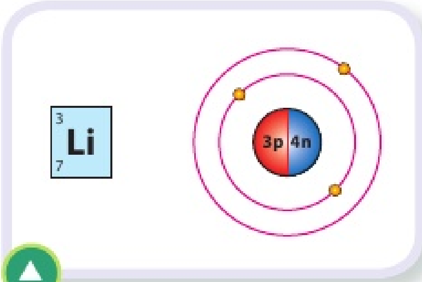


Figure 1.13 *Lithium*

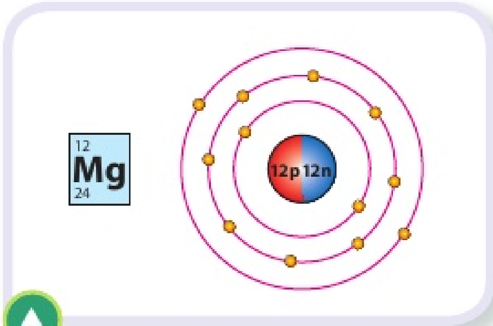


Figure 1.14 *Magnesium*

1.12 An atom is made up of three types of particle. Name them.

1.13 See in **Figure 1.14** the electron configuration for magnesium. Find the electron configuration of the first 20 elements.

1.14 What is the atomic number and mass number of an atom?

1.15 Complete this sentence: The protons and neutrons are found in the ______________ and the ______________ spin around in orbits or ______________.

Q R4 Research

1.16 Use this table to help you answer the questions below.

Element	Atomic number	Mass number
Oxygen	8	16
Chlorine	17	35
Gallium	31	70
Zinc	30	65
Tungsten	74	184

(a) How many protons would you expect to find in an atom of chlorine?

(b) Where in the atom would the protons be found?

(c) How many neutrons would you expect to find in an atom of zinc?

(d) Which element in the table above has atoms that contain the same number of protons and neutrons?

Q U1 Understanding

1.17 Give the number of protons, electrons and neutrons in the atoms in **Figure 1.15**.

(a)	(b)	(c)
10 Ne 20	9 F 19	27 Co 60

(d)	(e)	(f)
92 U 235	7 N 14	19 K 39

Figure 1.15

Q U1 Understanding

1.18 Draw fully labelled diagrams of the atoms in **Figure 1.16**.

(a)	(b)	(c)	(d)
2 He 4	4 Be 9	13 Al 27	20 Ca 40

Figure 1.16

Q U1 Understanding C1 Communicating

1.19 Divide into groups. Each group is to design and construct a model of an atom (element of your choice) showing the exact number of protons, neutrons and electrons.

Ionic bonding

Most elements in the periodic table are either metals or non-metals.

- Metals tend to want to lose electrons.
- Non-metals usually want to gain electrons.

So, when a metal atom meets a non-metal atom they react.

All atoms have equal numbers of protons and electrons and so have no charge. However, when an atom loses or gains electrons it becomes an **ion**.

When an atom loses electrons it becomes positively charged. Hence, when an atom gains electrons it becomes negatively charged.

These opposite charges attract and a bond called an **ionic bond** is formed.

Let's look at two examples of ionic bonding.

Figure 1.17 *A sodium chloride model. Each sodium ion (silver) is surrounded by 6 chlorine ions (green)*

1 How sodium and chlorine react

Sodium atom

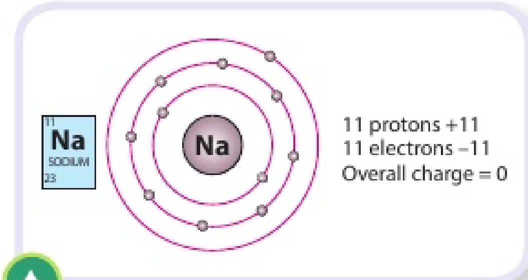


Figure 1.18 *Sodium, 11 electrons, electronic configuration (2, 8, 1) Wants to lose one electron*

Chlorine atom

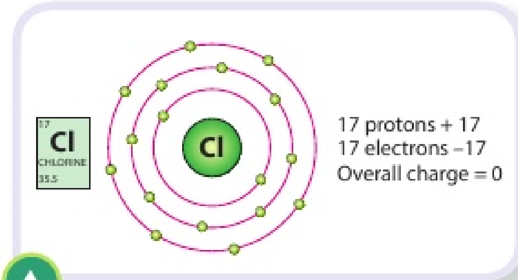


Figure 1.19 *Chlorine, 17 electrons, electronic configuration (2, 8, 7) Wants to gain one electron*

- On losing an electron the sodium atom becomes positively charged. It is now a sodium ion.
- On gaining an electron the chlorine atom becomes negatively charged. It is now a chlorine ion.

These opposite charges are attracted to each other and an ionic bond results.

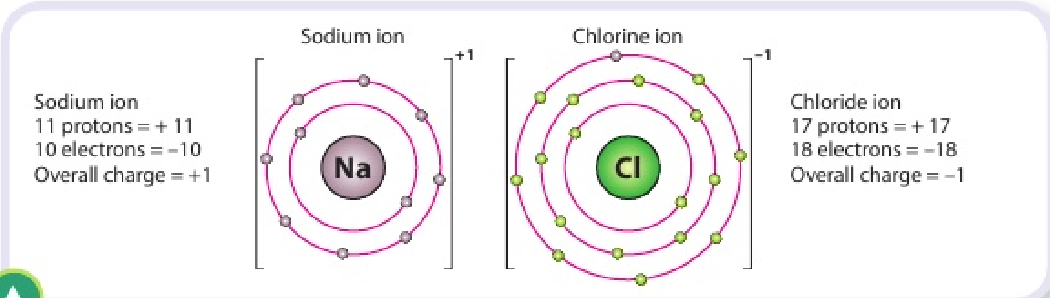


Figure 1.20 *Na^{+1} (2, 8) – Cl^{-1} (2, 8, 8)*

2 How magnesium and oxygen react

- Magnesium is a metal which has 12 electrons (2, 8, 2).
- Oxygen is a non-metal with 8 electrons (2, 6).
- Magnesium will donate 2 electrons and form an ion with a +2 charge.
- Oxygen ends up with a -2 charge when it accepts the 2 electrons.

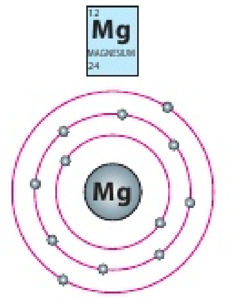


Figure 1.21 *Magnesium and oxygen atoms*

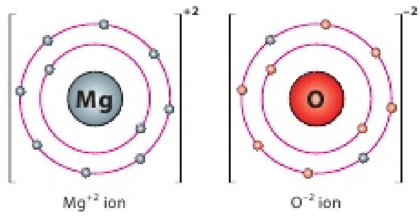


Figure 1.22 *Ionic compound called magnesium oxide*

Q U1

1.20 Why do most elements react?

1.21 What is an ion?

1.22 A negative ion is an atom that has ______________ electrons.

1.23 A positive ion is an atom that has ______________ electrons.

1.24 Sodium chloride, NaCl, is common salt.

(a) Draw Bohr structure diagrams showing the arrangement of electrons in a sodium and in a chlorine atom.

(b) Describe how a sodium atom and a chlorine atom combine to produce sodium chloride. You may use a diagram if it helps.

(c) What word is used to describe the type of bond formed between sodium and chlorine in sodium chloride?

Formulas for ionic compounds

A formula is a combination of chemical symbols used to represent a compound. An ionic compound is composed of ions and in the formulas for such compounds, the charges of the ions must add up to zero.

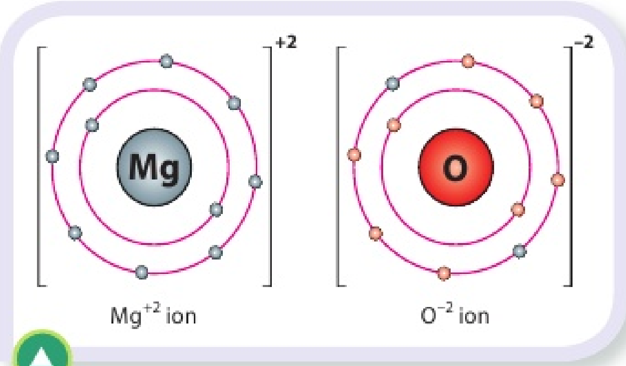


Figure 1.23 *magnesium oxide*

In formulas for ionic compounds the positive ion is usually written first. We can see from Figure 1.23 how the relative electrical charge is indicated: plus (+) for a positive charge and (–) for a negative charge together with a number. Where the number is 1, it can be left out, so a sodium ion can be represented as: Na^{+}

All ionic compounds have a net charge of zero, so the compound sodium chloride (see Figure 1.20) is made up of the sodium ion (+1) and the chloride ion (–1) which combine to form NaCl. In cases where more than one atom of a compound is used in a formula, the formula is written like this: $CaCl_2$ (calcium chloride).

Q U1 Understanding

1.25 The symbol for the Chloride ion is Cl^{-1}. From the formula $CaCl_2$

(a) work out the charge of the Calcium ion.

(b) say how many calcium ions and how many chloride ions the formula represents.

1.26 The following is a list of common ions. When ionic compounds are formed the overall charge is usually zero. Predict the formula of the following compounds.

Na^{+1}	Mg^{+2}	H^{+1}	F^{-1}	Cl^{-1}	O^{-2}

(a) Sodium chloride ___________________________

(b) Magnesium oxide ___________________________

(c) Magnesium chloride ___________________________

(d) Hydrochloric acid ___________________________

(e) Sodium fluoride ___________________________

Research R3 Communicating C1 Communicating C2 Communicating C3 Society S1 Society S2

1.27 Scientists discovered in about 1939 that a powerful explosion might be possible by splitting an atom.

Albert Einstein had many theories that helped scientists in developing the atomic bomb. He was so frightened about what might happen if Hitler in Germany made the bomb first that he wrote to US President Franklin Roosevelt telling him about the bomb. This led to the Manhattan Project being set up, which involved over two hundred scientists researching and developing the bomb. The first atomic bomb exploded in 1945 in the New Mexico desert. The temperature at the centre of the explosion was three times hotter than the centre of the Sun.

On 6 August 1945 an atomic bomb named 'Little Boy' was dropped on Hiroshima, Japan.

Research the history of the first atomic bomb and answer the following questions:

(a) Who was the United States president who made the decision to use the atomic bomb against Japan?

(b) On 9 August 1945 a second bomb was dropped on Japan.

  (i) Where was it dropped?

  (ii) What was its nickname?

(c) Which elements were used to make the two atomic bombs?

(d) What were the devastating effects of the atomic bombs?

(e) Did the atomic bombs have any long-lasting effects on our society?

(f) Do you think scientists regret creating such a bomb?

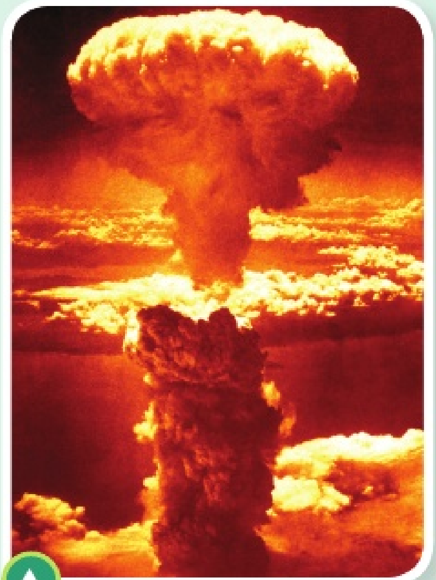


Figure 1.24 *'Little Boy', the bomb dropped on Hiroshima*

Scientist Biography

Watch the video to find out more about Einstein and his discoveries.

MODULE 2

Observing change

Learning outcomes

At the end of this module you will be able to:

- Distinguish between a physical change and a chemical change 8.2.2.2
- Outline the changes that may occur during a chemical reaction 8.2.2.2
- Describe the law of conservation of mass 8.2.3.4
- Calculate mass fractions of elements in a substance 8.2.3.1
- Determine the ratio of substances by means of experiment 8.2.3.2
- Describe chemical reactions in nature and daily human activities 8.2.2.2
- Write word equations for reactants and products in chemical reactions 8.2.3.3
- Classify chemical reactions by number and composition of initial and formed substances 8.2.2.1

Keywords

✔ chemical change ✔ physical change ✔ hydrochloric acid ✔ sodium hydroxide ✔ spatula ✔ conservation ✔ combination ✔ decomposition

In science there are many changes, but can they always be reversed?

Physical change

Our study of chemistry leads us to a better understanding of our world and the processes by which materials can change and be changed. We know that matter exists in three states (see Figure 2.1 for a recap).

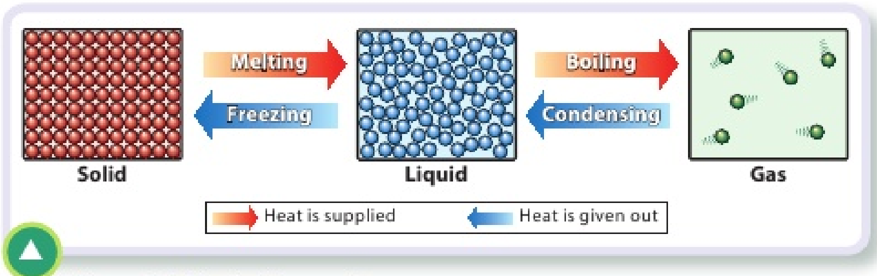


Figure 2.1 *Physical changes in matter*

The changes shown in Figure 2.1 are reversible and no new substance is formed; there is just a change of state. This is called a **physical change**, where there is no change in particles, just their arrangement and their energy.

2.1 Look at the following pictures and discuss in groups why these may be physical changes.

For each of them, discuss the questions:

(a) Are any new substances made?

(b) Can you reverse the changes (that is, go back to what it looked like at the start)?

Blowing up a balloon

Melting chocolate

Melting ice

Sharpening a pencil

Figure 2.2 *Different matter undergoing physical changes*

2.2 Now consider what might happen when a match is ignited and burning.

Discuss these questions:

(a) What changes occur to the match when it has been lit? List them.

(b) Is energy produced?

(c) What forms of energy have been produced?

(d) Can you reverse this change?

(e) Is this a physical change?

Figure 2.3 *A burning match*

Chemical change

A burning match is an example of a chemical change, where a chemical reaction takes place and a new substance is formed. During a chemical change energy may be released or absorbed. This means chemical changes are usually very difficult to reverse.

Here are some everyday examples of chemical change:

Figure 2.4 *Iron reacts with oxygen and water to form rust (consisting of iron(III) oxides $Fe_2O_3 \cdot nH_2O$)*

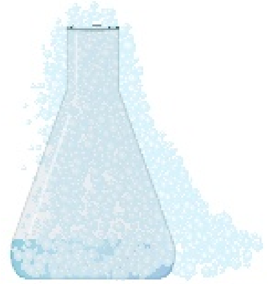
Figure 2.5 *When you add a dissolvable tablet to water you may get bubbles*

2.3 List two differences between physical change and chemical change.

2.4 State which of the following changes are physical changes and which are chemical changes:

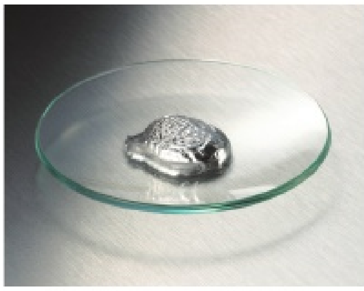
Burning wood	Dissolving sugar in tea	Baking a cake
Making toast	Dicing potatoes	Cutting paper

2.5 Look at the following pictures within your group and discuss what changes occur.



(a) *Raw liver in hydrogen peroxide*

(b) *Ice hotel in Sweden, which is built every year*



(c) *Melting gallium*

(d) *Combustion engine of a space ship*

Figure 2.6 *What changes are taking or have taken place?*

2.6 Look at this candle.
Which label (A or B) shows a physical change and which shows a chemical change? Explain your answer.

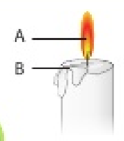


Figure 2.7 *Changes taking place to a candle*

Chemical reactions

Many useful materials are produced from chemical changes (also known as chemical reactions), such as plastics, which are made from reactions with oil.

The chemicals that react together are called **reactants** and the substances formed are called the **products**.

Observing change

During a chemical reaction the atoms rearrange to form a new substance. There are some signs that indicate this has occurred:

- Colour change
- A change in temperature
- Light is emitted
- Bubbles of gas or precipitate are produced.

Figure 2.8 *A chemical reaction takes place when a firework is lit*

Figure 2.9 *When you add a dissolvable tablet to water and you get bubbles then a chemical change occurs*

Did you know?

Have you ever smelled a rotten egg? It smells completely different from fresh eggs. A chemical change during spoilage of eggs causes this odour.

Research R2 Research R3 Research R4 Research R5 Portfolio 10

Activity 2.1

Question

What changes are taking place in chemical reactions?

Equipment needed

4 test tubes	Thermometer	Baking soda
Test tube rack	Sodium hydroxide (0.1 M)	Copper oxide
Spatula	Sulfuric acid (0.1 M)	Water
Iron nails	Vinegar	Universal indicator

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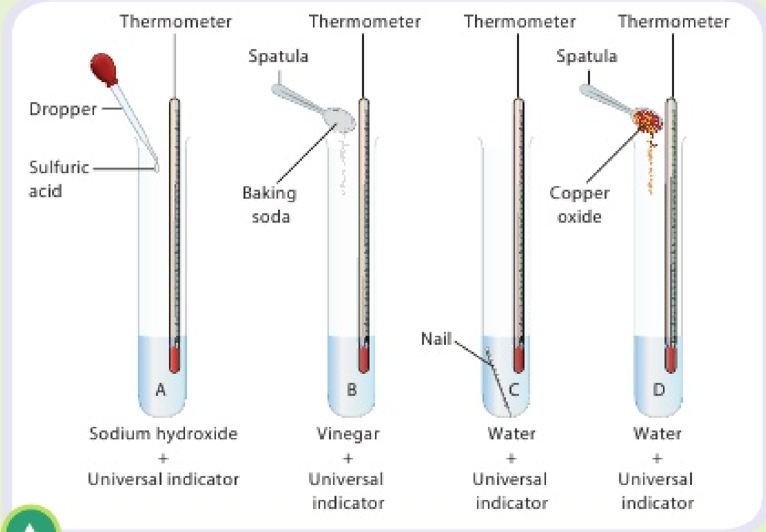


Figure 2.10 *Investigating changes*

Conducting the activity

You are going to add substances to four different test tubes. For each test tube, record:

(a) Any change in temperature
(b) Any change in colour
(c) Any noticeable odour
(d) If there is formation of bubbles.

1. *Test tube A* Place 3 cm^3 of sodium hydroxide into the test tube. Add a few drops of universal indicator. Measure the temperature of the solution.
   Add 3 cm^3 of sulfuric acid. Measure the temperature again. Note any changes.
2. *Test tube B* Place 3 cm^3 of vinegar into a test tube. Add a few drops of universal indicator. Measure the temperature. Add half a spatula of baking soda. Measure the temperature again. Note any changes.
3. *Test tube C* Place 3 cm^3 of water and a few drops of universal indicator into a test tube. Measure the temperature. Add an iron nail. Measure the temperature again. Note any changes.
4. *Test tube D* Place 3 cm^3 of water into a test tube and add a few drops of universal indicator. Add half a spatula of copper oxide. Measure the temperature again. Note any changes.

Results

Look at your results for each test tube and answer the following questions:

(a) In which test tubes did a physical change take place, and in which did a chemical change take place?
(b) Write down the properties that tell you if there has been a chemical reaction (chemical change).

Conservation of mass

We know that matter can be changed from one form into another, but during physical and chemical changes, is there a change in the overall mass of the matter, or does it remain constant? We will investigate this in the next activity.

Research R2 R3 R4 Portfolio 12

Activity 2.2

Question

Does mass change during physical changes and chemical changes?

Equipment needed

Electronic balance	Water	Hydrochloric acid (0.1 M)
4 beakers	Sugar	Universal indicator
Spatula	Sodium hydroxide (0.1 M)	

Conducting the activity

Reaction A

1. Place two beakers (50 ml) on a balance and press the zero button.
2. Place 10 cm^3 of water in one beaker and add a few drops of universal indicator.
3. Place one spatula of sugar into the other beaker.

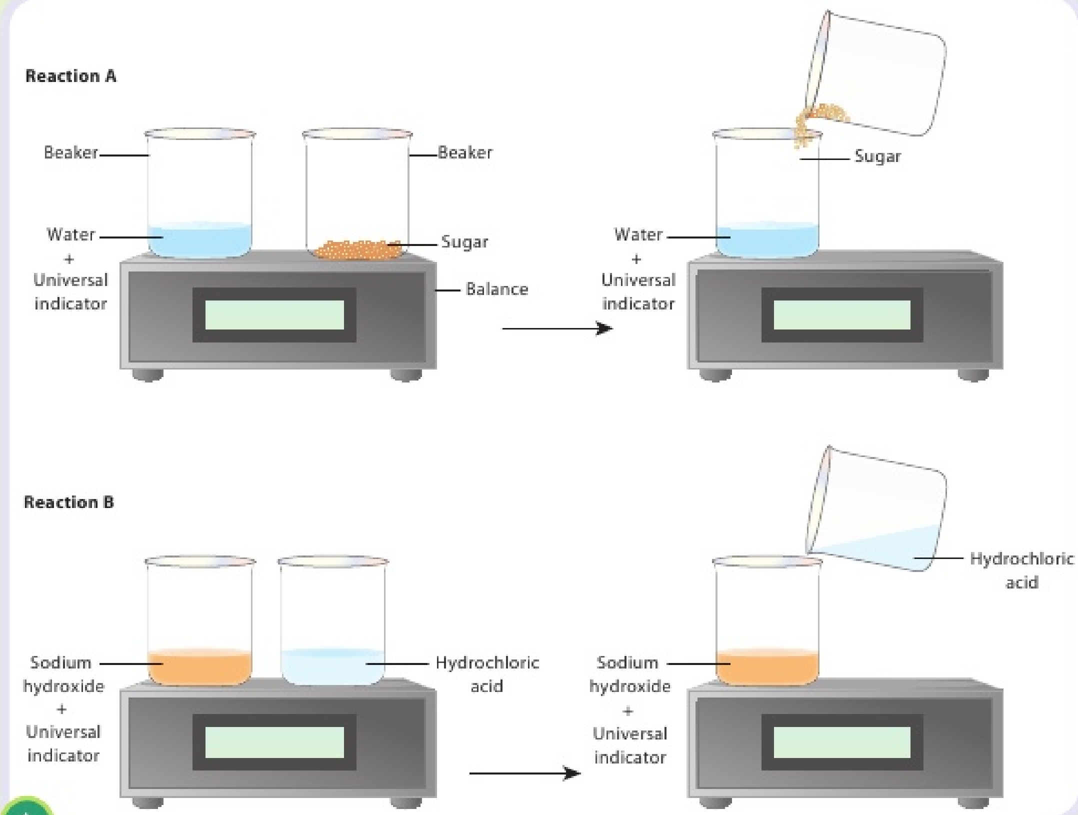


Figure 2.11 *Investigating changes in mass during physical changes and chemical changes*

4. Record the total mass on the balance.
5. Add the sugar to the water and replace the empty beaker on the balance.
6. Record the total mass.

Reaction B

1. Place two beakers (50 ml) on a balance and press the zero button.
2. Add 10 cm^3 of sodium hydroxide into one beaker and add a few drops of universal indicator.
3. Add 10 cm^3 of hydrochloric acid into a second beaker on the balance.
4. Record the total mass on the balance.
5. Add the hydrochloric acid solution to the sodium hydroxide solution and replace the empty beaker on the balance.
6. Record the total mass.

Q Research R5 Communicating C2

Divide into groups. Within your group, analyse your results and answer the following questions:

2.7 Did you notice any changes during reaction A and reaction B? Which reaction, A or B, showed a chemical reaction? What evidence do you have to support this?

2.8 Was the initial mass the same as the final mass for each reaction?

2.9 During a physical or chemical change is there a change in the overall mass?

Law of conservation of mass

Antoine Lavoisier, a French chemist in the eighteenth century, discovered the law of the conservation of mass. He discovered that the mass of a substance cannot be created or destroyed, so during a physical and chemical change there is no change in the overall mass.

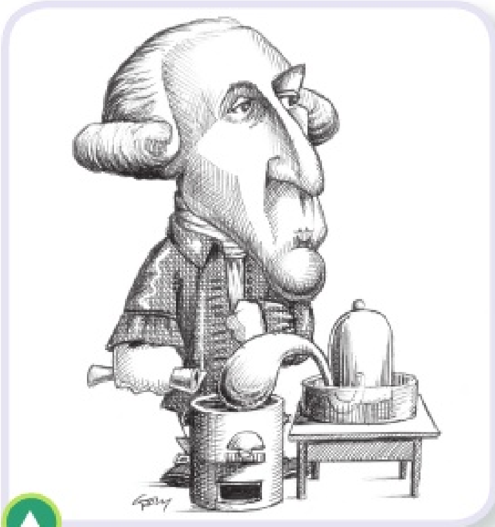


Figure 2.12 *Antoine Lavoisier*

Did you know?

If you leave a can of soda open for a long time the carbonic acid (fizz) decomposes into carbon dioxide and water. It loses its bubbles. This is another example of a chemical change.

Particle model diagrams

A particle model diagram shows how particles (atoms) rearrange to form a new substance during a chemical reaction. Figure 2.13 and Figure 2.14 are examples of these diagrams.

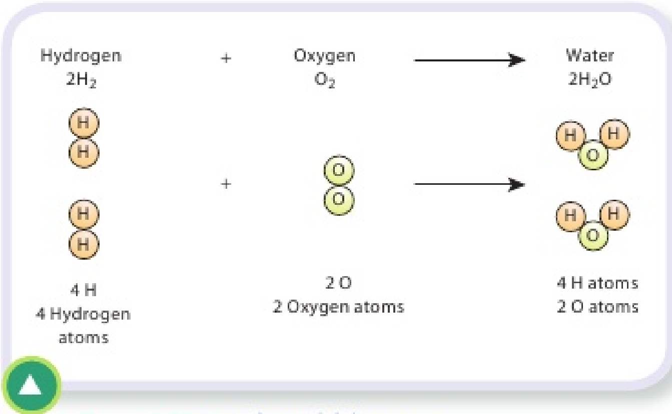


Figure 2.13 *Particle model diagrams*

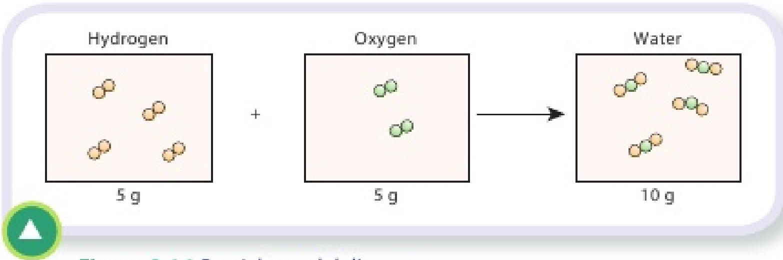


Figure 2.14 *Particle model diagrams*

Q

2.10 Draw a particle model diagram for Figure 2.15:

Copper + Oxygen → Copper oxide

Figure 2.15

2.11 According to the law of the conservation of mass, how much copper was produced? (See Figure 2.16.)

Zinc +	Copper Sulfate →	Zinc Sulfate +	Copper
65 g	160 g	161 g	x g

Figure 2.16

2.12 If 50 grams of sodium reacts with chlorine to form 128 grams of salt, how many grams of chlorine reacted? (See Figure 2.17.)

Sodium +	Chlorine →	Sodium chloride (salt)
23 g	x g	58.5 g

Figure 2.17

Types of Reaction

There are four basic types of chemical reaction.

- A chemical reaction in which **two** reactants react to form just **one** product is called a **combination** reaction.
- A chemical reaction in which **one** reactant breaks down to give **two** or more products is called a **decomposition** reaction. In this decomposition reaction:

  calcium carbonate $\longrightarrow$ calcium oxide + carbon dioxide

  heat is needed to break down the more complex molecule into the two products.
- A **single replacement** reaction is where **one** element replaces another element in a compound. This happens because this element is more reactive than the element in the compound.

  sodium + water $\longrightarrow$ sodium hydroxide + hydrogen

- A **double replacement** reaction occurs when **two** compounds exchange their ionic partners. The positive ion in the first compound combines with the negative ion in the second compound and the negative ion in the first compound combines with the positive ion in the second.

  calcium nitrate + sulfuric acid $\longrightarrow$ calcium sulfate + nitric acid

Q Understanding U2

2.13 What types of reaction are shown in Figures 2.15, 2.16 and 2.17?

2.14 What types of reaction are these ?

(a) zinc + hydrochloric acid $\longrightarrow$ hydrogen + zinc chloride

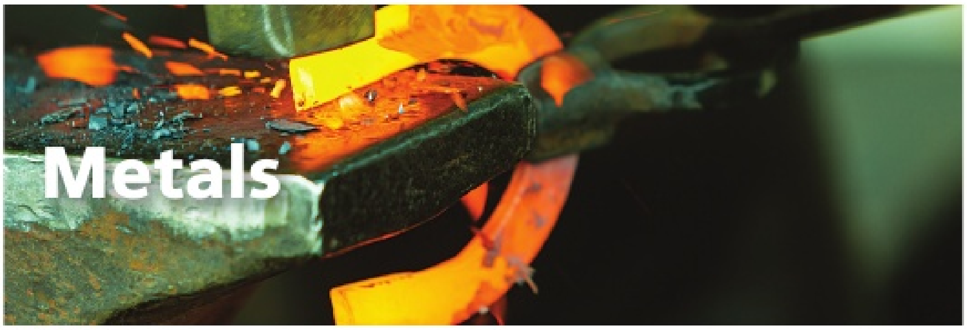
(b) sulfur + oxygen $\longrightarrow$ sulfur dioxide

(c) water $\longrightarrow$ hydrogen + oxygen

(d) hydrochloric acid + sodium hydroxide $\longrightarrow$ sodium chloride + water

[see **Module 4** for more on different types of reaction]

MODULE 3



Metals

Learning outcomes

At the end of this module you will be able to:

- Identify that some metals oxidise faster than others 8.2.4.1
- Explain conditions which effect corrosion 8.2.4.3
- Describe reactions of very reactive metals with water 8.2.4.2
- Explain reactions of metals with acid solutions 8.2.4.4
- Identify that different metals have different rates of reactivity 8.2.4.7
- Make word equations of reactions between metals and acids 8.2.4.5
- Plan and carry out displacement reactions of metals in salt solutions 8.2.4.6

Keywords

✔ physical properties ✔ melting point ✔ conductivity
✔ density ✔ shiny ✔ reactivity ✔ word equation
✔ corrosion ✔ rust ✔ alkali metal ✔ metal oxide

The way a metal is used depends on its properties. For example:

- Aluminium is quite light yet strong, so it is used to make aircraft.
- Copper is a brown-coloured metal which is a good electrical conductor and it stretches easily, so it is used for electrical wiring.

How is an element classed as a metal?

Figure 3.1 shows some of the typical properties of metals

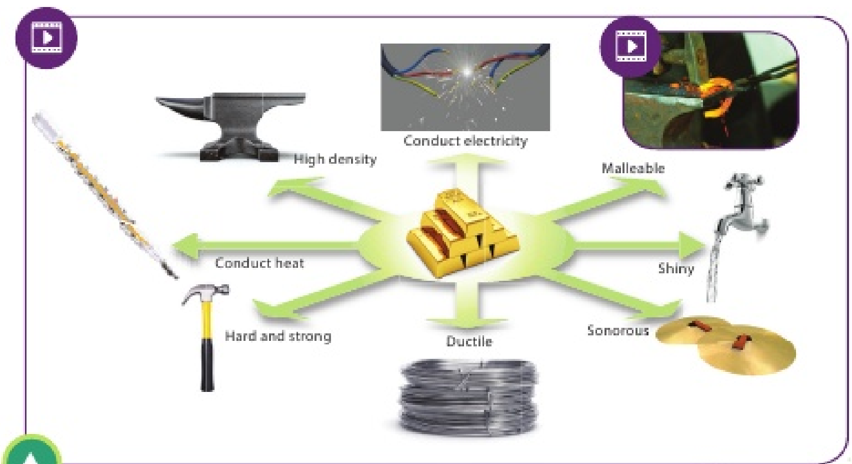


Figure 3.1 *The physical properties of metals*

Not all metals have all these properties. For example:

- The alkali metals are soft with low densities, e.g. sodium and potassium.
- Mercury is a liquid at room temperature.

The alkali metals

Alkali metals are very reactive and behave similarly. Elements in this group include:

- Lithium (Li)
- Sodium (Na)
- Potassium (K).

Understanding U2 Research R3 Portfolio 18

Activity 3.1

Question

What do alkali metals have in common?

Equipment needed

Water basin	Battery	Knife	Bunsen burner
Crocodile clips	Bulb	Combustion spoon	Sodium

Conducting the activity

1. Watch as your teacher demonstrates each property of the alkali metal.
2. Make notes in the table below.

How is the metal stored?	
Does it look like a metal?	
What do you notice when the metal is cut?	
Is it a good conductor of electricity?	
Does the metal melt easily?	
What happens when the metal reacts with water?	
What colour does litmus paper turn in the water?	

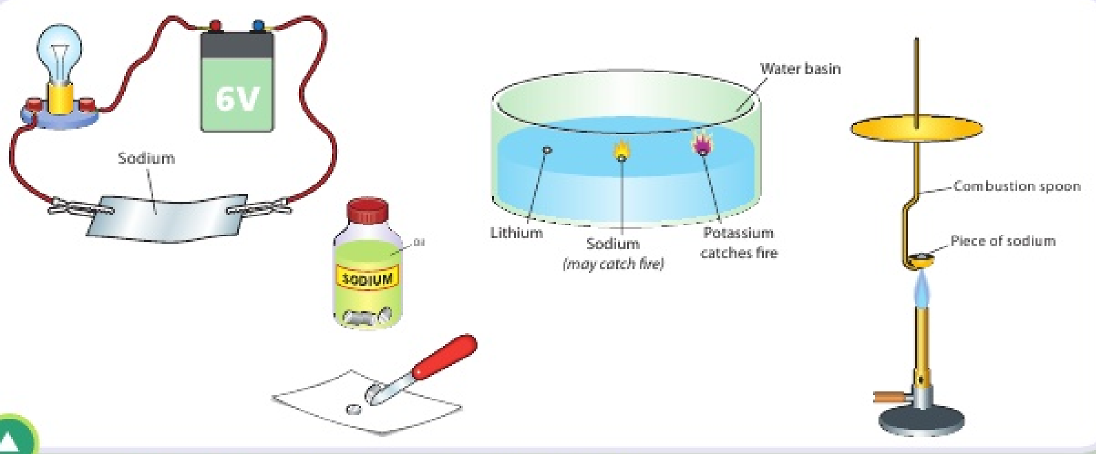


Figure 3.2 *Examining the properties of alkali metals*

Q Understanding U2 Research R5

3.1 Now complete the mind map:

From this demonstration you learned that alkali metals have the following in common:

Alkali metals

- _________ and easily cut
- Quickly lose their _________
- Stored in _________
- Conduct _________
- Low density (float on _________)
- Low _________ point

These are some of the physical properties that all alkali metals share. Grouping elements together makes life easier for chemists.

By learning how sodium reacts and behaves, we can also predict how the other alkali metals will behave.

Did you know?

There are ninety-two elements in the periodic table that occur in nature. All of the other elements are strictly synthetic. Technetium was the first element to be made artificially.

Reactions of alkali metals with water

When an alkali metal such as **lithium** is added to water it floats on the surface of the water and forms a ball shape. Hydrogen gas is released and lithium hydroxide is formed, which dissolves in the water.

lithium + water $\longrightarrow$ lithium hydroxide + hydrogen

Sodium is more reactive than lithium. It may even catch fire (orange flame) as it fizzes about on the surface.

sodium + water $\longrightarrow$ sodium hydroxide + hydrogen

Potassium is even more reactive than sodium.

Figure 3.3 *Potassium reacting with water*

Q Understanding U2

3.2 Can you predict the products of the reaction between potassium and water?

Reactions of alkali metals with air

All metals are shiny. Alkali metals look dull, but when they are cut a shiny surface is revealed. This shiny surface quickly goes dull again as the metal reacts with oxygen in the air, forming the metal oxide.

For lithium the reaction with oxygen is:

$$\boxed{\text{lithium} + \text{oxygen} \longrightarrow \text{lithium oxide}}$$

For sodium the reaction with oxygen is:

$$\boxed{\text{sodium} + \text{oxygen} \longrightarrow \text{sodium oxide}}$$

Q Understanding U2

3.3 Complete the following sentence: The elements in group ________________ are called the alkali metals. Because they react vigorously with water they are stored under ________________. Lithium, ________________ and ________________ are examples of alkali metals.

Corrosion of metals

Most metals are reactive. They react with air, water and other substances and corrode. We have seen that when a piece of sodium is cut, a shiny surface can be seen. This will quickly lose its shine as it reacts with oxygen.

Other metals such as gold and silver are unreactive and never corrode. They remain shiny and are ideal for making jewellery. In general, the more reactive a metal is, the more quickly it corrodes.

Rusting

The corrosion of iron is called **rusting**.

Iron (or steel) reacts with both water and oxygen (air) and forms rust. This orange/brown substance flakes away and eventually all the metal is corroded.

Figure 3.4 *Rusting experiment*

R2 R3 C2

Portfolio 19

Activity 3.2

Question

Are oxygen and water necessary for rusting?

Equipment needed

Test tubes	Oil
Steel (iron) nails	Drying agent e.g. calcium chloride
Stoppers	Water
Test-tube rack	

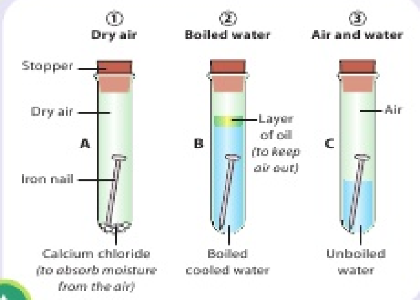


Figure 3.5 *Testing conditions for rusting*

Conducting the activity

1. Set up three test tubes as shown in Figure 3.5.
2. Examine the test tubes after one week.
3. Note your observations.

U2 R5

3.4 What can you conclude about corrosion from this experiment?

Rust prevention

Rusting of steel and iron costs us millions of Tenge each year. Any object containing iron will rust if exposed to water and oxygen. Rusting can be prevented by coating the metal with a material to prevent water and oxygen coming into contact with it. Below are some examples.

S2

3.5 Can you say what the missing method is in Figure 3.6?

To prevent rusting

- ______________
- Plastic coating
- Galvanising (i.e. coating with zinc)
- Tin plating (coat with a layer of tin)
- Greasing (coat with grease or oil)

Figure 3.6 *Ways to prevent rusting*

Reactivities of metals

Most metals want to lose electrons in order to have a full outer shell of electrons and become more stable. They will react with substances like water and air to achieve this. Some metals are more reactive than others.
In the activities below we look at how four different metals react with water and with acid.

Figure 3.7 *Copper reacts with nitric acid but aluminium does not*

Q U2 R1

3.6 What do you need to keep the same to make a fair test?

R2 R3 R5

Portfolio 20

Activity 3.3

Question

How do four different metals react with water?

Equipment needed

Test-tube rack	Water	Magnesium
Four test tubes	Calcium	Zinc and Copper

Conducting the activity

1. Set up the apparatus as shown.
2. Study each metal in the test tubes and note your observations.

Four metals	Observations

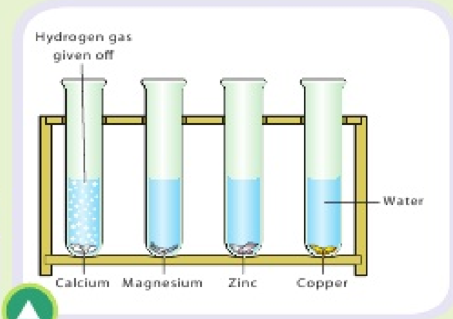


Figure 3.8 *The reactivity of some metals with water*

Q U2 R5

3.7 What conclusions from the experiment can you draw about the reactivities of these metals?

3.8 Which of these metals do you predict will be most reactive in acid solutions?

3.9 What kind of chemical changes do you predict you will see?

Research R2 Research R3

Activity 3.4

Question

What are the different reactivities of four metals with dilute acid?

Equipment needed

Test-tube rack
Four test tubes
Dilute acid
Calcium
Magnesium
Zinc and Copper

Conducting the activity

1. Set up the apparatus as shown (remember to ensure fair testing).
2. Study each test tube and compare the amount of bubbles of hydrogen gas produced.

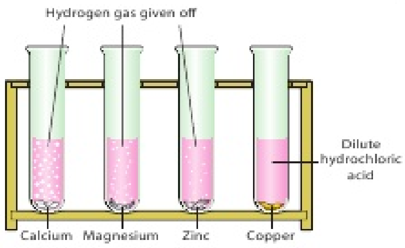


Figure 3.9 *The reactivity of some metals with dilute hydrochloric acid*

Q Understanding U2 Research R4 Research R5

3.10 What conclusions from the experiment can you draw about the reactivities of these metals?

Q Understanding U2

3.11 Can you write the word equations for the reactions in Activity 3.4?

Relative reactivities

Q Understanding U2 Research R5

3.12 Based on your observations in Activities 3.3 and 3.4 put the four chemicals on the reactivity ladder.

Ca Mg Zn Cu

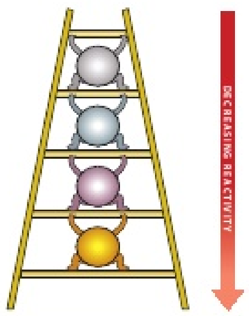


Figure 3.10 *Metals in decreasing order of reactivity*

Displacement reactions between metals and salts.

We have seen from Activity 3.3 and Activity 3.4 that some metals are more reactive than others. Another way to show this is by carrying out a displacement reaction involving a metal and a compound of a different metal.

Q U2 Understanding

3.13 Look at Table 3.1. From this table can you predict what will happen when a strip of magnesium is placed in a test tube of iron sulphate solution?

Table 3.1 **Reactivity series for common metals**

Element	Reactivity
Potassium	Most
Sodium	
Calcium	
Magnesium	
Aluminium	
Carbon	
Zinc	
Iron	
Tin	
Copper	Least

R3 Research R4 Research

Activity 3.5

Question

What do you observe when your teacher places each metal in each solution?

Equipment needed

3 x test tube racks
3 x test tubes with magnesium sulphate solution
3 x test tubes with zinc sulphate solution
3 x test tubes with iron sulphate solution
3 strips of zinc
3 strips of magnesium
Iron fillings

Conducting the activity

1. Place three samples of each solution in three test tubes in each rack.
2. Then place a strip of each metal in each test tube and allow time for any reaction to be observed.
3. Complete the table below with a tick or x if you observe a reaction when each metal is placed in each solution.

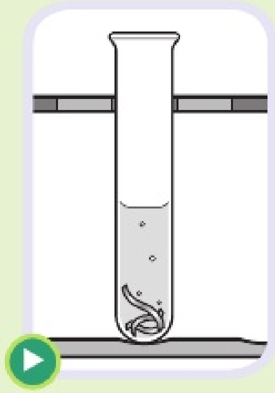


Figure 3.11 *Displacement reactions*

	Magnesium	Zinc	Iron
Magnesium sulphate			
Iron sulphate			
Zinc sulphate			

Q Understanding U2 Research R5

3.14 In what order of reactivity would you put these three metals?

3.15 Can you write the word equation for each reaction you observed?

MODULE 4

Reactions, Equations and Calculations

Learning outcomes

At the end of this module you will be able to:

- Explain exothermic and endothermic reactions 8.1.3.1
- Explain thermal change of energy in terms of the theory of particles 8.1.1.5
- Describe the processes involved in the combustion of hydrocarbons 8.3.1.1
- Explain the greenhouse effect and potential solutions 8.3.2.1
- Assess the potential of the use of different fuels and their environmental impact 8.3.1.4
- Explain mole as the unit of measurement of the amount of a substance 8.1.1.1
- Perform a range of calculations relating to the mass of substances, chemical equations and chemical reactions 8.1.1.2 8.2.3.5
- Perform a range of calculations relating to Avogadro's law and volume and density of gases 8.2.3.6 8.2.3.7 8.2.3.8

Keywords

✔ exothermic ✔ endothermic ✔ bond energy ✔ activation energy
✔ energy profile diagrams ✔ reactants and products ✔ hydrocarbons
✔ combustion ✔ mole ✔ molecular mass ✔ balanced equation

What energy changes occur during a chemical reaction?

One piece of evidence to show that a chemical change or reaction has taken place is to identify a change in temperature. In a chemical reaction, heat energy is usually released or taken from its surroundings.

What is an exothermic reaction?

An exothermic reaction is a reaction where energy is transferred from the chemicals to the surroundings. This is where the temperature of the reaction mixture usually rises. For example, combustion, such as the burning of fuels or fireworks is an exothermic reaction.

Figure 4.1 *An exothermic reaction is taking place on this barbecue*

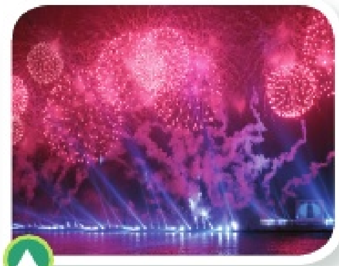


Figure 4.2 *An exothermic reaction takes place when fireworks burn*

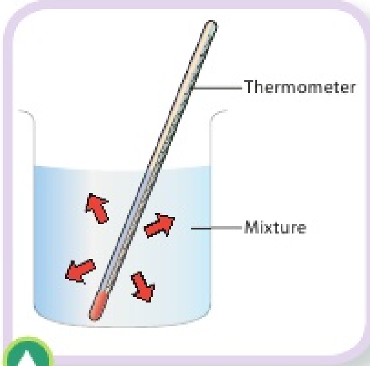


Figure 4.3 *Heat is transferred from the chemicals to the surroundings*

What is an endothermic reaction?

An endothermic reaction is where energy is absorbed by the chemicals from the surroundings in order for the reaction to take place. The reaction mixture usually shows a fall in temperature. For example, ammonium chloride dissolving in water is an endothermic reaction.

Figure 4.4 *Solid barium hydroxide and solid ammonium chloride in a flask on a damp piece of wood. As the two chemicals mix, an endothermic reaction occurs causing a drop in temperature which freezes the water between the conical flask and the block of wood*

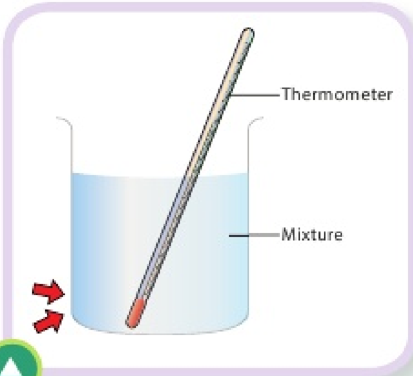


Figure 4.5 *Heat is absorbed by the chemicals from the surroundings*

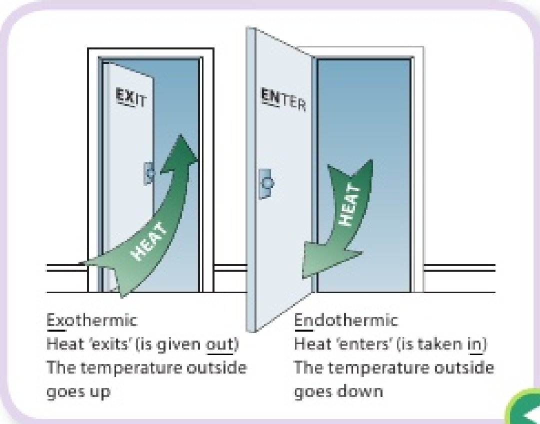


Figure 4.6 *Heat can be lost or gained during a reaction*

Q U3 Understanding

4.1 Distinguish between an exothermic reaction and an endothermic reaction.

Activity 4.1

Question

What energy changes are occurring in chemical reactions?

Equipment needed

4 polystyrene cups (labelled A–D)
Graduated cylinder
Thermometer
Sodium hydroxide (0.1 M)
Hydrochloric acid (0.1 M)
Sodium hydrogen carbonate solution
Citric acid
Ammonium nitrate
Water
Dilute sulphuric acid (0.1 M)
Magnesium ribbon

Conducting the activity

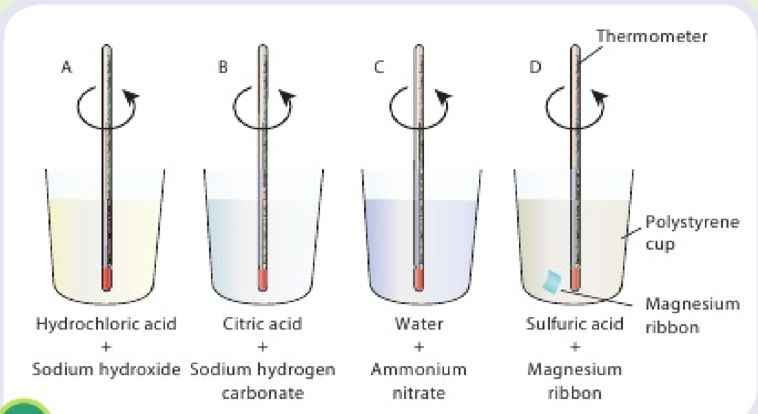


Figure 4.7

Reaction A

1. Measure 10 cm^3 of sodium hydroxide solution using a graduated cylinder and place it in polystyrene cup A.
2. Measure the temperature with the thermometer and record the result in your table (see below).
3. Add 10 cm^3 of hydrochloric acid. Stir with the thermometer; measure the temperature and record the result in your table.

Reaction B

4. Add 10 cm^3 of sodium hydrogen carbonate solution to polystyrene cup B. Measure the temperature and record the results in your table.
5. Add 10 cm^3 of citric acid. Stir with the thermometer; measure the temperature and record the result in your table.

Reaction C

6. Place 10 cm^3 of water in polystyrene cup C. Measure the temperature and record the result in your table.
7. Add 1 g of ammonium nitrate. Stir with the thermometer; measure the temperature and record the result in your table.

Reaction D

8. Add 10 cm^3 of sulphuric acid to polystyrene cup D. Measure the temperature and record the result in your table.
9. Add 1 g of magnesium ribbon. Stir with the thermometer; measure the temperature and record the result in your table.

Reaction	Initial temperature (°C)	Final temperature (°C)	Endothermic or exothermic
A			
B			
C			
D			

4.2 Did you find more exothermic or endothermic reactions?

4.3 The first reaction is between an acid and a base. What do you call this type of reaction?

4.4 What gas is produced in reaction B?

4.5 What gas is produced in reaction D?

4.6 Why did you use polystyrene cups and not glass beakers?

Did you know?

Gunpowder was discovered about a thousand years ago by the Chinese, and fireworks were invented in China to scare off evil spirits.

4.7 What is the clue in the terms 'exothermic' and 'endothermic' that indicates that they are linked to heat energy?

How are bonds made and broken?

During a chemical reaction, old bonds are broken (reactants) and new bonds are formed (products). Breaking bonds requires energy, so it is an endothermic process; but energy is released when new bonds are formed, so bond formation is an exothermic process.

Bond breaking – endothermic

Figures 4.8, 4.9 and 4.10 illustrate endothermic reactions.

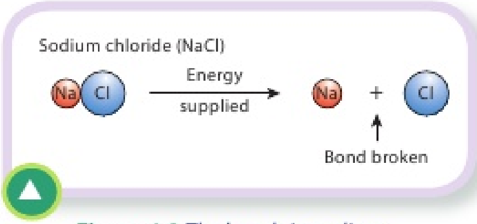


Figure 4.8 *The bonds in sodium chloride are broken*

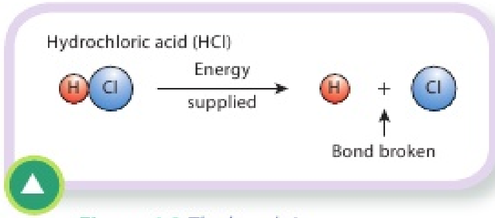


Figure 4.9 *The bonds in hydrochloric acid are broken*

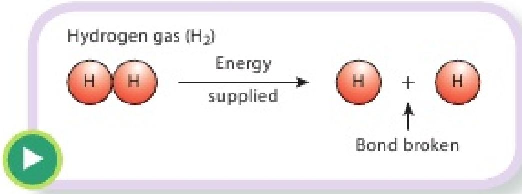


Figure 4.10 *The bonds in the hydrogen molecule are broken*

Bond formation – exothermic

Figure 4.11 and Figure 4.12 illustrate exothermic reactions.

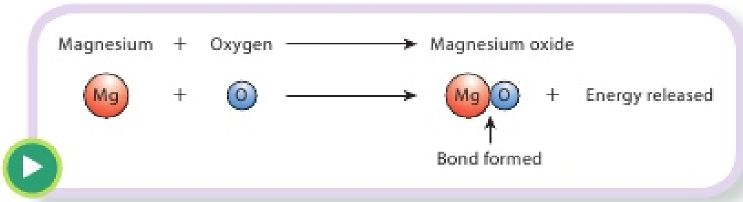


Figure 4.11 *The magnesium and oxygen atoms bond to form magnesium oxide and energy is released*

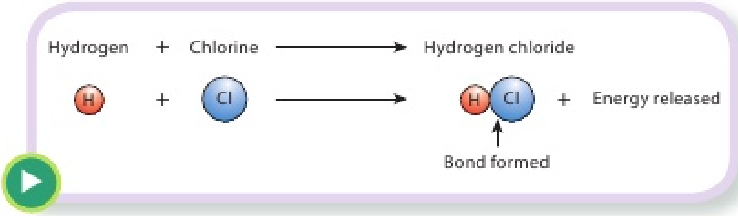


Figure 4.12 *The hydrogen and chlorine atoms bond to form hydrogen chloride and energy is released*

What is bond energy?

Bond energy is the energy required to break different bonds. The units are **kilojoules per mol** (**KJ/mol**). A mole (symbol **mol**) is the unit used to measure the amount of a substance. Table 4.1 shows the amount of energy required to break different bonds.

Table 4.1 The amount of energy required to break different bonds

Bond	Bond energy (KJ/mol)
H—H	436
Cl—Cl	242
H—Cl	431
C—H	413
C—C	347
C—O	335

4.8 Sodium reacts with water in an exothermic reaction. What does this suggest about the energy needed to break the bonds compared with the energy released when they are made?

4.9 Look at this table and answer the questions below.

Reaction	Starting temperature (°C)	Final temperature (°C)
A + B	19	27
C + D	20	25
E + F	19	17

(a) For each reaction say whether it is exothermic or endothermic. How can you tell?

(b) The volume of solution was the same in each reaction. Which had the largest energy change?

What are energy profile diagrams?

We can show the energy transfer in reactions on energy profile diagrams. The diagrams show us the energy stored in the reactants compared to the energy stored in the products, so they tell us if a reaction is exothermic or endothermic.

Activation energy, shown in energy profile diagrams, is the minimum energy that colliding particles must have for a reaction to occur.

Energy profile diagram for an exothermic reaction

Figure 4.13 shows an exothermic reaction because the products are at a lower energy than the reactants so energy has been given out. The difference in height (ΔH) is the symbol for the 'change in energy' in a reaction. ΔH is negative for an exothermic reaction. The difference in energy is given out as heat, so the temperature of the surroundings rises.

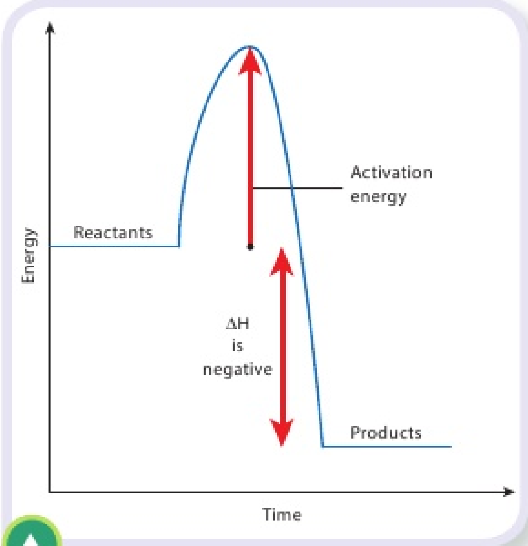


Figure 4.13 *An exothermic reaction energy profile diagram*

Energy profile diagram for an endothermic reaction

In Figure 4.14 the products are at a higher energy than the reactants so this shows an endothermic reaction. Extra energy was required to form the products, which was taken in from the surroundings; therefore, the temperature of the surroundings falls. ΔH (change in energy) is positive for endothermic reactions.

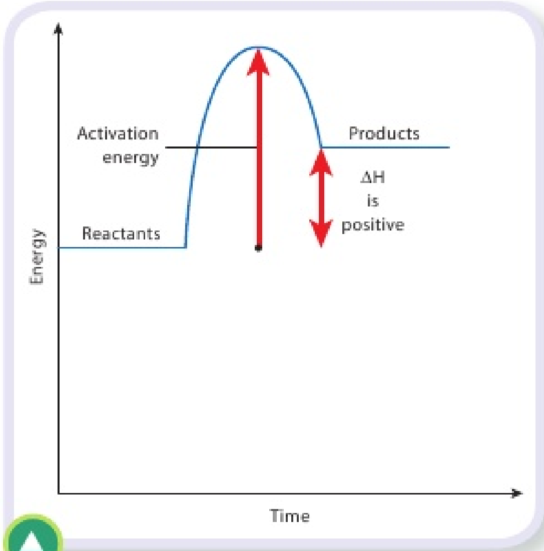


Figure 4.14 *An endothermic reaction energy profile diagram*

Q U3

4.10 Draw energy profile diagrams for exothermic and endothermic reactions.

4.11 When hydrochloric acid and sodium hydroxide react in a beaker, the temperature rises:

$$HCl + NaOH \longrightarrow NaCl + H_2O \qquad \Delta H = -58 \text{ KJ/mol}$$

(a) Is this reaction exothermic or endothermic?

(b) Draw an energy profile diagram to show this change.

Q U3 C1 C2

4.12 In your group, decide whether each of the following reactions are endothermic or exothermic and give reasons for your decision.

Reaction A Two chemicals are mixed together in a fume cupboard at room temperature. The reaction starts to fizz and a gas is produced, and the mixture soon catches fire with a purple flame.

Reaction B When two chemicals are combined in a beaker the outside of the beaker frosts up and it gets stuck to the desk due to the formation of ice.

Reaction C When a piece of sodium is added to chlorine gas in a flask a bright light is seen and a drop of water on the flask soon vaporises.

Reaction D Butane from the Bunsen burner is ignited to heat some water.

Fossil fuels and combustion

Energy comes from chemical reactions. We burn oil and gas to provide heat energy for our homes and many types of industry. We have seen that combustion is an exothermic reaction, it involves the rapid burning of a fuel with oxygen. We will now look at such reactions involving fossil fuels.

Fossil fuels

Fossil fuels are fuels formed from the remains of plants and animals that lived millions of years ago.

- Coal was formed from trees and ferns that died and were buried and compressed under swamps.
- Oil and natural gas were formed from animals and plants that lived in the sea. Their remains were buried under layers of sand on the seabed. As the pressure increased, they eventually formed oil and natural gas.

Fossil fuels are organic compounds, which means they originated in living things. They are composed mainly of two elements, **hydrogen** and **carbon**. Such compounds are called **hydrocarbons**.

Figure 4.15 *Biofuel produced from sunflowers*

Figure 4.16 *Burning fossil fuels releases carbon dioxide and water vapour into the atmosphere*

4.13 Identify four ways in which humans burn the fossil fuels.

Research R2 Research R3 Research R4 Portfolio 26

Activity 4.2

Question

What are products of combustion of a hydrocarbon?

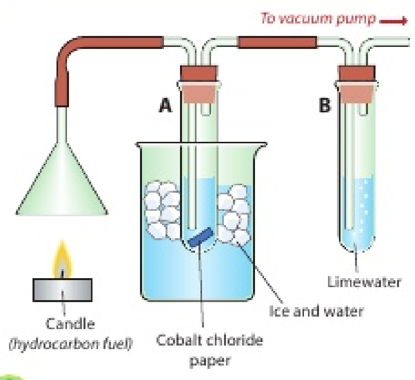


Figure 4.17 *Examining the products of the combustion of a hydrocarbon*

Equipment needed

Graduated cylinder
Basin
Steel wool
Water
A candle (or any fossil fuel)
Cobalt chloride paper
Limewater
Ice and water

Conducting the activity

1. Set up the apparatus as shown.
2. Allow the fuel to burn until a change is observed in each test tube.
3. Record your observations for test tube A and test tube B.

Understanding U4 Research R5

4.14 Why is test tube A placed in a beaker of iced water?

4.15 What conclusions can you draw from your observations?

Complete and incomplete combustion

The experiment in Figure 4.17 was an example of **complete combustion**. When a fuel burns with a plentiful supply of air, the different elements of the fuel and the oxygen react fully. A fuel such as natural gas – methane CH_4 – is a hydrocarbon. Hydrocarbons are compounds of hydrogen and carbon only. In reactions where hydrocarbons burn fully, carbon oxidises to form carbon dioxide and hydrogen oxidises to form water.

Incomplete combustion happens as a result of there not being sufficient air [oxygen] for full combustion to take place. Water, carbon monoxide and carbon in the form of soot are produced in such reactions. Carbon monoxide is a poisonous gas.

Complete combustion

hydrocarbon fuel + oxygen ⟶ water + carbon dioxide

Incomplete combustion

hydrocarbon fuel + oxygen ⟶ carbon monoxide + carbon + water

The increasing dependence of societies on the burning of fossil fuels such as oil, natural gas and coal to meet their energy needs has many disadvantages. Not only are these fuels non-renewable but increasing levels of carbon dioxide and other gases released into the air have very serious consequences for the environment.

The greenhouse effect

In a greenhouse, heat from the sun is trapped by the glass, causing the temperature in the greenhouse to rise. In much the same way, heat from the sun is trapped by carbon dioxide and other gases. This causes the temperature of the Earth to rise.

The levels of carbon dioxide in the atmosphere are increasing mainly due to the burning of fossil fuels. This causes the Earth to become warmer and leads to global warming. This in turn has many effects such as the melting of the polar ice caps and increased flooding of land.

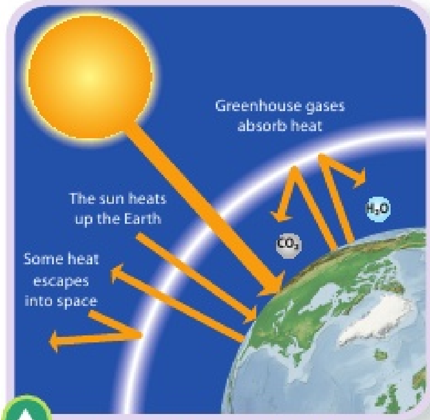


Figure 4.18 *How carbon dioxide affects the atmospheric temperature*

Figure 4.19 *The greenhouse effect*

Acid rain

Rainwater has always been mildly acidic. Carbon dioxide dissolves in water in the atmosphere to form carbonic acid, which has a pH of about 5.5.

Nowadays rain in parts of the world can have a pH as low as 3 or 4. The reason for this lowering of pH is mainly due to the burning of fossil fuels.

Sulphuric acid

Many fossil fuels, especially coal, contain sulphur. When burned the sulphur turns into sulphur dioxide. This gas dissolves in rainwater, forming sulphuric acid which is a strong acid.

Nitric acid

Car exhausts emit oxides of nitrogen. When they dissolve in water vapour in the atmosphere they form nitric acid which has a low pH.

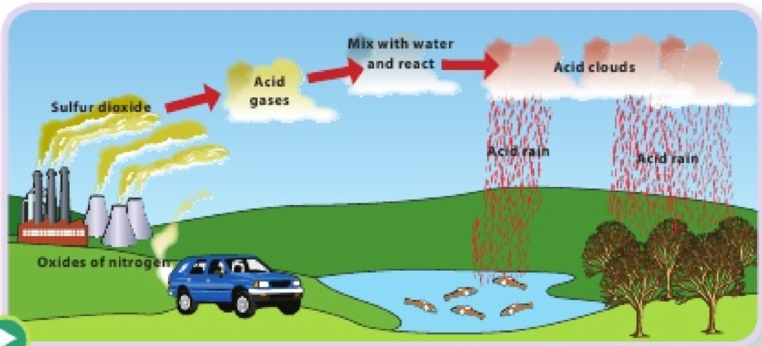


Figure 4.20 *How acid rain is formed*

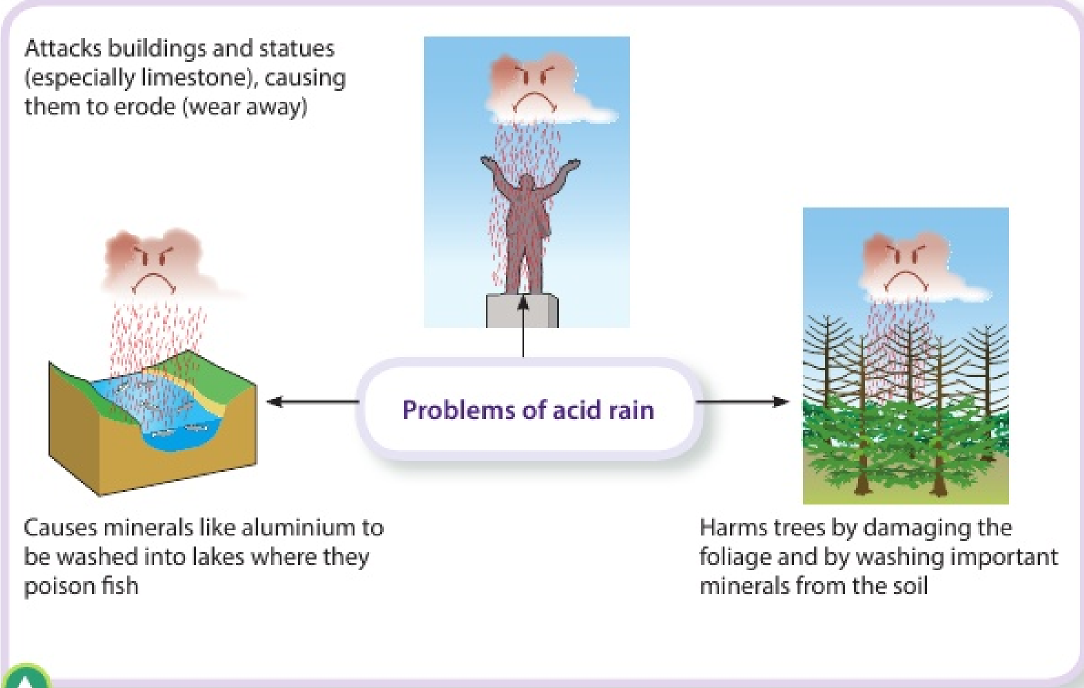


Figure 4.21 *Problems caused by acid rain*

Global warming and the effects of acid rain could be reduced in the following ways:

- Decreasing the use of fossil fuels as an energy source
- Encouraging the use and development of alternative energy sources such as solar and wind energy
- Using low-sulphur fuels such as natural gas (methane) instead of coal, turf and oil which have higher sulphur levels
- Fitting catalytic converters in cars to remove pollutants such as oxides of nitrogen
- Planting trees; this is essential as trees absorb carbon dioxide from the air and produce oxygen.

Figure 4.22 *An impression of the statue of liberty flooded as a result of global warming*

Q

4.16 Name a gas that can lead to acid rain. What is its chemical formula?

4.17 Explain the process that could cause the changes in **Figure 4.22**

Calculations

Chemists need to make a range of calculations relating to the mass of substances involved in chemical reactions.

Calculating relative molecular mass

The relative molecular mass of a substance can be calculated by adding the relative atomic masses of all the atoms in the molecule.

Table 4.2 Relative atomic mass values for use in calculations (You can find this number in the top left corner of each element in your periodic table.)

Element	A_r	Element	A_r	Element	A_r
Hydrogen	1	Phosphorous	31	Manganese	55
Carbon	12	Sulphur	32	Iron	56
Nitrogen	14	Chlorine	35.5	Copper	63.5
Oxygen	16	Potassium	39	Zinc	65
Fluorine	19	Argon	40	Bromine	80
Sodium	23	Calcium	40	Silver	108
Magnesium	24	Vanadium	51	Iodine	127
Aluminium	27	Chromium	52	Lead	207

The relative molecular mass of a substance, shown in grams, is one mole of that substance. So

Q Understanding U1

4.18 Calculate the relative molecular mass of

(a) carbon dioxide (CO_2)

(b) sulphuric acid (H_2SO_4)

one mole of carbon dioxide has a mass of 44 and one mole of sulphuric acid has a mass of 98.

Q Understanding U1

4.19 Calculate the relative molecular mass of

(a) sulphur dioxide (SO_2)

(b) nitric acid (HNO_3)

The Mole

The mole is the SI unit of amount of substance.

A mole of any substance is defined as the amount of substance that contains as many particles (atoms or molecules or ions) as there are atoms of ^{12}C in 12 g of ^{12}C.

The number of atoms of the ^{12}C isotope in 12 g of ^{12}C can be measured, and is found to be approximately 6×10^{23}. This is the number of particles per mole for all substances, and is called the Avogadro Constant (L):

$$L = 6 \times 10^{23}\ mol^{-1}$$

Understanding U1 Understanding U2 Society S1

4.20 Research the scientist Amadeo Avogadro and prepare a short fact file about his life and contribution to Chemistry.

Figure 4.23 *Amadeo Avogadro*

In one mole of neon, which is composed of atoms, there are 6×10^{23} atoms. In one mole of oxygen, which is composed of molecules, there are 6×10^{23} molecules. In one mole of water, which is composed of molecules, there are 6×10^{23} molecules. In one mole of sodium chloride (NaCl), which is composed of sodium ions and chloride ions, there are 6×10^{23} sodium ions and 6×10^{23} chloride ions.

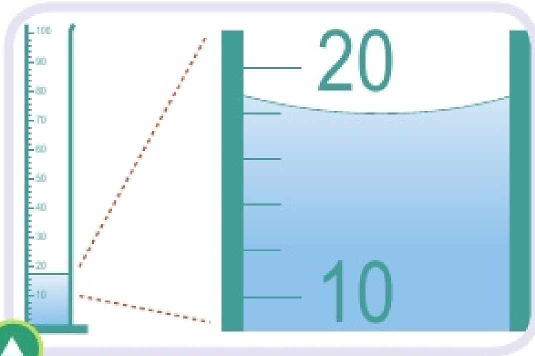


Figure 4.24 *One mole of water has a mass of 18 g and a volume of 18 cm³*

In a chemical equation, the ratio which holds between molecules, atoms or ions of the substances also holds in terms of moles. For example, consider the equation

$2H_{2(g)}$	+	$O_{2(g)}$	$\longrightarrow$	$2H_2O_{(g)}$
2 molecules		1 molecule		2 molecules

Multiplying by the Avogadro number:

	$2 \times 6 \times 10^{23}$	+	6×10^{23}	$\longrightarrow$	$2 \times 6 \times 10^{23}$
	molecules		molecules		molecules
=	2 moles		1 mole		2 moles

Example calculation 1

How many atoms are there in 4 moles of neon gas? (Neon gas is composed of neon atoms.)

Answer:

1 mole = 6×10^{23} atoms

4 moles = $4 \times 6 \times 10^{23}$ atoms

= 24×10^{23} atoms

= 2.4×10^{24} atoms

Q U1 Understanding

4.21 **(a)** How many molecules are there in 0.05 moles of carbon dioxide gas? (Carbon dioxide is composed of molecules.)

(b) How many atoms are there in 0.15 moles of sulphur dioxide gas? (Sulphur dioxide gas is composed of SO_2 molecules.)

(c) A sample of oxygen gas contains 3×10^{22} molecules. How many moles is this?

Mole calculations

From this basic equation a range of calculations can be made:

number of moles = mass ÷ relative molecular mass

Calculating the number of moles

Example calculation 2

How many moles are in 8 g of oxygen?

Answer:

$M_r(O_2) = 32$

32 g = mass of 1 mole of oxygen

8 g = mass of 8 / 32 moles

= 0.25 moles

Q U1 Understanding

4.22 How many moles are there in 14 g of nitrogen?

Calculating the mass

mass = number of moles x relative molecular mass

Example calculation 3

What is the mass of 7.5 moles of water?

Answer:

$M_r(H_2O) = 18$

Mass of 1 mole of water = 18 g

Mass of 7.5 moles of water = 7.5 x 18 g

= 135 g

Q U1 Understanding

4.23 What is the mass of

(a) 4 moles of iron?

(b) 0.3 moles of carbon dioxide?

(c) 0.125 moles of nitrogen?

(d) 0.75 moles of benzene (C_6H_6)?

(e) 0.5 moles of sodium hydroxide?

Calculating relative molecular mass

6 mol of carbon dioxide has a mass of 264 g. What is the relative molecular mass?

Q U1 Understanding

4.24 What is the equation you need to solve this problem?

4.25 What is the solution to the problem?

4.26 12 mol of water has a mass of 216g. What is the relative molecular mass?

4.27 3 mol of sulphuric acid has a mass of 294g? What is the relative molecular mass?

Balancing chemical equations

In balancing an equation, formulas cannot be altered in any way, but can only be multiplied by an appropriate number.

Example calculation 4

Methane reacts with oxygen, forming carbon dioxide and water vapour only. Write a balanced chemical equation for the reaction.

Answer:

The unbalanced equation is $CH_4 + O_2 = CO_2 + H_2O$

To balance the equation, each element is checked in turn to see whether or not there are equal numbers of atoms of that element on the two sides of the equation.

(a) **Carbon:** There is one carbon atom on each side of the equation, so CH_4 and CO_2 are left as they are.

(b) **Hydrogen:** There are four hydrogen atoms on the left-hand side, and only two on the right-hand side. H_2O must therefore be multiplied by two:
$CH_4 + O_2 = CO_2 + 2H_2O$

(c) **Oxygen:** There are now two oxygen atoms on the left-hand side, and four on the right-hand side. O_2 must therefore be multiplied by two:
$CH_4 + 2O_2 = CO_2 + 2H_2O$

Q U2 Understanding

4.28 Balance these equations:

(a) $H_2 + Cl_2 = HCl$ **(b)** $ZnS + O_2 = ZnO + SO_2$

Calculation of masses of reactants or products from balanced chemical equations

We can use the unit of the mole to calculate the masses of products and reactants in chemical reactions.

Example calculation 5

Magnesium reacts with water vapour to form magnesium oxide and hydrogen, according to the equation:

$$Mg + H_2O \longrightarrow MgO + H_2$$

If 60 g of magnesium is reacted with excess water vapour, calculate the mass of magnesium oxide produced.

Answer:

60 g Mg = 60 / 24 moles Mg = 2.5 moles Mg

Form the equation:

	Mg	+	H_2O	$\longrightarrow$	MgO	+	H_2
	1 mole		1 mole		1 mole		1 mole
Therefore:	2.5 moles		2.5 moles		2.5 moles		2.5 moles

1 mole MgO = 40 g MgO

2.5 moles MgO = 2.5 x 40 g = 100 g

Understanding U2

4.29 Hydrogen reacts with chlorine to form hydrogen chloride, according to the equation:

$$H_2 + Cl_2 \longrightarrow 2HCl$$

If 11.2 litres of hydrogen (measured at s.t.p.) are used to react with excess chlorine, what mass of hydrogen chloride is formed?

s.t.p. = standard temperature and pressure

Calculating the volume of gases

Avogadro's law

Equal volumes of gases, under the same conditions of temperature and pressure, contain equal numbers of molecules.

According to Avogadro's law, equal volumes of all gases, under the same conditions of temperature and pressure, have the same number of molecules. Therefore, one mole of any gas, for example nitrogen, should occupy the same volume as one mole of any other gas, for example carbon dioxide, under the same conditions of temperature and pressure. It is usual for volumes of gases to be compared at a standard temperature and pressure.

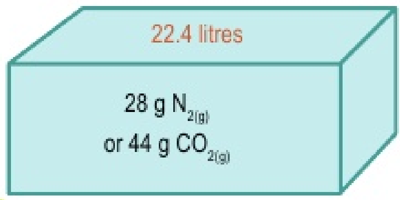


Figure 4.25 *One mole of any gas at s.t.p. occupies 22.4 litres*

The volume occupied by one mole of a gas at s.t.p. is called the molar volume at s.t.p. of the gas, and can be taken to be the same for all gases.

Molar volume at s.t.p.

- 22,4 L
- 22,400 cm^3
- 2,24 x 10^{-2} m^3

For example, one mole of nitrogen and one mole of carbon dioxide each occupy a volume of 22.4 at s.t.p.

Example calculation 6

What is the volume in litres at s.t.p. of 0.25 moles of chlorine gas?

Answer:

1 mole = 22.4 L at s.t.p.

0.25 moles = 0.25 x 22.4 L at s.t.p.

= 5.6 L

Understanding U1

4.30 How many moles are there in 280 cm^3 of nitrogen gas at s.t.p.?

4.31 What is the volume in litres at s.t.p. of

(a) 0.5 moles of oxygen gas

(b) 0.025 moles of nitrogen gas

(c) 3 moles of hydrogen sulphide gas?

MODULE 5
Oxygen, Nitrogen and Hydrogen

Learning outcomes

At the end of this module you will be able to:

- Identify the gases which make up the air 8.4.2.2
- Produce a sample of oxygen and study its properties 8.4.2.3
- Explain the difference between oxygen and ozone and describe the ozone layer in the atmosphere 8.4.2.4 8.4.2.5
- Produce a sample of hydrogen and study its properties 8.4.2.1

Keywords

✔ composition ✔ syringe ✔ neutral gas ✔ magnesium oxide
✔ odourless ✔ atmosphere ✔ CFCs ✔ ozone layer
✔ hydrochloric acid ✔ taper

Gases in the air

The atmosphere is the layer of gases around the Earth. These gases are known as air. Air is a mixture of gases which we need to keep us warm and to burn substances in.

Figure 5.1 *The Earth's atmosphere. The blue colour is caused by the reflection of sunlight as it passes through the atmosphere*

Two gases – nitrogen and oxygen – make up 99 per cent of air. The composition of air is not exactly the same all the time:

- There is more water vapour in the air on a damp day than a dry day.
- Carbon dioxide levels are higher over busy cities and industrial areas than in remote areas.

Nitrogen

Nitrogen gas makes up about four-fifths of air. It is an unreactive gas which we breathe in and out all the time. If nitrogen gas is cooled to almost -200°C it condenses to a liquid and is used to freeze things quickly.

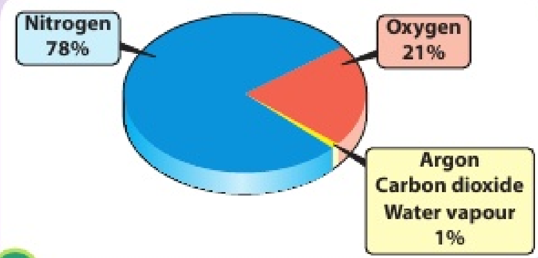


Figure 5.2 *The gases that make up air*

R2 R3 R4

Activity 5.1

Question

What is the percentage of oxygen in the air?

Equipment needed

Graduated cylinder
Steel wool
Basin
Water

Conducting the activity

1. Pack some steel wool into the bottom of a graduated cylinder or test tube.
2. Invert the cylinder or test tube in a basin containing water.
3. Leave for about one week.
4. Measure the height that the water rises up the cylinder.
5. Note your observations.

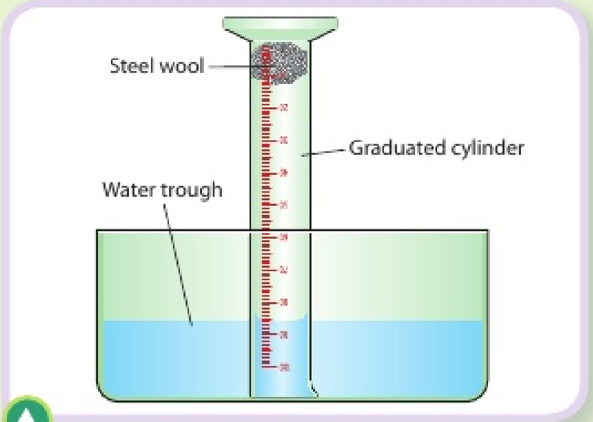


Figure 5.3 *To determine the percentage of O_2 in air*

Q U4

5.1 What do you predict will happen to the level of the water in test tube?

R2 R3 R4

Activity 5.2

Question

What is the percentage of oxygen in the air?

Equipment needed

Two retort stands
Silica glass tube
Copper turnings
Two gas syringes
Connectors

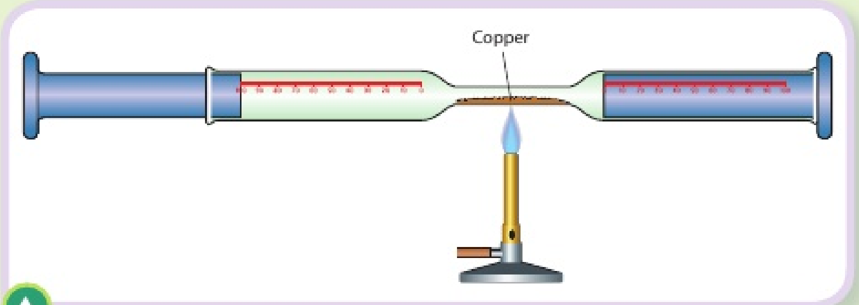


Figure 5.4 *To measure the percentage of O_2 in air*

CHEM Gr8 Ss Mod 5.indd 56 09/06/2020 14:17

Conducting the experiment

1. Set up the apparatus as shown in Figure 5.4.
2. Before you begin heating the copper metal turnings make sure that one syringe contains 100 cm^3 of air and the other syringe is completely closed.
3. Heat the copper turning and slowly pass the 100 cm^3 of air from one syringe to the other.
4. Pass the air from syringe to syringe several times or until the volume does not change.
5. Let the apparatus cool down and read the final volume of air left.

Q Understanding U4

5.2 What can you conclude about the amount of oxygen in air?

5.3 What happens to the copper turnings?

5.4 Why wait until the apparatus cools before taking the final readings on the syringe?

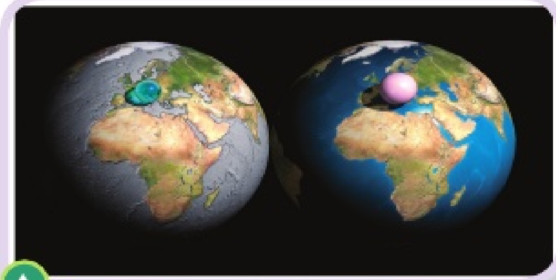


Figure 5.5 *Global water and air volumes. The blue sphere represents the volume of all water on Earth and the pink represents the atmosphere*

Oxygen

Oxygen can be prepared in the lab by the decomposition (break-down) of hydrogen peroxide. This is a slow process so a chemical called manganese dioxide is used to speed it up. Chemicals like these are called catalysts. Catalysts are usually written above the arrow in chemical equations.

Did you know?

Catalysts

Substances that alter the rate of a chemical reaction but are not used up themselves.

The word equation for the preparation of oxygen is:

$$\text{hydrogen peroxide} \xrightarrow{\text{manganese dioxide}} \text{oxygen} + \text{water}$$

The chemical equation is:

$$2H_2O_2 \xrightarrow{MnO_2} O_2 + 2H_2O$$

R2 R3 R4

Portfolio 34

Activity 5.3

Question

What are the properties of oxygen?

Equipment needed

Tap funnel
Conical flask
Tubing
Two-holed stopper
Basin
Beehive shelf
Gas jars and lids
Wooden splints
Combustion spoon
Hydrogen peroxide
Manganese dioxide
Water
Red and blue litmus papers
Magnesium ribbon
Charcoal
Limewater

Conducting the activity

1. Set up the apparatus as shown in Figure 5.6.
2. Slowly open the tap on the tap funnel and allow a few cm^3 of the hydrogen peroxide to fall into the black manganese dioxide powder.
3. Oxygen gas will be seen bubbling through the water in the gas jar. When the jar is full of oxygen (empty of water) remove it and place a lid on it. Replace the jar with another inverted jar full of water.

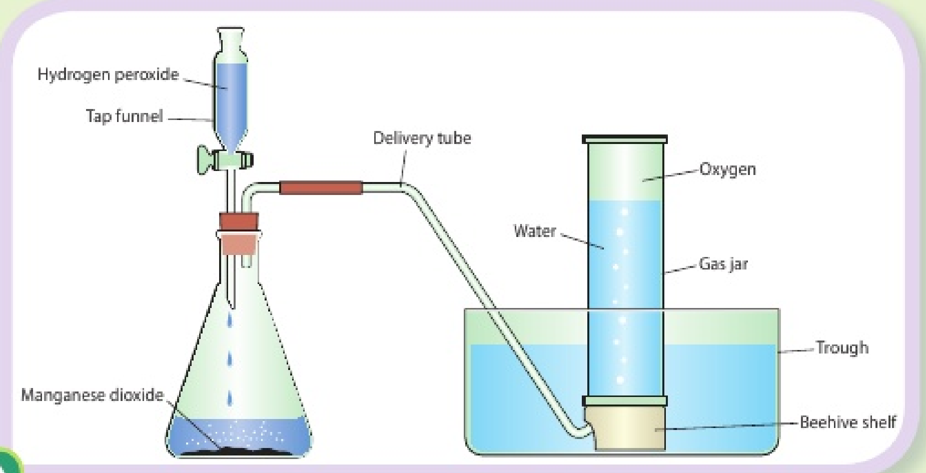


Figure 5.6 *To prepare oxygen gas*

4. Collect 4-5 jars of gas. (The first jar collected will be impure as it contains air from the conical flask and delivery tubing.)
5. Carry out the following tests.
6. Record your results in a table like the table in Table 5.1.

Table 5.1

	Observation 1	Observation 2
Jar 1		
Jar 2		
Jar 3		
Jar 4		

Jar 1

Note the colour and smell of oxygen. Test its pH using pH paper or moist red and blue litmus papers.

Jar 2

Place a glowing splint into a jar of oxygen.

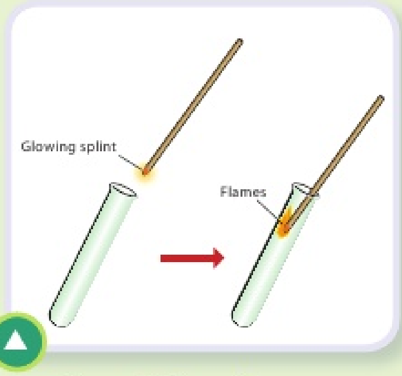


Figure 5.7 *To test for oxygen*

Jar 3

Heat a piece of charcoal (carbon) in a combustion spoon over a Bunsen burner until it glows red. Quickly transfer to a jar of oxygen.

- Test the products of the reaction using moist litmus papers.
- Add limewater to the gas jar.

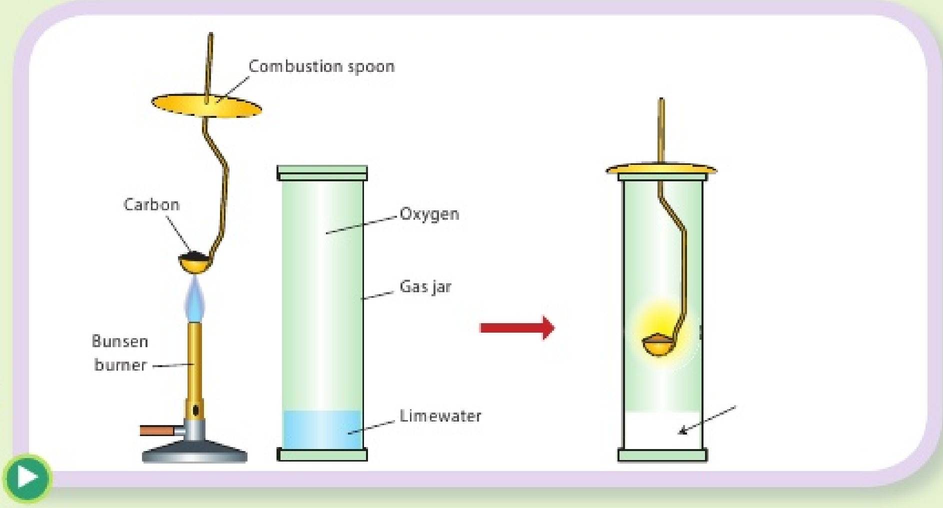


Figure 5.8 *Combustion of carbon*

Jar 4

Hold a piece of magnesium ribbon in a metal tongs over a Bunsen burner until it catches fire and burns. Then quickly transfer to a jar of oxygen. Add pieces of moist red and blue litmus papers to the jar when the flames have gone out.

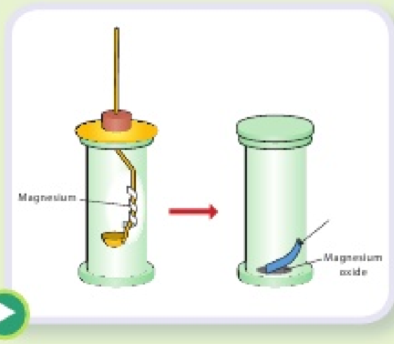


Figure 5.9 *Magnesium burning in a jar of oxygen produces a basic oxide*

Q U4 R4 R5

5.5 What conclusion about the properties of oxygen can you draw from each of your observations?

Ozone

The ozone layer in the upper atmosphere, which is where the greatest concentrations of ozone are found, acts as a sunscreen by absorbing ultraviolet radiation, which is damaging to plants and animals when it reaches Earth. Ultraviolet radiation causes sunburn and, ultimately, skin cancer in humans.

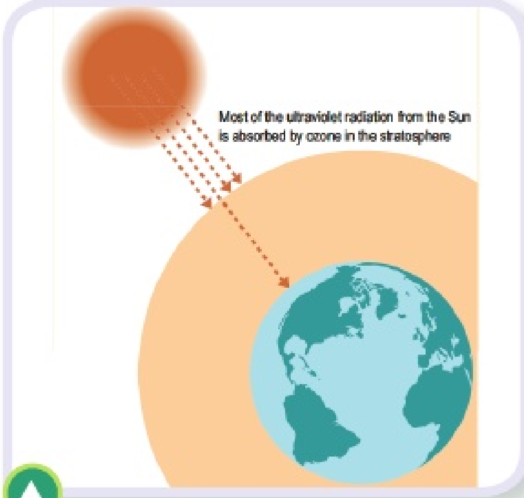


Figure 5.9 *The ozone layer absorbs most of the harmful ultraviolet radiation coming from the Sun*

Ozone, O_3 is a form of oxygen with three oxygen atoms in each molecule. It is formed in the stratosphere by the following photochemical reactions (reactions caused by light energy)

$$O_{2(g)} + \text{energy} \longrightarrow O\bullet_{(g)} + O\bullet_{(g)}$$
$$O_{2(g)} + O\bullet_{(g)} \longrightarrow O_{3(g)}$$

In the first of the two reactions, radiation from the sun splits oxygen molecules into high-energy oxygen atoms. In the second reaction, oxygen atoms, $O\bullet$, combine with oxygen molecules, O_2, to form ozone molecules, O_3.

Ozone is being created and destroyed all the time and a balance exists to maintain the ozone concentration. The balance is upset when certain substances get into the stratosphere [upper atmosphere].

CFCs

Because of their unreactive nature, CFCs, Chlorofluorocarbons, are extremely useful compounds.

Apart from their original use as refrigerants, they are also useful as air-conditioning gases, as aerosol propellant and as fire extinguishers.

However, they cause damage to the ozone in the ozone layer when they escape into the atmosphere. They react with ozone and lessen the amount of ozone in the stratosphere. This has led to many countries banning or limiting their uses.

Q Communicating C1 Communicating C2 Society S2

5.6 When, according to the UN environment statistics division, did Kazakhstan achieve zero use of ozone-depleting CFCs?

Hydrogen

Hydrogen is not one of the gases in the air, but is the most common element in the universe. It is the lightest element and has an atomic number of 1.

Hydrogen, H_2, is a colourless, odourless gas, which is almost insoluble in water.

Discovered in 1776, it is the least dense of all gases and was at one time used in balloons.

Hydrogen has considerable potential as a motor fuel. It burns more efficiently than petrol and is non-polluting. Water is the only product of the reaction, apart from trace quantities of nitrogen oxides.

Hydrogen is widely used as a fuel for propelling spacecraft, some of the water produced being used as drinking water for those manning the craft.

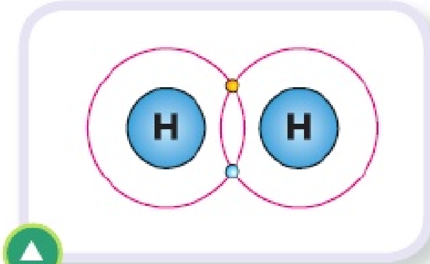


Figure 5.10 *Hydrogen molecule single bond*

Figure 5.11 *A stamp showing a satellite. The world's first satellite was launched from Baikonur in Kazakhstan on 4 October 1957.*

Test for hydrogen

To confirm the bubbles of gas in the experiment below are hydrogen you must place a lighted splint into the test tube. Hydrogen is easily ignited because it is flammable and you will hear a distinctive 'squeaky' pop sound if hydrogen is present.

Research R2 Research R3 Research R4

Activity 5.4

Question

What are the properties of hydrogen?

Equipment needed

Conical flask	Test tube	Dilute hydrochloric acid
Tap funnel	Taper	Zinc
Tubing	Stopper	Water
Trough	Beehive shelf	

Conducting the activity

1. Set up the apparatus as shown.
2. Fill the test tube with water and invert it on the beehive shelf.

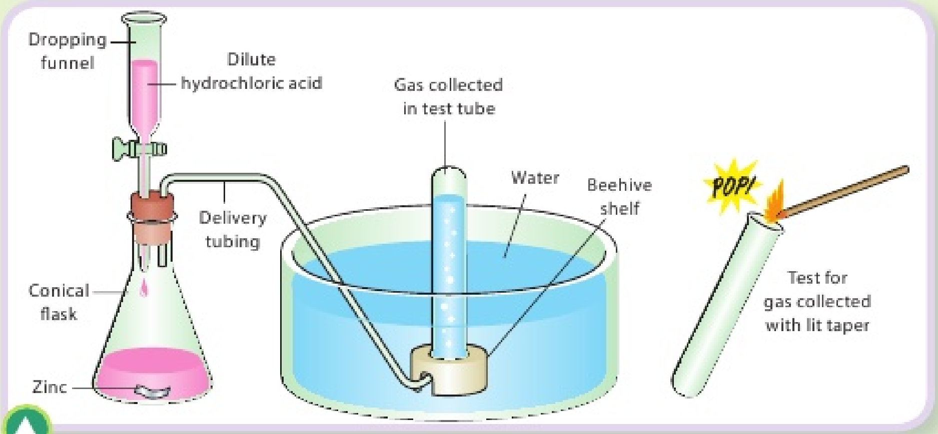


Figure 5.12

3. Add the hydrochloric acid to the zinc.
4. Stopper the test tube while under water when it is filled with hydrogen gas.
5. Remove the stopper and immediately bring a lit taper to the mouth of the test tube.

Q Research R5

5.7 From your observations which of the following are true. Mark ✓ / ✗.

Hydrogen is colourless.

Hydrogen is soluble.

Hydrogen is odourless.

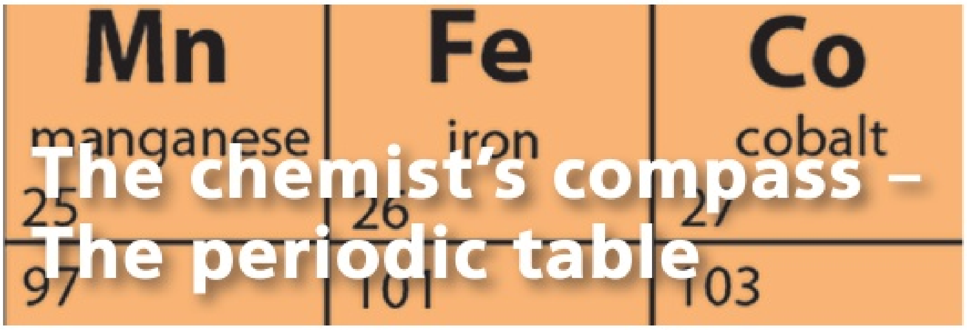
Hydrogen is flammable.

Hydrogen burns with a pop when lit.

Understanding U4 Society S1

5.8 Research how and when hydrogen was discovered. What similarities were there between its discovery and the test you have performed?

MODULE 6



The chemist's compass – The periodic table

Learning outcomes

At the end of this module you will be able to:

- Explain the meaning of atomic number, group and period. 8.2.1.1
- Recognise that elements of one group have the same number of electrons on their outer shell. 8.2.1.2
- Predict characteristic differences of properties of elements in groups and periods. 8.2.1.3 8.2.1.7
- Characterise an element by its place in the periodic system. 8.2.1.4
- Recognise that elements with similar chemical properties belong to one group. 8.2.1.5
- Name chemical families and give examples of elements in different groups. 8.2.1.6
- Explain the mechanism of the formation of covalent polar and non-polar bonds. 8.2.4.1
- Describe the mechanism of the formation of ionic bonds and predict properties of ionic compounds. 8.2.4.2
- Explain the relation between lattice type and the properties of a substance. 8.2.4.3

Keywords

✔ periodic table ✔ periods ✔ groups ✔ alkali ✔ halogens ✔ noble gases ✔ ions ✔ chemical formula

What is the periodic table?

The periodic table is considered to be one of the most important tools of the chemist. It shows a list of all the **elements**.

In the 1880s, scientists discovered a number of elements. They realised that some of these elements behaved in similar ways to others. A Russian chemist called Dmitri Mendeleev arranged them in order of the mass of the atoms of each element. He also lined up elements that behaved similarly.

Only about half of the elements had been discovered by this time. Mendeleev left gaps in his table for elements that had yet to be discovered.

Understanding U2 Research R3

6.1 Dmitri Mendeleev was passionate about chemistry. Research him and make notes on:

(a) His life

(b) His periodic table

(c) His other achievements.

The modern periodic table

The modern periodic table has no gaps in it. The elements are not arranged according to atomic mass but by **atomic number**. The table is arranged into:

- Vertical columns, called **groups**, where the elements with similar physical and chemical properties (behaviours) are grouped together
- Horizontal rows, called **periods**, which are in order of increasing proton number.

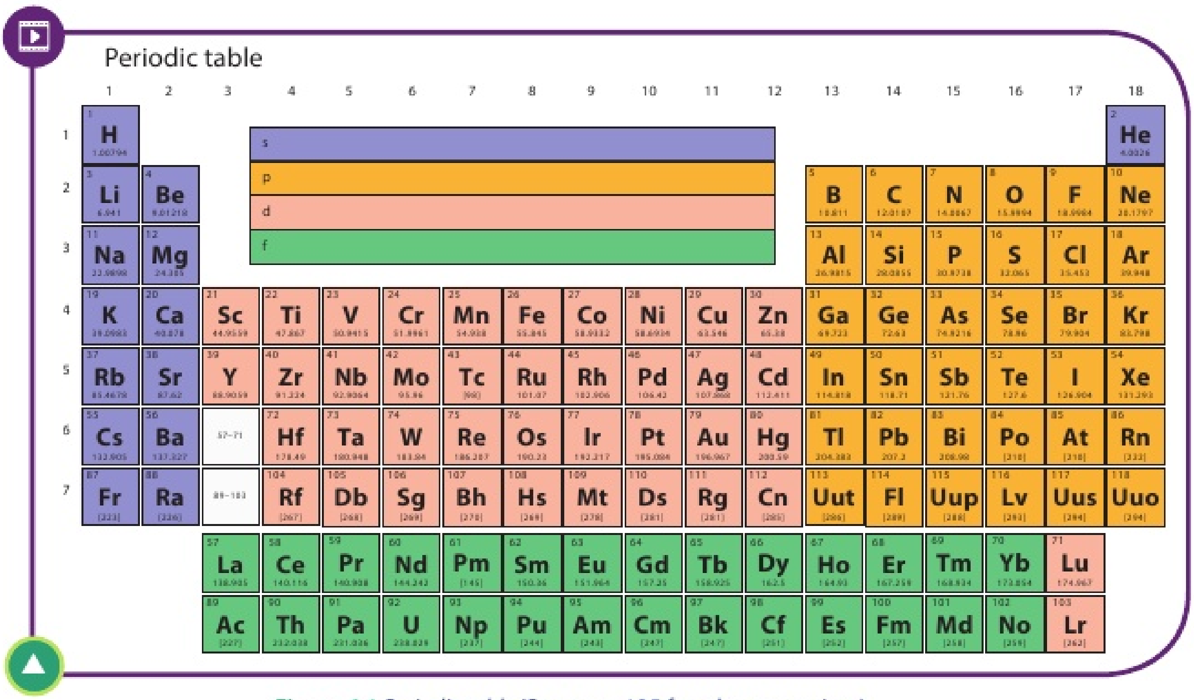


Figure 6.1 *Periodic table* (See page 125 for a larger version.)

Understanding U2

6.2 Complete the following sentence: Vertical columns of elements are called ______________ and horizontal rows of **elements** are known as ______________.

6.3 From the periodic table name:

(a) Three elements from group 2.

(b) Three elements from period 2.

(c) The element in group 1 period 3.

Scientist Biography

Marie Curie was a scientist who discovered two new elements: radium and polonium. Watch a video animation to find out more about Marie Curie and her investigations.

Q U2 Understanding

6.4 Refer to Figures 6.1 and 6.2. Complete this table to show that you understand how the periodic table is arranged.

Element	Symbol	Period	Group	Metal	Non-metal
Carbon					
	H				
		2	3		
Barium					
Chlorine					
	He				

Group names

The periodic table has four named groups. They are:

- **Group 1** – the alkali metals
- **Group 2** – the alkaline earth metals
- **Group 7** – the halogens
- **Group 8 (group 0)** – the noble gases.

We already looked at group 1 in Module 3. We will look at the other groups in this module.

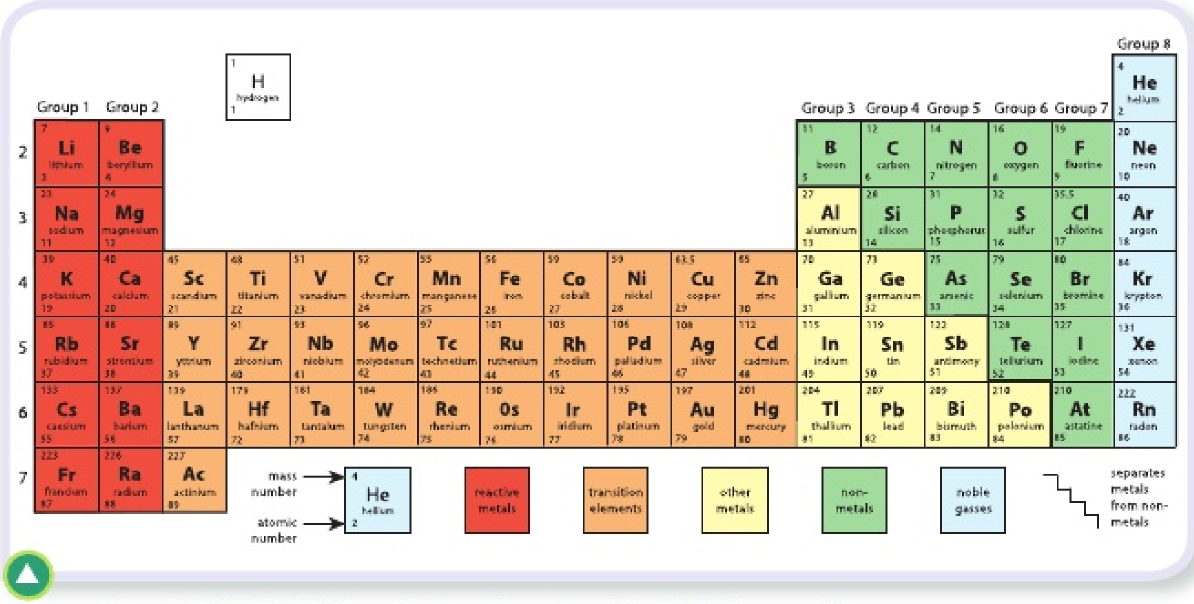


Figure 6.2 *Periodic table – each colour shows elements with similar properties.*

Group 2 – the alkaline earth metals

Elements in this group include:

- Beryllium (Be)
- Magnesium (Mg)
- Calcium (Ca).

Figure 6.3 *Alkaline earth metals – beryllium, magnesium, calcium, strontium and barium*

All elements in this group have two electrons in their outer shell. Group 2 elements are not as reactive as group 1 elements.

Group 7 – the halogens

The elements in group 7 have seven electrons in their atom's outer shell. (The word halogen also contains seven letters!) Elements in this group include:

- Fluorine, in the form of compounds such as sodium fluoride, which is sometimes added to water to strengthen enamel in our teeth.
- Chlorine, which is sometimes added to water to kill bacteria.

Figure 6.4 *The halogens (chlorine, bromine and iodine)*

Elements in this group are trying to gain an electron because they want a full outer shell.

Group 8 (group 0) – the noble gases

The elements in group 8 are all gases and have full outer shells of electrons in their atoms. As a result, the noble gases are unreactive and thus very stable. Members of the noble gases include:

- Helium (He)
- Neon (Ne)
- Argon (Ar), which is used in filament lamps (light bulbs).

Figure 6.5 *Helium, one of the noble gases, is sometimes used in balloons*

6.5 Complete the following sentence: The elements of group 7 are called the ________________. They get ________________ reactive as you go down the group; therefore ________________ is the most reactive.

Understanding U2 Communicating C1 Society S2

6.6 Find out how these names were assigned to each of these groups:

(a) The alkali metals

(b) The alkaline earth metals

(c) The halogens

(d) The noble gases.

6.7 Select one element from each group and find out where it may be used.

6.8 How many naturally occurring elements are there?

6.9 Look at **Figure 6.6** and answer these questions using the letters A–E *and* the elements' names.

- **(a)** Which two elements are in the same group?
- **(b)** Which two elements are in the same period?
- **(c)** Name the group to which elements A and B belong.

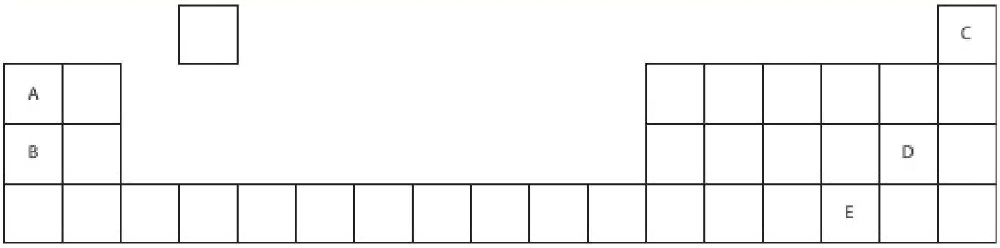


Figure 6.6

6.10 From the periodic table name:

- **(a)** Three elements from group 3.
- **(b)** Three elements from period 3.
- **(c)** The element in group 2, period 3.

How are chemical bonds formed?

A compound forms when two or more elements chemically combine. The question is: why do these elements react and combine with each other? They react to have a full outer shell of electrons. There are two ways for atoms to achieve this:

- **Ionic bonding** This occurs when one atom *loses* and another atom *gains* electrons.
- **Covalent bonding** This occurs when atoms *share* electrons.

Electron shell rules

- Electrons are found in shells or orbits.
- The shells nearest the nucleus are filled first.
- The first shell can hold up to two electrons.
- The second shell can hold up to eight electrons. Atoms of elements occupying the A groups can also hold up to eight electrons in the outer shell.
- Atoms that have a full outer shell of electrons are stable and do not react.
- In most atoms the outer shell is not full and this makes them want to react.

What are ions?

All atoms would like full outer shells and they will either lose or gain electrons in order to achieve this. When atoms lose or gain electrons they become charged particles and are called ions. The elements that readily form ions make up groups 1, 2, 6 and 7:

- **Group 1** – alkali metals. These have only one electron in the outer shell, which they lose to form a positive ion (+1), e.g. Na^{+1}.
- **Group 2** – alkaline earth metals. These have two electrons in their outer shell so will lose two electrons to form a positive ion (+2) e.g. Mg^{+2}.
- **Group 6** – non-metals. These have six electrons in their outer shell and so will gain two electrons for a full outer shell to form a negative ion (–2) e.g. O^{-2}.
- **Group 7** – halogens. These have seven electrons in their outer shell and so will gain one electron to form a negative ion (–1) e.g. Cl^{-1}.

Ionic bonding does not result in the formation of molecules. The oppositely charged ions attract each other to form a rigid three-dimensional lattice. Each ion in the lattice is surrounded by others of opposite charge.

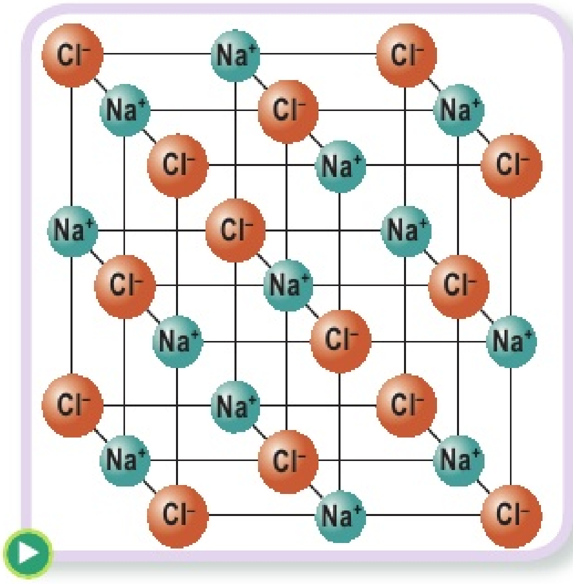


Figure 6.7

The structure of crystal lattices is determined by an X-ray technique as in Figure 6.7. This shows the structure of a small part of the sodium chloride crystal. Many millions of sodium and chloride ions are arranged in this way in a single crystal of sodium to make up a giant crystal structure.

Figure 6.7 shows that in the model of NaCl crystal each sodium ion in the lattice is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions. However not all ionic substances have the same lattice structure as sodium chloride.

Mg^{2+} and O^{2-} ions have a greater number of charges which means they form stronger ionic bonds than the Na^+ and Cl^- ions in sodium chloride.

In lattice structure models for magnesium oxide, the ions are typically represented closer together, indicating that more energy is required to break the bonds. Magnesium oxide has a much higher melting point, for example, than sodium chloride.

Properties of ionic compounds

- Strong forces of attraction between the ions means a lot of energy required to break bonds
- High melting points
- Do not conduct electricity in solid state
- Conduct electricity in molten state or when dissolved in water.

Understanding U1 Communicating C1

6.11 Find out why magnesium oxide is used in electrical insulation.

Forming compounds – chemical formulas

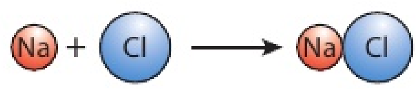
You can use the periodic table to predict the ratio of atoms in compounds.

Combine metal and non-metal

When you combine a metal and a non-metal the name of the metal comes first followed by the name of the non-metal, and the ending is changed to **ide**.

sodium	+	chlorine	=	sodium chloride
Metal		*Non-metal*		*Compound*

Sodium is in group 1 of the periodic table so it needs to lose one electron. Chlorine is in group 7 of the periodic table so it needs to gain one electron. So the ratio of atoms in the compound sodium chloride is 1:1. One sodium atom bonds with one chlorine atom.



Understanding U1

6.12 What are the names of the compounds formed between:

(a) Calcium and chlorine?

(b) Lithium and bromine?

(c) Magnesium and oxygen?

Combine non-metal and non-metal

Again you can use the periodic table to predict the ratio of atoms in compounds. You need to work out the valency of each atom. This is the number of bonds each atom can form. For example:

- Hydrogen is in group 1 of the periodic table, which shows it has only one electron and needs one more electron in order to be stable so that it can form one bond; so its combining power = 1.
- Carbon is found in group 4 of the periodic table, which shows it has four electrons in its outer shell and it requires four more electrons in order to be stable so it can form four bonds; so its combining power = 4.

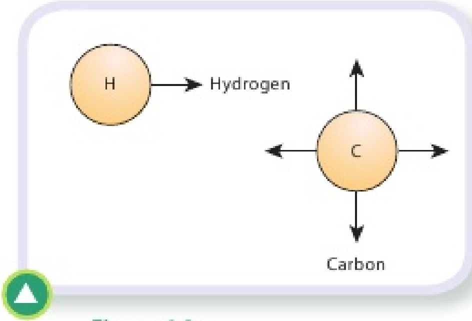


Figure 6.8

Carbon has combining power = 4 (4 bonds) whereas hydrogen has combining power = 1 (1 bond), so you will need four hydrogen atoms to bond with one carbon. This is methane. Natural gas is mainly methane.

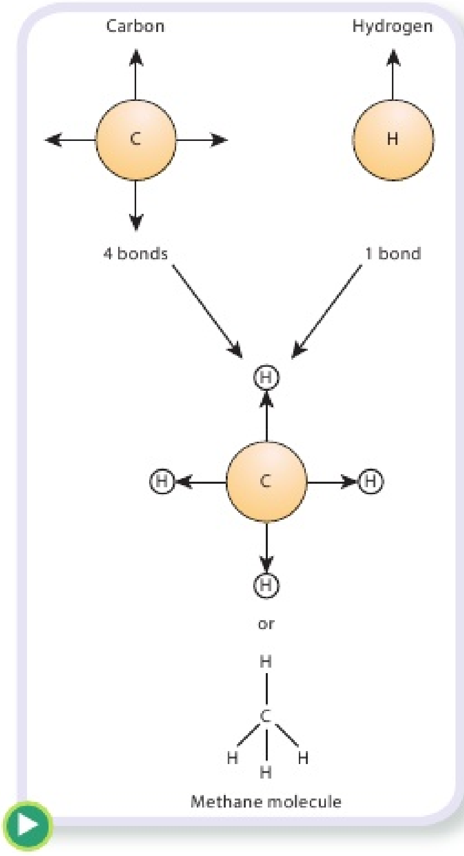
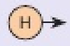
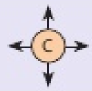


Figure 6.9 *Chemical formula for a compound containing carbon and hydrogen*

Note: Atoms do not always obey the **octet** rule (eight electrons in the outer shell). For example, hydrogen, beryllium and boron have too few electrons in the outer shell to form an octet.

6.13 In your group, copy and complete the following table using the periodic table.

Element	Symbol	Group number	Electrons needed	Combining power	Show bonds
Hydrogen	H	1	1	1	
Carbon	C	4	4	4	
Nitrogen					
Oxygen					
Sulphur					
Chlorine					

Did you know?

Carbon is a very interesting element:

- It can form about 10 million compounds with almost all other elements in the periodic table.
- The human body's mass is 18.5% carbon.

6.14 Using the periodic table work out the formula and the name of the compound formed between:

(a) Hydrogen and chlorine

(b) Nitrogen and hydrogen

(c) Carbon and oxygen.

6.15 Draw diagrams and show the bonding in a water molecule.

How to combine a metal atom, a non-metal atom and oxygen

The name of the metal comes first. The name of the non-metal comes next with the ending **ate**. This occurs only when an oxygen atom is present. For example:

potassium + phosphorus + oxygen → potassium phosphate

sodium + carbon + oxygen → sodium carbonate

Q U1 Understanding

6.16 Write the name of the compounds formed between the following:

- **(a)** Potassium + nitrogen + oxygen
- **(b)** Magnesium + carbon + oxygen
- **(c)** Calcium + sulphur + oxygen
- **(d)** Lithium + phosphorous + oxygen.

6.17 What are the names of the different types of atoms in the following compounds:

- **(a)** Calcium nitrate?
- **(b)** Potassium carbonate?
- **(c)** Sodium bromide?
- **(d)** Sodium carbonate?

6.18 Refer to the periodic table and work out the formula for the following compounds:

- **(a)** Sodium chloride
- **(b)** Magnesium chloride
- **(c)** Sodium oxide
- **(d)** Magnesium oxide
- **(e)** Potassium oxide.

Hydrogen molecule (H_2)

Each hydrogen atom has one electron in its shell, so each atom has a combining power = 1, so two hydrogen atoms will overlap their shells and bond forming a hydrogen molecule (H_2).

Figure 6.10

Oxygen molecule (O_2)

Each oxygen atom has six electrons in the outer shell so it needs to gain two electrons to fill the outer shell. So each oxygen atom has a combining power = 2; in an oxygen molecule two oxygen atoms are bonded together and there is a **double bond**.

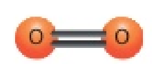


Figure 6.11

Q U1 Understanding

6.19 Refer to the periodic table to work out what a nitrogen molecule looks like.

MODULE 7

Particles in solution

Learning outcomes

At the end of this module you will be able to:

- Classify substances according to their level of solubility in water 8.3.4.1
- Describe solutions and their importance in nature and everyday life 8.3.4.2
- Explain impact of temperature on the solubility of a substance 8.3.4.3
- Calculate the solubility of a substance in 100 g of water using evaporation and compare results with reference data 8.3.4.4
- Perform percentage mass concentration calculations 8.3.4.5
- Calculate molarity of substances in solution 8.3.4.6

Keywords

✔ solute ✔ solvent ✔ solution ✔ soluble ✔ insoluble ✔ solubility ✔ crystallisation ✔ concentration ✔ molarity ✔ dissolve

What are solutions?

When we think about solutions we usually think of them as a solid dissolved in a liquid, e.g. seawater contains salt dissolved in water. However, solutions can be very different to this. For example, oxygen dissolves in water to form a solution that allows fish to breathe.

Dissolving

When a substance dissolves in a liquid its particles slip into the gaps between the particles of the liquid and they become completely mixed up together.

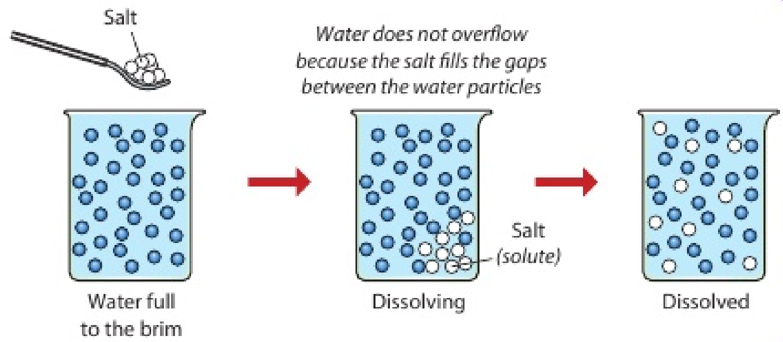


Figure 7.1 *Salt dissolving in water*

Did you know?

If you pour a handful of salt into a full glass of water, the water level will actually go down rather than overflowing the glass. Why do you think this is?

Q U3 C1

7.1 Look at Figure 7.2. There are certain words we need to know. Explain each of the following words:

(a) Solvent
(b) Solute
(c) Solution.

Figure 7.2 *Coffee dissolving in water*

Q

7.2 Write out and complete the following:

Most solutions are made up of a substance called the _______________, which dissolves in the liquid called the _______________. Substances that dissolve are said to be _______________, while those that don't are _______________ . A cup of coffee is a _______________, the hot water is the _______________ and the coffee granules are the _______________ .

Concentrated and Dilute solutions

Have you ever tried tasting fruit squash without adding water first? It tastes unbearably sweet. These types of drinks are **concentrated solutions**. This means there is a lot of solute dissolved in the solvent.

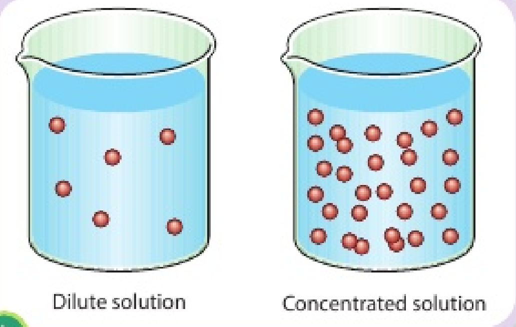


Figure 7.3 *Dilute and concentrated solutions*

Q

7.3 What is the solute in the concentrated solution of squash?

By adding water to the concentrated solution of squash, you are making the solution more dilute and nicer to taste.

If you add too much water, the drink may taste too weak (dilute).

7.4 Briefly explain the difference between a dilute solution and a concentrated solution.

7.5 If you make your glass of squash too dilute, what could you do to make it more concentrated again?

Did you know?

There is about ½ lb or 250 g of salt (NaCl) in the average human body.

7.6 Give one everyday example of each of the following:

- **(a)** A gas dissolved in a liquid
- **(b)** A liquid dissolved in a liquid
- **(c)** A solid dissolved in a liquid
- **(d)** A gas dissolved in a solid.

7.7 Name a solvent other than water and a substance that dissolves in it.

Did you know?

One bucketful of water contains more atoms than there are bucketfuls of water in the ocean.

7.8 Divide into groups. Within your group outline how you could investigate if there is a change in the overall mass when a solid dissolves in a solvent.

Soluble/Insoluble

Solids that dissolve in a liquid are said to be **soluble** and solids that do not dissolve are called **insoluble**.

7.9 Make a list of five solids that are soluble in water and five solids that are insoluble in water.

7.10 In a salt solution you cannot tell if there is any solute present. Draw a particle diagram for a salt solution and for pure water.

7.11 Find out:

- **(a)** What name is given to two liquids that mix, e.g. alcohol and water?
- **(b)** What name is given to two liquids that do not mix, e.g. oil and water?

Solubility

How many spoons of sugar will dissolve in a cup of tea? The amount depends on two factors:

1 The volume of tea (solvent)

2 The temperature of the tea (solvent)

The solubility of a substance is the number of grams of the solute that dissolves in 100 g of the solvent (usually a liquid) at a particular temperature.

The solubility of solutes usually increases with temperature. A notable exception is oxygen dissolved in water. As water is heated, less and less oxygen can be dissolved in it.

Global warming is causing the Earth's waters to heat up. What effect do you think this will have on fish life?

Research R2 Research R3 Research R4 Portfolio 45

Activity 7.1

Question

What is the solubility of a variety of substances in water?

Equipment needed

Test tubes
Test-tube rack
Spatula
Water

Variety of chemicals, e.g. iodine, potassium permanganate, sodium chloride, sulfur, wax

Conducting the activity

1. Half fill a test tube with water.
2. Add a spatula half full of the substance to be tested to the water.
3. Stopper the test tube and shake for 15–20 seconds.
4. Examine the contents of the test tube and decide whether the substance being tested is soluble or insoluble. How will you know?
5. Rinse out the test tube and repeat the experiment with the next substance to be tested.
6. Try as many substances as are available and record your results as shown in the table.

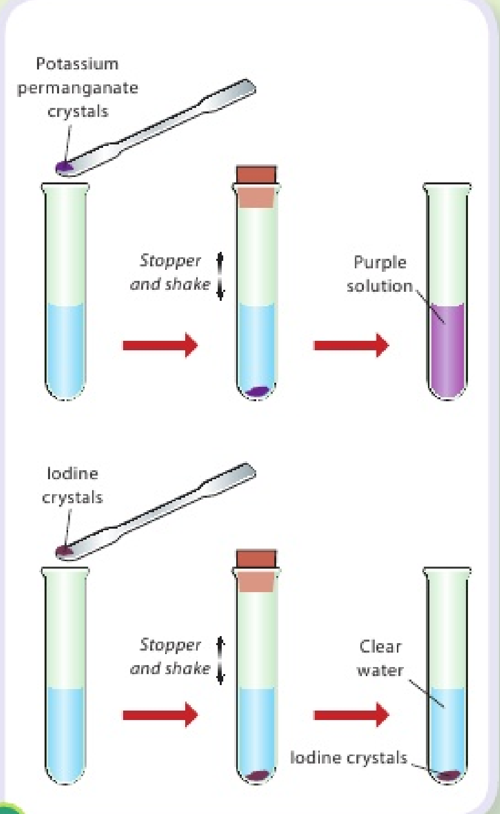


Figure 7.4 *Solubility testing*

Substances being tested	Soluble in water
Sodium chloride (salt)	Yes
Sulfur	
Wax	
Sand	
Copper sulphate	

Understanding U3 Research R5

7.12 How from the experiment could you determine if the substances were soluble in water?

Research R2 Research R3 Research R5 Portfolio 47

Activity 7.2

Question

Does the solubility of water change with temperature?

Equipment needed

Beaker	Stirring rod	Copper sulphate crystals
Spatula	Pestle and mortar	
Thermometer	Water	

Safety

- Use tongs to remove the beaker from the hot plate.

Conducting the activity

1. Place approximately 100 cm³ of water in a beaker and note its temperature.
2. Using a pestle and mortar, grind up a sample of copper sulphate crystals.
3. Slowly and with constant stirring add the powdered copper sulphate to the water. Continue until the powder will no longer dissolve but instead settles to the bottom.
4. Heat the beaker to 50°C and note what happens to the undissolved copper sulphate.
5. Add more copper sulphate until no more will dissolve.
6. Heat the beaker to near boiling (80–90°C) and again try to dissolve more solute. What do you notice about how temperature affects solubility?
7. Allow the solution to cool. This can be quickened by holding the beaker under running tap water. What do you notice?

Figure 7.6 *A copper sulphate solution*

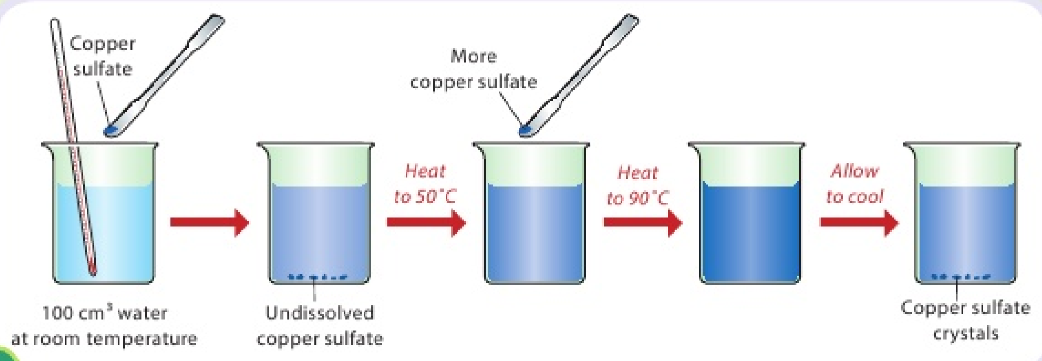


Figure 7.5 *Investigating the effect of temperature on solubility*

Communicating C1 Society S2

7.13 Research why water is such a good solvent and why certain substances will dissolve in water and why other substances will not.

Solubility curves

A solubility curve plots the mass of solute dissolved in a saturated solution at different temperatures. The graph shows how the solubility of copper sulphate (sulphate) increases as the temperature of the solvent increases.

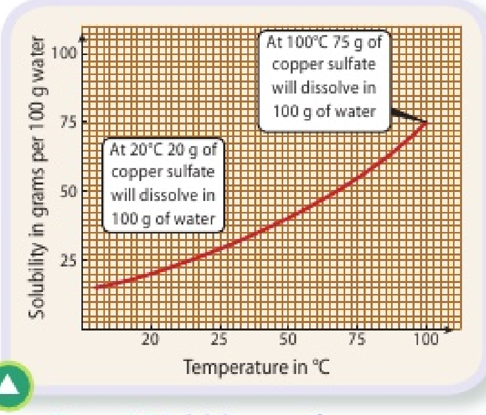


Figure 7.7 *Solubility curve for copper sulphate (solute) in water (solvent)*

Crystallisation

Crystals are tiny particles built in a regular manner, e.g. salt, diamonds and snowflakes. When a hot, saturated solution of copper sulphate is cooled down, some of the copper sulphate will come out of solution and form crystals. This is because a cooler solution will not dissolve as much solute as a warmer one.

Let's refer back to our solubility curve. If the 100°C solution of copper sulphate containing 75 g of dissolved solute is cooled to 20°C, only 20 g of copper sulphate will remain dissolved. The rest comes out of solution and forms crystals. We should end up with 55 g (75 g–20 g) of copper sulphate crystals.

Figure 7.8 *A variety of salt crystals*

Research R2 Research R3 Portfolio 48

Activity 7.3

Question

How can we grow crystals using copper sulphate solution?

Equipment needed

- Hot plate (or Bunsen burner)
- Thermometer
- Beaker
- Stirring rod
- Evaporating dish
- Pestle and mortar
- Spatula
- Copper sulphate or aluminium sulphate
- Water

Safety

- Use tongs for removing the beaker from the hot plate.

Conducting the activity

1. Using the pestle and mortar, grind up a sample of copper sulphate (or alum) crystals.
2. Add 100 cm^3 of water to a beaker and heat on the hot plate.
3. Slowly and with constant stirring, add the powdered copper sulphate to the water.
4. Heat the water to 70°C and add in more copper sulphate until no more will dissolve.
5. Pour approximately half of the solution into a pre-warmed evaporating dish and allow to cool slowly.
6. The remaining solution in the beaker can be quickly cooled by standing it in iced water or by allowing running water from the tap to pour over the outside of the beaker.

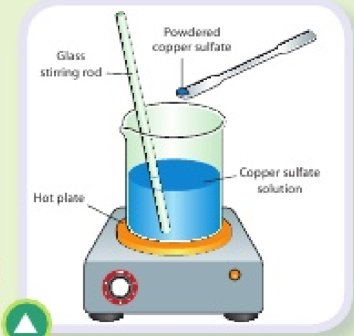


Figure 7.9

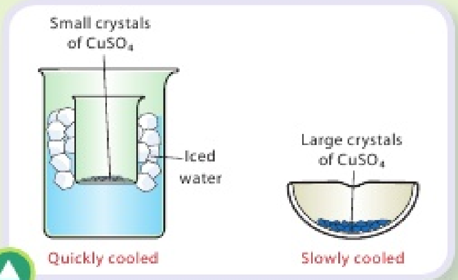


Figure 7.10 *Crystallisation*

Figure 7.11 *Sugar crystals*

Q R4 R5

7.14 What can you conclude about cooling saturated solutions?

7.15 What was the effect of using different cooling methods on the formation of the crystals?

Q Understanding U3 Research R1

7.16 **Investigation:** In your group, read the following and determine if you can solve the problem using separation methods.

Brand A crisps states that it has less salt than Brand B crisps. Design and carry out an investigation to test and analyse the crisps to see if this is true.

Understanding U3 Communicating C1 Communicating C2

7.17 Research what foods contain large amounts of salt.

7.18 What problems can arise from excess salt in your diet?

7.19 Research the solutes in a fizzy soda drink.

Q Understanding U3 Communicating C1 Society S2

7.20 Water is the most important solvent as it dissolves very many solutes. Make a list of all the different jobs that water is used for around the home. Note how it is being used as a solvent in each job.

Were you surprised at how much we use water?

Calculating percentage concentration by mass

We know from Module 2 and the law of the conservation of mass that

$$\text{mass of solution} = \text{mass of solute} + \text{mass of solvent}$$

A solution's percent concentration by mass, % m/m, tells us how many *grams of solute* we get for every 100 g of solution.

The calculation for working out the percentage concentration of a solute in a solution is relatively simple and can be worked out using the formula

$$\text{mass percent} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

Example problem 1

What is the percent by mass of 275 g of a solute in a 500 g solution?

Answer:

275 g ÷ 500 g × 100 % = 55.0 % (m/m)

Example problem 2

The mass of a solute is 22.4 g and its mass percentage of a solution is 14%. Calculate the mass of the solvent and the total mass of the solution.

Answer:

Total mass of solution = 22.4 x 100 ÷ 14 = 160 g

Total mass of solvent = 160 - 22.4 = 137.6 g

Q R4 U3

7.21 A solution is made by dissolving 125 g of sodium chloride in 1.5 kg of water. What is the percent by mass?

7.22 A solution is made by dissolving 15.0 g of Na_2CO_3 in 235 g water. What is the mass percent (m/m) of Na_2CO_3 in the solution?

7.23 How many grams of sodium chloride are needed to prepare 225 g of a 10.0% (m/m) sodium chloride solution?

7.24 The mass of a solute is 3.7 g and its mass percentage of a solution is 9%. Calculate the mass of the solvent and the total mass of the solution.

Calculating molarity

Molarity is the term used for the measurement:

number of moles of a *solute* per litre of solution

Molarity is a unit of concentration measuring the number of moles of a solute per litre of solution. The strategy for solving molarity problems is quite simple.

The key to calculating molarity is to remember the units of molarity: moles per litre. Find the number of **moles** of the solute dissolved in **litres** of solution.

Example problem 3

Calculate the molarity of a solution prepared by dissolving 23.7 grams of $KMnO_4$ into enough water to make 750 mL of solution.

Answer:

This example has neither moles or litres needed to find molarity. Find the number of moles of the solute first.

To convert grams to moles, the **molar mass** of the solute is needed. From the **periodic table**:

Molar mass of K = 39.1 g

Molar mass of Mn = 54.9 g

Molar mass of O = 16.0 g

Molar mass of $KMnO_4$ = 39.1 g + 54.9 g + (16.0 g x 4)

Molar mass of $KMnO_4$ = 158.0 g

Use this number to convert grams to moles:

moles of $KMnO_4$ = 23.7 g $KMnO_4$ x (1 mol $KMnO_4$/158 grams $KMnO_4$)
moles of $KMnO_4$ = 0.15 moles $KMnO_4$

Now the litres of solution is needed. **Remember** this is the total volume of the solution, not the volume of solvent used to dissolve the solute. This example is prepared with 'enough water' to make 750 mL of solution.

Convert 750 mL to litres.

Litres of solution = mL of solution x (1 L/1000 mL)
Litres of solution = 750 mL x (1 L/1000 mL)
Litres of solution = 0.75 L

This is enough to calculate the molarity.

Molarity = moles solute/Litre solution
Molarity = 0.15 moles of $KMnO_4$/0.75 L of solution
Molarity = 0.20 M

The molarity of this solution is 0.20 M.

The calculation is quite simple as long as you remember that you are calculating how many moles of the solute are dissolved in the litres of the solution.

Example problem 4

What is the molarity of a 0.350 L solution that contains 46 g of potassium chloride?

Answer:

1 mole of potassium chloride = 74.6 g
46 g of potassium chloride = 0.62 mol
molarity (M) = 0.62 mol KCl ÷ 0.350 L = 1.8 (M) KCl

Q Research R4 Research R5

7.25 What is the molar concentration of 1.0 mol of KCl dissolved in 750.0 mL of solution?

7.26 How many litres of a 2.00 M sodium chloride solution are needed to provide 67.3 g of NaCl?

7.27 How many moles of NaCl are contained in 0.500L of a 1.5M solution.

Quick review how to calculate molarity

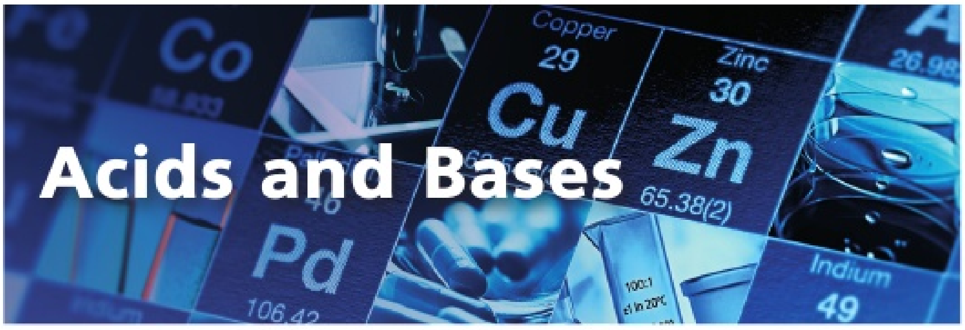
To calculate molarity

- Find the number of moles of solute dissolved in solution.
- Find the volume of solution in litres.
- Divide moles solute by litres solution.

Make certain to use the correct number of significant figures when reporting your answer.

MODULE 8

Acids and Bases



Learning outcomes

At the end of this module you will be able to:

- Understand the classification and properties of acids and make word and chemical equations for their reactions 8.3.4.8
- Understand the classification and properties of bases and make word and chemical equations for their reactions 8.3.4.9
- Explain and apply different methods for the production of salts and write word and chemical equations for their reactions 8.3.4.10
- Understand the classification and properties of salts and write word and chemical equations for their reactions 8.3.4.11
- Explore the genetic relation between principal classes of non-organic compounds 8.3.4.12

Keywords

✔ acid ✔ base ✔ alkali ✔ neutral ✔ litmus paper ✔ universal indicator ✔ pH scale ✔ titration ✔ neutralisation ✔ particle theory ✔ dissociate ✔ salts

What are acids?

What springs to mind when you hear the word 'acid'? Most people think of acids as dangerous, corrosive liquids. Not all acids are like this, however. For example:

- Lemon juice contains an acid called citric acid
- Vinegar is a dilute solution of ethanoic acid.

The word 'acid' comes from a Latin word meaning **sour**. The acids in food such as those listed above tend to have a sour, sharp taste. The acids you use in the laboratory are strong acids that are corrosive. So, be careful when using them! Table 8.1 shows some everyday acids and some laboratory acids.

A substance that is an acid is said to be **acidic**.

Table 8.1 Some everyday acids and some laboratory acids

Everyday acids	Laboratory acids
Lemon juice (citric acid)	Hydrochloric acid (HCl)
Rain water (carbonic acid)	Sulfuric acid (H_2SO_4)
Vinegar	

Communicating C1 Society S2

8.1 Name and give the formulas of two acids and two bases used in the lab.

8.2 What are alkalis?

8.3 We come across acids and bases every day and maybe don't realise it. The food industry uses acids in many products, such as fizzy drinks, and bases are used in many cleaning products.
Examine the labels on different food and cleaning products and see if you can identify any acids or bases.

How can you tell an acid from a base?

You cannot tell by looking at a substance whether it is an acid or a base or a neutral substance. You have to test them with an indicator.

What are indicators?

An indicator is a chemical that shows, by means of a colour change, whether a substance is an acid or a base. In the laboratory we use:

- Litmus indicator
- Phenolphthalein
- Universal indicator
- Methyl organge.

What is litmus indicator?

Litmus is paper that has been treated with a water-soluble mixture of different dyes from lichens, which grow on the bark of trees or on rocks.

Litmus indicator will change colour from blue to red in acids and from red to blue in bases (or alkalis).

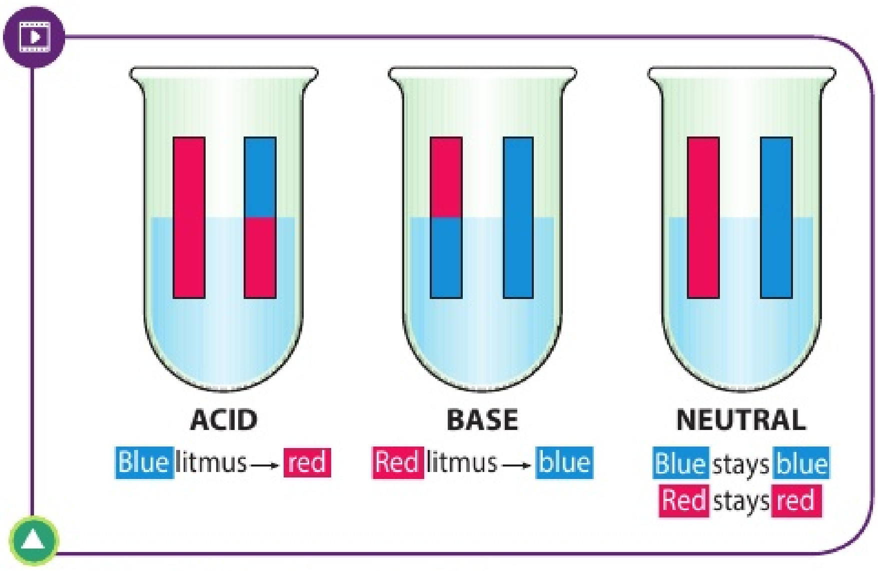


Figure 8.3 *Colour changes of litmus indicator*

Remember: *b* for base, *b* for *b*lue litmus.

R2 R3 R4 Portfolio 55

Activity 8.1

Question

What household substances are acids, bases or neutral?

Equipment needed

Test tube rack — Litmus paper

Test tubes — Variety of household substances, laboratory chemicals and water

Conducting the activity

1. Place a selection of household and laboratory chemicals in different test tubes. If the substance is a solid, try dissolving a spatulaful of it in a small volume of water first.
2. Dip pieces of blue and red litmus papers into the liquid of each test tube.
3. Examine the papers and observe any colour changes.
4. Enter your results into a table like the one shown below.

Substances being tested	Acidic	Basic	Neutral
Vinegar	✓		
Water			✓

Q R5

8.4 What effect do acids have on:

(a) Blue litmus paper? (b) Red litmus paper?

What is universal indicator?

Universal indicator is a mixture of dyes that change to different colours according to how strong the acid or base is. Universal indicator gives a range of colours, as shown in Figure 8.4, which can be used to give a value on the pH scale, which we look at next.

What is the pH scale?

Some acids are safe to handle and even eat, whereas others such as car battery acid are very dangerous and corrosive. In order to compare the strengths of acids and bases chemists devised a scale called the pH scale. The scale is from 0 to 14:

- 0–7 is an acid (the lower the number, the more acidic it is)
- 7 is neutral (pure water has a pH at almost 7)
- 7–14 is an alkali (the higher the number, the greater the alkalinity).

The pH can be measured using a pH meter or universal indicator.

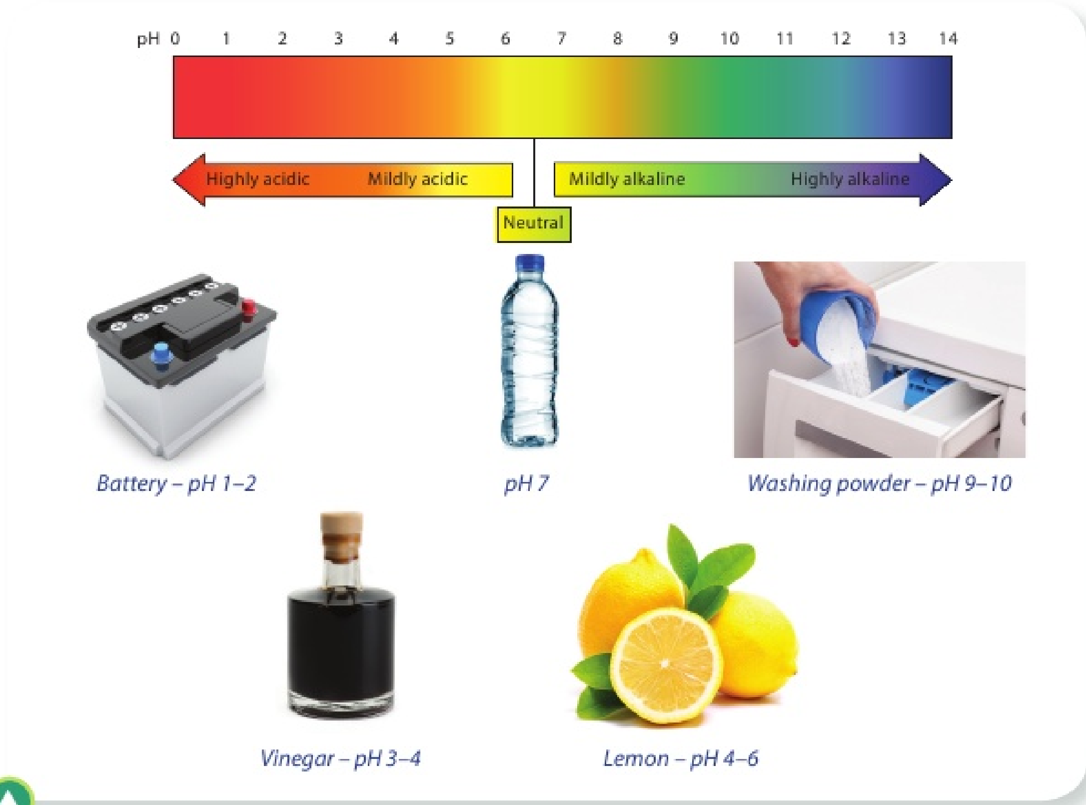


Figure 8.4 *A variety of household substances and their position on the pH scale*

8.5 What advantage does universal indicator have over litmus paper?

Figure 8.5 *A pH meter*

Research R2 Research R3 Research R4

Portfolio 55

Activity 8.2

Question

What is the pH of a variety of household substances?

Equipment needed

Test tube rack
Test tubes
Universal indicator or pH meter
Variety of household substances

Conducting the activity

1. Place the substance being tested into a test tube. If the substance is solid, dissolve it in water first.
2. Place a piece of universal indicator paper into the liquid in the test tube.
3. Note the colour of the indicator and, using a colour chart, find its pH.
4. Draw a pH scale in your report by drawing a number line from 0 to 14.
5. On the pH scale, write the names of the substances tested and indicate their pHs as shown in Figure 8.7.

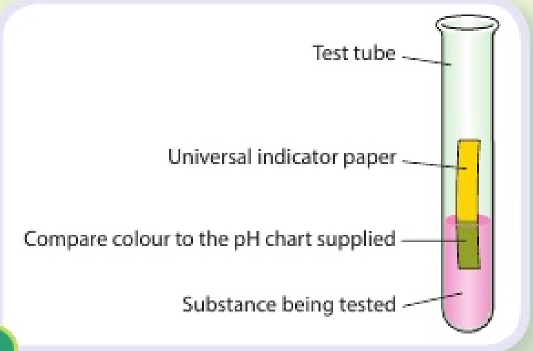


Figure 8.6 *Investigating the pH of a substance*

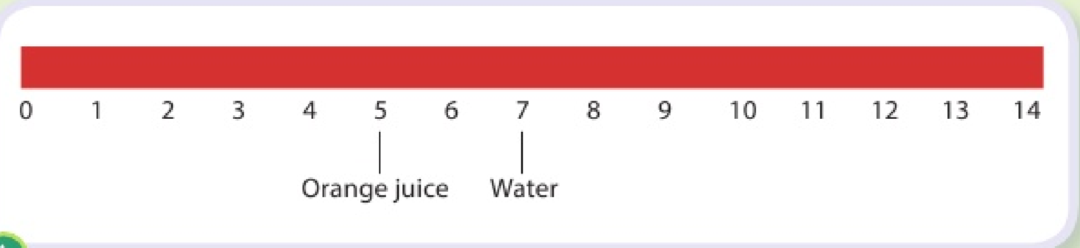


Figure 8.7 *A pH number line*

Did you know?

When you are stung by a nettle, hairs on the nettle leaf inject methanoic acid into your skin. A dock leaf contains chemicals that when rubbed over the sting neutralise this acid and cool down the skin.

Q U3 S2

8.6 What is the pH scale?

8.7 Name a substance that would have:

(a) A pH below 7

(b) A pH of 7

(c) A pH above 7.

Say which is acidic, basic or neutral.

8.8 Some people add lemon to tea instead of milk. When the lemon is added, the tea changes colour slightly. Why does this happen?

Making indicators

R2 R3 R5

Portfolio 55

Activity 8.3

Question

How do you make your own indicator using red cabbage?

Equipment needed

Bunsen burner	Glass rod	Water
Tripod	Test tube rack	A range of acid, base and neutral substances
Wire gauze	Test tubes dropper	
Beaker	Red cabbage	

Conducting the activity

1. Tear up some red cabbage leaves and add them to a beaker containing approximately 100 cm^3 of water.
2. Heat the water to boiling and stir the leaves using a glass rod. The water turns purple as the indicator is extracted from the leaves.
3. Allow the solution to cool.
4. Add acidic, neutral and alkaline solutions to different test tubes.
5. Using a dropper add 5–6 drops of the red cabbage indicator to each test tube. Stopper the tubes and shake.
6. Note the colour in each test tube and summarise your results in a table similar to the one below.

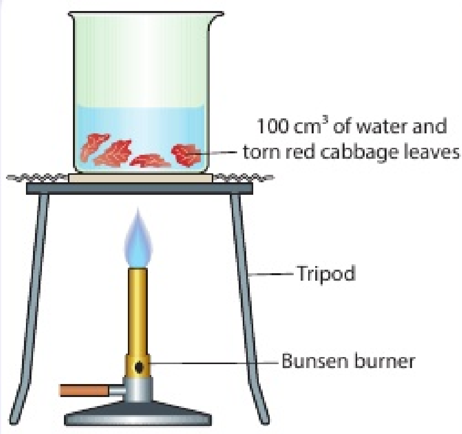


Figure 8.8 *Extracting the indicator from red cabbage*

Substance tested	Indicator colour
Acidic substance	
Alkaline substance	
Neutral substance	

Note: If you store the red cabbage indicator in the fridge it will last longer.

Q R3 Research

8.9 State the colours of the red cabbage indicator in acid, base and neutral solutions.

Reactions of acids

Acids are part of our everyday life and we must understand their key reactions. Acids react with many substances but one product they all have in common is a **salt**.

Acid and metal

An acid usually reacts with a metal to produce a salt and hydrogen gas. For example, the reaction between zinc and hydrochloric acid produces a salt and hydrogen gas.

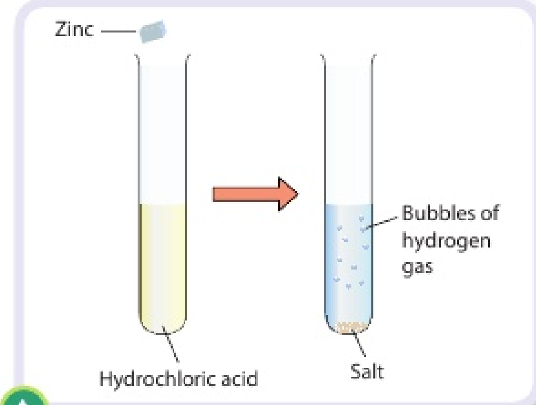


Figure 8.9 *Acid and metal*

Acids react with most metals and a salt is produced. But unlike the reaction between acids and bases we do not get water. Instead hydrogen gas is produced. The word equation for the reaction is:

metal + acid = salt + hydrogen

Note: When naming the salt the first part is the name of the metal (zinc) and the second is based on the acid used (hydrochloric acid). So:

Zinc + hydrochloric acid = zinc chloride + hydrogen

Test for hydrogen

To confirm the bubbles of gas are hydrogen you must place a lighted splint into the test tube. Hydrogen is easily ignited because it is flammable and you will hear a distinctive 'squeaky pop' sound if hydrogen is present.

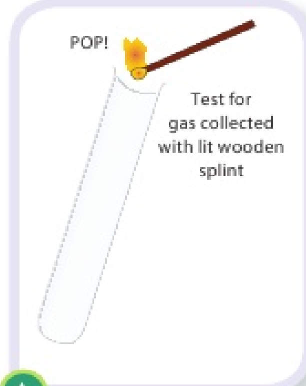


Figure 8.10 *Testing for hydrogen*

Research R2 Research R3 Research R4

Portfolio 57

Activity 8.4

Question

What happens when calcium reacts with hydrochloric acid?

Equipment needed

Hydrochloric acid (0.1 M)	Test tube
Calcium	Test tube rack
Cotton wool	Wooden splint

Safety

- Keep the test tube in the test tube rack at all times because it can become very hot.

Conducting the activity

1. Place 10 cm^3 of hydrochloric acid in the test tube.
2. Add a few granules of calcium.
3. Stopper the test tube with some cotton wool and allow gas to collect inside it for about 30 seconds.

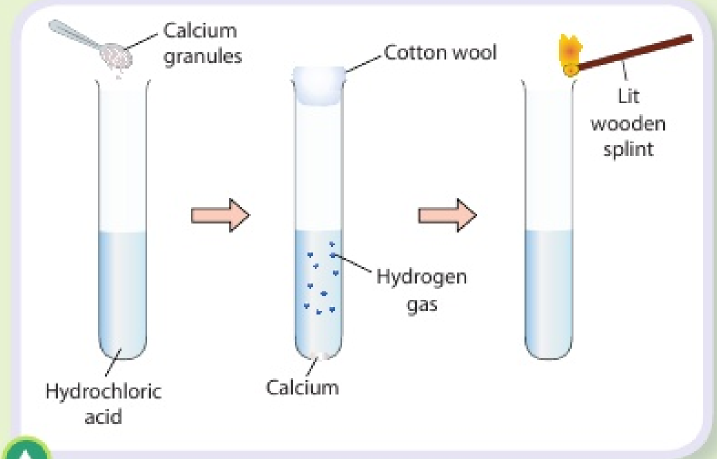


Figure 8.11 *Reacting calcium with hydrochloric acid*

4. Remove the stopper and place a lighted splint into the test tube. Record what happens.

Q Research R5

8.10 What is the name of the salt formed?

8.11 Name the gas produced.

8.12 Outline what happened when you tested the gas with a lighted splint.

What is neutralisation?

We have already mentioned that an acid is the chemical opposite of a base. So, what would happen if an acid and a base were mixed together?

The answer is they react together and **neutralise** or cancel each other. They react and produce a salt and water. Both salt and water are neutral. These reactions are known as **neutralisation reactions**.

$$\boxed{\text{an acid} + \text{a base} \longrightarrow \text{salt} + \text{water}}$$

Neutralisation can be done in a very accurate way by **acid–base titration**. The acid is placed in a burette and the base and indicator are placed in a conical flask. The conical flask is placed under the burette and the acid is added slowly to the base until the indicator changes colour when the salt and water are formed (neutralisation).

Titrations allow you to find out exactly how much acid is needed to neutralise a base.

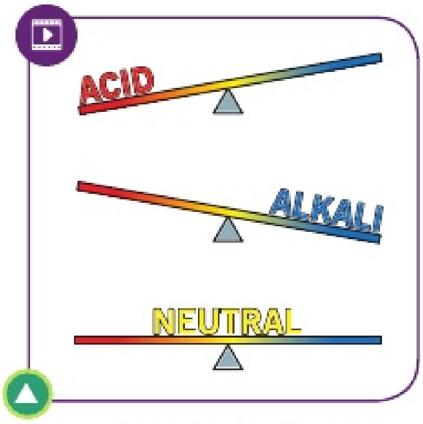


Figure 8.12 *Colour changes at different pHs*

Figure 8.13 *A student reading a burette*

R2 R3 R4

Activity 8.5

Question

How do you titrate hydrochloric acid (HCl) against sodium hydroxide (NaOH) and prepare a sample of sodium chloride (NaCl)?

Equipment needed

Burette	Evaporating dish	Dilute hydrochloric acid (0.1 M)
Pipette	Bunsen burner	Dilute sodium hydroxide
Pipette filter	Tripod	(0.1 M)
Conical flask	Wire gauze	An indicator
White tile	Beakers	Electronic balance
Retort stand	Funnel	

Conducting the activity

1. Using a pipette filter measure 25 cm^3 of dilute sodium hydroxide into a conical flask.
2. Add 3–4 drops of indicator, e.g. litmus, and place the flask on a white tile.
3. Using a funnel, fill the burette to the 0 cm^3 mark with the dilute hydrochloric acid.
4. Slowly and with continuous swirling of the conical flask add the acid from the burette until the indicator just changes colour. Note the volume of acid added.
5. Repeat the titration, taking care to add the acid drop by drop near the end point.
6. Calculate the average volume of acid from these two titrations.

7. Repeat the experiment without using any indicator and add the volume of acid calculated in step 6.
8. The conical flask contains a solution of sodium chloride in water.
9. Place a sample of the sodium chloride solution into an evaporating dish and evaporate off the water to leave the sodium chloride behind.
10. Use an electronic balance and find the mass of sodium chloride salt produced.

Remember: *Acid* is *a*bove in a burette. *Base* is *below* in a flask.

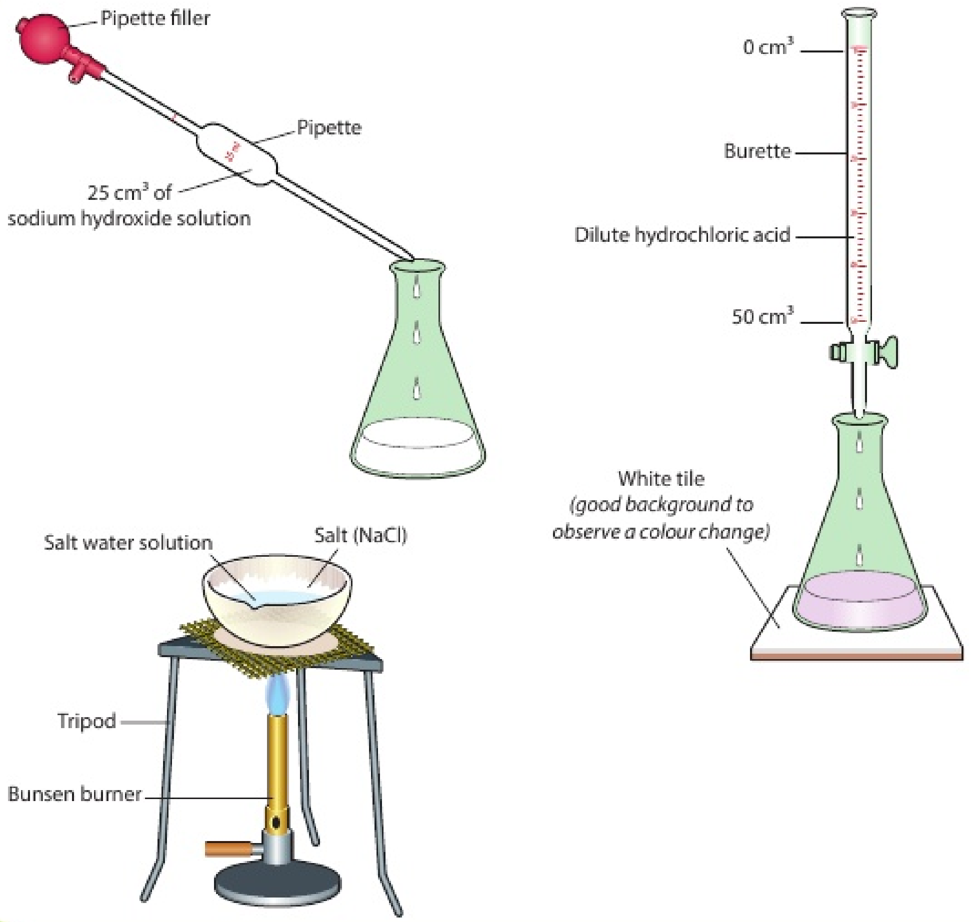


Figure 8.14 *Stages involved in a titration*

Q R 4 Research

From the results of the titration, answer the following:

8.13 What was the colour of the indicator when you added it to sodium hydroxide in the conical flask?

8.14 What was the colour of the universal indicator at the end point?

8.15 What was the average volume of acid needed to neutralise the base?

8.16 What is the mass of the sodium chloride salt produced?

Q Understanding U3

8.17 What is neutralisation?

8.18 Explain acid-base titration.

What is particle theory?

What happens during a neutralisation reaction is explained by particle theory

The common laboratory acids all have one atom in common: the hydrogen atom.

- Hydrochloric acid (HCl)
- Nitric acid (HNO_3)
- Sulfuric acid (H_2SO_4).

What happens to acids in water?

When laboratory acids are dissolved in water, the hydrogen atoms separate from the other atoms; it is these free hydrogen particles (**ions**) that make a solution acidic. The more hydrogen ions (H^+) there are, the lower the pH will be.

Acids are compounds that **dissociate** (break) into their ions when placed in water. The strong acid hydrogen chloride is one example – it breaks up into H^+ ions and Cl^- ions when placed in water:

$$HCl \longrightarrow H^{+1} + Cl^{-1}$$

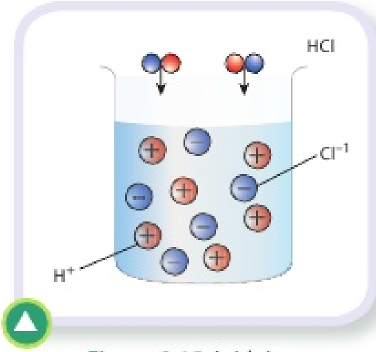


Figure 8.15 *Acids in water*

What happens to bases in water?

The common laboratory base sodium hydroxide (NaOH) will also dissociate when added to water.

$$NaOH \longrightarrow Na^{+1} + OH^{-1}$$

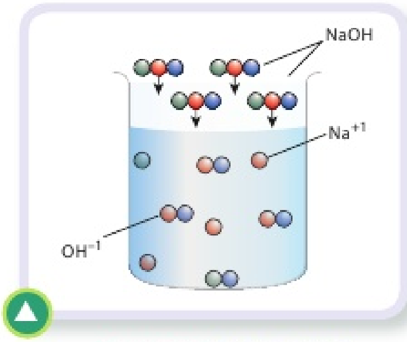


Figure 8.16 *Bases in water*

Neutralisation

When you add an acid to a base they neutralise – all the ions will rearrange and the products salt and water are formed.

$$H^{+1} + OH^{-1} \longrightarrow H_2O$$
$$Na^{+1} + Cl^{-1} \longrightarrow NaCl$$

Did you know?

If soil is too acidic, most crops will not grow well. Farmers can spread powdered limestone or lime on the soil to neutralise it.

Q Understanding U3 Society S2

8.19 A bee sting is acidic and a wasp sting is alkaline. The bee sting can be neutralised by rubbing a mild base such as baking soda or toothpaste on it. How would you treat a wasp sting?

Acid and Carbonate

Acids can also be neutralised by reacting them with carbonates (which also have pH more than 7). In this type of neutralisation, carbon dioxide gas is also produced.

an acid + a carbonate $\longrightarrow$ salt + water + carbon dioxide

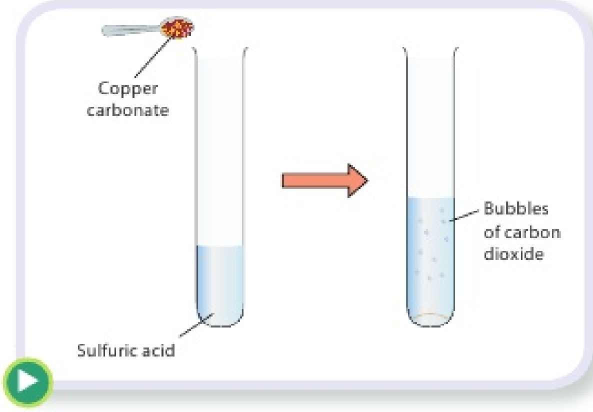


Figure 8.17 *The reaction of an acid and a carbonate*

Did you know?

Have you ever had indigestion? It is caused by too much hydrochloric acid in your stomach. The pain can be relieved by taking an antacid (anti-acid) chemical, which is a base.

Research R2 Research R3 Research R4 Portfolio 59

Activity 8.6

Question

Which brand of indigestion tablet is the most effective at neutralising an acid?

Equipment needed

Conical flask	Mortar and pestle
Burette	De-ionised water (purified water with no dissolved salt ions)
White tile	Universal indicator
Retort stand	Two brands of indigestion tablets
Pipette	Hydrochloric acid (0.1 M)
2 beakers	

Conducting the activity

1. Grind to powder one tablet from the first brand using a mortar and pestle.
2. Place the powder into the conical flask. Wash out the mortar with de-ionised water and transfer the washings to the conical flask.

3. Add 25 cm^3 of de-ionised water to the conical flask and swirl to mix after adding a few drops of universal indicator with a pipette.
4. Fill a burette with dilute hydrochloric acid (0.1 M).
5. Place the conical flask on the white tile under the burette (the white tile makes it easier to see any colour changes). Add the hydrochloric acid from the burette into the conical flask slowly, until the indicator changes colour (end point).
6. Record the total volume of acid added.
7. Repeat steps 1 to 6 with a second tablet from the first brand.
8. Repeat steps 1 to 6 twice with the second brand of indigestion tablet.
9. Note your results in a table similar to the one below.
10. Draw a bar chart or line graph showing the results for each brand of tablet.

Brand	Volume of HCl neutralised by tablet (cm^3)		
	Try 1	Try 2	Average

Q Understanding R2 Communicating R4 Society R5

8.20 Why do you think it is important to repeat each experiment twice for each brand of tablet?

8.21 Why did you place the conical flask on a white tile while carrying out titration?

8.22 From your observations during the activity:

(a) What colour was the universal indicator when added to the indigestion tablet solution?

(b) What was the colour change at the end point?

(c) Which brand of indigestion tablet was the most effective at neutralising the acid?

Q Understanding U3 Research R1

8.23 Inna has a solution that she thinks might be acidic. Describe three tests she could do to check whether or not it is acidic.

8.24 Nurlan tested some solutions with universal indicator. He wrote down their pHs: 1, 5, 7 and 14.

But he forgot to write down the names of the solutions. Can you help him by matching the pHs to the correct solutions?

Solutions tested	pH (1, 5, 7, 14)
Distilled water	
Sulfuric acid	
Sodium hydroxide	
Vinegar	

MODULE 9

Carbon and the Carbon cycle

Learning outcomes

At the end of this chapter you will be able to:

- Explain why carbon forms four bonds in many compounds 8.4.3.1
- Describe carbon spread in nature in form of simple substance and in composition of minerals 8.4.3.2
- Compare structure and properties of allotropic varieties of carbon; 8.4.3.3
- Explore uses of the different allotropic forms of carbon 8.4.3.4
- Explore the physical and chemical properties of carbon 8.4.3.5
- Describe the conditions of the formation of carbon dioxide and monoxide during combustion and their physiological effects 8.4.3.6
- Produce carbon dioxide to prove its presence and study its properties. 8.4.3.7
- Explain the different stages of the carbon cycle 8.4.3.8

Keywords

✔ acidic ✔ organic ✔ cycling ✔ photosynthesis ✔ respiration ✔ transpiration ✔ decomposition ✔ methane ✔ allotropes ✔ carbon monoxide ✔ limewater

Carbon Compounds

Most of the compounds that all life requires contain carbon. There are over ten million compounds of carbon that are known. It is the primary building block of all matter and comprises around one fifth of the human body.

As we have seen in Module 4, the organic compounds known as fossil fuels are hydrocarbons and are currently used to provide for most of the world's energy needs. We shall see in this module that all life is implicated in the natural carbon cycle. Carbon dioxide is converted through the various life processes of animals and plants into different carbon compounds which then decay and form fossil fuels.

Carbon dioxide also has a very wide range of everyday uses as shown in Figure 9.1:

Uses of carbon dioxide

Refrigeration

At temperatures below -78°C carbon dioxide forms a solid called dry ice. It is ideal for cooling substances as it is much colder than ice and does not melt into a liquid; hence the name 'dry ice'/

Fizzy drinks

Carbon dioxide is slightly soluble in water. It forms an acidic solution and puts the fizz in drinks.

Stage effects

When water is added to dry ice it produces thick dense clouds of water vapour which are used in special effects.

Fire extinguishers

Carbon dioxide does not support combustion and is more dense than air. This cuts off the oxygen supply to fires.

Figure 9.1

Figure 9.2 *Dry ice is solid carbon dioxide*

9.1 Which property of carbon dioxide leads to its use in fizzy drinks?

9.2 Which properties of carbon dioxide might simple lab tests for its presence target?

Carbon is the only element with a whole branch of chemistry -organic chemistry - based solely on studying it and its reactions.

The Carbon Atom

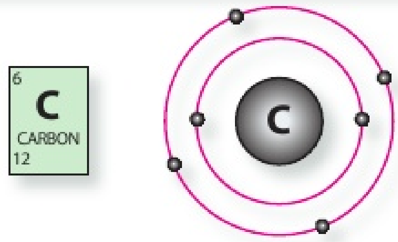
Look at the electron arrangement for carbon atoms.

To become stable the carbon atom wants to gain four more atoms to fill its second shell.

Four hydrogen atoms are needed to each share their single electrons with the four outer shell electrons of carbon.

Look at each atom in the methane molecule.

- The hydrogen atoms have two electrons and have a full first shell.
- The carbon atom has eight electrons in the second shell and is also full.



Carbon, 6 electrons,
electron configuration (2, 4)

Figure 9.3

Covalent compounds are not charged because electrons are not lost or gained but are shared.

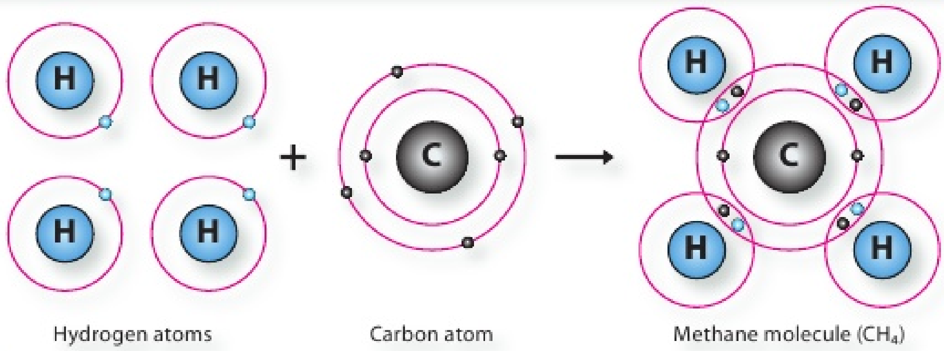


Figure 9.4 *Covalent bonding in methane (natural gas)*

The allotropes of carbon

Graphite, diamond and Buckminster fullerene are the three allotropes of carbon; that is, they are three different physical forms of the same element. In each form, carbon atoms are joined by strong covalent bonds but the different structures of the bonds give each allotrope very different properties.

Graphite

As Figure 9.5 shows, graphite is made up of giant two-dimensional molecules. It is a layered structure and these layers move relatively easily when force is applied. Graphite is thus quite soft and can be used as a lubricant in machinery and as pencil lead. Think about what happens when you press a pencil gently on paper – you are leaving a trace of graphite on the paper. Graphite is insoluble, has a high melting point and is a very good conductor of electricity. These properties make it ideal for use as electrodes in equipment we use in electrolysis experiments in the lab.

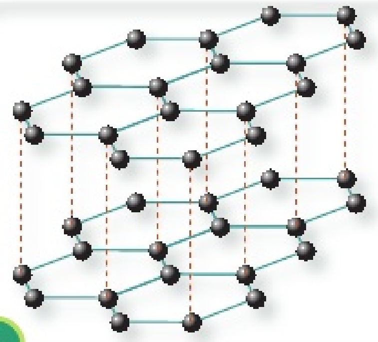


Figure 9.5 *The structure of graphite*

Diamond

Diamond has very different properties to graphite. It is colourless and is the hardest and has the highest melting point of any know material. It has many uses in drilling and cutting tools because of these properties. Its lustrous appearance also makes it very popular in different forms of jewellery.

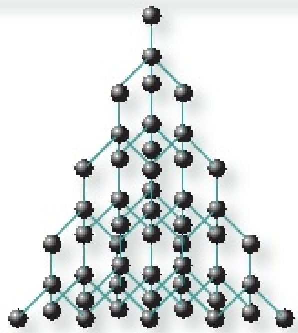


Figure 9.6 *The crystal structure of diamond*

Buckminster Fullerene

Discovered in the 1980's by English Chemist Harry Kroto, Buckminster Fullerene is a form of carbon with the formula C_{60}.

The fullerene structure is incredibly light and strong and these properties give it a wide range of uses such as medical drug delivery and applications in strengthening and coating products from sports equipment to textiles. Tiny 'pipe-like' structures made from fullerene are called 'nanotubes' and their properties as good conductors of electricity and heat mean they have a wide range of applications in electronic devices.

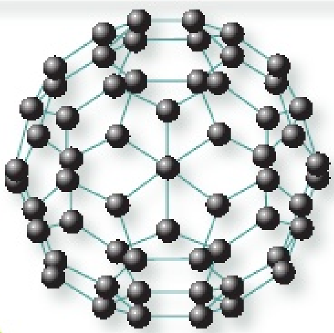


Figure 9.7 *Buckminster Fullerene*

Q Understanding U4

9.3 Name two differences in the properties of carbon allotropes: graphite and diamond which account for their different uses.

9.4 What properties make graphite ideal as a material for electrodes?

Activity 9.1

Question

How can we prepare a sample of carbon dioxide gas to examine its properties?

Equipment needed

- Conical flask
- Tap funnel
- Gas jars
- Delivery tubing
- Cardboard
- Wooden splint
- Candle
- Dilute hydrochloric acid
- Calcium carbonate (marble chips)
- Limewater
- Moist red and blue litmus papers

Carbon dioxide can be produced by reacting an acid with a carbonate. They react and produce carbon dioxide, water and a salt.

Word equation

hydrochloric acid + calcium carbonate → calcium chloride + water + carbon dioxide

Chemical equation

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

Conducting the activity

1. Set up the apparatus as shown in Figure 9.8.
2. Slowly allow the dilute hydrochloric acid to fall on the marble chips. The chips fizz as they react with the acid, producing carbon dioxide.
3. Continue to add the acid as necessary to maintain the reaction.
4. Collect three jars of carbon dioxide gas. To check when a jar is full of carbon dioxide, hold a lighted splint at the mouth of the jar. When the jar is full of the gas the splint will be extinguished.
5. Note your observations in the table below.

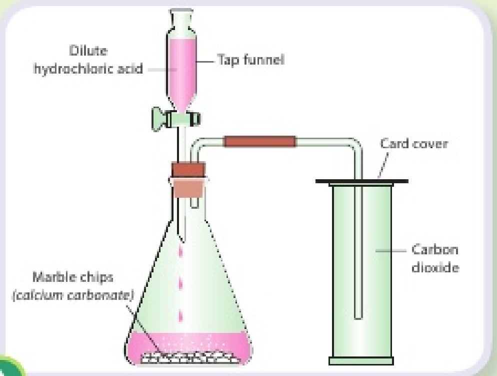


Figure 9.8 *Apparatus for the preparation of carbon dioxide*

Table 9.1

Table 1	Observations
Jar 1	
Jar 2	
Jar 3	

Carry out the following tests.

Jar 1

Note the colour and smell of carbon dioxide. Test its pH by adding pieces of moist red and blue litmus papers.

Jar 2

Pour a small volume of limewater into a jar of carbon dioxide gas. Place a lid on the jar and shake for 10 seconds.

Jar 3

Place a lit candle in an empty beaker. Invert a gas jar of carbon dioxide above the beaker and wait.

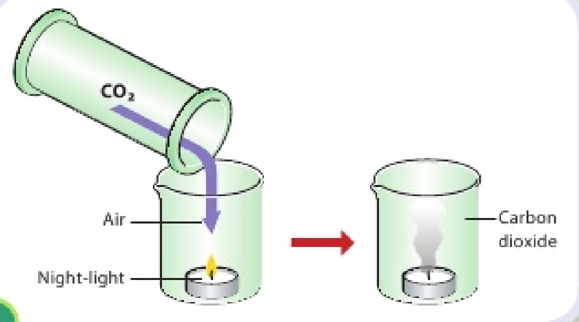


Figure 9.9 *Carbon dioxide gas being poured onto a candle*

Test for carbon dioxide

When carbon dioxide comes into contact with limewater (which is calcium hydroxide) they react together and form calcium carbonate and water.

The word equation for this reaction is

$$\text{calcium hydroxide} + \text{carbon dioxide} \longrightarrow \text{calcium carbonate} + \text{water}$$

The chemical equation is

$$Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$$

The calcium carbonate which is produced is insoluble in water and causes the solution to change colour from clear to milky. This is the characteristic test for carbon dioxide gas.

Carbon dioxide turns limewater milky.

Q Understanding U4 Communicating R4 Society R5

9.5 What did you observe in Jar 1 and what can you conclude?

9.6 What did you observe in Jar 2 and what can you conclude?

9.7 What did you observe in Jar 3 and what can you conclude?

R2 R3 R4

Activity 9.2

Question

How can we prepare a carbon dioxide balloon?

Equipment needed

Test tube

Balloon

Vinegar

Baking soda

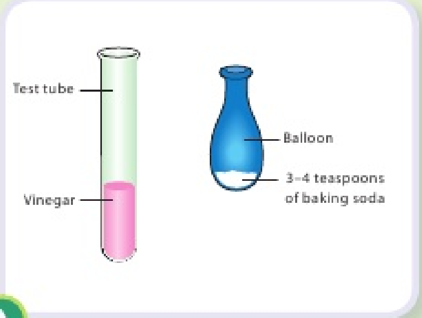


Figure 9.10

Conducting the activity

1. Set up as shown in Figure 9.10.
2. Place the top of the balloon over the mouth of the test tube and then invert the balloon allowing the baking soda to come into contact with the vinegar.
3. Allow the balloon to inflate with the carbon dioxide produced and then remove it and tie a knot.
4. Throw the balloon in the air and see what happens.

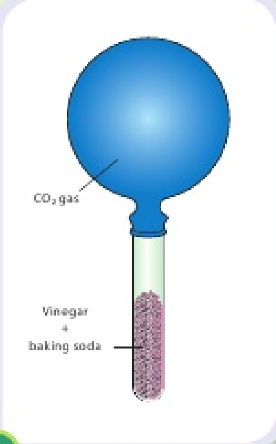


Figure 9.11

U4 R5

9.8 What conclusion does this experiment confirm about carbon dioxide?

9.9 What type of reaction is the reaction between the acid and the base?

Combustion

We have seen in Module 4 that when fuels such as hydrocarbons are burned with a good supply of oxygen combustion takes place. Complete combustion is the process whereby the fuel burns fully and the carbon is oxidised and becomes carbon dioxide and the hydrogen is oxidised to form water.

$$\text{hydrocarbon} + \text{oxygen} \longrightarrow \text{carbon dioxide} + \text{water}$$

When the supply of oxygen, however, is restricted incomplete combustion occurs. The products of this reaction are carbon monoxide, carbon and water.

$$\text{hydrocarbon} + \text{oxygen} \longrightarrow \text{carbon monoxide} + \text{water}$$

The carbon is produced in the form of soot. Carbon monoxide is a poisonous gas and all the more dangerous to humans because it has no taste or smell so is very hard to detect. Incomplete combustion occurs when the supply of air or oxygen is poor. Water is still produced, but carbon monoxide and carbon are produced instead of carbon dioxide. It can kill if people are exposed to it in large quantities and even small amounts can make people feel unwell. When carbon monoxide is breathed in, it enters the bloodstream. Here it mixes with haemoglobin to form carboxyhaemoglobin which prevents the blood from being able to carry oxygen to cells and tissues.

Research R2 Research R3 Portfolio 68

Activity 9.3

Question

What does incomplete combustion look like?

Equipment needed

Bunsen burner
Methane gas

Conducting the activity

1. Your teacher will demonstrate the flame from the Bunsen burner when the air hole is fully open and closed.
2. Note your observations.

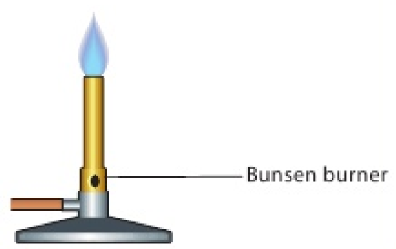


Figure 9.12

Understanding U4 Research R5

9.10 What would you expect to see form if the Bunsen burner burned with the air closed for a long time?

Does nature recycle?

For many households and businesses it is a daily routine to recycle materials such as plastic, aluminium, paper and glass. We do this in order to conserve and protect important resources in our environment.

Figure 9.13 *Our daily recycling*

But what about recycling in nature? How does nature recycle material and nutrients (chemicals that organisms need in order to live)?

Scientists do not use the term 'recycling' in terms of nature. Instead they use the term **cycling** or **cycles**.

Nature excels at cycling vital resources. It has to do so!

Earth has a fixed amount of matter. Matter is anything that occupies space and has a mass. Therefore nature is continually cycling matter to ensure that it can be used time and time again.

Figure 9.14 *Cycling in nature*

Communicating C1

In pairs, complete the following research:

9.11 Why it is important for Earth to cycle matter?

9.12 How do cycles in nature maintain the stability of Earth's climate?

9.13 Identify key nutrients that nature cycles.

Two of the key resources that nature cycles are water and carbon. We will be looking at carbon in this module and water in the next and considering their importance to us and our planet.

What is the carbon cycle?

Scientists know that carbon is the basis of all life. Carbon is the second most abundant substance in living things.

9.14 What substance is found in the greatest quantities in living things?

Any substance that contains carbon is called an **organic** substance.

We all need carbon for survival. Earth has only a certain amount of carbon and so it is very important that nature continues to cycle it to maintain its supplies.

However, a fine balance must be kept between too much carbon and too little carbon in the atmosphere.

Understanding U4

9.15 Carbon is the fourth most abundant element in the entire universe. What do you think are the top three elements in our universe?

What are the stages of the carbon cycle?

Carbon, like water, is constantly moving in nature from one process to the next. It is this constantly moving carbon that results in the carbon cycle. This is shown in Figure 9.15.

Within the carbon cycle there are six key stages. In these stages carbon moves from:

1. The atmosphere to plants
2. Plants to animals
3. Plants and animals to the ground
4. Living things into the atmosphere
5. Fossil fuels back into the atmosphere
6. The atmosphere into the oceans.

We will look at each stage.

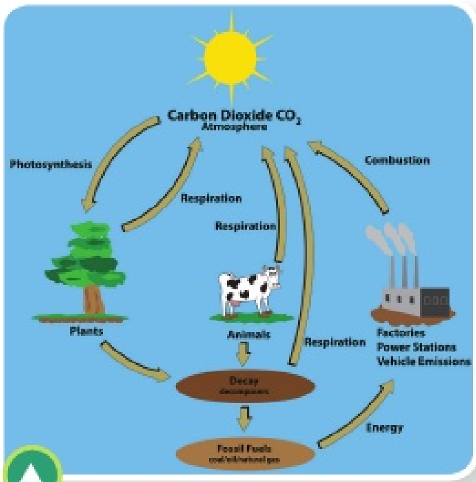


Figure 9.15 *The carbon cycle*

understanding U4

9.16 Carbon appears in many forms, such as coal, diamonds and graphite in pencils. Research these three forms of the same element, carbon, and clearly distinguish between them in terms of:

(a) Appearance

(b) How they are formed

(c) What they can be used for.

Carbon moves from the atmosphere to plants

Carbon, in the form of carbon dioxide gas, is taken from the atmosphere by plants and is converted into food containing carbon. This process, that allows plants to uptake carbon dioxide and make food, is called **photosynthesis**.

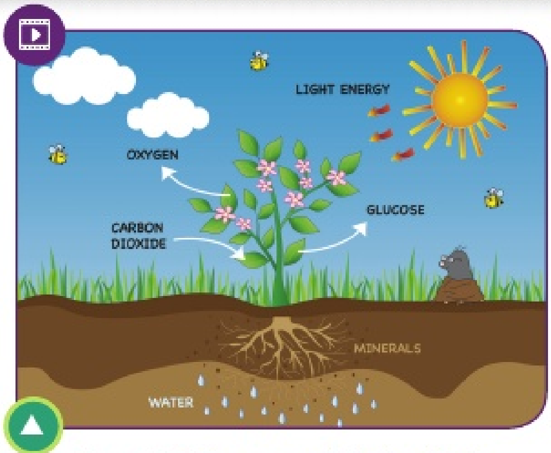


Figure 9.16 *The process of photosynthesis*

Carbon moves from plants to animals

Animals eat plants and other animals in order to get their supply of carbon.

Figure 9.17 *Animals get carbon by eating plants and other animals*

Carbon moves from plants and animals to the ground

When plants, animals and micro-organisms die the remains of their structures will decompose into the ground. Over millions of years the remains may become fossil fuels.

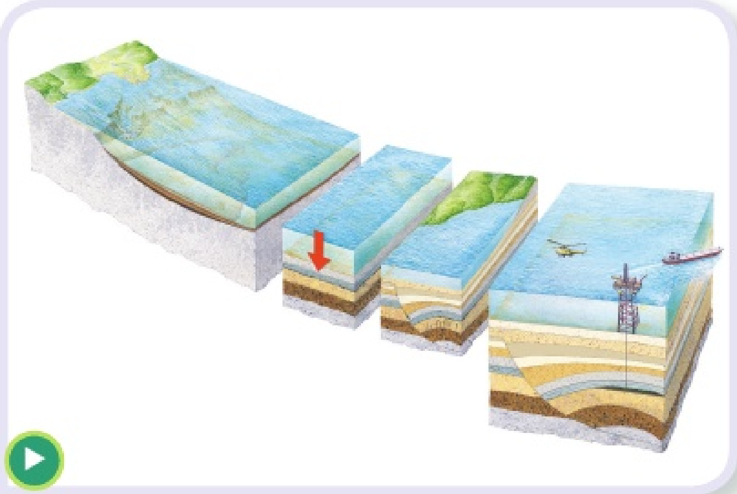


Figure 9.18 *Fossil fuels are the remains of plants and animals*

9.17 What does the word 'fossil' mean?

9.18 What are the main fossil fuels on Earth?

Carbon moves from living things into the atmosphere

Plants, animals and micro-organisms break down food containing carbon in a process called **respiration**. This process releases energy to the organism but produces both water and carbon dioxide as waste products. The carbon dioxide is released into the atmosphere.

Carbon moves from fossil fuels back into the atmosphere

Fossil fuels are fuels formed from the remains of plants and animals that lived millions of years ago.

- Coal was formed from trees and ferns that died and were buried and compressed under swamps.
- Oil and natural gas were formed from animals and plants that lived in the sea.

Carbon moves from the atmosphere into the oceans

Most of the carbon that is absorbed from the atmosphere is absorbed by the oceans. Carbon is involved in many processes that take place in the oceans.

Did you know?

The amount of carbon in the oceans is fifty times greater than the amount contained in the atmosphere.

Figure 9.19 *Organisms in the oceans use carbon*

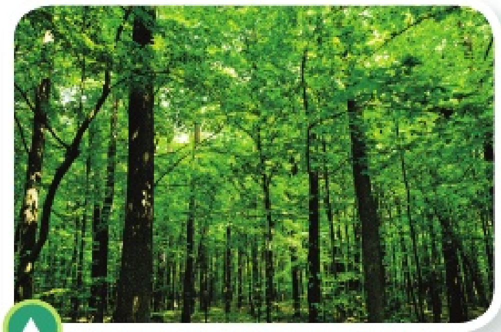


Figure 9.20 *A forest is a carbon sink*

What are carbon sinks and sources?

In nature, places that store carbon are called **carbon sinks** or **reservoirs**. An example of a carbon sink would be a forest full of trees.

Q Understanding U4

9.19 What structures within the trees contain carbon?

Where carbon is given off or released in great quantities, this is called a **carbon source**. An example of a carbon source is a volcano.

Figure 9.21 *A volcano is a carbon source – carbon is released when it erupts*

Q Understanding U4

9.20 Draw a simple diagram to show the carbon cycle.

Did you know?

By the time a person feels thirsty, his or her body has lost over one percent of their total water volume in their body.

Water can exist in solid (ice), liquid (water) or gas (water vapour) form.

Figure 10.3 *Water can exist in solid (ice), liquid (water) or gas (water vapour) form*

Q Understanding U4

10.1 Give examples from nature of water in solid, liquid and gas form.

Water in nature is always in constant motion.

The **water cycle** is the motion or movement of water from the oceans to the atmosphere and back to the oceans.

What are the stages of the water cycle?

Within the water cycle there are six key stages:

- Stage 1 – evaporation
- Stage 2 – condensation
- Stage 3 – precipitation (rainfall)
- Stage 4 – infiltration (soaking)
- Stage 5 – runoff/collection
- Stage 6 – transpiration.

We will look at each stage individually.

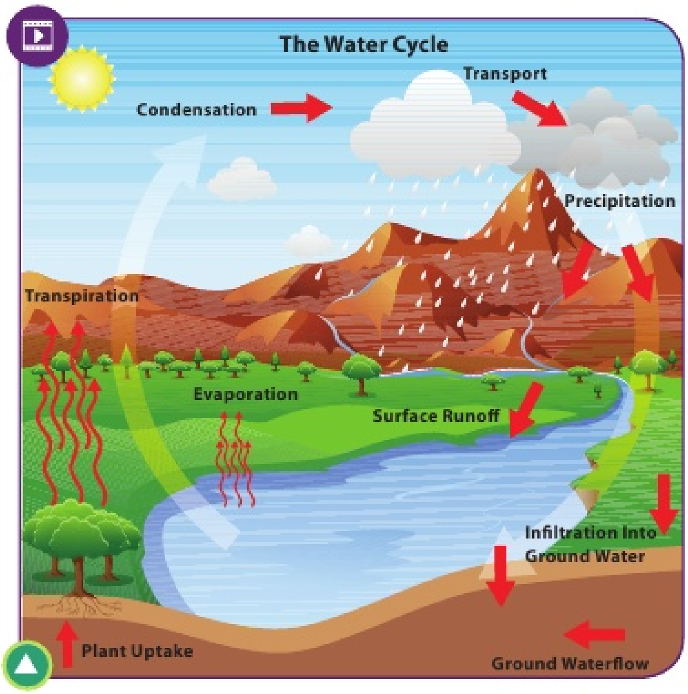


Figure 10.4 *The water cycle*

Stage 1 – evaporation

Evaporation is the changing of water in the form of a liquid to water in the form of a gas. Most of the evaporation that takes place on Earth comes from the oceans. Only a small amount of evaporation takes place from inland water (rivers and lakes) and plants.

Evaporation requires energy in the form of heat. It is the Sun's heat energy that drives this process causing water to change from a liquid to a vapour/gas.

Figure 10.5 *A boiling kettle, where water is turned into water vapour (a gas)*

Stage 2 – condensation

Condensation takes place when a gas changes to liquid. Think of steam from the shower hitting a mirror – what do you notice happening?

The higher up you go on Earth the colder it gets – think of snow caps on the tops of mountains. So as water vapour rises it starts to get colder and condenses. The water droplets are still light enough to remain suspended or held in the air. The droplets come together in vast numbers and clouds are formed.

Figure 10.6 *Condensation on glass*

Understanding U4 Research R2 Research R5 Portfolio 75

Activity 10.1

Question

How can we demonstrate evaporation?

Equipment needed

A piece of material about 30 cm square, cut in half
Zip-lock bags
Basin
Water
Stop watch or timer
Beaker
Open container
Top-pan balance/weighing scales

Safety

- Ensure that water does not spill on the floor, making another student's area unsafe.

Conducting the activity

1. Measure 50 cm^3 of water into a beaker.
2. Place your two pieces of material in a basin and pour the water from the beaker onto the material.

3. Leave the material to soak in the water for 5 minutes.
4. Remove the material and twist the pieces once for 10 seconds.
5. Place one piece of material on a top-pan balance/weighing scales and record the mass of the material soaked with water. Do the same with the second piece.
6. Place one piece of material into a zip-lock bag.
7. Place the second piece of material into an open container.
8. Leave both the bag and the container in an open, airy place for two to three days.
9. After the two to three days record the mass of each piece of material and note the 'dryness' of the material.

Stage 3 – precipitation/rainfall

As the water droplets held within the clouds continue to rise they carry on getting colder. When the water droplets become too heavy to be held in the cloud, the water falls back to Earth as rain, sleet or snow.

Figure 10.7 *Rain falls when water droplets become too heavy to be held in the cloud*

Q U4 Understanding

10.2 When might the level of ground water increase or decrease?

Did you know?

Water can spend thousands of years locked up in ground water or frozen within the polar ice caps. Yet it can remain within the atmosphere only for a few days.

Stage 4 – infiltration

Infiltration occurs when the rain hits the ground and infiltrates or soaks through the soil and the rocks. This water then collects and forms **ground water.** Ground water is water that is held in the ground and supplies wells and springs.

Figure 10.8 *Groundwater runs off the land into rivers and lakes*

Stage 5 – runoff/collection

The water eventually runs off the land or out of the soil and collects into the rivers and lakes.

Stage 6 – transpiration

Plants are constantly taking up water from the soil through their roots. To continue this uptake of water from the soil, the plants must lose water into the air from their leaves in the form of water vapour. This loss of water is called transpiration.

The water has now completed its full cycle.

Q Understanding U4

10.3 What do the words evaporation and condensation mean?

10.4 Draw a simple diagram to show the water cycle.

Causes of water pollution

The main causes of water pollution are:

- Badly treated sewage (toilet) waste
- Oil spills
- The dumping of household, farming and industrial wastes.

Q Communicating C3 Society S2

10.5 When nutrients enter a water supply (such as a stream, river, pond or lake) they allow more simple plants called algae to grow in the water. This happens because the nutrients are rich in minerals such as phosphates. The extra minerals may result in an overgrowth of algae, called an algal bloom.

When the algae die, bacteria cause them to decompose. The bacteria use up all the oxygen in the water as they grow. This results in the death of all the animals and plants in the water.

Nutrients are present in wastes such as animal manure (slurry), fertilisers that are washed off the soil, dumped milk, sewage and washing detergents.

(a) What chemical element is mainly responsible for the death of plants, as described above?

(b) When algae grow in water they carry out photosynthesis. What gas do they add to the water?

(c) The gas you named in part (b) has benefits for fish and other water-based animals. Why do animals benefit from this gas?

(d) Why are farmers advised not to apply fertilisers when there is heavy rain?

(e) Untreated sewage or waste water from kitchen sinks can be examples of poor conservation practice. Explain why this is so.

Figure 10.9 *An algal bloom caused by excess fertiliser*

Figure 10.10 *Fish killed by lack of oxygen in the water*

Research R2 Research R4 Research R5 Society S2

Portfolio 76

Activity 10.2

Question

Can we purify water?

Our drinking water is purified in five steps:

1. It is aerated to add oxygen to the water and to remove unwanted gases (this is called *aeration*).
2. Small dissolved dirt particles are made to clump together (this is called *flocculation*).
3. It is allowed to settle so that large particles sink to the bottom (*sedimentation*).
4. It is filtered to remove tiny particles from the water (*filtration*).
5. Chemicals are added to kill micro-organisms (*disinfection*).

We will carry out the first four steps. Disinfectants are dangerous so we will not use them in the classroom. As a result the water we purify is **not** safe to drink.

Equipment needed

- About 1 litre of muddy water (or 1 litre of water with two cups of soil mixed into it)
- Empty 2 litre plastic soft drinks bottle with its cap
- Empty 2 litre plastic bottle cut about halfway down
- Two large beakers
- Filter funnel
- Alum (aluminium potassium sulphate) solution
- Small graduated cylinder
- 250 ml beaker
- Fine sand
- Coarse sand
- Small pebbles
- Filter paper
- Rubber band
- Stirring rod or spoon
- Stopwatch or timer

Safety

- Take care when cutting the plastic bottle – make sure your knife or scissors do not slip.
- The filtered water is **not** safe to drink.

Conducting the activity

1. Pour half of the dirty water into a beaker and leave it aside. This acts as a control.
2. Pour the other half of the dirty water into the empty bottle. Close the lid and shake vigorously for 30 seconds.
3. Pour the water into a large empty beaker and back into the bottle ten times. (Steps 2 and 3 aerate the water and remove any smelly gases.)
4. Pour the water into the base of the bottle with the top cut off.
5. Add 30 ml of alum solution to the water and slowly stir it for five minutes. What do you notice about the water as you stir it? (This is flocculation.)
6. Let the water sit without stirring it or moving it for twenty minutes. Observe it every five minutes. What do you notice is happening in the water? (This is called sedimentation.)

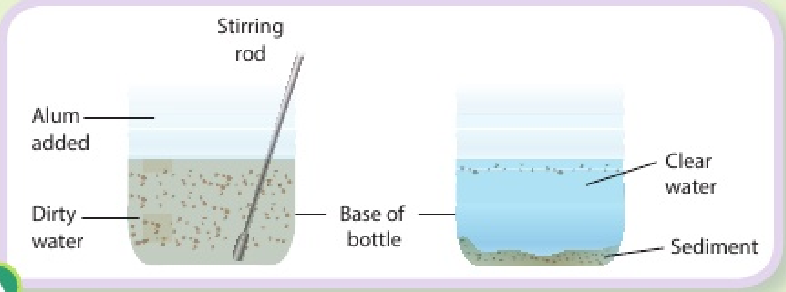


Figure 10.11

7. Place the filter paper over the mouth of the bottle with the bottom cut off and secure it firmly in place using the rubber band. Place it upside down in the beaker.
8. Pour a 250 ml beaker of pebbles into the bottle. Then pour a 250 ml beaker of coarse sand on top of the pebbles. Then pour two 250 ml beakers of fine sand on top of the coarse sand.
9. Pour about 2 litres of clean tap water onto the sand in the upturned bottle. Be careful not to disturb the sand too much. Pour away the water that passes through the filter paper. (This step cleans the filter materials, i.e. the sand and pebbles.)
10. Pour about two-thirds of the dirty water through the filter system. Do not pour any of the sediment into the filter bottle.
11. When all the water has filtered through, compare the filtrate to the original dirty water. What do you notice? How different are the two water samples?
12. Note that the filtrate is **not** safe to drink. It may contain micro-organisms that are harmful.

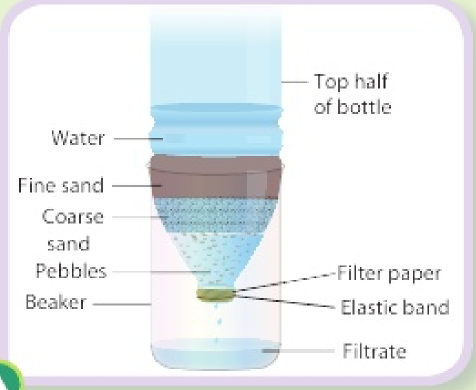


Figure 10.12

Communicating C1 Society S2

10.6 How is water aerated naturally?

10.7 How is water naturally filtered?

10.8 Find out what chemicals are added to our drinking water to disinfect it.

10.9 Why is rainwater that is collected in a barrel not safe to drink but water in a well (or spring water) is safe to drink?

What is hard water?

Water has many uses. One of these is washing. Soap is added to water to make washing easier. How the soap reacts with the water depends on whether it is **hard water** or **soft water**.

Hard water: Does not lather easily with soap.
Soft water: Lathers easily with soap.

Hard water

Water is an excellent solvent. As a result, water almost always has solutes dissolved in it. Many bottled waters taste different due to the dissolved solids which the water picks up as it flows over and through different rocks.

- Soap is added to water to help it dissolve even more substances like grease and oil.
- The soap usually forms a lather on top of the water.
- In some regions a 'scum' will form on the water instead of a lather.
- More soap must be added to the water in order to get lather.

Rainwater

Rainwater is mildly acidic due to dissolved carbon dioxide in it forming carbonic acid.

- Rainwater dissolves minerals as it passes through the ground and rocks.
- Hard water is always found in regions where there is a lot of limestone rock. When water passes through limestone, calcium and magnesium ions are dissolved in it.
- These ions react with soap to form an insoluble compound: scum which floats on top of the water

Figure 10.13 *Soap scum formed by hard water*

$$\text{calcium ions} + \text{soap} \longrightarrow \text{scum}$$

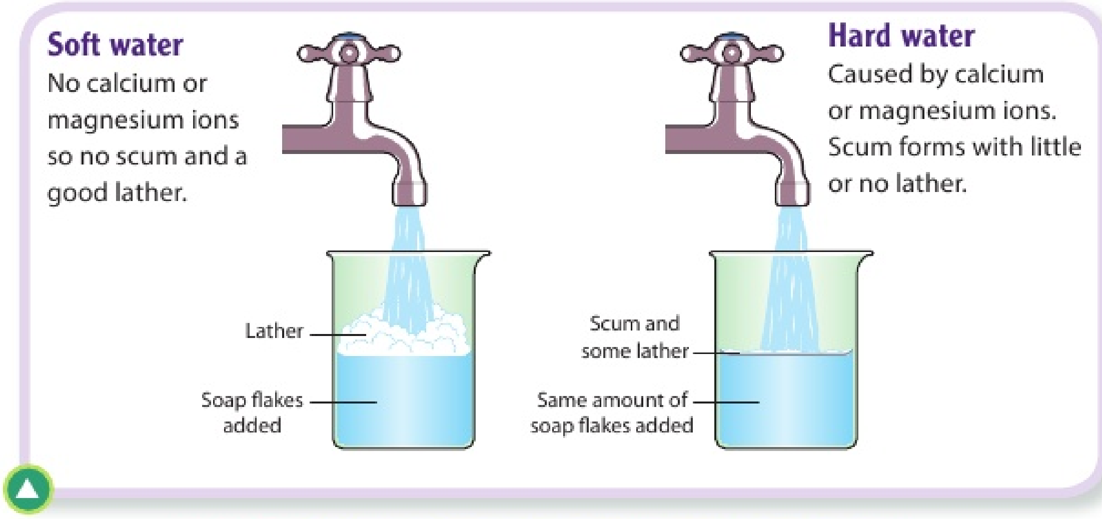


Figure 10.14 *Soft water lathers well with soap, whereas hard water does not*

Problems associated with hard water

- When hard water is heated in a kettle a white insoluble solid know as limescale is deposited.
- Limescale sticks to heating elements and pipes making them less efficient.
- Kettle descalers are acidic solutions which react with and dissolve the limescale (calcium carbonate).

Figure 10.15 *Thick deposits of limescale on the interior of a hot water pipe*

Problems caused by hard water mean it is often necessary to 'soften' it before it can be used. Softening water involves removing the calcium and magnesium ions through:

- boiling
- distillation
- ion exchange.

Ion exchange

One of the easiest and cheapest ways to remove hardness from water is to pass it through an ion exchanger.

As water passes through the ion exchanger, calcium and magnesium ions are swapped for sodium ions. Once calcium and magnesium ions are removed, the water has been softened. Eventually the exchanger will fill with calcium ions. This can be reversed by adding salt (sodium chloride) to the column and 'washing out' the calcium ions.

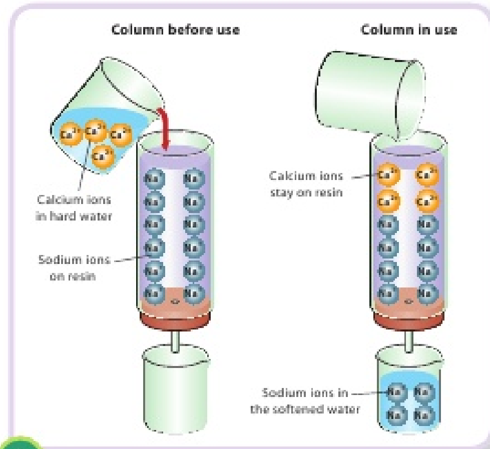


Figure 10.16 *Ion exchange*

U4

10.10 Complete the table with one word in each gap.

Advantages	Disadvantages
Good source of calcium which is needed for strong bones and ____________ .	Limescale blocks ____________ and leaves a scale on ____________ .
Most people prefer the taste of hard water.	It wastes ____________ .
It is good for brewing and tanning.	It produces a ____________ with soap.

U4 R2 R5

Portfolio 77

Activity 10.3

Question

How can we test for water hardness?

Equipment needed

Test tubes
Test-tube rack
Burette or Dropper
Soap solution or Soap flakes
Water samples

Conducting the experiment

1. Place equal volumes of different water samples in the test tubes as shown.
2. Add 1 cm^3 of the soap solution (or a soap flake) to each sample of water using a burette or a dropper.
3. Stopper and shake the test tube for five seconds and see if a lather (lasting more than 10 seconds) has formed.
4. If not, add another 1 cm^3 of soap solution and stopper and shake the test tube again. Continue doing this until a lather has formed.
5. Repeat this process with the water samples in the other test tubes.
6. Note your results in a table similar to the one below.

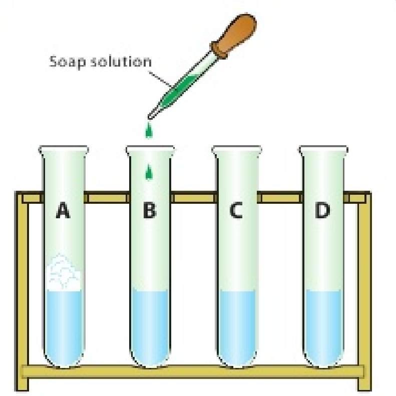


Figure 10.17 *Testing water samples for hardness*

Water Sample	A	B	C	D
Volume soap added [cm^3] to form lather				

U4 R5

10.11 What conclusions can you draw about water hardness in each sample?

Common tests for water

Either of the following tests can be used to test for the presence of water.

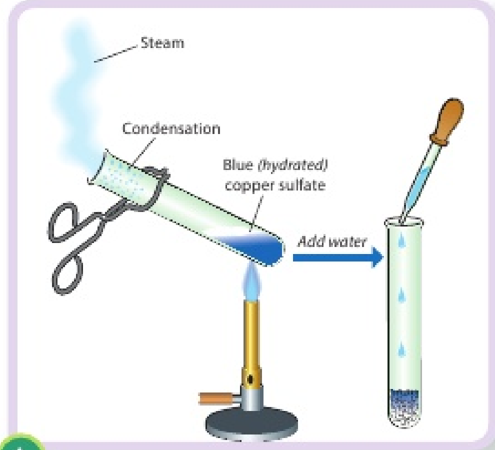


Figure 10.18 *Water will cause anhydrous (dry) copper sulfate to change colour from white to blue*

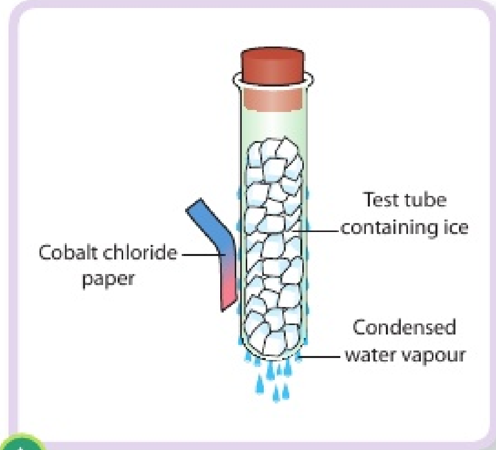


Figure 10.19 *Water will cause cobalt chloride paper to change colour from blue to pink*

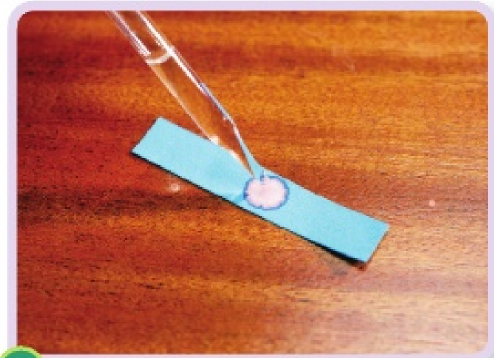


Figure 10.20 *A piece of cobalt chloride paper with water being dropped onto it*

Glossary

A

acid A substance which turns blue litmus red and has a pH less than 7. Acids have a sour taste and a sharp feel.

acid rain Rainwater with a pH of less than 5.5.

activation energy The least amount of energy required by atoms or molecules so they can undergo a chemical reaction.

reactivity series A list of metals placed in order of how reactive they are.

alkali A substance which reacts with an acid to form a salt and water. Alkalis are bases that dissolve in water.

alkali metals The elements of Group 1 of the periodic table, e.g. sodium.

alkaline earth metals The elements of Group II of the periodic table, e.g. calcium.

alloy A mixture of metals.

atom The smallest part of an element that still has the properties of that element.

atomic number The number of protons in an atom of that element.

B

base A substance that neutralises an acid to form a salt and water. Bases turn red litmus blue.

bioplastics Biodegradable plastics made from renewable biological sources.

boiling point The temperature at which a liquid boils.

bond Force of attraction that holds atoms together

bond energy The energy required to break bonds of a molecule into individual atoms.

C

carbon footprint The amount of carbon dioxide released that results from the activities of a person, persons or community.

carbon reservoir A place on earth where there is a buildup of carbon dioxide.

carbon sink Part of the earth that can absorb carbon dioxide from the atmosphere.

catalysts Substances that alter the rate of a chemical reaction but are not used up themselves.

ceramic A non metallic synthetic material

chemical change A change (reaction) in which a new substance is formed.

chemical energy Energy stored inside materials in the chemical bonds holding them together.

chemical reactions Many materials are produced from chemical changes and these are known as chemical reactions.

chlorination The addition of chlorine to water to kill bacteria.

chromatography Separation techniques used to separate a mixture of dissolved substances in a solution.

composites A material which is made up of at least two different materials.

composting Where organic waste such as food waste decomposes in the presence of oxygen, with no production of methane gas.

compounds Made up of two or more different types of atom chemically combined.

concentrated solution A solution which contains a large amount of solute in a small amount of solvent.

condensation Where cooling a gas causes it to change into a liquid.

conservation of mass When matter is changed from one form into another but there is no change in the overall mass.

convection The transfer of heat through liquids and gases by the mass movement of particles.

corrosion An undesirable process whereby a metal changes to its oxide or some other compound by combining with oxygen from air.

covalent bond A bond that consists of a pair of electrons shared between two non-metal atoms.

crystallisation The formation of crystals by cooling a saturated solution or by evaporating off solvent.

D

diffusion The name used to describe the way particles in gases and liquids spread throughout the space in which they are placed.

dilute solution A solution containing a small amount of solute in a large amount of solvent.

distillation A process used to separate two miscible liquids with different boiling points, e.g. alcohol and water.

E

electrolysis Where an electric current is passed through a molten sample of a metal compound causing it to split up and release the metal.

electron A negatively charged particle in an atom.

electronic configuration The arrangement of electrons in an atom.

elements Substances made up of only one type of atom.

endothermic A reaction where energy is absorbed by the chemicals from its surroundings in order for the reaction to take place.

energy The ability to do work. It is measured in joules (J).

energy profile diagram Shows the energy transfer in a chemical reaction.

evaporation The changing of a liquid to a gas.

exothermic A reaction where energy is transferred from the chemicals to the surroundings.

exploration Investigations that are undertaken in an unfamiliar area.

F

filtration A method used to separate insoluble solids from liquids.

fossil fuel A natural fuel, coal, oil, peat and gas, that was formed over millions of years from the remains of dead plants and animals.

freezing Where a liquid is cooled and changes into a solid.

fuel A substance that burns in oxygen and produces heat.

G

global warming An increase in the temperature in the earth's atmosphere due to the greenhouse effect.

greenhouse effect When the heat energy of the sun is trapped within the earth's atmosphere.

H

halogens The elements of Group VII of the periodic table, e.g. chlorine.

heat A form of energy that is transferred from warmer bodies to colder ones.

hydrocarbons Compounds made up of hydrogen and carbon.

I

indicators Chemicals that show by changing colour whether a substance is acidic, alkaline or neutral.

insoluble Solids that do not dissolve in a liquid.

ion An atom that has lost or gained electrons and is then charged.

isotopes Atoms that have the same number of protons but different numbers of neutrons.

L

latent heat Heat that is needed to change the state of a substance.

limewater A chemical that turns from clear to cloudy (or milky) if carbon dioxide is present.

M

mass The amount of matter in an object. This amount never changes.

mass number The number of protons and neutrons in an atom of that element.

matter Anything that occupies space and has mass.

melting point The temperature at which solid materials turn to a liquid.

miscible liquids Liquids that mix, e.g. alcohol and water.

mixture Contains two or more different substances mingled together but not chemically combined.

molecule Made up of two or more atoms chemically combined.

N

nanotechnology The use of (tiny) nanoparticles, which are measured in nanometres.

neutron An uncharged subatomic particle found in the nucleus of an atom.

noble gases The elements of Group VIII of the periodic table, e.g. helium.

non-renewable Sources of energy that cannot be replaced once they are used.

nuclear energy Energy from making and breaking nuclear bonds.

nuclear fusion Atoms that fuse together releasing energy in the form of heat and light.

nucleus The central part of an atom that is made up of protons and neutrons. It also controls the cell's activities.

P

periodic table A table that is an arrangement of elements in order of increasing atomic number arranged in rows called periods and columns called groups.

pH scale A scale, which runs from 0 to 14, which indicates the level of acidity or basicity of a solution.

physical change A change in which no new substance is formed.

plastics Man-made materials made from crude oil.

pollution The addition of harmful or unwanted materials to an environment.

polymerisation The process involving the joining together of many small molecules called monomers to form a large molecule called a polymer.

proton A positively charged subatomic particle found in the nucleus of an atom.

S

salt A compound formed when the hydrogen of an acid is replaced with a metal.

saturated solution A solution that contains as much dissolved solute as possible at that temperature.

soluble Solids that dissolve in a liquid.

solute A substance that dissolves.

solution A mixture of a solute and a solvent.

solvent A liquid in which a solute dissolves.

surface area The measure of how much surface of reactants is exposed.

synthetic materials Made from man-made materials rather than natural materials.

T

temperature A measure of how hot or cold an object is.

time It is measured by a basic unit called the *second* (symbol *s*).

titration A process used to determine the quantity of an acid that is required to exactly neutralise a fixed quantity of base.

V

valency of an element The number of electrons an atom of that element wants to gain, lose or share in order to have a full outer shell (to be chemically stable).

volume A measure of the space taken up by an object.

W

waste management The way we control the amount of unwanted material we produce and how we dispose of it.

weight The force of gravity.

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1.00794																	2 He 4.0026
2	3 Li 6.941	4 Be 9.01218											5 B 10.811	6 C 12.0107	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.1797
3	11 Na 22.9898	12 Mg 24.305											13 Al 26.9815	14 Si 28.0855	15 P 30.9738	16 S 32.065	17 Cl 35.453	18 Ar 39.948
4	19 K 39.0983	20 Ca 40.078	21 Sc 44.9559	22 Ti 47.867	23 V 50.9415	24 Cr 51.9961	25 Mn 54.938	26 Fe 55.845	27 Co 58.9332	28 Ni 58.6934	29 Cu 63.546	30 Zn 65.38	31 Ga 69.723	32 Ge 72.63	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.798
5	37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.224	41 Nb 92.9064	42 Mo 95.96	43 Tc [98]	44 Ru 101.07	45 Rh 102.906	46 Pd 106.42	47 Ag 107.868	48 Cd 112.411	49 In 114.818	50 Sn 118.71	51 Sb 121.76	52 Te 127.6	53 I 126.904	54 Xe 131.293
6	55 Cs 132.905	56 Ba 137.327	57–71	72 Hf 178.49	73 Ta 180.948	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.217	78 Pt 195.084	79 Au 196.967	80 Hg 200.59	81 Tl 204.383	82 Pb 207.2	83 Bi 208.98	84 Po [210]	85 At [210]	86 Rn [222]
7	87 Fr [223]	88 Ra [226]	89–103	104 Rf [267]	105 Db [268]	106 Sg [269]	107 Bh [270]	108 Hs [269]	109 Mt [278]	110 Ds [281]	111 Rg [281]	112 Cn [285]	113 Uut [286]	114 Fl [289]	115 Uup [288]	116 Lv [293]	117 Uus [294]	118 Uuo [294]
			57 La 138.905	58 Ce 140.116	59 Pr 140.908	60 Nd 144.242	61 Pm [145]	62 Sm 150.36	63 Eu 151.964	64 Gd 157.25	65 Tb 158.925	66 Dy 162.5	67 Ho 164.93	68 Er 167.259	69 Tm 168.934	70 Yb 173.054	71 Lu 174.967	
			89 Ac [227]	90 Th 232.038	91 Pa 231.036	92 U 238.029	93 Np [237]	94 Pu [244]	95 Am [243]	96 Cm [247]	97 Bk [247]	98 Cf [251]	99 Es [252]	100 Fm [257]	101 Md [258]	102 No [259]	103 Lr [262]	

s
p
d
f