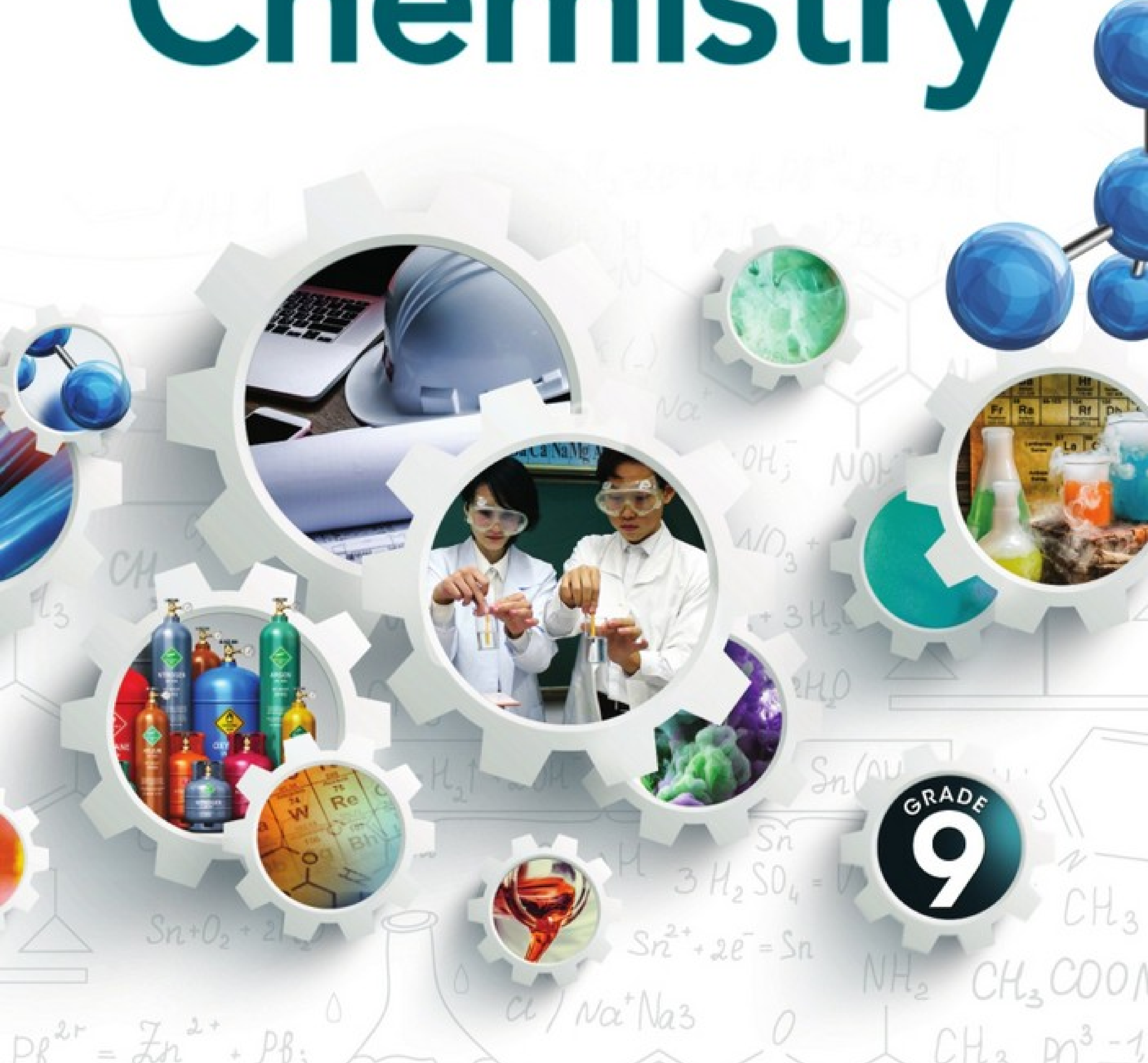
Chemistry



GRADE
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Chemistry

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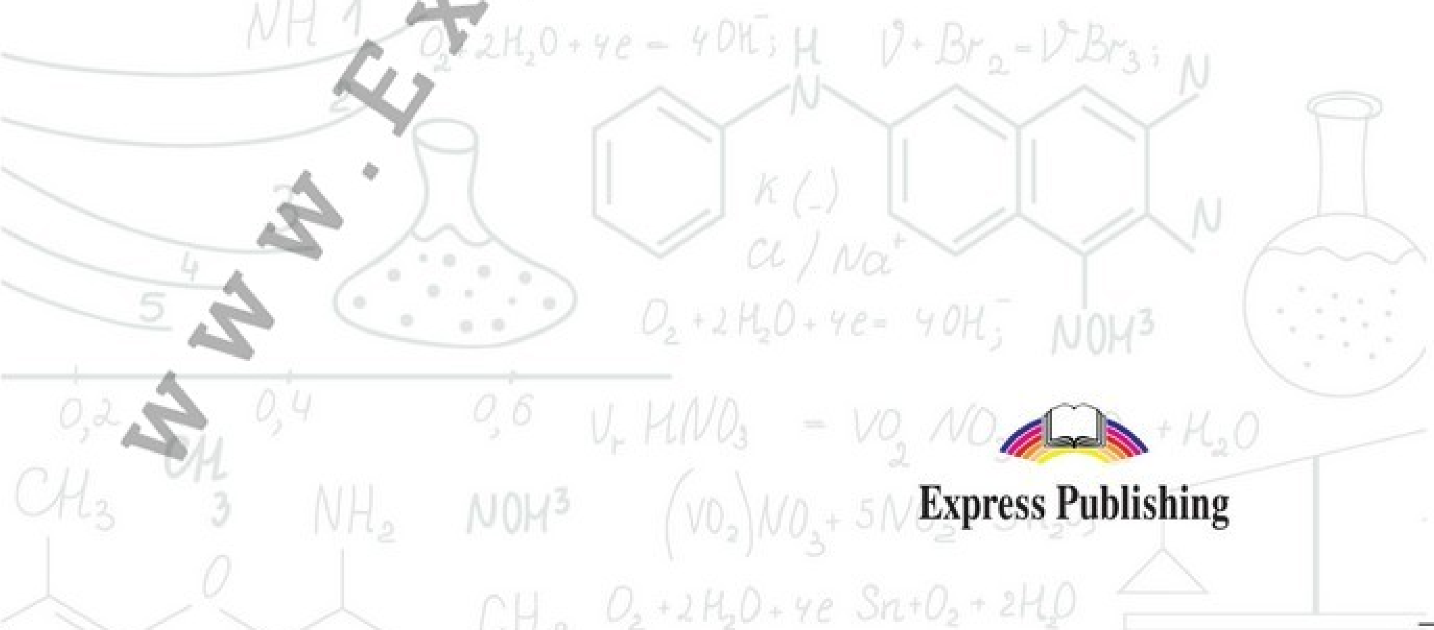

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Contents

Introduction

For the student

Welcome to your new Chemistry textbook, *Grade 9 Chemistry*. Your textbook comes with a **Grade 9 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 9 Chemistry subject programme in Kazakhstan, *Grade 9 Chemistry* aims to give students a sense of enjoyment and an interest in the learning of science. The book is based on the Grade 9 Learning Outcomes 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

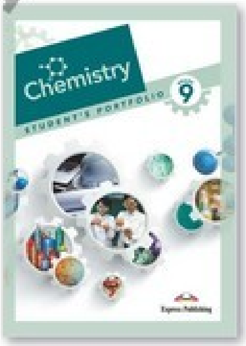
Key fact

Question

Group work

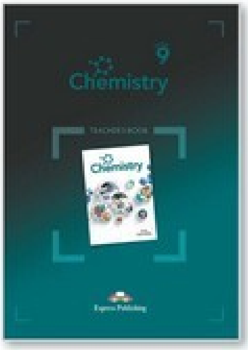
Research

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 Book



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

- An outline of the Grade 9 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

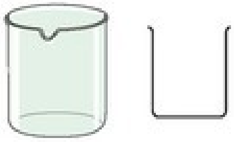
The *Grade 9 Chemistry* **digital resources** will further enhance classroom learning. These resources have been designed to integrate with the Textbook and to complement lessons suggested in the Teacher's Book. Following the principles of the new national Chemistry subject programme, material is provided to suit a range of learner types and to encourage participation and engagement on the part of the student.

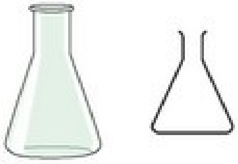
A series of **videos** allows students to observe science in action across all modules. These videos reinforce the topic at hand and allow for other perspectives, which may be discussed in class. Similarly, a series of **videos** about **scientist biographies** presents a lively gateway to develop students'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

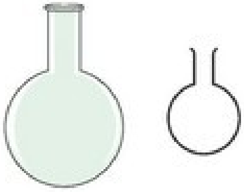
Further classroom discussion and participation is opened up through **PowerPoint presentations**, including a thematic presentation of information from the Textbook. **Experiment videos** allow for a visual review of activities carried out in the classroom. **Extra assessment material** is provided to 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 Book.

Laboratory equipment


Beaker

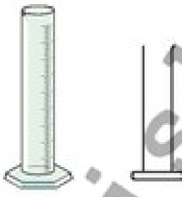

Conical flask

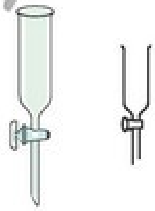

Round-bottomed flask


Test tube


Burette

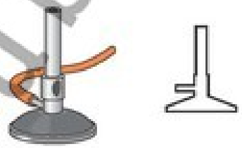

Pipette


Graduated cylinder


Tap funnel


Filter funnel


Evaporation dish


Bunsen burner


Stand


Tripod

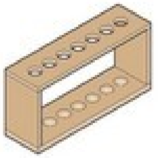

Gauze


Spatula

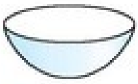

Tongs


Test tube holder

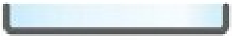

Thermometer


Test tube rack


Balance


Crucible


Pipe clay triangle


Petri dish

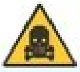
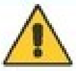
Laboratory safety rules for pupils

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

Electrolytic dissociation

Learning outcomes

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

- Define and give examples of electrolytes and non-electrolytes 9.4.1.1
- Explain the dependence of electrical conductivity on the type of chemical bond 9.4.1.2
- Understand the basic principles of the theory of electrolytic dissociation 9.4.1.3
- Explain the mechanism of electrolytic dissociation of substances with ionic and polar covalent types of bonds in water solution 9.4.1.4
- Distinguish between acidity and alkalinity of solutions 9.4.1.5
- Make equations of electrolytic dissociation of acids, alkalis, normal and acid salts 9.4.1.6
- Define degree of dissociation and distinguish between strong and weak electrolytes 9.4.1.7
- Make equations of reactions in molecular and ionic form 9.2.2.1
- Explain ion exchange reactions 9.2.2.2
- Make equations of reactions for the chemical properties of acids, soluble and insoluble bases and neutral salts in molecular and ionic form 9.3.4.1
- Draw conclusions from experiments about the chemical properties of acids, bases and medium salts 9.3.4.2
- Determine, through experimentation, the medium of the solution of salts 9.3.4.3
- Predict reaction of medium in neutral salt solution; make molecular and ionic equations for hydrolysis 9.3.4.5

Keywords

✓ electrolyte ✓ non-electrolyte ✓ dissociation ✓ polar
✓ non-polar ✓ neutralisation ✓ proton donor ✓ proton acceptor
✓ amphoteric substance ✓ ion exchange ✓ anion ✓ cation
✓ hydrolysis ✓ conjugate ✓ species

Electrolytes and Non-electrolytes

The term electrolyte is given to any salt or ionisable molecule that conducts electricity in solution. When a salt dissolves, for example, the dissociated ions in the salt solution move freely and this allows a charge to flow.

A typical example of an electrolyte solution is when common table salt is placed into water. Sodium chloride will dissolve into its component ions thus:

$$\boxed{NaCl_{(s)} \longrightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}}$$

In this ionic compound the ions in the solid separate and disperse uniformly throughout the solution. This process of physical change is known as dissociation. In this solvate form the strong electrostatic forces between the ions are reduced. Ionic compounds generally dissociate almost completely when they dissolve and for this reason are considered to be strong electrolytes.

Covalent compounds are electrolytes in solutions where their molecules react chemically with the solvent to produce ions. A typical example of this is pure hydrogen chloride HCl. The gas has no ions but when dissolved in water, the solution is a very good conductor. The water molecules allow ions to form in the solution.

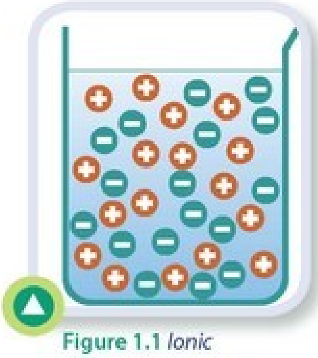


Figure 1.1 *Ionic compound in solution*

The term non-electrolyte refers to compounds that do not ionise in solution and thus will not conduct electricity. Such compounds are typically formed through covalent rather than ionic bonding. These polar and non-polar covalent compounds dissolve in water as molecules instead of ions. Most carbon compounds are non-electrolytes. Fats, sugars, and alcohols are typically non-electrolytes. A common example of a non-electrolyte is glucose (sugar) $C_6H_{12}O_6$. Although it readily dissolves in water, its ions do not dissociate in the solution.

1.1 Look at the question and equipment needed section of **Activity 1.1**. Cover the rest of the activity.

- **(a)** Can you predict how you will conduct the experiment?
- **(b)** Can you predict how you will determine whether a solution is an electrolyte or non-electrolyte?
- **(c)** Can you predict how you will determine which are stronger and weaker electrolytes?

Activity 1.1

Question

How can we show whether a solution is an electrolyte or non-electrolyte?

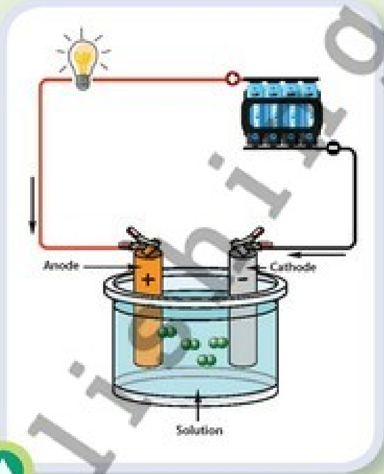


Figure 1.2

Equipment needed

4 x 1.5v batteries
3.6 w bulb
Wires and alligator clips
6 glass beakers
2 x electrodes

Safety

- Take care when pouring solutions as some are corrosive.

Conducting the activity

1. Set up the equipment to test for electrolytes as shown in **Figure 1.2**.
2. Place 20 mL of each solution in a beaker and test each solution by placing the electrode in the beaker.
3. Record observations of changes that occur in the table below.

Solution	Observations
Sodium chloride solution	
Distilled water	
Sugar solution	
Hydrochloric acid solution	
Sulfur dioxide solution	
Acetic acid solution	

4. Wash the electrodes between each test and then dry thoroughly with tissue.

1.2 Did you notice any trends in the chemical formulas of the compounds tested relating to your results of their conductivity?

1.3 Were any ionic substances non-electrolytes?

Electrolytic dissociation theory

In the section that follows, we will consider two well-known theories of electrolytic dissociation.

Arrhenius theory

Most acids commonly used are aqueous solutions where the acid molecules are dissolved in water. Acids usually need water before they can show their acidic properties. The table compares the properties of pure citric acid and citric acid solution.

Test	Pure citric acid	Citric acid solution
Blue litmus paper	Remains blue	Turns red
Sodium hydrogencarbonate	Does not react	Reacts
Magnesium	Does not react	Reacts
Conduction of electricity	Non-conductor	Conductor

In 1887, the Swedish chemist, Arrhenius, developed a theory of acids and bases. The Arrhenius theory helps to explain how water allows acids to show their acidic properties.

It was known at that time that pure acids, such as anhydrous sulfuric acid and glacial ethanoic acid, are non-conductors of electricity, while their aqueous solutions conduct electricity. It was also known that acids contain at least one hydrogen atom per molecule of acid. Arrhenius stated that:

- An acid is a neutral molecule (HX), which dissociates in water to form a hydrogen ion and an anion:

$$\boxed{HX \longrightarrow H^+ + X^-}$$

- The acidic properties of the solution are due to the presence of H_+ ions.
- A base (MOH) dissociates in water to form a hydroxide ion and a cation:

$$\boxed{MOX \longrightarrow OH^- + M^+}$$

- The basic properties of the solution are due to the presence of OH^- ions.
- In solution strong acids and bases are fully dissociated.
- In solution weak acids and bases are only slightly dissociated.

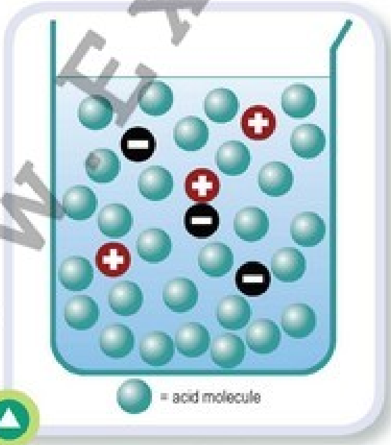


Figure 1.3 *A weak acid in aqueous solution, according to the Arrhenius theory*

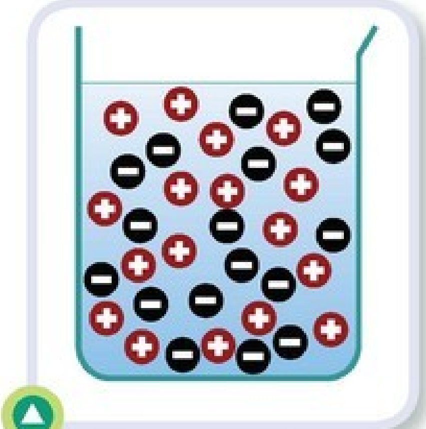


Figure 1.4 *A strong acid in aqueous solution, according to the Arrhenius theory*

An Arrhenius acid is a substance that dissociates in aqueous solution, forming hydrogen ions.

An Arrhenius base is a substance that dissociates in aqueous solution, forming hydroxide ions.

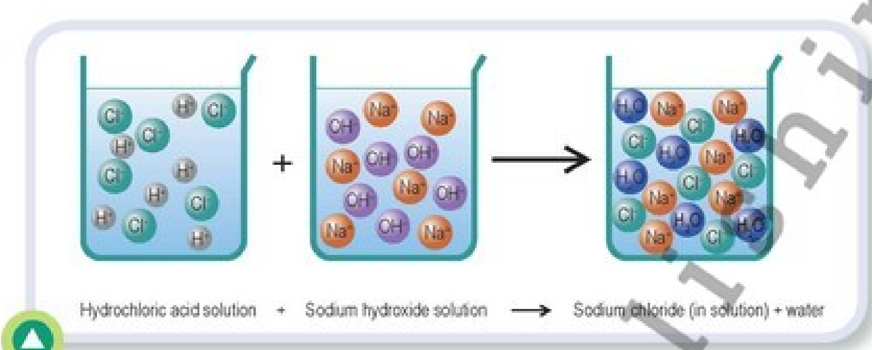


Figure 1.5 *Neutralisation of a solution of hydrochloric acid according to the Arrhenius theory*

According to the Arrhenius theory, neutralisation of a solution of a strong acid such as HCl by a solution of a strong base such as NaOH (**Figure 1.5**), occurs as follows:

- In solution, the strong acid HCl is fully dissociated into H^+ and Cl^- ions, while the strong base NaOH is fully dissociated into Na^+ and OH^- ions.
- When the two solutions are mixed, the H^+ ions react with the OH^- ions:

$$\boxed{H^+ + OH^- \longrightarrow H_2O}$$

- The two other ions present, Na^+ and Cl^-, do not react, as they are more attracted to water molecules than to each other.

Evidence for the Arrhenius theory

- The Arrhenius theory explains why strong acids conduct electricity only when dissolved in water; there are no ions present in a pure acid (**Figure 1.6**).
- The Arrhenius theory also explains why a solution of a weak acid does not conduct electricity as well as a solution of a strong acid of a similar concentration; there are fewer ions present in the weak acid solution.
- For similar reasons, a solution of a weak base does not conduct electricity as well as a solution of a strong base of a similar concentration.
- Acids usually need water before they can show their acidic properties.

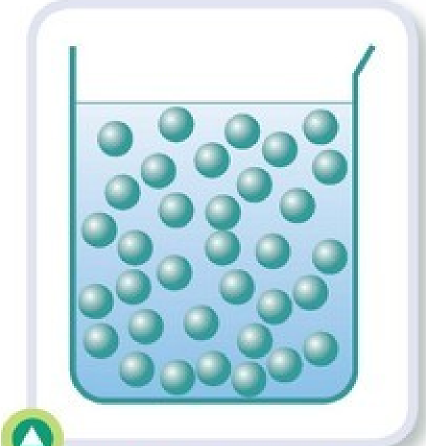


Figure 1.6 *In a pure acid, such as concentrated sulfuric acid, none of the acid molecules dissociate into ions*

Limitations of the Arrhenius theory

- The Arrhenius theory is limited to acid-base reactions that take place in water.
- When an acid dissolves in water, the hydronium ion, H_3O^+, is formed, rather than the hydrogen ion.
- The Arrhenius theory does not explain how certain substances can be amphoteric, i.e. act as an acid with a base **and** as a base with an acid.

Q U4 R1

1.4 What is an acid according to the Arrhenius theory?

1.5 What is the difference between a strong acid and a weak acid according to the Arrhenius theory?

1.6 What causes the basic properties of a solution according to the Arrhenius theory?

1.7 What evidence is there for the Arrhenius theory of acids and bases?

The Bronsted-Lowry Theory of acids and bases

In 1923, Bronsted and Lowry proposed a new and broader theory of acids and bases, the **Bronsted-Lowry theory**. According to this theory:

- An acid is a proton donor.
- A base is a proton acceptor.
- An acid–base reaction involves transfer of a proton (H^+) from the acid to the base.
- The stronger the acid, the more readily it transfers a proton.
- The stronger the base, the more readily it accepts a proton.
- The weaker the acid, the less readily it transfers a proton.
- The weaker the base, the less readily it accepts a proton.

Figure 1.7 *Johannes Bronsted was a Danish chemist who developed a theory of acids and bases in 1923*

Figure 1.8 *Thomas Lowry was an English chemist who independently developed the same theory of acids and bases as Bronsted*

Acid–base reactions, according to the Bronsted–Lowry theory, involve proton transfer. When HCl is added to water, the acid, HCl, transfers a proton to water, which in this case, being a proton acceptor, is a base:

$$\underset{\text{hydrogen chloride}}{HCl_{(g)}} \overset{H^+}{+} H_2O_{(l)} \rightleftharpoons \underset{\text{hydronium ion}}{H_3O^+_{(aq)}} + Cl^-_{(aq)}$$

A Bronsted-Lowry acid is a proton donor.

A Bronsted-Lowry base is a proton acceptor.

Advantages of the Bronsted-Lowry theory

- The Bronsted-Lowry theory shows how the hydronium ion is formed when an acid reacts with water.
- The Arrhenius theory only deals with reactions in water. The Bronsted-Lowry theory broadens the range of reactions that can be regarded as acid–base reactions. For example, ammonia reacts with hydrogen chloride gas in the absence of water.

$$NH_{3(g)} \overset{H^+}{+} HCl_{(g)} \rightleftharpoons NH_4{}^+{}_{(g)} + Cl^-_{(g)} \rightleftharpoons NH_4Cl_{(s)}$$

  Ammonia acts as a Bronsted-Lowry base, because it accepts a proton from the Bronsted-Lowry acid, hydrogen chloride.
- The Bronsted-Lowry theory explains how substances can be amphoteric. For example when water reacts with ammonia

$$NH_{3(g)} \overset{H^+}{+} H_2O_{(l)} \rightleftharpoons NH_4{}^+{}_{(aq)} + OH^-_{(aq)}$$

  it acts as an acid, donating a proton to ammonia.
- When it reacts with nitric acid

$$HNO_{3(aq)} \overset{H^+}{+} H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + NO_3{}^-{}_{(aq)}$$

  it acts as a base, accepting a proton from nitric acid. (Note that the symbol $\rightleftharpoons$ indicates that the reactions are reversible.)

An amphoteric substance is one which can react as an acid with a base and as a base with an acid.

Sample question 1

Identify the Bronsted-Lowry acids in the following reaction:

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

Answer:

When the reaction goes from left to right, H_2O acts as a Bronsted-Lowry acid, as it transfers a proton to NH_3.

When the reaction goes from right to left, NH_4^+ acts as a Bronsted-Lowry acid, as it transfers a proton to OH^-.

Sample question 2

Identify the Bronsted-Lowry bases in the following reaction:

$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$

Answer:

When the reaction goes from left to right, H_2O acts as a Bronsted-Lowry base, as it accepts a proton from CH_3COOH.

When the reaction goes from right to left, CH_3COO^- acts as a Bronsted-Lowry base, as it accepts a proton from H_3O^+.

Q Understanding U4

1.8 What is an acid according to the Bronsted-Lowry theory?

1.9 What is a base according to the Bronsted-Lowry theory?

1.10 What happens in an acid–base reaction according to the Bronsted-Lowry theory?

1.11 What is the difference between a strong acid and a weak acid according to the Bronsted-Lowry theory?

1.12 What is the difference between a strong base and a weak base according to the Bronsted-Lowry theory?

1.13 State one advantage of the Bronsted-Lowry theory compared to the Arrhenius theory.

1.14 Identify the two species acting as an acid in each of the following reactions:

- **(a)** $C_2H_5COOH + H_2O \rightleftharpoons C_2H_5COO^- + H_3O^+$
- **(b)** $HCO_3^- + H_2O \rightleftharpoons H_2CO_3 + OH^-$
- **(c)** $HNO_3 + H_2O \rightleftharpoons H_3O^+ + NO_3^-$
- **(d)** $H_2SO_3 + H_2O \rightleftharpoons HSO_3^- + H_3O^+$

1.15 Identify the two species acting as a base in each of the following reactions:

- **(a)** $HCOOH + H_2O \rightleftharpoons HCOO^- + H_3O^+$
- **(b)** $NH_3 + H_2O \rightleftharpoons NH_3^+ + OH^-$
- **(c)** $NH_3 + CH_3COOH \rightleftharpoons NH_4^+ + CH_3COO^-$

The pH scale

The pH scale was devised by Sorensen, a Danish chemist, in 1909.

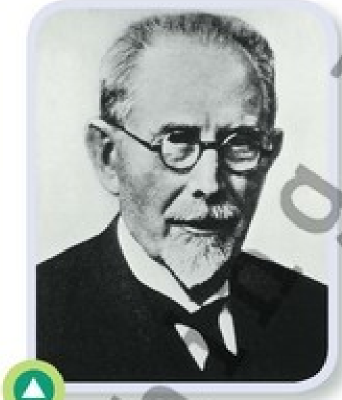


Figure 1.9 *Soren Sorensen*

The pH of a solution is defined as the negative logarithm to the base 10 of the hydrogen ion concentration.

- Since H^+ exists as H_3O^+ in solution, $pH = -\log_{10} [H_3O^+]$
- The square brackets denote concentration in mol l^{-1}.
- At 298 K, the pH of pure water $= -\log_{10} [H_3O^+]$
  $= -\log_{10} (10^{-7}) = -(-7) = 7.$

Sample question 3

Calculate the pH of a solution with a H_3O^+ concentration of 0.01 M.

Answer:

$$pH = -\log_{10} [H_3O^+]$$
$$= -\log_{10} (0.01)$$
$$= 2$$

Sample question 4

Calculate the H_3O^+ concentration of a solution with a pH of 3.

Answer:

$$pH = -\log_{10} [H_3O^+] = 3$$
$$-pH = -\log_{10} [H_3O^+] = -3$$
$$[H_3O^+] = \text{antilog} (-3)$$
$$= 0.001 \text{ M}$$

Measuring pH

Universal indicator solution, or universal indicator paper, can be used to find the pH of a solution. Universal indicator is a mixture of indicators that changes colour as the pH is changed. The pH of a solution is measured after adding universal indicator by comparing the colour of the solution with a colour chart.

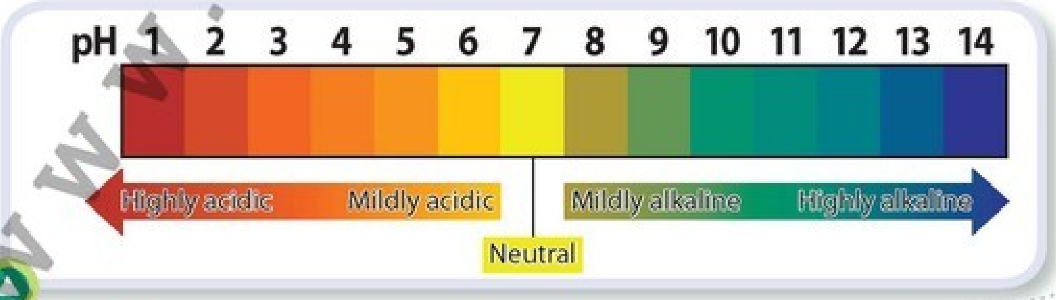


Figure 1.10

Lab Work 1

Activity 1.2

Question

How can we show that a difference of 1 in the pH number of a solution relates to a difference in concentration by a factor of 10?

Equipment needed

test tube racks
dropping pipettes x 2
beakers x 2
measuring cylinders x 2
test tubes x 13
distilled water
universal indicator
0.1 M sodium hydroxide
0.1 M hydrochloric acid

Safety

- Be careful when pouring not to spill solutions as they can be corrosive.

Conducting the activity

Group A

1. Take 7 tests tubes and label them 1-7.
2. Pour 10 cm^3 of **distilled water** into test tube 7 and place in a test tube rack.
3. Half-fill test tube 1 with hydrochloric acid solution.
4. Pour 1 cm^3 of the hydrochloric acid solution into a measuring cylinder and add distilled water to the 10 cm^3 mark. Place some of this solution in test tube 2. Fill to the same level of solution as in test tube 1.
5. Pour the remaining solution in the beaker away, except for 1 cm^3. Again, add distilled water to the 10 cm^3 mark. Add this solution to test tube 3 [same level as test tube 1 and 2] and with the remaining solution in the beaker repeat the exact process of dilution. Continue like this until test tubes 4, 5 and 6 contain solutions.

Group B

1. Number test tubes 8-13.
2. Half-fill test tube 8 with sodium hydroxide solution.
3. Follow the exact process of dilution [as group A] for the solution each time for test tubes 9-13.

Results

Place test tubes 1-13 in order in test tube racks. The test tubes now contain solutions with pH 1-13. Using a pipette, add a drop of universal indicator to each test tube. Gently move each test tube to allow the contents to mix. Add another drop of universal indicator to each test tube; should it be necessary to distinguish colours more clearly use a pH indicator chart to compare your range of colours.

1.16 What was the significance of diluting the solution by the power of 10 each time?
1.17 What factors might make it difficult to get an exact colour chart match for test tube 7?

Ion exchange reactions

Ion exchange reactions are reversible reactions involving two substances, each consisting of positively and negatively charged ions, which exchange one or more of their ionic components. This type of reaction is typically represented by the following generic equation:

$$\boxed{AX + BY \longrightarrow AY + BX}$$

In aqueous solution, the ions in ionic substances are much freer to move around the solution than within the more rigid structure of a crystal. The type of double displacement reaction indicated in the equation above in thus common when mixing two aqueous solutions.

Certain insoluble materials also have positive or negative charges on their surfaces and react with ionic solutions to remove ions and replace them with different ions. These types of reaction are typically used for the softening of water, the separation of substances or the purification of chemicals.

Ion exchange reactions involve three different types of reaction:

- precipitation reactions which result in the formation of an insoluble solid
- gas forming reactions which result in the formation of a gas that escapes from the solution
- acid-base reactions – also known as neutralisation reactions- which typically result in the formation of a salt and water

Q U2

1.18 Can you predict which of these reactions will be gas forming reactions?

(a) $NaCL + NH_4NO_3$

(b) $CaCO_3 + HCL$

(c) $Na_2SO_4 + HCl$

(d) $NaOH + AgNO_3$

R2 R3 R4 R5 **Practical Work 1**

Activity 1.3

Question

Can we experimentally identify the type of reaction that occurs when two aqueous solutions are mixed through visual observation and temperature changes?

Equipment needed

Test tube rack

11 x test tubes

0.1 ml of: $NaCl$, NH_4NO_3, Na_2SO_4, $Ba(NO_3)_2$, $NaOH$, $AgNO_3$, HCL

0.5 g of $Mg(OH)_2$ and $CaCO_3$

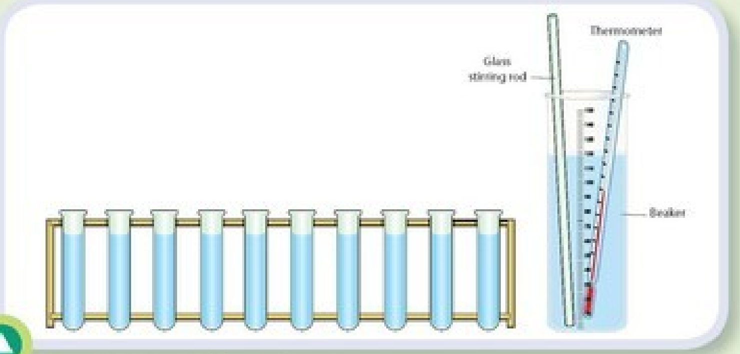


Figure 1.11 *Ion Exchange Reactions*

25.0 ml of 1.0 M HCL
150 ml Beaker
Deionised water
Glass stirring rod
Thermometer

Procedure

In 1-5 below add the two substances from the marked test tubes into the mixing test tube, stir with glass rod and observe.

Record your observations in the table below

Dispose of contents appropriately and rinse mixing test tube and glass rod after each experiment.

1. Add 0.1 ml of deionised water and 0,1ml each of NaCl and NH_4NO_3 solution
2. Add 0.1 ml of deionised water and 0,1ml each of Na_2SO_4 and $Ba(NO_3)_2$ solution
3. Add 0.1 ml of deionised water and 0,1ml each of NaOH + $AgNO_3$ solution
4. In a dry test tube, add a small amount of $CaCO_3$ (size of a pencil rubber), 0.1 ml of deionised water and 0.1 ml HCl.
5. Add 0.5g of $Mg(OH)_2$ to the beaker. Then add 25,0 ml of deionised water. Swirl the mixture. To this mixture now add a mixture of 25.0 ml of 1.0 M HCL.

	Observations	Type of ion exchange reaction
1		
2		
3		
4		
5		

Q Understanding U2 Research R5

1.19 How can you explain your observations in stage 5?

1.20 You are presented with two clear aqueous solutions. One is sodium carbonate and the other is sodium nitrate. You do not know which is which. Can you think of a test that will help you distinguish them?

Ion exchange reactions have many laboratory, medical and practical applications. They are used, for example, in a wide range of diagnostic clinical tests on blood and in dialysis where ion exchange filtering performs the functions of failing kidneys.

Did you know?

Indigestion is caused by too much acid in the stomach. The discomfort can be cured by taking stomach powders such as magnesium hydroxide or sodium hydrogencarbonate, which are bases.

Perhaps one of the best known uses of ion exchange, however, is in the treatment of hard water described below.

The chlorides and sulfates of calcium and magnesium cause permanent hardness in water.

Permanent hardness may be removed by ion exchange and by distillation. Unlike temporary hardness, it cannot be removed by boiling.

Ion exchange involves the water being passed through a cation-exchange resin. The resins used are usually complex sodium compounds, represented here as RNa. The calcium and magnesium ions in the hard water swap places with the sodium ions in the resin.

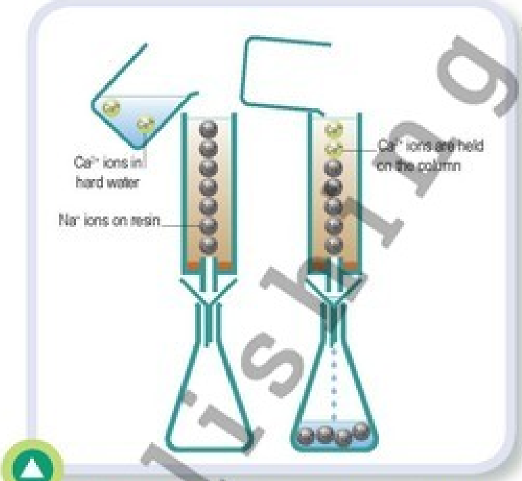


Figure 1.12 *Removing hardness using an ion exchange resin*

$$\boxed{2RNa_{(s)} + Ca^{++}_{(aq)} \longrightarrow R_2Ca_{(s)} + 2Na^+_{(aq)}}$$

In this way, the hardness causing calcium and magnesium ions are removed from the water and replaced by sodium ions, which do not cause hardness. The anions are unaffected by this process.

Deionisation

More advanced ion exchange resins remove all ions from the water. The resin used is a mixture of a cation exchanger, which replaces metal ions in water by hydrogen ions, H^+, and an anion exchanger, which replaces anions in water by hydroxide ions, OH^-. The H^+ and OH^- ions then react with each other as follows:

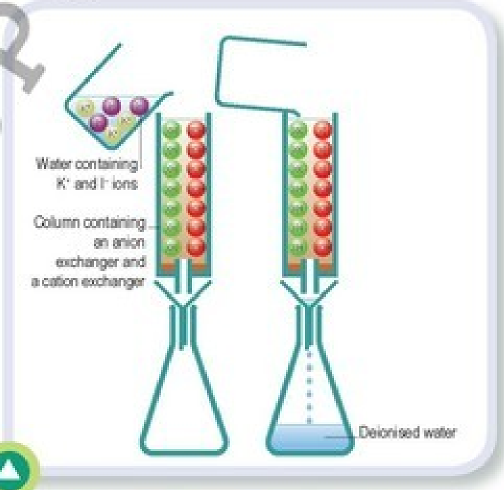


Figure 1.13 *When water is deionised, metal ions and anions are removed from the water*

$$\boxed{H^+_{(aq)} + OH^- \longrightarrow H_2O}$$

The effect of the resin is that dissolved salts are replaced with the water molecules formed by the combination of hydrogen and hydroxide ions. The water is said to have been **deionised**.

Formulae of molecular and ionic equations

Molecular equations

Molecules consist of two or more atoms bonded by covalent bonds. The relationship between these molecules can be expressed through two kinds of formula: their empirical formula or their molecular formula.

The empirical formula represents the most simple ratio of atoms in the molecule. For example:

benzene CH water H_2O

The molecular formula, however, expresses not only the ratio of the atoms in the molecule but also their exact number. Compare the two types of formula in the table below.

	Empirical formula	Molecular formula
benzene	CH	C_6H_6
ethane	CH_3	C_2H_6
water	H_2O	H_2O
acetic acid	CH_2O	$C_2H_4O_2$
glucose	CH_2O	$C_6H_{12}O_6$

We can work out the formula of an ionic compound when we know the number of ions in it.

In ionic compounds the atoms or molecules bond together through the mutual attraction of negatively charged 'cations' and positively charged 'anions'. Cations can attract more than one anion and anions can attract more than one cation.

For example, the compound formed from Na^+ and SO_4^{2-} will consist of two Na^+ ions to every one SO_4^{2-} ion so that the compound is neutral overall. The formula is therefore Na_2SO_4.

The table below shows the formula for common positive ions and negative ions.

Positive ions (cations)

Name	Formula	Name	Formula	Name	Formula
potassium	K^+	ammonium	NH_4^+	copper (II)	Cu^{2+}
hydrogen	H^+	calcium	Ca^{2+}	lead	Pb^{2+}
aluminium	Al^{3+}	barium	Ba^{2+}	iron (II)	Fe^{3+}
sodium	Na^+	silver	Ag^+	iron (III)	Fe^+
zinc	Zn^{2+}	lithium	Li^+	magnesium	Mg^{2+}

Negative ions (anions)

Name	Formula	Name	Formula	Name	Formula
fluoride	F^-	nitrate	NO_3^-	hydrogen carbonate	HCO_3^-
chloride	Cl^-	oxide	O^{2-}	hydroxide	OH^-
bromide	Br^-	phosphate	PO_4^{3-}	sulfide	S^{2-}
Iodide	I^-	carbonate	CO_3^{-2}	sulfate	SO_4^{2-}

U2 Understanding

1.21 What is the difference between the empirical and molecular formula?

1.22 The molecular formula for the compound is C_8H_{18}. What is the empirical formula?

1.23 What is the ionic formula for the following compounds:

- **(a)** magnesium sulfate
- **(b)** silver nitrate
- **(c)** iron (II) sulfate
- **(d)** potassium hydroxide
- **(e)** sodium fluoride
- **(f)** calcium hydroxide

Hydrolysis of salts

An acid reacts with a base to give water and "salt" as products in a process called neutralisation. The term used to describe the reverse of neutralisation – reaction of a salt with water – is known as hydrolysis. Salts are typically weakly acidic to weakly basic on the pH spectrum.

> Acidic salts are formed from the reaction of a weak base with a strong acid.
> Basic salts are formed from the reaction of a strong base and a weak acid.

We know in Brønsted-Lowry theory, that acids are proton donors and bases are proton acceptors. In hydrolysis reactions an acid can donate a proton to water to form aqueous H_3O^+ and the conjugate base; a base can accept a proton from water to form OH^- and the conjugate acid. Anions that are the conjugate bases of weak acids react with water to form OH^-. Cations that are the conjugate acids of weak bases can undergo a proton transfer reaction with water to form H_3O^+.

Salts from strong acids and strong bases do not hydrolyse. The ions in these salt compounds e.g. NaCl do not react with water, so soluble salt solutions of such compounds will give a neutral pH reading.

R2 R3 R4 R5 Research

Lab Work 2

Activity 1.4

Question

How can we determine whether salt solutions in hydrolysis reactions are acidic, basic or neutral?

Equipment needed

Beaker
Test tube
Litmus paper
Deionised water
(100 mL) solution of following salts:
NaCl, $NaHCO_3$, LiCl, Na_3PO_4, $NaC_2H_3O_2$, NaI, KNO_3, KCl

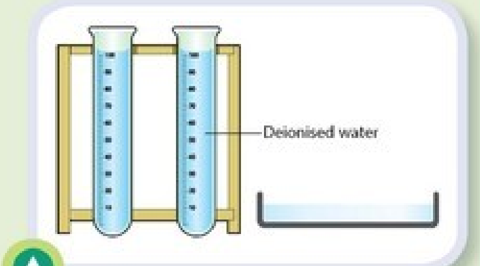


Figure 1.14 *Testing pH of salt solutions*

Procedure

1. Place a small amount of each salt into a test tube with 100 mL of deionised water.
2. Test each solution with litmus paper and record your results in the table in your student portfolio.
3. After each test, dispose of solutions down a drain and rinse the test tube thoroughly.

1.24 Predict the pH of the solutions of these salts (>7, =7, <7). Write the formula of each one first

- **(a)** potassium sulfate
- **(b)** ammonium chloride
- **(c)** sodium hydrogen carbonate
- **(d)** sodium hydroxide
- **(d)** sodium chloride
- **(f)** copper sulfate

Activity 1.5

Question

Can we predict the relative acidity and basicity of salt solutions in hydrolysis reactions and verify predictions through experiment?

Equipment needed

six test tubes
measuring cylinder
pipette
potassium sulfate
ammonium chloride
sodium hydrogen carbonate
sodium hydroxide
sodium chloride
copper sulfate
distilled water
universal indicator
pH meter

Safety

- Be careful when pouring not to spill solutions as they can be corrosive.

Conducting the activity

1. Label six test tubes in a test tube rack with the name of one of the salts.
2. Place a quantity of roughly 0.1 g of each salt in the labelled test tube and 10.0 mL of distilled water in each test tube.
3. Test the pH using both universal indicator – 3-4 drops from a pipette – and a pH meter. Clean the electrode with distilled water between each reading.
4. Record your results in the table.

	predicted pH (>7, =7, <7)	recorded pH
potassium sulfate		
ammonium chloride		
sodium hydrogen carbonate		
sodium hydroxide		
sodium chloride		
copper sulfate		

1.25 How did your predictions compare to the results? Discuss any differences with members of other groups.

1.26 Discuss in groups the acid and base from which each salt is derived and whether they are considered 'strong' or 'weak'.

1.27 State two chemical properties of (i) an acid (ii) a base.

1.28 Explain what is meant by neutralisation.

1.29 Explain how neutralisation is useful in treating indigestion.

1.30 What is the difference between a strong acid and a weak acid in Arrhenius theory?

1.31 Why is Bronsted-Lowry theory considered to be an advance on the Arrhenius theory?

1.32 Write the formulae for the following compounds:

- **(a)** sodium nitrate
- **(b)** iron (II) hydroxide
- **(c)** calcium chloride
- **(d)** ammonium chloride.

1.33 In each of the following reactions indicate which species are acting as acids and which are acting as bases:

$SO_3^{-2} + HCN \rightleftharpoons HSO_3^- + CN^-$

$H_2SO_4 + NH_3 \rightleftharpoons HSO_4^- + NH_4^+$

1.34 Write down one conjugate acid–base pair from each of the reactions.

Conjugate acids and bases

In the reaction of ammonia with water

$$\boxed{NH_3 + H_2 \rightleftharpoons NH_4^+ + OH^-}$$

ammonia is a base, and accepts a proton to form the ammonium ion, NH_4^+.

Water acts as an acid, donating a proton to ammonia; the hydroxide ion, OH^-, is formed. Because this is a reversible reaction, the ammonium ion, NH_4^+, is an acid, since it can donate a proton. The hydroxide ion, OH^-, is a base, since it can accept a proton. The acids and bases in the above reaction are related as follows:

$$NH_3 \underset{-H^+}{\overset{+H^+}{\rightleftharpoons}} NH_4^+$$

$$H_2O \underset{+H^+}{\overset{-H^+}{\rightleftharpoons}} OH^-$$

Species such as these **that differ by the presence or absence of a proton** are called **conjugate acid–base pairs**. H_2O and OH^- form one conjugate pair, while NH_3 and NH_4^+ form another. NH_4^+ is called the **conjugate acid** of NH_3, while OH^- is the **conjugate base** of H_2O.

Sample question 5

What is the conjugate acid of (a) NH_3 (b) O^{2-}?

Answer:

(a) The conjugate acid of NH_3 is formed when a proton (H^+) is added to it. Therefore, the conjugate acid of NH_3 is NH_4^+.

(b) The conjugate acid of O^{2-} is formed when a proton (H^+) is added to it. Therefore, the conjugate acid of O^{2-} is OH^-.

A conjugate base is formed when a proton (H^+) is removed from a Bronsted-Lowry acid.

Sample question 6

What is the conjugate base of **(a)** NH_3 **(b)** HNO_3?

Answer:

(a) The conjugate base of NH_3 is formed when a proton (H^+) is removed from it. Therefore, the conjugate base of NH_3 is NH^{2-}.

(b) The conjugate base of HNO_3 is formed when a proton (H^+) is removed from it. Therefore, the conjugate base of HNO_3 is NO_3^-.

A conjugate acid–base pair is an acid and base that differ by the presence or absence of a protonacid.

Sample question 7

In the reaction of ammonia with hydrogen chloride:

$$NH_3 + HCl \rightleftharpoons NH_4^+ + Cl^-$$

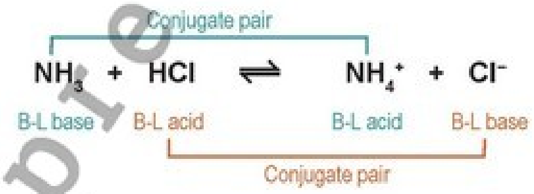
identify the conjugate acid–base pairs.

Answer:

When NH_3 accepts a proton from HCl, it forms its conjugate acid NH_4^+.

When HCl donates a proton to NH_3, it forms its conjugate base Cl^-.

The conjugate acid–base pairs are **(a)** NH_3 and NH_4^+ and **(b)** HCl and Cl^-.



Relative strengths of acids and bases

In the reaction

$$HNO_3 + H_2O \rightleftharpoons H_3O^+ + NO_3^-$$

the product H_3O^+ is acidic and the product NO_3^- is basic. An acid-base reaction such as this tends to lead preferentially to the formation of the weaker acid and base. HNO_3 is a stronger acid than H_3O^+, while H_2O is a stronger base than NO_3^-. As a result, the reaction goes mainly to the right.

In the reaction

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

the reaction goes mainly to the left. This is because CH_3COOH is a weaker acid than H_3O^+, and H_2O is a weaker base than CH_3COO^-.

$$\boxed{\underset{\text{weaker acid}}{CH_3COOH} + \underset{\text{weaker base}}{H_2O} \rightleftharpoons \underset{\text{stronger base}}{CH_3COO^-} + \underset{\text{stronger acid}}{H_3O^+}}$$

When a strong acid, such as HNO_3, donates a proton, the conjugate base that is formed has little tendency to accept a proton and form HNO_3 again, and so is a weak base. The opposite is true in the case of a weak acid, such as CH_3COOH. When it donates a proton, forming its conjugate base, CH_3COO^-, this has a much greater tendency to accept a proton and form CH_3COOH again. Because of this, CH_3COO^- is a relatively strong base. In general, **the stronger an acid, the weaker its conjugate base, and the stronger a base, the weaker its conjugate acid.**

1.35 What is a conjugate acid-base pair?

1.36 What is the conjugate acid of

(a) H_2O **(b)** OH^- **(c)** HS^- **(d)** CH_3COO^- **(e)** HCO_3^- **(f)** Cl^-

1.37 What is the conjugate base of **(a)** H_2O **(b)** CH_3COOH **(c)** HCl **(d)** H_2SO_4

1.38 Identify the conjugate acid–base pairs in each of the following reactions:

(a) $CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$

(b) $HCOOH + H_2O \rightleftharpoons HCOO^- + H_3O^+$

(c) $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

(d) $HNO_3 + H_2O \rightleftharpoons H_3O^+ + NO_3^-$

MODULE 2

Qualitative tests

Learning outcomes

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

- Describe and carry out flame tests to identify metal cations: Li^+, Na^+, K^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, Cu^{2+} 9.4.1.8
- Carry out qualitative analysis to test for cations: Fe^{2+}, Fe^{3+}, Cu^{2+} 9.4.1.9
- Carry out a range of anion tests for chloride, bromide, iodide, sulfate, carbonate, orthophosphate, nitrate, silicate-ions 9.4.1.10
- Devise and carry out an experiment to identify the cations and anions of unknown substances 9.4.1.11
- Perform calculations involving the excess of one reagent, when initial quantities and the balanced chemical equation are known 9.2.3.1

Keywords

✓ corrosive ✓ cross-contamination ✓ transition metals
✓ precipitates ✓ reagent ✓ excess ✓ limiting reactant
✓ non-organic ✓ cation ✓ anion ✓ aqueous solution

Qualitative analysis of inorganic compounds

In this module, we shall work almost exclusively in the lab, planning and carrying out analysis of a range of non-organic compounds. Lab safety procedures should be strictly adhered to at all times.

Flame tests

When salts of the metals lithium, barium, potassium, copper, strontium and sodium are heated in the flame of a Bunsen burner, colours characteristic of the particular element are given off. The test procedure to be followed in **Activity 2.1**, involves introducing a sample of the salt using a platinum wire to a non-luminous flame and observing the colour of the flame that results.

Lab Work 3

Activity 2.1

Question

How does the colour of light of Li^+, Na^+, K^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, Cu^{2+} vary when they are held in a non-luminous flame?

Equipment needed

Sodium chloride	Platinum (or nichrome) wire held in glass rod
Potassium chloride	Bunsen burner
Barium chloride	Six small beakers
Strontium chloride	Labels
Copper(II) chloride	Test tubes
Concentrated hydrochloric acid	Pestle and mortar

Safety

- Hold the platinum wire in a glass rod when heating over the flame. Show caution when handling substances as some are corrosive.

Conducting the activity

1. Clean the platinum wire using concentrated hydrochloric acid in a test tube. Do this in the fume cupboard.
2. Crush the salt to be tested with a pestle and mortar, and transfer it to a labelled beaker.
3. Dip the platinum wire in concentrated hydrochloric acid and then in the salt to be tested.
4. Place the platinum wire in the blue flame of the Bunsen burner as in the diagram, and note the colour given off.
5. Repeat the experiment for each of the other salts. Again, note the colour in each case.

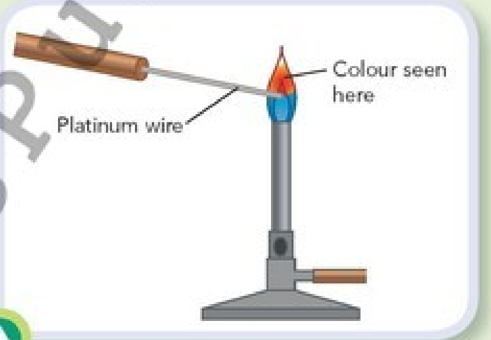


Figure 2.1

Results

Metal	Salt being tested	Flame colour
Barium		
Copper		
Lithium		
Potassium		
Sodium		
Strontium		

2.1 Cross-contamination could be a problem in this experiment. What can you do to avoid cross-contamination?

2.2 When a sample of an unknown metal salt was heated in a flame, a blue–green colour was observed. What metal was present in the salt?

2.3 Why do metal salts give off a characteristic colour in a flame?

2.4 You are given a sample of an unknown metal salt. Explain how you would find out what metal is present in the salt.

Tests for transition metal ions

Transition metals form coloured compounds with other elements. Most of these compounds are soluble in water and form coloured solutions. Mixing these solutions with a sodium hydroxide solution gives a transition metal hydroxide. As these hydroxides are insoluble, they form different coloured precipitates in these reactions and this provides the basis for determining which ions are present.

Lab Work 4

Activity 2.2

Question

How can we test whether an unknown substance contains Cu^{2+} Fe^{2+} or Fe^{3+} ions?

Equipment needed

Test tubes
Beaker containing sodium hydroxide solution
Test tube rack
Iron (II) nitrate solution
Iron (III) nitrate solution
Copper sulfate solution

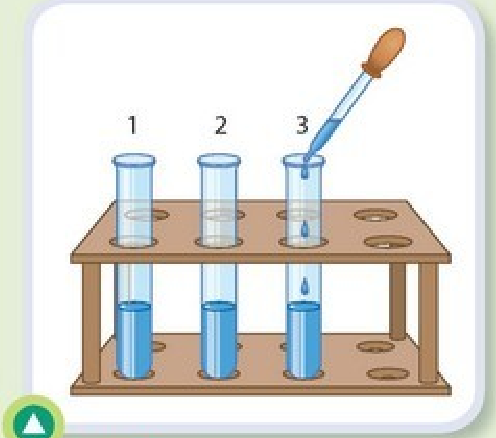


Figure 2.2

Safety

- Be careful when handling sodium hydroxide as it is highly corrosive.

Conducting the activity

1. You will start with three solutions which are in three unmarked test tubes.
2. Label the test tubes 1, 2 and 3.
3. These three test tubes contain iron (II) nitrate solution iron (III) nitrate solution copper sulfate solution. You need to determine which is which.
4. Using a pipette, add a few drops of sodium hydroxide.
5. Note your observations for each test tube and note the number of the test tube that corresponds to the colour description in the following table.

Metal ion	Colour
Copper(II), Cu^{2+} [substance no. __________]	Blue
Iron(II), Fe^{2+} [substance no.__________]	Green - then orange-brown after time
Iron(III), Fe^{3+} [substance no. __________]	Orange-brown

6. Record your observations in the Student Portfolio.

Research Understanding

Q R5 U4

2.5 What can you conclude from your observations?

2.6 Do you think you would have obtained similar results if you had used different salts of these metals e,g, iron (II) nitrate or copper chloride?

2.7 Complete the equation for the reaction of $CuSO_{4(aq)} + 2NaOH_{(aq)} \rightarrow$ ________________

Anion tests

Positively charged ions, cations, in ionic substances may be identified by carrying out flame tests as in **Activity 2.1**. Negatively charged ions, anions, may be identified by tests carried out on their aqueous solutions as in **Activity 2.3**.

Easily identifiable results, which may include the production of characteristically coloured precipitates, are obtained when the aqueous solutions react with certain reagents.

All of the tests described in **Activity 2.3** follow similar procedures using the same equipment. Rinse all equipment well between tests to avoid cross-contamination.

Research Research Research

R2 R3 R4

Lab Work 5

Portfolio 21

Activity 2.3

Question

How can we identify negatively charged ions in aqueous solutions?

Equipment needed

Test tube rack
Test tube holder
Test tubes
Stoppers for test tubes fitted with plastic delivery tubing
Labels
Droppers
Beakers
Wash bottle
Dilute solutions of the salts being tested
Deionised water
Bunsen burner, tripod and gauze
Thermometer

Safety

- Take care when pouring all solutions and handling beakers, pipettes and test tubes containing solutions as some may be corrosive.

Conducting the activity

Test A: Testing for sulfite and sulfate anions

1. Add 2 cm^3 of a sulfite solution to a clean test tube and add some barium chloride solution using a dropper. Record what happens.
2. Add a few cm^3 of hydrochloric acid and mix gently. Record what happens.
3. Repeat steps 1 and 2, using a sulfate solution instead of a sulfite solution.
4. Record your results in the Student Portfolio.

Figure 2.3

2.8 What can you conclude from your observations?

2.9 Complete the equations for the reactions:

(a) $Na_2SO_{3(aq)} + BaCl_{2(aq)} \longrightarrow 2NaCl_{(aq)} +$ ________

$Na_2SO_{4(aq)} + BaCl_{2(aq)} \longrightarrow 2NaCl_{(aq)} +$ ________

(b) $BaSO_{3(s)} + 2HCl_{(aq)} \longrightarrow BaCl_{2(aq)} +$ ________

Test B: Test for the chloride (Cl^-) anion

1. Add 2 cm^3 of chloride solution to a clean test tube.
2. Add a few drops of silver nitrate solution using a dropper. Record what happens.
3. Add a few cm^3 of dilute ammonia solution.
4. Record your results in the Student Portfolio.

2.10 What can you conclude from your observations?

2.11 Complete the equations for the reaction: $NaCl_{(aq)} + AgNO_{3(aq)} \longrightarrow$ ________________

Test C: Test for the nitrate NO_3 anion

1. Add 2 cm^3 of potassium nitrate solution to a clean test tube.
2. Add 3 cm^3 of cold saturated iron (II) sulfate solution using a dropper. Record what happens.
3. Carefully add 2 cm^3 of concentrated sulfuric acid slowly down the wall of the test tube using a dropper. Do not mix the contents of the test tube.
4. Record what happens after a few minutes in your Student Portfolio.

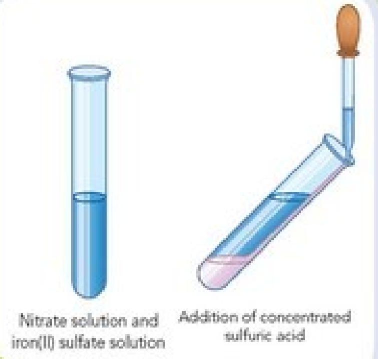


Figure 2.5

2.12 What can you conclude from your observations?

Test D: Test for the phosphate PO_4^{3-} anion

1. Add 2 cm^3 of disodium hydrogen phosphate (V) solution to a clean test tube.
2. Add about 6 cm^3 of the clear ammonium molybdate reagent to the test tube using a dropper.
3. Add a few drops of concentrated nitric acid.
4. Warm gently by placing in a beaker of water at a temperature not exceeding 40°C. Record what happens.
5. Add an equal volume of ammonia solution to the contents of the test tube at the end of step 3.
6. Record your observations in the Student Portfolio.

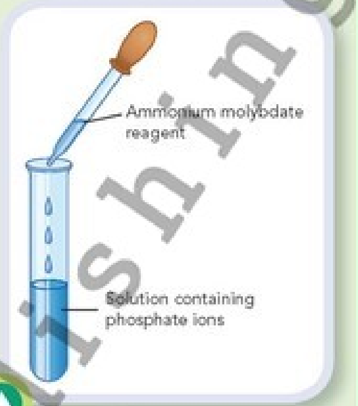


Figure 2.6

2.13 What can you conclude from your observations?

Test E: Test for the carbonate (CO_3^{2-}) and hydrogencarbonate (HCO_3^-) anions

1. Add 2 cm^3 of a carbonate solution to a test tube.
2. Add some dilute hydrochloric acid using a dropper. Record what happens.
3. Repeat steps 1 and 2 using the arrangement shown in the diagram above. Record what happens.
4. Add a few cm^3 of magnesium sulfate solution to some fresh carbonate solution in a clean test tube.
5. Record your results.
6. Repeat steps 1 to 4, using a hydrogen carbonate solution instead of a carbonate solution.

Figure 2.7

2.14 What conclusions can you draw from your observations?

2.15 Complete the equations for the reactions in the different phases of the test:

(a) $Na_2SO_{3(aq)} + 2HCl_{(aq)} \longrightarrow$ ______________

$NaHCO_{3(aq)} + HCl_{(aq)} \longrightarrow$ ______________

(b) $Na_2CO_{3(aq)} + MgSO_{4(aq)} \longrightarrow$ ______________

(c) $2NaHCO_{3(aq)} + MgSO_{4(aq)} \longrightarrow$ ______________

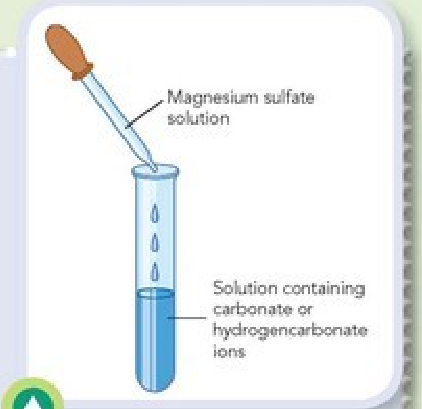


Figure 2.8

Practical Work 2

Activity 2.4

Question

How can we plan a series of tests to identify the ions in an unknown single ionic compound?

Equipment needed

All equipment and chemicals used in Activities 2.1, 2.2 and 2.3 are available to you.

Safety

- Take care when pouring all solutions and handling beakers, pipettes and test tubes containing solutions as some may be corrosive.

Conducting the activity

1. You will be given an 'unknown' salt solution.
2. Plan a series of tests to determine the cation and anion in the unknown compound.
3. Show the stages of your plan to your teacher.
4. After discussing amendments to your plan, carry out the range of tests to identify your unknown substance.
5. Record all observations for each test.


Figure 2.9 *unknown salt*

2.16 Were you able to identify the unknown salt?

2.17 Are there any steps in your plan you would have done differently or in a different order?

Calculations involving excess of one reagent

In chemical reactions it often occurs that one of the reactants is present in excess. We can calculate which reactant is present in excess if the balanced chemical equation is known, together with the initial quantities of all reactants. In such calculations, the substance that is not present in excess is known as the limiting reactant. It is the amount of this substance that dictates how much of each of the products is formed.

Sample question 1

Sodium sulfite reacts with sulfuric acid as follows:

$$Na_2SO_{3(s)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + SO_{2(g)} + H_2O_{(l)}$$

(a) If a solution containing 4.9 g sulfuric acid is added to 50.4 g sodium sulfite, show that the sodium sulfite is present in excess.

(b) Calculate the mass of sodium sulfate and the volume of sulfur dioxide (measured at s.t.p.) formed.

Answer:

(a) Moles of sulfuric acid present initially = 4.9/98 = 0.05

$Na_2SO_{3(s)}$ +	$H_2SO_{4(aq)}$ →	$Na_2SO_{4(aq)}$ +	$SO_{2(g)}$ +	$H_2O_{(l)}$
1mole	1mole	1mole	1mole	1mole
0.05 moles	0.05 moles	0.05 moles	0.05 moles	0.05 moles

Since the 0.05 moles of sulfuric acid that are present initially will react fully with 0.05 moles of sodium sulfite, and there are 0.4 moles of sodium sulfite present initially, it is clear that it is the sodium sulfite that is present in excess. Sulfuric acid is the limiting reagent.

(b)

$Na_2SO_{3(s)}$ +	$H_2SO_{4(aq)}$ →	$Na_2SO_{4(aq)}$ +	$SO_{2(g)}$ +	$H_2O_{(l)}$
1mole	1mole	1mole	1mole	1mole
0.05 moles	0.05 moles	0.05 moles	0.05 moles	0.05 moles

Amount of sodium sulfate formed = 0.05 moles.

Mass of sodium sulfate formed = 0.05 x 142 g = 7.1 g

Amount of sulfur dioxide formed = 0.05 moles.

Volume (at s.t.p.) of sulfur dioxide formed = 0.05 x 22.41 = 1.121

2.18 If 4.95 g of ethylene (C_2H_4) are combusted with 3.25 g of oxygen:

(a) What is the limiting reagent?

(b) How many grams of CO_2 are formed?

2.19 a) Complete the balanced chemical equation for the reaction between copper (II) chloride and sodium nitrate.

$CuCl_2 + 2NaNO_3 \rightarrow Cu(NO_3)_2 +$ _________

(b) If 15 grams of copper (II) chloride react with 20 grams of sodium nitrate, how much sodium chloride can be formed?

(c) What is the limiting reagent in this reaction?

2.20 Zinc reacts with sulfuric acid according to the equation

$Zn_{(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$

A 250 cm^3 aqueous solution containing 9.8 g sulfuric acid is added to 13 g zinc.

(a) Show that the zinc is present in excess.

(b) Calculate the mass of zinc sulfate formed.

(c) Calculate the volume of hydrogen (measured at s.t.p.) formed.

MODULE 3

Rates of reaction and Chemical equilibrium

Learning outcomes

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

- Explain the concept of reaction speed 9.3.2.1
- Determine the factors that affect the rate of reactions and explain them in terms of the kinetic theory of particles 9.3.2.2
- Explain the difference between catalysts and reagents and the effect on reaction rates 9.3.2.3
- Explain the effect of inhibitors on reaction rates 9.3.2.4
- Distinguish between reversible and irreversible reactions 9.3.3.1
- Describe chemical equilibrium as a dynamic process and use the Le Chatelier-Brown principle to predict shifts in chemical equilibrium 9.3.3.2
- Recognise the effect of changing conditions on the rate of chemical reactions and the state of chemical equilibrium 9.3.3.3
- Explain chemical equilibrium in terms of the kinetic theory of particles 9.3.3.4

Keywords

✓ collide ✓ concentration ✓ temperature ✓ surface area ✓ catalyst
✓ activation energy ✓ equilibrium ✓ reversible ✓ decompose
✓ constant ✓ counteract

What is the 'rate of reaction'?

The rate of a chemical reaction tells us how quickly a chemical reaction happens. One of the slowest rates of reaction is the rusting of iron. A really fast reaction is an explosion, where it is all over in a fraction of a second.

Figure 3.1 *A nail rusting has a slow rate of reaction*

Figure 3.2 *An explosion has a very fast rate of reaction*

In industry it is very important to know how long it will take for a reaction to occur. Reactions can take varying amounts of time from start to finish, so chemists measure how much reactants are used up and how much product is formed in a certain time.

Figure 3.3 *Rate at which the reactants are used up or the product is formed*

What is particle theory?

During a chemical reaction the particles in the reactants are moving all the time. The rate of a reaction simply depends on how often and how hard the reacting particles **collide** with each other. For the particles to react they must collide with enough energy to break the bonds in the molecules. Once the bonds are broken the atoms can rearrange to make the products.

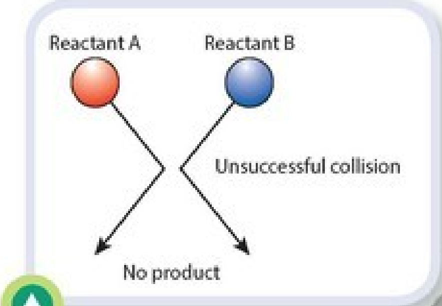


Figure 3.4 *Particle theory – an unsuccessful collision*

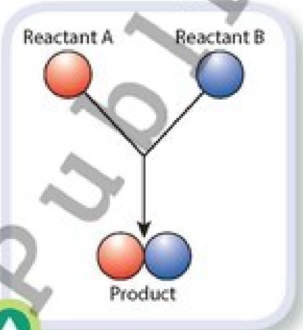


Figure 3.5 *Particle theory – a successful collision*

Increasing the rate of a reaction involves increasing the number of successful collisions between the reacting particles. There are a number of factors that will affect the behaviour of particles and will have an overall effect.

What is the effect of raising the temperature on the rate of a reaction?

If you change the temperature of the reaction, will this change the rate of reaction?

Lab Work 6

Activity 3.1

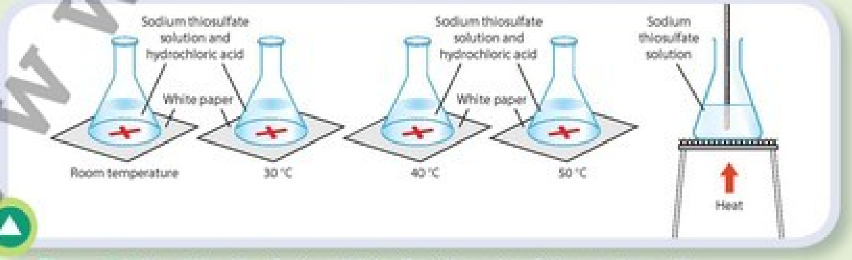


Figure 3.6 *Investigating factors that affect the rate of chemical reaction*

Question

What is the effect of temperature on rate of reaction?

We will be using a solution that turns cloudy with the addition of hydrochloric acid. We will test this at different temperatures by placing it in a flask over a piece of paper with a cross marked on it and seeing how long it is before we are unable to see the cross because of the cloudiness.

Equipment needed

4 conical flasks
Tripod
Bunsen burner
Wire gauze
Graduated cylinder
Dilute hydrochloric acid (3 M)
Sodium thiosulfate (0.05 M)
Thermometer
Paper
Marker
Stopwatch or timer
Metal tongs

Safety

- Use metal tongs when removing the conical flask from the Bunsen burner.

Conducting the activity

1. Mark a cross on a piece of white paper.
2. Measure 50 cm^3 of sodium thiosulfate solution into a conical flask.
3. Place the flask on a piece of white paper marked with a cross.
4. Record the temperature of the solution using the thermometer.
5. Add 5 cm^3 of hydrochloric acid (3 M) and swirl. Start the timer.
6. Record the time it takes for the solution to obscure the cross.
7. Repeat steps 1 to 6 at 30°C, 40°C and 50°C, heating the sodium thiosulfate solution in the flask over a Bunsen burner. Record all your results in a table like the following and note any conclusions you make from the results.

Temp °C	Time taken for cross not to be seen (seconds)
Room temperature	
30	
40	
50	

3.1 What happens to the rate of reaction as you increase the temperature?

When you raise the temperature the particles have more energy so they move around more quickly. This results in more collisions in a certain time, and so reactions get faster as we raise the temperature.

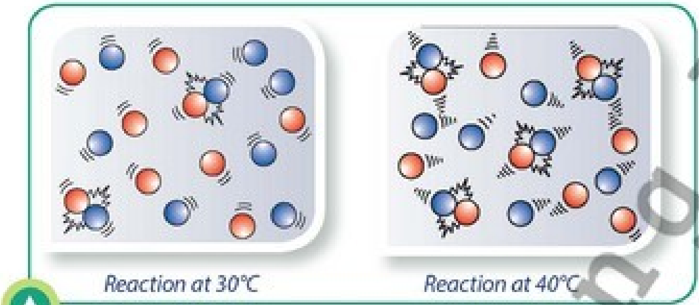


Figure 3.7 *As particles get more energy they move at a greater speed, and so there are more collisions and reactions get faster*

Q C2 S2

3.2 Have you ever tasted milk that has gone off? Some foods react when exposed to air, and turn into acids. Why do we store milk in a fridge, and why does refrigeration make it stay fresh for longer?

3.3 You are asked to make jelly for a party starting in three hours. What will you do to dissolve the jelly cubes in water as quickly as possible?

Did you know?

Light sticks produce light because of a chemical reaction. Dropping a light stick into hot water makes it glow more intensely, demonstrating that the reaction occurs faster at higher temperatures.

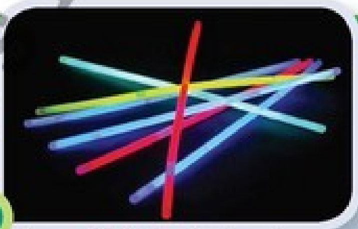


Figure 3.8 *A chemical reaction makes these light sticks glow*

Q

3.4 Define rate of reaction.

3.5 Explain how a change in temperature can affect the rate of reaction.

3.6 Plot the data below on a graph, with time on the x-axis and temperature on the y-axis. Describe what effect an increase in temperature has on the volume of carbon dioxide produced.

Volume of CO_2 produced

Time (s)	20°C	30°C
0	0	0
20	7	9
40	14	16
60	18	19
80	21	24
100	25	25

What is the effect of concentration?

Activity 3.2

Question

What is the effect of concentration on rate of reaction?

Equipment needed

Conical flask
Bung
Glass tubing
Basin
Graduated cylinder
Hydrochloric acid (0.1 M)
Water
Marble chips
Stopwatch or timer

Conducting the activity

1. Set up the apparatus as shown in **Figure 3.9**.
2. Place 5 g of marble chips in the conical flask.
3. Add a solution containing 10 cm^3 of hydrochloric acid (0.1 M) and 40 cm^3 of water.
4. Start the timer as soon as you place the bung in the flask.
5. Time how long it takes to collect 20 cm^3 of gas.
6. Repeat steps 2 to 5 using different concentrations of acid as shown in the table.
7. Record all your results in the table.

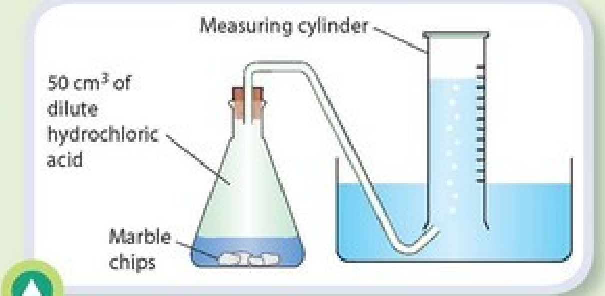


Figure 3.9

Acid (cm^3)	Water (cm^3)	Time to collect 20 cm^3 of carbon dioxide (s)
10	40	
20	30	
30	20	
40	10	
50	0	

3.7 What happens to the rate of reaction as you increase the concentration of the acid?

When you increase the concentration, you increase the number of particles, which results in more collisions in a certain time and so you get an increase in rate of reaction (that is, the reaction gets faster).

3.8 As you increase the concentration of acid there are more acid particles in the same volume. Use the particle theory and diagrams to explain why you think this may cause an overall increase in the rate of a reaction.

What is surface area?

Surface area is the measure of how much surface (of the reactants) is exposed. When a large piece of solid material is cut into smaller pieces its surface area is increased. What effect will an increase of surface area have on the rate of a chemical reaction?

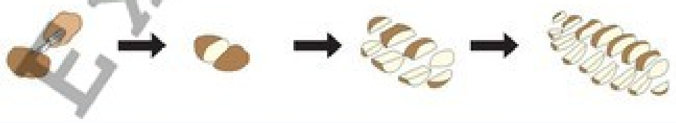


Figure 3.10 *You increase the surface area of the potato each time you cut it into smaller pieces*

R2 R3 R4

Activity 3.3

Question

What is the effect of increasing surface area on the rate of a reaction?

Equipment needed

Conical flask
Electronic balance
Cotton wool
Dilute hydrochloric acid
Marble chips (5 mm)
Marble chips (10 mm)
Stopwatch or timer

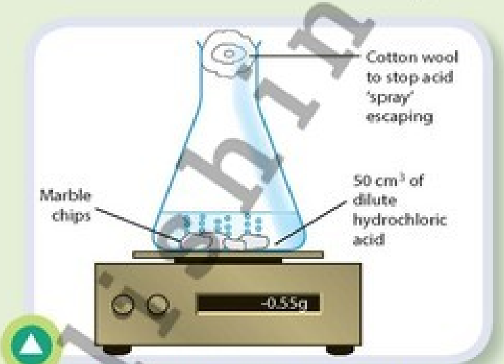


Figure 3.11 *Testing the effect of increasing the surface area on the rate of a reaction*

Conducting the activity

1. Set up the apparatus as shown in **Figure 3.11**.
2. Place 5 g of marble chips (5 mm) into a conical flask. Add 50 cm^3 of dilute hydrochloric acid.
3. Place cotton wool in the top of the flask.
4. Start the timer.
5. Record the loss in mass every 30 seconds.
6. Record your results in a table like the one below.
7. Repeat steps 2 to 6 using the larger marble chips (10 mm).

Time (s)	Mass (g)	Loss in mass
0		
30		
60		
90		
120		
150		
180		

Q R4 R5

3.9 Which marble chips have the larger surface area?

3.10 Why is it important to keep all other factors in the activity the same?

3.11 Plot both sets of results on a graph. Put loss of mass (mass of gas given off) on the y-axis and time on the x-axis.

3.12 Which size of marble chips reacts faster?

3.13 What happens to the rate of reaction as we increase the surface area?

The smaller the particle size of a solid, the greater the surface area on which the reaction can take place, and so the faster the reaction.

Understanding U3

Figure 3.12 *Types of iron*

3.14 Place the three types of iron shown in **Figure 3.12** in order of increasing surface area.

3.15 If each type of iron was heated over a Bunsen burner, which would show the greatest rate of reaction?

What is the effect of using a catalyst?

Transition metals lie in between group 2 and group 3 in the periodic table. Some well-known transition metals are iron, copper, chromium, nickel and gold. Transition metals and their compounds are important catalysts. (A catalyst makes something happen without being part of what happens.) How will a catalyst affect the rate of reaction?

Research R2 Research R3 Research R4

Practical Work 3

Portfolio 36

Activity 3.4

Question

What is the effect of a catalyst on the rate of a chemical reaction?

Equipment needed

2 test tubes
Hydrogen peroxide (6%)
Manganese dioxide (catalyst)
Spatula
Wooden splints

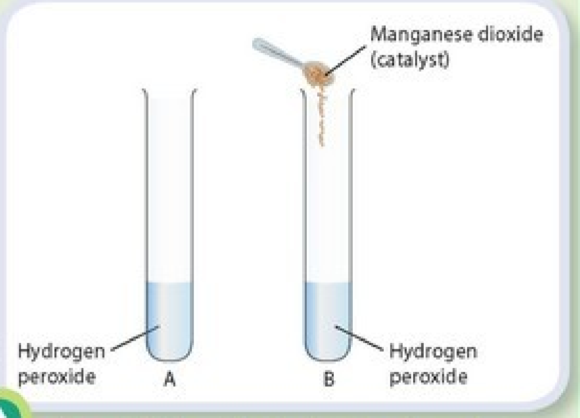


Figure 3.13 *Investigating the effect of a catalyst*

Safety

- Wear safety goggles and gloves at all times when using the hydrogen peroxide.

Conducting the activity

1. Place 5 cm^3 of hydrogen peroxide into each of test tubes A and B.
2. Add a little manganese dioxide to test tube B.
3. Test each test tube with a glowing splint.
4. Note your results.

Make sure that in your results you note answers to the following:

(a) Did bubbles form in test tube A?
(b) What happened as soon as you added the catalyst manganese dioxide?
(c) Did the glowing splint relight in each test tube?
(d) Which test tube showed the greater rate of reaction?
(e) How did you judge the rate of reaction?

3.16 The catalyst manganese dioxide speeds up the reaction in the previous activity without getting used up. How can you get the manganese dioxide back to be used again? (Hint: it is insoluble in water.)

How do catalysts increase the rate of a reaction?

Activation energy is the minimum amount of energy that reactant particles must have in order for them to react. When you add a catalyst it lowers the activation energy; this makes it easier for the particles to react, so there is an increase in the rate of reaction.

Did you know?

Enzymes are natural catalysts in the human body. Without the enzyme amylase in our saliva it would take starch several weeks to break down.

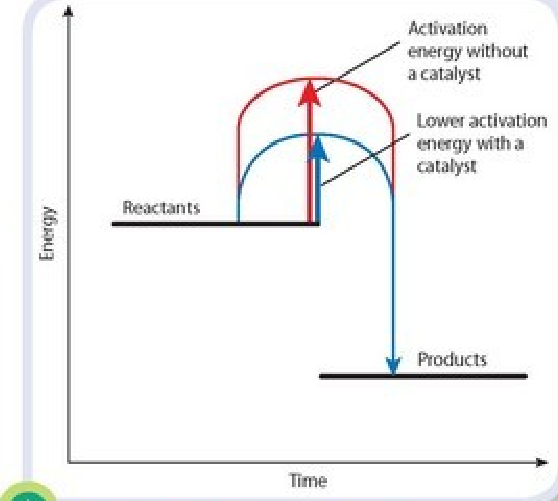


Figure 3.14 *How a catalyst works*

3.17 What are catalysts?

3.18 Explain how catalysts increase the rate of a chemical reaction.

3.19 Place the reactions in **Figure 3.15** in order of increasing rate of reaction.

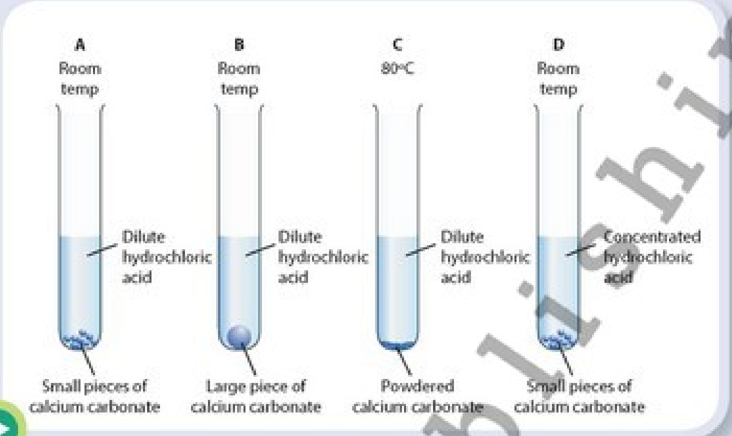


Figure 3.15

Reversible and irreversible reactions

There are many reactions – like combustion- which are irreversible. In these reactions reactants form products which cannot be or are not easily turned back into the reactants. The products of combustion – water and carbon dioxide – are stable and, therefore, to do not react with each other to reconvert into the reactants.

When we write chemical equations for such reactions, the irreversible nature of the reaction is represented by the unidirectional arrow: (→).

Reversible reactions

Reversible reactions behave differently. In a reversible reaction, the products can react to produce the original reactants again. In chemical equations for such reversible reactions, we use two arrows, as shown below in reaction of hydrogen and iodine:

$$\boxed{H_{(2)} + I_{(2)} \rightleftharpoons 2HI}$$

The concept of dynamic equilibrium

The concept of dynamic equilibrium refers to reversible reactions which occur continuously in both directions at the same time. Such reactions are referred to as dynamic reactions and when the rate of the reactions in both directions is equal, the reaction is in a state of **dynamic equilibrium**.

This does not necessarily mean, however, that the concentrations of the reactants and products are equal. We say that they are at equilibrium when their concentrations are constant. In the reaction

$$H_{(2)} + I_{(2)} \rightleftharpoons 2HI$$

as much hydrogen iodide is decomposing at equilibrium as is being formed, and so the concentrations of all three substances remain constant. This can be represented on a graph as follows:

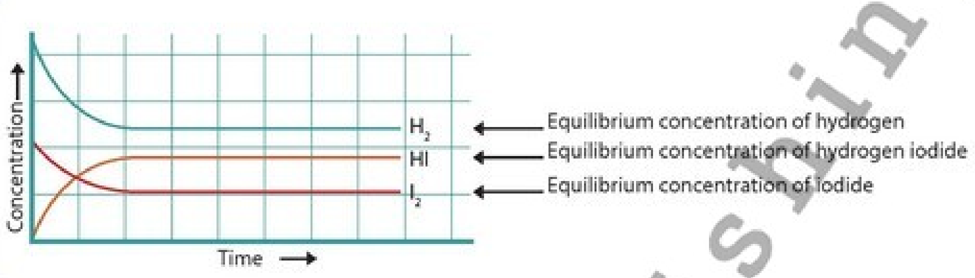


Figure 3.16

Only the reactants are found to be present in the initial stages, but as the reaction progresses, the concentrations of the hydrogen and iodine decrease as that of the product hydrogen iodide increases. Eventually a point will be reached where changes in the concentrations of the reactants are no longer observed and equilibrium is reached. The process is said to be dynamic because neither reaction goes to the end. The reactant amount that remains in such reversible reactions will depend on these factors:

- Temperature
- Concentration of the reactants
- Pressure within the system
- Type of reaction

The rate of the reaction to the left of the equation is called the forward rate, the rate of the reaction to the right is called the reverse rate.

$$H_2 + I_2 \rightleftharpoons 2HI$$

3.20 Consider these graphs. Some show **concentration of reactants v time**, others **rate of reaction v time**. Which graphs represent a state of chemical equilibrium?

Graph A

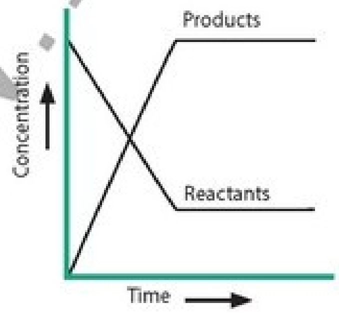


Figure 3.17

Graph B

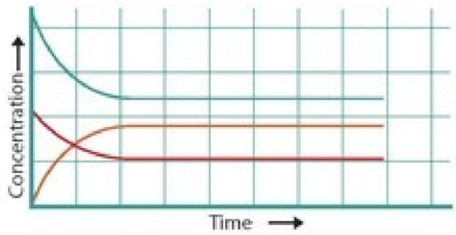


Figure 3.18

Graph C

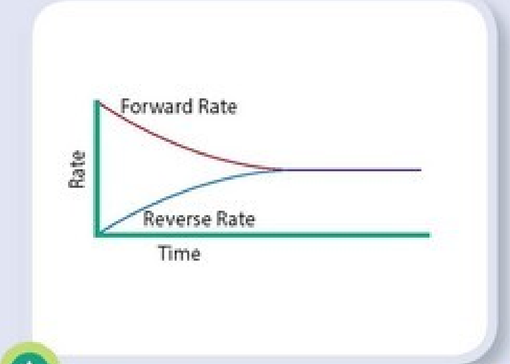


Figure 3.19

Graph D

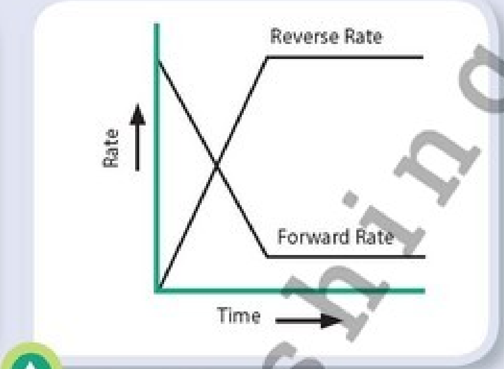


Figure 3.20

Graph E

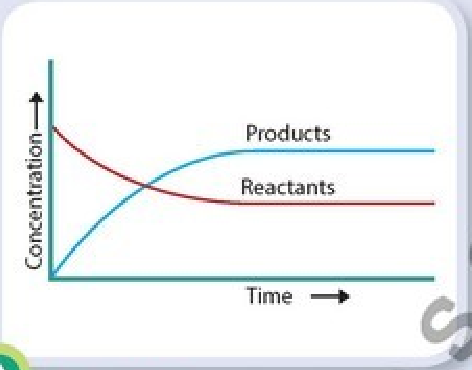


Figure 3.21

Graph F

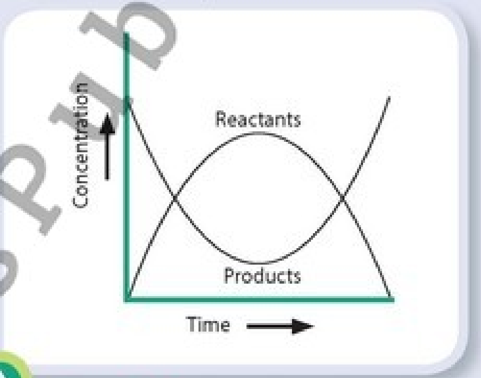


Figure 3.22

3.21 Which lines in graph b represent the reactants of the forward reaction and which line represents the products?

3.22 At which point in graph b is equilibrium reached ?

Equilibrium constant

Taking our example of a reversible reaction from above

$$H_2 + I_2 \rightleftharpoons 2HI$$

we can consider the equilibrium concentrations at a constant temperature of 764 K of the three species: hydrogen, iodine and hydrogen iodide. The table below shows a number of sets of different starting concentrations at this temperature. The square brackets denote concentrations in mol l^{-1}.

$[H_2] \times 10^3$	$[I_2] \times 10^3$	$[HI] \times 10^3$	$\frac{[HI]^2}{[H_2] \times [I_2]}$
1.894	1.896	12.83	45.9
2.265	2.840	17.15	45.7
1.699	4.057	17.79	45.7
2.484	2.514	16.95	46.0
1.971	1.981	13.42	46.1

From the table it can be seen that although the concentrations at equilibrium are very different there is a constant relationship between them [fourth column in the table] within a reasonable margin of error:

$$\boxed{\frac{[HI]^2}{[H_2] \times [I_2]} = \text{constant}}$$

The square brackets denote concentrations in mol L_{-1}.

This relationship could be presented on a graph like this.

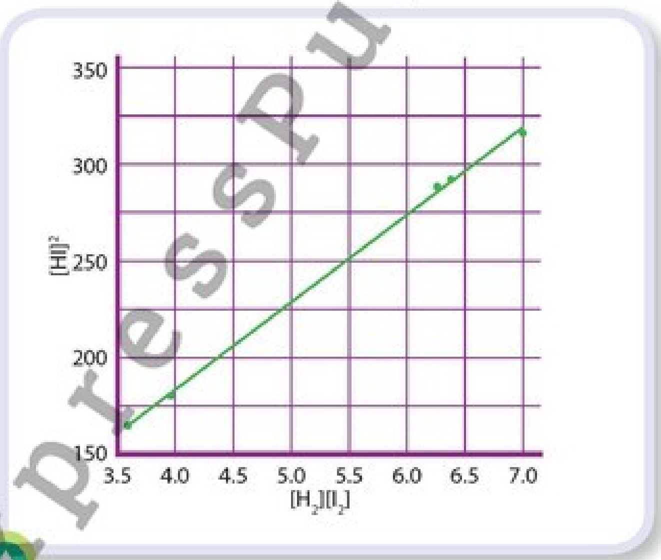


Figure 3.23

The straight line is indicative of:

$$\boxed{\frac{[HI]^2}{[H_2] \times [I_2]} = \text{constant}}$$

This constant value is known as the equilibrium constant [K_c] representing the constant relationship in terms of molar concentration.

It is important to remember that this is a constant for different concentrations at a specific temperature. If the temperature of the reaction changes then the value of K_c changes.

Look at the sample calculation for the same reaction at a completely different temperature.

Sample question 1

$H_2 + I_2 \rightleftharpoons HI$

If the equilibrium [HI] is 0.75 M and the equilibrium $[H_2]$ is 0.20 M, what is the equilibrium $[I_2]$ if the K_c is 0.40?

Answer:

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

[HI] is squared because of the coefficient 2 in the balanced chemical equation.

$$0.40 = \frac{(0.75)^2}{(0.20)[I_2]}$$

Rearrange as:

$$[I_2] = \frac{(0.75)^2}{(0.20)(0.40)}$$

$$[I_2] = 7.0\ M$$

The magnitude of the equilibrium constant gives an indication of the **position** of equilibrium. A large value of K_c, as in the table data , indicates that at equilibrium the concentration of products is large compared to the concentration of reactants.

From our table, we saw that in the reaction at 764 K $H_2 + I_2 \rightleftharpoons 2HI$ K_c is 45.9. This means that HI predominates at equilibrium at this temperature.

In the example question the K_c is low at 0.40 suggesting that the reactants predominate at this temperature.

The value of the equilibrium constant depends on temperature. If the forward reaction is exothermic, the equilibrium constant decreases as the temperature rises. If the forward reaction is endothermic, the equilibrium constant increases as the temperature rises.

Q Understanding U_4

3.23 Write the equilibrium constant expressions for the following reactions:

(a) $2SO_2 + O_2 \rightleftharpoons 2SO_3$ **(b)** $2NO + O_2 \rightleftharpoons 2NO_2$ **(c)** $2NH_3 \rightleftharpoons N_2 + 3H_2$

3.24 In the reaction $H_2 + I_2 \rightleftharpoons 2HI$ if the equilibrium [HI] is 0.060 M and the equilibrium $[I_2]$ is 0.90 M, what is the equilibrium $[H_2]$ if the K_c is 0.40?

3.25 In a reaction at a particular temperature between sulfur dioxide and oxygen forming sulfur trioxide $2SO_2 + O_2 \rightleftharpoons 2SO_3$ it was found that at equilibrium the concentrations of SO_2, O_2 and SO_3 were 0.07 mol l_{-1}, 0.035 mol l_{-1} and 0.03 mol l_{-1} respectively. Calculate the value of the equilibrium constant (K_c) for this reaction at this temperature.

3.26 What does a small value of (K_c) indicate about relative concentrations of species at equilibrium?

3.27 If the value of (K_c) decreases as the temperature increases, what does this indicate?

Le Chatelier's Principle

We use a principle known as the Le Chatelier's principle, named after the French chemist, Henry-Louis Le Chatelier, (1850 -1936), to predict what happens in chemical reactions where equilibrium is disturbed. This could be due, for example, to a change in temperature, pressure or the concentration of one of the species.

Le Chatelier's principle:

> If a dynamic equilibrium is disturbed by changing conditions, the position of equilibrium moves to counteract the change.

We shall look at the example reaction for the production of ammonia from nitrogen and hydrogen for each type of change.

Change in concentration

In the reaction

$$\boxed{N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}}$$

If we add either more N_2 or H_2 or both when the reaction is at equilibrium, the balance of the equation shifts to the right towards the product as more product is produced to return the reaction to equilibrium.

$$\boxed{N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}}$$

If more NH_3 is added when the reaction is at equilibrium, the balance of the equation would to the left as more nitrogen and hydrogen would be produced to return the reaction to equilibrium.

$$\boxed{N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}}$$

The reverse is true for a decrease in either the reactants or the products.

Change in temperature

In our reaction

$$\boxed{N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}}$$

the forward reaction is exothermic. An increase in temperature in the reaction at equilibrium will shift the equilibrium to the left (the endothermic direction). The reaction producing the nitrogen and hydrogen will take place preferentially. This will decrease the value K_c.

$$\boxed{N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}}$$

A decrease in temperature will favour the exothermic reaction, shifting the equilibrium to the right. This will result in the value of K_c. increasing.

$$\boxed{N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}}$$

Change in pressure

This change relates only to reactions where the reactants or products are gases. Equilibrium reactions involving only liquids and solids will not be affected by changes in pressure.

When a gas is involved in the equilibrium reaction, a change in pressure may affect the position of the equilibrium.

In the reaction

$$\boxed{N_2 + 3H_2 \rightleftharpoons 2NH_3}$$

there are four molecules on the left and only two on the right. If a mixture of nitrogen, hydrogen and ammonia at equilibrium is subjected to **increased pressure**, the equilibrium shifts to the right, forming more ammonia. This counteracts the effect of the increase in pressure, as the number of molecules present is reduced.

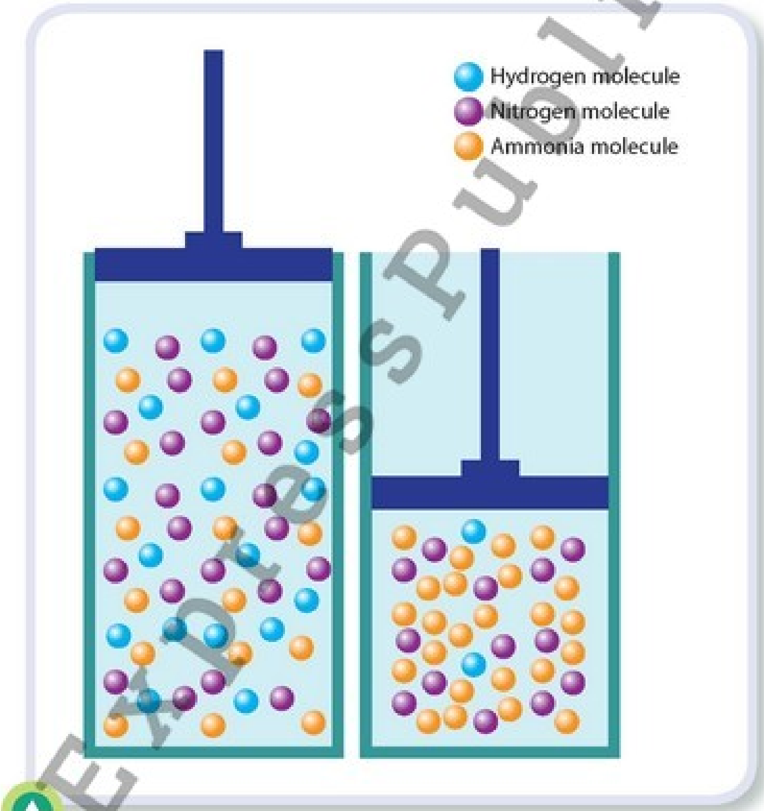


Figure 3.24

In **Figure 3.24** more ammonia molecules are produced to counteract the increase in pressure.

A decrease in pressure will have the opposite effect, causing more hydrogen and nitrogen molecules to be formed.

It should be noted that in our earlier example reaction

$$\boxed{H_2 + I_2 \rightleftharpoons 2HI}$$

a change in pressure will not alter the equilibrium as there are the same number of molecules on each side of the equation.

3.28 In the reaction $2SO_2 + O_2 \rightleftharpoons 2SO_3$ will increasing the concentration of oxygen move the equilibrium position to the left or the right?

3.29 If the forward reaction is exothermic, how does increasing the temperature of the reaction affect the value of K_c?

3.30 In the reversible reaction $COCl_2 \rightleftharpoons CO + Cl_2$ the forward reaction is exothermic. At equilibrium, what is the effect of:

(a) decreasing the temperature

(b) increasing the pressure

(c) adding some $COCl_2$

(d) removing some Cl_2

3.31 Why does a change in pressure not affect equilibrium in the reaction: $H_2 + I_2 \rightleftharpoons 2HI$?

MODULE 4

Oxidation and Reduction

Learning outcomes

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

- Understand and apply the rules for oxidation state 9.2.2.3
- Understand how the processes of oxidisation and reduction (redox-process) are interrelated and run simultaneously 9.2.2.4
- Recognise redox processes as involving a change in oxidation state 9.2.2.5
- Recognise that oxidation involves the loss of electrons and reduction involves the gaining of electrons 9.2.2.6
- Balance redox equations using the electron balancing method 9.2.2.7

Keywords

✓ redox ✓ oxidation ✓ reduction ✓ respiration ✓ photosynthesis
✓ chemical species ✓ enzymes ✓ oxidising agent ✓ reducing agent
✓ electronegative ✓ bleach ✓ half equation

Redox reactions

Reactions involving transfer of electrons are known as **oxidation–reduction** (or redox) reactions. In the following redox reaction:

$$Zn_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$$

Electrons are lost by the zinc atom and gained by the copper ion. Everyday examples of redox reactions include the rusting of iron and biological processes, such as respiration and photosynthesis.

Not all reactions involve electron transfer. In the reaction that occurs in the test for chloride ions, positive silver ions are attracted to negative chloride ions, and solid silver chloride forms.

$$Ag^{+}_{(aq)} + Cl^{-}_{(aq)} \longrightarrow AgCl_{(s)}$$

There is no transfer of electrons; the positive silver ions are simply attracted to the negative chloride ions, and solid silver chloride forms.

Oxidation and Reduction

Oxidation is the loss of electrons.

A chemical species is said to be oxidised if it loses electrons, for example, Zn in the reaction:

$$Zn_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$$

However, a species is reduced if it gains electrons, for example, Cu^{2+} in the above reaction.

Reduction is the gain of electrons.

In redox reactions, both oxidation, or loss of electrons, and reduction, or gain of electrons, occur.

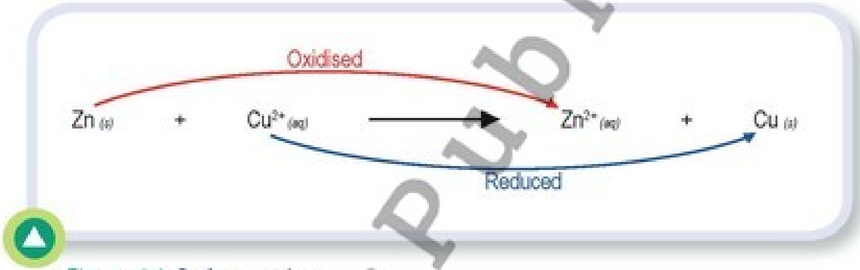


Figure 4.1 *Redox reaction*

It is relatively easy to work out what has been oxidised and what has been reduced in reactions such as:

$$Zn_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$$

where the charges on ions are clearly shown.

In reactions such as: $2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$

the fact that metals when forming compounds are usually oxidised, since they form positive ions and therefore lose electrons, can be used. Non-metals, on bonding with metals, usually form negative ions by gaining electrons, and therefore are reduced. In the above example, the metal sodium loses electrons and is therefore oxidised. The non-metal chlorine gains electrons and is therefore reduced.

$$\begin{array}{ll}
2Na - 2e^- \longrightarrow 2Na^+ & \text{oxidation} \\
Cl_2 + 2e^- \longrightarrow 2Cl^- & \text{reduction} \\
2Na + Cl_2 \longrightarrow 2Na^+ + 2Cl^- &
\end{array}$$

Rusting of iron

In the rusting of iron, iron reacts with oxygen and water, forming first Fe^{2+} ions and eventually Fe^{3+} ions. In this reaction the iron atoms lose electrons and are oxidised. Oxygen atoms gain electrons and are reduced.

Figure 4.2 *Iron is oxidised when rusting occurs*

Sample question 1

What is oxidised and what is reduced in the following reaction?

$2NaCl_{(s)} \longrightarrow 2Na_{(s)} + Cl_{2(g)}$

Answer:

Sodium ions in sodium chloride gain electrons to form sodium atoms, and are therefore reduced. Chloride ions lose electrons to form chlorine molecules, and are oxidised.

Did you know?

Apples turn brown when cut because certain enzymes and a class of chemicals known as phenols react in the presence of oxygen. This is an oxidation reaction.

4.1 What happens in a redox reaction?

4.2 Give two everyday examples of redox reactions.

4.3 In the reactions, **(i)** $2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)}$ **(ii)** $CuSO_{4(aq)} + Mg_{(s)} \longrightarrow Cu_{(s)} + MgSO_{4(aq)}$ **(iii)** $2KCl_{(l)} \longrightarrow 2K_{(s)} + Cl_{2(g)}$ identify:

- **(a)** what is oxidised?
- **(b)** what is reduced?

Oxidising agents

An oxidising agent is a substance that allows oxidation to happen by gaining electrons itself.

In a redox reaction, the oxidising agent is itself reduced. For example, in the reaction:

$$Cu^{2+}_{(aq)} + Mg_{(s)} \longrightarrow Cu_{(s)} + Mg^{2+}_{(aq)}$$

the copper ion, which is itself reduced, is the oxidising agent.

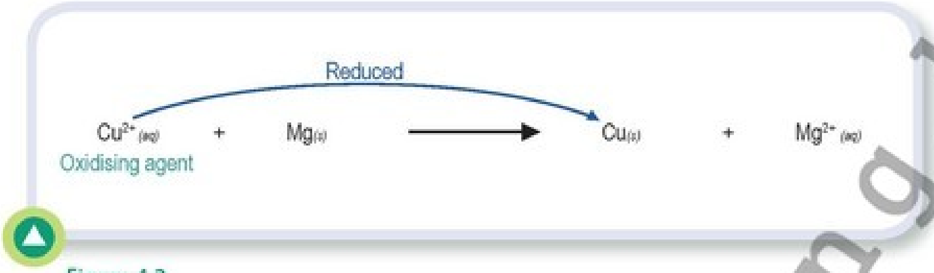


Figure 4.3

Figure 4.4 *Use of dilute hydrogen peroxide*

In a number of everyday processes, such as rusting of iron, combustion of fuels and respiration in living organisms, oxygen acts as an oxidising agent. Hydrogen peroxide is a powerful oxidising agent; it is released by some detergents in water, and is quite effective at removing stains, by oxidising the substances causing the stains.

Dilute solutions of hydrogen peroxide are used to lighten hair. The bleach reacts with the melanin in hair, removing the color in an irreversible chemical reaction. The bleach oxidises the melanin molecule.

Reducing agents

A reducing agent is a substance that allows reduction to happen by losing electrons itself.

In a redox reaction, the reducing agent is itself oxidised in the reaction. For example, in the reaction:

$$\boxed{Cu^{2+}_{(aq)} + Mg_{(s)} \longrightarrow Cu_{(s)} + Mg^{2+}_{(aq)}}$$

magnesium, which is itself oxidised, is the reducing agent.

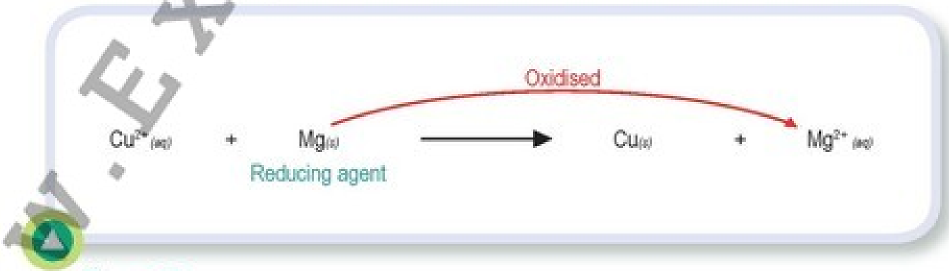


Figure 4.5

Sulfur dioxide, which is used in industry to bleach wool and paper, for example, acts as a reducing agent when it bleaches materials. So we can see that a bleaching agent can be either be an oxidizing or reducing agent. This will depend on the nature of chromophores that are responsible for the colour of the substance it is reacting with. The chromophores are either oxidized or reduced to produce a colorless or whitened substance.

Sample question 2

What is **(a)** the oxidising agent **(b)** the reducing agent in the following reaction?

$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)}$

Answer:

(a) Since zinc gains electrons in the reaction, it is reduced, and therefore is the oxidising agent.

(b) Since magnesium loses electrons in the reaction, it is oxidised, and therefore is the reducing agent.

4.4 What is an oxidising agent?

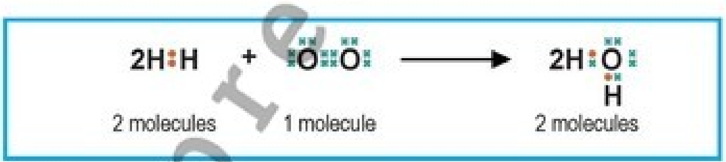
4.5 What is a reducing agent?

4.6 In the reactions, **(i)** $2CaO_{(s)} \longrightarrow 2Ca_{(s)} + O_{2(g)}$ **(ii)** $2Fe^{2+}_{(aq)} + Cl_{2(g)} \longrightarrow 2Fe^{3+}_{(aq)} + 2Cl^{-}_{(aq)}$ identify:

(a) the oxidising agent **(b)** the reducing agent.

Oxidation numbers

It is usually easy to work out what is oxidised and what is reduced in redox reactions involving metal ions, because the gain and loss of electrons is readily seen. Redox reactions involving covalent compounds are more difficult to keep track of in this way. Consider the reaction between hydrogen and oxygen, forming water:



At the beginning of the reaction, the hydrogen atoms share a pair of electrons equally; the oxygen atoms also share electrons equally. After the reaction, each hydrogen atom shares a pair of electrons with an oxygen atom. Since hydrogen is less electronegative than oxygen, the sharing is unequal – there is a partial transfer of electrons from the hydrogen to the oxygen. To keep track of electron shifts like this, and of oxidation–reduction in general, oxidation numbers are very useful.

The oxidation number of an atom is the charge that the atom appears to have when the electrons are distributed according to certain rules.

According to these rules, electrons shared between unlike atoms are counted with the more electronegative atom, while electrons shared between like atoms are divided equally between them.

The main rules are as follows:

1. **In free elements, each atom has an oxidation number of 0.** For example, in a nitrogen molecule (N_2) each nitrogen atom has an oxidation number of 0.
2. **The sum of the oxidation numbers of all the atoms in a molecule is 0.** For example, in the water molecule (H_2O), where hydrogen has an oxidation number of +1 and oxygen has an oxidation number of –2, the oxidation numbers add up to 0:
   H_2O
   $2(+1) - 2 = 0$
3. **The oxidation number of a simple ion containing one atom is equal to the charge on the ion.** For example, in a chloride ion (Cl^-) the oxidation number of chlorine is –1.
4. **The sum of the oxidation numbers of all the atoms in a complex ion is equal to the charge on the ion.** For example, in a nitrate ion (NO^{3-}) where nitrogen has an oxidation number of +5 and oxygen has an oxidation number of –2, the oxidation numbers add up to –1:
   NO^{3-}
   $+5 + 3(-2) = -1$
5. **Hydrogen has an oxidation number of +1 in its compounds, except in metallic hydrides where it is –1.**
6. **Oxygen has an oxidation number of –2 in its compounds, except in hydrogen peroxide where it is –1, and when bonded to fluorine it is +2.**
7. **The oxidation number of Group 1 elements in their compounds is always +1, and the oxidation number of Group 2 elements in their compounds is always +2.**
8. **The oxidation number of a halogen when bonded to a less electronegative atom is –1.** Fluorine is the most electronegative element, and always has an oxidation number of –1 in its compounds.

Chlorine has an oxidation number of –1 in compounds where it is not bonded to oxygen or fluorine.

Oxidation states

Oxidation numbers represent the oxidation state of an element. As indicated in rule 7, Group 1 and 2 elements in their compounds have just one normal oxidation state each (1 and 2 respectively). From rule 6, oxygen in its compounds usually has just one oxidation state. With the exception of fluorine, the halogens in their compounds have several oxidation states.

The transition metals each have a number of different oxidation states in their compounds. **Table 4.1** indicates the common oxidation states of a number of transition metals.

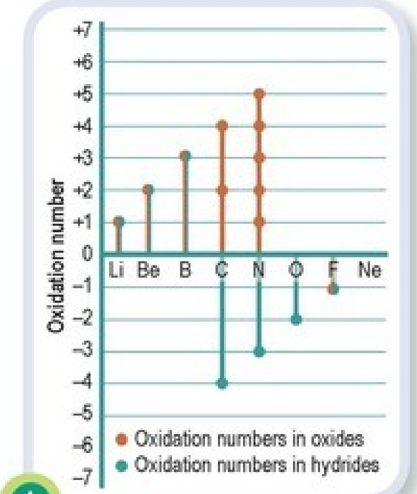


Figure 4.6 *Oxidation numbers of second period elements in their oxides and their hydrides*

Figure 4.7 *Transition metals*

Table 4.1

Element	Oxidation states
Chromium	2, 3, 6
Manganese	2, 3, 4, 6, 7
Iron	2, 3, 6
Copper	1, 2

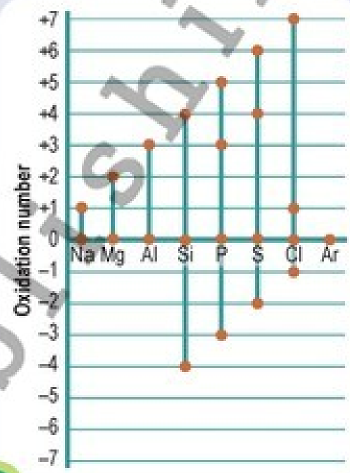


Figure 4.8 *Some oxidation states of elements in the third period. Note that all elements have an oxidation state of 0 when uncombined.*

Calculation of oxidation numbers

The rules for oxidation numbers are very useful in calculating oxidation numbers of atoms whose oxidation numbers are not given directly by the rules.

Sample question 3

What is the oxidation number of the phosphorus atom in the H_3PO_4 molecule?

Answer:

Let the oxidation number of P = x

The oxidation number of H = +1 (from rule 5)

The oxidation number of O = −2 (from rule 6)

Therefore 3(+1) + x + 4(−2) = 0 (from rule 2)

Therefore x = −3 + 8 = +5

Sample question 4

What is the oxidation number of Mn in $KMnO_4$?

Answer:

Let the oxidation number of Mn = x

The oxidation number of K = +1 (from rule 7)

The oxidation number of O = −2 (from rule 6)

Therefore +1 + x + 4(−2) = 0 (from rule 2)

Therefore x = −1 + 8 = +7

Sample question 5

What is the oxidation number of C in $C_6H_{12}O_6$?

Answer:

Let the oxidation number of C = x

The oxidation number of H = +1 (from rule 5)

The oxidation number of O = −2 (from rule 6)

Therefore 6(x) + 12(+1) + 6(−2) = 0 (from rule 2)

Therefore x = −12 + 12 = +0

4.7 Use the oxidation number rules to find the oxidation number of:

(a) Cl in $KClO_4$
(b) S in Na_2SO_4
(c) C in CO_2
(d) H in CaH_2
(e) Zn in $ZnCl_2$
(f) Cr in $K_2Cr_2O_7$

4.8 Assign oxidation numbers to each atom in the following:

(a) Al_2O_3
(b) NaH
(c) HNO_3
(d) K_2CrO_4
(e) Co_2O
(f) FeO
(g) NH_3
(h) S_8
(i) $Mg(NO_3)_2$
(j) PO_4^{3-}

Oxidation and Reduction in terms of oxidation numbers

Reduction is a decrease in oxidation number.

Oxidation numbers may be used to find out what is oxidised and what is reduced in a redox reaction. **Oxidation**, in terms of oxidation numbers, is **an increase in oxidation number**, while **reduction is a decrease in oxidation number**.

In working out what is oxidised and what is reduced in a reaction, it is important to remember that oxidation number is not an actual charge. To avoid confusion, and errors in calculations, the oxidation number should be written below the atom to which it applies, as shown in the following examples:

H_2O	$Cr_2O_7^{2-}$	OH^-
+1 −2	+6 −2	−2 +1

(Note that when doing this the oxidation number for only a single atom of each element should be written.)

Reduction is a decrease in oxidation number.

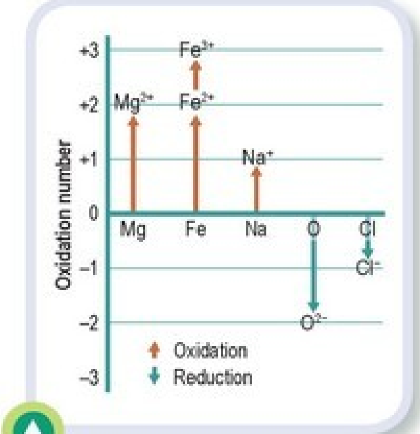


Figure 4.9 *Oxidation is an increase in oxidation number: reduction is a decrease*

Sample question 6

What is **(a)** oxidised **(b)** reduced in the following redox reaction?

$2SO_2 + O_2 \longrightarrow 2SO_3$

Answer:

Rules 1 and 6 respectively can be used to find the oxidation numbers of oxygen, while rule 2 can then be used to calculate the oxidation numbers of sulfur.

$2SO_2 + O_2 \longrightarrow 2SO_3$
+4 –2 0 +6 –2

Note that in all questions of this kind the number of moles given in the balanced equation for each substance should be ignored.

(a) The oxidation number of sulfur increases from +4 to +6. Therefore, sulfur is oxidised.

(b) The oxidation number of oxygen in O_2 decreases from 0 to –2. Therefore, the oxygen in O_2 is reduced.

Sample question 7

What is **(a)** oxidised **(b)** reduced **(c)** the oxidising agent **(d)** the reducing agent in the following reaction?

$2MnO_4^- + 16H^+ + 5C_2O_4^{2-} \longrightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$

Answer:

Assign oxidation numbers to all of the atoms in the equation, using the rules:

$2MnO_4^- + 16H^+ + 5C_2O_4^{2-} \longrightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$
+7 –2 +1 +3 –2 +2 +4 –2 +1 –2

(a) The oxidation number of carbon increases from +3 to +4. Therefore, carbon is oxidised.

(b) The oxidation number of manganese decreases from +7 to +2. Therefore, manganese is reduced.

(c) Manganese is the oxidising agent, because it is reduced itself.

(d) Carbon is the reducing agent, because it is oxidised itself.

Everyday examples of oxidising and reducing agents

Bleaching is a redox reaction. In bleaching, unwanted colours are removed by either oxidation or reduction. Many domestic bleaches contain sodium hypochlorite (NaOCl). This is an example of an oxidising bleach. When it bleaches by oxidation, it is reduced to sodium chloride:

$NaOCl \longrightarrow NaCl$
+1–2 +1 +1–1

Reducing bleaches include sodium sulfite (Na_2SO_3), used, for example in the paper industry, and sulfur dioxide (SO_2), used in solution where it forms sulfite ions (SO_3^{2-}). When sulfite ions bleach by reduction, they are converted to sulfate ions:

$SO_3^{2-} \longrightarrow SO_4^{2-}$
+4–2 +6–2

Figure 4.10
An oxidising bleach

4.9 What is **(i)** oxidised **(ii)** reduced in each of the following redox reactions?

(a) $C + O_2 \longrightarrow CO_2$

(b) $2Na + F_2 \longrightarrow 2NaF$

(c) $2H^+ + Zn \longrightarrow Zn^{2+} + H^2$

(d) $PCl_3 + Cl_2 \longrightarrow PCl_5$

4.10 Using oxidation numbers, determine which of the following are redox reactions. For each redox reaction, state what is: **(i)** oxidised **(ii)** reduced **(iii)** the oxidising agent **(iv)** the reducing agent

(a) $Sn + 4HNO_3 \longrightarrow SnO_2 + 4NO_2 + 2H_2O$

(b) $N_2O_4 \longrightarrow 2NO_2$

(c) $2Fe^{2+} + Cl_2 \longrightarrow 2Fe^{3+} + 2Cl^-$

(d) $Cl_2 + 2OH^- \longrightarrow Cl^- + ClO^- + H_2O$

Balancing redox equations using oxidation numbers

It is often not possible to correctly balance redox equations by inspection. A redox equation such as:

$$Cr_2O_7^{2-} + Fe^{2+} + H^+ \longrightarrow Cr^{3+} + Fe^{3+} + H_2O$$

may be balanced by a method using oxidation numbers. The method is as follows:

1. **Assign oxidation numbers to all the atoms in the equation.** In the example above:

$Cr_2O_7^{2-} + Fe^{2+} + H^+ \longrightarrow Cr^{3+} + Fe^{3+} + H_2O$
+6 −2   +2   +1   +3   +3   +1 −2

2. **Identify atoms that are oxidised or reduced, and in each case write down the number of electrons lost or gained per atom:**

$Cr + 3e^- \longrightarrow Cr$ Reduced
+6   +3
$Fe - e^- \longrightarrow Fe$ Oxidised
+2   +3

3. Attach subscripts (if any) as in the original equation to the oxidised and reduced atoms, and balance the resulting half equations:

$Cr_2 + 6e^- \longrightarrow 2Cr$ (1)
+6   +3
$Fe - e^- \longrightarrow Fe$ (2)
+2   +3

4. **Balance electron loss and gain by multiplying one or both of the half equations, and then add the resulting half equations:**

Leave half equation (1) as it is, and multiply half equation (2) by 6

$$\begin{array}{l} Cr_2 + 6e^- \longrightarrow 2Cr \\ \underline{Fe - e^- \longrightarrow Fe} \\ Cr_2 + 6Fe \longrightarrow 2Cr + 6Fe \end{array}$$

5. **Attach species and/or charges that were attached to the oxidised or reduced atoms in the original equation:**

$$Cr_2O_7^{2-} + 6Fe^{2+} \longrightarrow 2Cr^{3+} + 6Fe^{3+}$$

6. **Include the remaining species from the original equation:**

$Cr_2O_7^{2-} + 6Fe^{2+} + H^+ \longrightarrow 2Cr^{3+} + 6Fe^{3+} + H_2O$

and complete the balancing by inspection:

$Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \longrightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$

Sample question 8

Balance the following redox equation:

$MnO_4^- + CH_3OH + H^+ \longrightarrow Mn^{2+} + HCHO + H_2O$

Answer:

(1) $MnO_4^- + CH_3OH + H^+ \longrightarrow Mn^{2+} + HCHO + H_2O$

+7 -2 -2+1-2+1 +1 +2 +10+1-2 +1 -2

(2) $Mn + 5e^- \longrightarrow Mn$ Reduced

+7 +2

$C - 2e^- \longrightarrow C$ Oxidised

−2 0

(3) $Mn + 5e^- \longrightarrow Mn$ (1)

+7 +2

$C - 2e^- \longrightarrow C$ (2)

−2 0

(4) Multiply half equation (1) by 2, and multiply half equation (2) by 5

$2Mn + 10e^- \longrightarrow 2Mn$

$\underline{5C - 10e^- \longrightarrow 5C}$

$2Mn + 5C \longrightarrow 2Mn + 5C$

(5) $2MnO_4^- + 5CH_3OH \longrightarrow 2Mn^{2+} + 5HCHO$

(6) $2MnO_4^- + 5CH_3OH + H^+ \longrightarrow 2Mn^{2+} + 5HCHO + H_2O$

Balancing by inspection

$2MnO_4^- + 5CH_3OH + 6H^+ \longrightarrow 2Mn^{2+} + 5HCHO + 8H_2O$

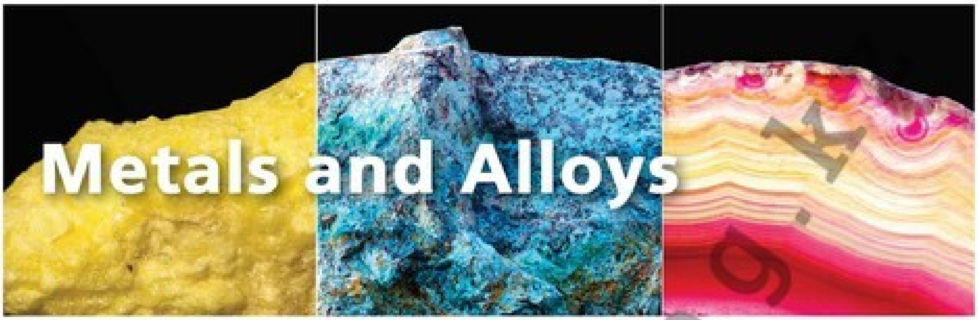
4.11 Using oxidation numbers, balance each of the following equations:

- **(a)** $MnO_4^- + Fe^{2+} + H^+ \longrightarrow Mn^{2+} + Fe^{3+} + H_2O$
- **(b)** $Cr_2O_7^{2-} + CH_3OH + H^+ \longrightarrow Cr^{3+} + HCHO + H_2O$
- **(c)** $ClO_3^- + C_6H_{12}O_6 \longrightarrow Cl^- + CO_2 + H_2O$
- **(d)** $NO_3^- + C \longrightarrow NO_2 + CO_2 + O^{2-}$

4.12 Balance each of the following equations, and name the substances acting as oxidising agent and as reducing agent in each case:

- **(a)** $Cr_2O_7^{2-} + C_2H_5OH + H^+ \longrightarrow Cr^{3+} + CH_3COOH + H_2O$
- **(b)** $MnO_4^- + H_2O_2 + H^+ \longrightarrow Mn^{2+} + O_2 + H_2O$

MODULE 5



Metals and Alloys

Learning outcomes

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

- Describe properties of metals using your knowledge of metallic bonds and crystal lattice structure 9.1.4.1
- Describe specific physical and chemical properties of metals and their reduction properties 9.1.4.2
- Explain the concept of an alloy and understand the benefits of alloys 9.1.4.3
- Distinguish between the composition and properties of different ferroalloys 9.1.4.4
- Identify forms of metal deposits in Kazakhstan; explain how they are mined and their environmental impact 9.4.2.4
- Describe the process of extracting metals from their ores 9.4.2.6
- Calculate theoretical and actual yields from equation reactions 9.2.3.6

Keywords

✓ conductivity ✓ density ✓ alloy ✓ metallic bonding
✓ molten ✓ rigidity ✓ corrosion ✓ extraction ✓ metal ore
✓ ball-bearings ✓ blast furnace ✓ coke ✓ charcoal
✓ precipitation ✓ theoretical yield ✓ percentage yield

Physical properties

Physical properties are the characteristics and features that define any material. Every material has a different set of properties that makes it perfect for some jobs and useless for other jobs. For example, the plastic used in the handle of a brush is hard, strong and stiff, whereas the plastic used in the bristles is soft and flexible.

Figure 5.1 *The handles of the brush and pan are made of hard plastic and the bristles of the brush are made of soft plastic*

Ninety-two elements are found naturally on Earth. These can be classified as:

- Metals
- Non-metals
- Semi-metals or metalloids (e.g. silicon).

There are just a few semi-metals, and in this module we will be looking at metals and non-metals.

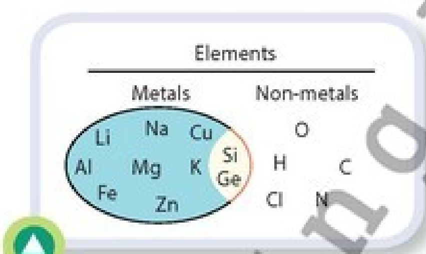


Figure 5.2 *Metals, semi-metals and non-metals*

Metals

Most of Earth's naturally occurring elements are metals. Their different properties make individual elements ideal for a large variety of uses in everyday life. Refer to the periodic table on page 166.

The properties of metals

Metals have a number of physical properties in common, but each metal is slightly different.

Figure 5.3 *A gold tooth – just one of the many uses of metal*

Metals form **metallic crystals**. Metallic crystals consist of closely-packed, regular arrangements of positive ions, which are surrounded by a sea of delocalised electrons. This sea of electrons binds the ions together. Most metallic crystal lattices are hard, though malleable and ductile. The melting points in metals are usually high but variable in different kinds of metal as can be seen in Table 5.1.

Table 5.1 Melting points of common metals in (°C)

aluminium	660	**potassium**	63.3
copper	1084	**tin**	232
pure gold	1063	**tungsten**	3400
lead	327.5	**zinc**	419.5
phosphorus	44		

Malleable means they can be hammered into different shapes; e.g. aluminium can be flattened into very thin sheets and sold as 'tinfoil'.

Ductile means they can be stretched out into wire, e.g. copper.

Metals are generally good conductors of electricity and are insoluble in all cases except for mercury.

Metals have high densities

Many metals have a high density. This means they feel very heavy for their size. The reason for this is that there are a lot of atoms packed into a small volume.

Metallic bonding

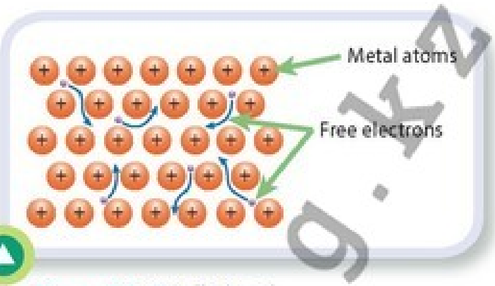


Figure 5.4 *Metallic bond*

Metal atoms are held tightly together by **electrons**. These electrons come from the outer shells of the atoms and hold the atoms together.

The metals are hard, strong and can conduct heat because the atoms are held tightly together. Metals can conduct electricity because the electrons are free to move.

Q Understanding U₁

5.1 Complete this sentence: Metals are good conductors of ________________ and ________________ .

5.2 Complete this sentence: Copper can be drawn out into thin wire. This means it is a ________________ metal.

5.3 Gold is a malleable metal. What does this mean?

5.4 What is unusual about the metal mercury?

Q Communicating C₂ Society S₂

5.5 Divide the class into groups. Each group is to choose and look at one metal. Within your group, discuss your chosen metal's properties and decide how it is best used.
An example is: Aluminium is strong and can be bent into shape but it is also light so it is used to make aeroplanes.

Understanding U₁ Communicating C₁ Communicating C₂ Society S₂

5.6 Titanium is a transition metal that is used in hip replacements. Research the properties of titanium and list three properties that make it suitable for this use.

5.7 Research the properties of the metals listed below, and explain why each has been chosen for its use:

- **(a)** Using copper to make wires.
- **(b)** Using iron to make bridges.
- **(c)** Using gold and silver to make jewellery.

Alloys

An alloy is a mixture of two or more elements, at least one of which is a metal.

An alloy can have very different properties from the elements from which it is formed. The elements are usually mixed by melting them, mixing the molten elements, and then allowing the mixture to solidify.

Sometimes metals are more useful when they are pure, e.g. pure copper is much better at conducting electricity than when it is mixed with other metals. However, many metals are better and more useful for certain jobs when they are combined with other metals or a non-metal (such as carbon). These are known as alloys. Some common alloys are listed in **Table 5.2**.

Table 5.2 Alloys, their uses and composition

Alloy	Composition	Use
Brass	Copper and zinc	Musical instruments, ornaments
Bronze	Copper and tin	Statues
Solder	Lead and tin	Soldering
Mild steel	Iron and carbon	Building reinforcement
Stainless steel	Iron, chromium and nickel	Knives, sinks
Alnico	Aluminium, nickel and cobalt	Powerful magnets

Steel

Pure iron has a regular arrangement of identical atoms. The layers of atoms can slide over each other and this makes the iron soft, which allows it to change shape easily. This makes it unsuitable for most uses.

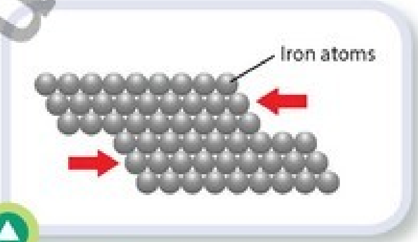


Figure 5.5 *Iron*

Most pure iron is changed into alloys of steel, where small amounts of carbon and sometimes other metals are added to the iron.

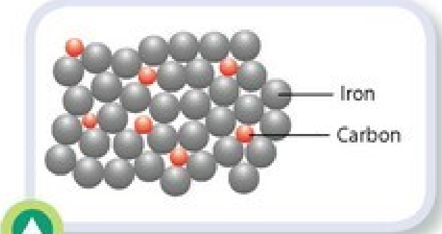


Figure 5.6 *Steel*

The carbon atoms upset the layers of iron atoms, making it more difficult for them to slide over each other. This makes the alloy harder.

Steel contains mainly iron with, typically, about 0.3% of carbon. Small, carefully measured amounts of transition metals such as chromium, nickel, tungsten, titanium, vanadium or manganese may also be added. These elements are chosen to give the steel particular properties to suit its use, such as hardness, rigidity, flexibility, and resistance to corrosion. Steel is much harder than iron and can be used in making cutting tools, girders for bridges and steel-framed buildings and reinforcements for building foundations, amongst many other uses.

Communicating C1 Communicating C2 Society S2

5.8 There are three different types of steel, as shown in the table below:

Type of steel	Properties
Low carbon steel	Easily shaped
High carbon steel	Very hard, inflexible
Stainless steel	Corrosion-resistant

(a) Which of these alloys are suitable for car bodies? Give your reasons.

(b) Stainless steel is corrosion-resistant. What does this mean?

(c) Give some everyday uses of stainless steel.

(d) High carbon steel is used for blades for cutting tools and for building bridges. Why is this?

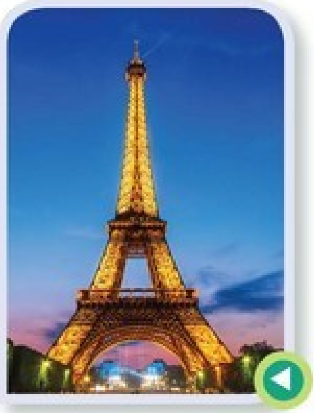


Figure 5.7 *The Eiffel Tower is made of iron and needs repainting every seven years*

Did you know?

The Eiffel Tower is made of iron. Iron goes rusty if air and water get on it. So the Eiffel Tower has to be painted every seven years and it takes an entire year of twenty-five painters working on it. That wouldn't happen with stainless steel!

Communicating C1 Communicating C2 Society S2

5.9 Some alloys are called **smart alloys**. One such smart alloy is **nitinol**. Research this alloy, looking particularly at:

(a) What metals are mixed together to form nitinol?

(b) Why is it called a smart alloy?

(c) Where is it used?

Redox reactions between metals and non-metals.

Reactions involving the transfer of electrons are known as **oxidation–reduction** (or redox) reactions. In redox reactions, both **oxidation**, or loss of electrons, and **reduction**, or gain of electrons, occur. The most common reducing agents are metals since they tend to lose electrons in their reactions with non-metals.

In the following reaction

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$

we can use the fact that metals when forming compounds are usually oxidised – since they form positive ions and therefore lose electrons – to help us write the oxidation and reduction equations. Non-metals on bonding with metals usually form negative ions by gaining electrons and, therefore, are reduced.

$2Na - 2e^- \longrightarrow 2Na^+$ oxidation
$Cl_2 + 2e^- \longrightarrow 2Cl^-$ reduction
$2Na + Cl_2 \longrightarrow 2Na^+ + 2Cl^-$

The rusting of iron is a common example of this 'reducing agent' quality of metals. In rusting, iron reacts with oxygen and water, forming first Fe^{2+} ions and eventually Fe^{3+} ions. In this reaction the iron atoms lose electrons and are oxidised. Oxygen atoms gain electrons and are reduced.

5.10 What feature of metals is it that makes them reducing agents in reactions with non-metals?

5.11 In the following reactions what is oxidised and what is reduced?
Write the oxidation and reduction equations out as above:

(a) $2Mg_{(s)} + O_{2(g)} \longrightarrow$ ____________________

(b) $2K_{(s)} + F_{2(g)} \longrightarrow$ ____________________

The extraction of metals.

Metals such as gold and metal compounds such as aluminium oxide, copper oxide and iron oxide are found in the Earth's crust. They are found in the form of **metal ores** which are rocks containing the metal or metal compound. Extraction is the term used to describe removing the metal from the other substances that it is mixed with when found in its metal ore state .

Did you know?

Kazakhstan is the world's leading producer of uranium. It provides a third of world output.

There are three methods typically used to extract metals from their ores:

- electrolysis
- reduction with carbon or carbon monoxide
- chemical reactions involving hydrogen.

The key factor in determining how a metal is extracted from its ore is its reactivity. A reactive metal such as aluminium, for example, is extracted by electrolysis, whereas a metal such as iron which is less reactive, is extracted through a reduction process involving carbon or carbon monoxide. Copper which is even less reactive is extracted through processes involving reactions with hydrogen.

The position of a metal in the reactivity series above or below carbon, dictates whether a metal is extracted using electrolysis or reduction with carbon. Those metals above carbon will be extracted by electrolysis, those below carbon will be extracted by reduction with carbon. During the reaction, the metal in the ore is displaced from its non-metal anion because carbon is more reactive than the metal. This process is cheaper than electrolysis because of availability of forms of carbon such as coke and charcoal.

The left-hand column of Table 5.3 shows metals in order of reactivity and the right-hand column shows the methods of extraction typically used.

Table 5.3 **Reactivity and extraction method**

Metals – most to least reactive	Method of extraction used
• potassium • sodium • magnesium • aluminium	Extracted by **electrolysis**
• iron • tin • lead	Extracted by reaction with **carbon** or **carbon monoxide**
• copper • gold	extracted by chemical reactions with **hydrogen**

Q Understanding U1

5.12 What is the difference between a mineral and an ore?

5.13 Which type of metal ores are extracted using carbon?

5.14 What advantage does carbon extraction have over electrolysis?

Manufacture of iron and steel

Iron occurs in ores such as haematite (impure Fe_2O_3), magnetite (impure Fe_3O_4) and iron pyrites (impure FeS). The extraction of iron from its ores is carried out in a **blast furnace**, in which the iron ore is reduced to molten iron, using coke at a high temperature. Unlike more reactive metals, this extraction is a **chemical reduction process** rather than an electrochemical process.

Manufacture of iron

The raw materials for the blast furnace process for the extraction of iron from ores such as haematite are: coke, limestone and the ore itself. In the blast furnace process, the ore, coke and limestone are fed regularly into the furnace from the top. Hot air is added from the bottom.

- Coke is mainly composed of carbon, and the ore is reduced by reaction with carbon:

$$\boxed{Fe_2O_{3(s)} + 3C_{(s)} \longrightarrow 2Fe_{(s)} + 3CO_{(g)}}$$

- The carbon monoxide causes further reduction in the upper, cooler regions of the furnace:

$$\boxed{Fe_2O_{3(s)} + 3CO_{(g)} \longrightarrow 2Fe_{(s)} + 3CO_{2(g)}}$$

  Most of the reduction of the iron ore is caused by carbon monoxide. The gases formed (CO, CO_2) are drawn off through ducts.

- The limestone decomposes into calcium oxide and carbon dioxide. The calcium oxide reacts with impurities in the iron ore, forming slag:

$$\boxed{\begin{array}{l} CaO + SiO_2 \longrightarrow CaSiO_3 \\ CaO + Al_2O_3 \longrightarrow CaAl_2O_4 \end{array}}$$

  The molten iron and slag trickle down into the hearth, where the less dense slag floats on the iron. The slag and iron are separately tapped off into large tanks.

- At this stage the iron is not very pure and, if allowed to solidify, forms pig iron. Pig iron is brittle, with a high carbon content, and is much less useful than steel.
  The slag is used in roadmaking.

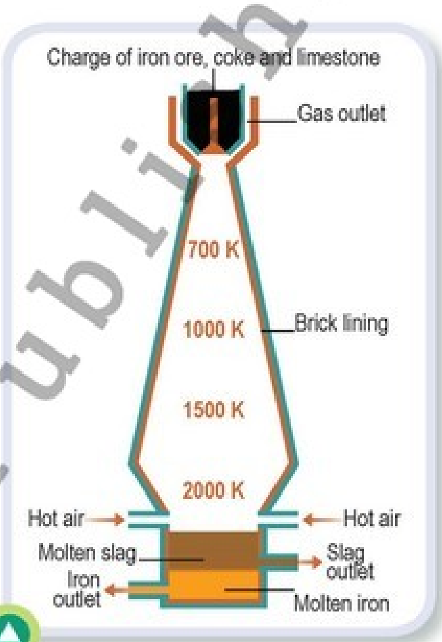


Figure 5.8 *The blast furnace*

Manufacture of steel

To make steel from iron, the iron must first be purified.

- Oxygen is blown onto the surface of the liquid iron at the base of the furnace. This oxidises silicon, manganese and some iron, and the products are removed as slag. Carbon is also oxidised and the gaseous products escape to the atmosphere.
- The molten iron, which is now sufficiently pure, is poured into a ladle and carefully measured amounts of carbon and other elements are added. This process is called the **basic oxygen process**.

Properties of different steels

Alloying carbon with iron to make steel increases the hardness and strength of the metal. Other metals may be added in controlled amounts as well. By a correct choice of metal, and by adding a suitable proportion of this metal to iron, steels can be made which are very hard, or very resistant to corrosion, or of great strength. Metals such as tungsten are used to harden steel. Corrosion-resistant steels are made using chromium and nickel. Manganese is used to make very tough steel.

Uses of iron and steel

Steel is used in the manufacture of car bodies, and in the construction of bridges and buildings. The uses of iron are more limited. It is used, for example, in manhole covers, chains and gates.

Q Understanding U4

5.15 Why is iron converted to steel?

5.16 In the blast furnace process for the extraction of iron from iron ore, what reducing agents are used?

5.17 How is iron from a blast furnace further purified?

5.18 How is steel made:

(a) harder **(b)** more corrosion-resistant **(c)** tougher?

5.19 Why are the reactions that take place in a blast furnace regarded as redox reactions?

Metal deposits in Kazakhstan

It has been estimated that 95% of the elements in the periodic table can be found in Kazakhstan. The vast mineral reserves have led to the country becoming a globally important producer of a wide range of metals. It is the world's largest producer of uranium and amongst the top ten producers in the world of: copper, magnesium, chromium, titanium and rhenium. It also has very large deposits of gold, bauxite (from which aluminium is produced), manganese and zinc.

The ferroalloy plant in Aktobe is an example of the huge scale of the metals industry in Kazakhstan. The plant has the capacity to produce 440,000 tons of high-carbon ferrochrome, which is used in the manufacture of stainless steel. High-carbon ferrochrome is particularly important in steels used to manufacture ball-bearings.

Figure 5.9

Environmental problems associated with the metals industry

There are a range of environmental problems associated with the mining, extraction and production of metals that organisations involved in the industry need to monitor closely. These include air pollution, water pollution and the production of undesirable wastes.

- A danger to air quality, for example, is when iron is being extracted from a sulfide ore, some sulfur dioxide will be formed in the blast furnace. When coke is being produced from coal, for use in the blast furnace, smoke and sulfur dioxide are formed. Emissions of these must be kept to a minimum. Fumes from a blast furnace must also be cleared of dust particles before being released into the atmosphere.
- The extraction of aluminium oxide from bauxite results in the formation of very large amounts of red mud, which can spoil the environment. This can be limited by the **recycling** of aluminium which is an economically viable process. It is also desirable from an environmental point of view, saving energy and conserving natural resources.

Figure 5.10 *Plant for the extraction of alumina from bauxite*

- Metals with high relative atomic masses such as mercury, cadmium and lead are known as heavy metals. Dipositive ions of these metals, i.e. Hg^{2+}, Cd^{2+} and Pb^{2+}, sometimes get into waterways from industrial effluent, and consequently into drinking water. These elements are cumulative poisons in that frequent exposure causes a build-up in the body, with consequent serious health damage. Before the effluent is run into a waterway, **the metal ions can be removed by means of precipitation**. For example, lead(II) hydroxide, $Pb(OH)_2$, is insoluble in water, so that if effluent containing Pb^{2+} ions is treated with calcium hydroxide solution, the lead will precipitate as $Pb(OH)_2$:

$$\boxed{Pb^{2+}_{(aq)} + 2OH^-_{(aq)} \longrightarrow Pb(OH)_{2(s)}}$$

Understanding U4

5.20 Name one use of high carbon ferrochrome.
5.21 Which gas associated with blast furnaces can cause air pollution?
5.22 Why is pollution of water by heavy metal ions a problem?
5.23 How can heavy metal ions be removed from water?

Calculating the mass of a substance in ore and yield given percentage impurities

In the above process of extracting iron, it is important to know how much iron can theoretically be obtained from the ore. To do this we have to work out the **theoretical yield**. This is a calculation using reacting masses based on balanced equations and it gives the maximum amount of product(s) possible from the reactants involved.

So if we asked the question: how much iron can be obtained from 1000 tonnes of pure haematite ore, formula Fe_2O_3 in a blast furnace? We could work out the theoretical yield as follows.

Sample calculation 1

Remember we saw in this process that iron(III) oxide ore (haematite) is reduced by carbon monoxide. The equation is:

$$\boxed{Fe_2O_{3(s)} + 3CO_{(g)} \longrightarrow 2Fe_{(l)} + 3CO_{2(g)}}$$

Atomic masses: Fe = 56 O = 16

Formula mass of ore from $Fe_2O_{3(s)}$: Fe: 2 x 56 O: 3 x 16

Reacting mass ratio is therefore 112/160 for iron.

So solving the ratio, $1000 \times \frac{112}{160} = 700$ **tonnes of iron extracted (theoretical maximum)**

However, when a chemical reaction occurs, the amounts of products isolated are often less than those calculated. This can be due, for example, to the reaction being reversible or to losses when purifying the products. In reactions where theoretical yields of products are not obtained, the calculation of **percentage yield** is useful.

$$\boxed{\text{Percentage yield} = \frac{\text{Actual yield x 100}}{\text{Theoretical yield}}\ \%}$$

Sample calculation 2

Imagine that the above sample of iron ore did not yield 700 tonnes of iron but only 630. Calculate the percentage % yield of the overall blast furnace process and how much iron is lost in waste?

Answer:

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

$$= 630 \times \frac{100}{700} = 90\%$$

which means

100% – 90% = 10% is lost in the slag of the furnace

5.24 Look at the reaction below. If 50 grams of Fe_3O_4 produces an 80% yield of Fe, how many grams are produced?

$$\boxed{Fe_3O_4 + 4H_2 \longrightarrow 3Fe + 4H_2O}$$

5.25 In the equation below 50 grams of magnesium reacts with an excess of nitric acid.

$$\boxed{Mg + 2HNO_3 \longrightarrow Mg(NO_3)_2 + H_2}$$

(a) What is the theoretical yield of hydrogen?

(b) If 2.5 grams of hydrogen is actually produced what is the percentage yield?

MODULE 6

Elements in Groups 1, 2 and 13

Learning outcomes

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

- Describe the general properties of alkaline metals using their atomic structure 9.2.1.1
- Write chemical equations, illustrating the general properties of oxides and alkali hydroxides 9.2.1.2
- Compare the general properties of metals of Groups 1 and 2 and make equations of their reactions 9.2.1.3
- Explain the properties of calcium oxides and metal hydroxides and describe their applications 9.2.1.4
- Describe the properties of aluminium using its atomic structure and describe the application of aluminium and its alloys 9.2.1.5
- Investigate amphoteric properties of aluminium and its oxide and hydroxide 9.2.1.6
- Plan and carry out an experiment on the interaction of metals in Groups 1, 2 and 13 and their simple and complex compounds 9.2.1.7

Keywords

✓ reactivity series ✓ density ✓ periodic trend ✓ respiratory equipment
✓ malleable ✓ amphoteric ✓ crystalline solid ✓ crystalline lattice
✓ physical state ✓ ionic oxide ✓ thermal conductivity ✓ anodising

The Alkali metals

The Group I elements are called the **alkali metals**. They are all very reactive metals, **increasing in reactivity down the group**.

Physical properties

- They are soft metals – they can be cut with a knife.
- They have low densities – lithium, sodium and potassium each float in water.

Chemical properties

- When they are freshly cut, they have a metallic shine, but this disappears rapidly due to reaction of the exposed metal surface with oxygen in the air. The metal oxide is formed, as it is when the alkali metals burn in air. The following **word equation** shows how sodium reacts with oxygen:

sodium + oxygen → sodium oxide

- They react vigorously with water, forming a basic solution and hydrogen. For example sodium reacts as follows:

sodium + water → sodium hydroxide + hydrogen

$$2Na_{(s)} + 2H_2O(l) \rightarrow 2NaOH_{(aq)} + H_{2(g)}$$

(In the above **chemical equation**, **chemical formulas** are used, and the subscripts following each formula give information about the physical states of reactants and products: (s) = solid; (l) = liquid; (aq) = dissolved in water [aqueous solution]; (g) = gas.)

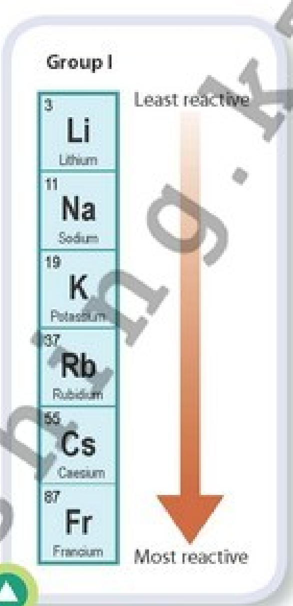


Figure 6.1 *Reactivity trend in Group I*

Alkali metals are stored in oil to prevent reactions with water and with atmospheric oxygen.

Activity 6.1

Question

What do we observe when Group 1 elements react with water?

Equipment needed

Lithium
Sodium
Potassium
Water
Litmus paper
Trough
Knife

Your teacher will conduct this experiment for you to observe. The chemical equations for the reactions between alkali metals and water in this demonstration are as follows:

$$2Li_{(s)} + 2H_2O_{(l)} \longrightarrow 2LiOH_{(aq)} + H_{2(g)}$$
$$2Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$$
$$2K_{(s)} + 2H_2O_{(l)} \longrightarrow 2KOH_{(aq)} + H_{2(g)}$$

Conducting the activity

1. Place a small piece of lithium in a trough of water. Describe and explain what happens.
2. Dip a piece of red litmus paper into the trough. Describe and explain what happens.
3. Place a small piece of sodium in a trough of water. Describe and explain what happens.

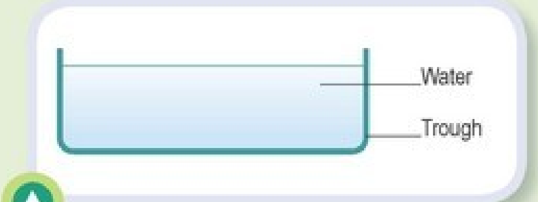


Figure 6.2

4. Dip a piece of red litmus paper into the trough. Describe and explain what happens.
5. Place a small piece of potassium in a trough of water. Describe and explain what happens.
6. Dip a piece of red litmus paper into the trough. Describe and explain what happens.

Q Research R4 Research R5

6.1 What metal shows the most energetic reaction?
6.2 What metal shows the least energetic reaction?

The increase in reactivity as you go down the group is called a periodic trend. It can be explained by looking at the electronic structure of the alkali metal atoms.

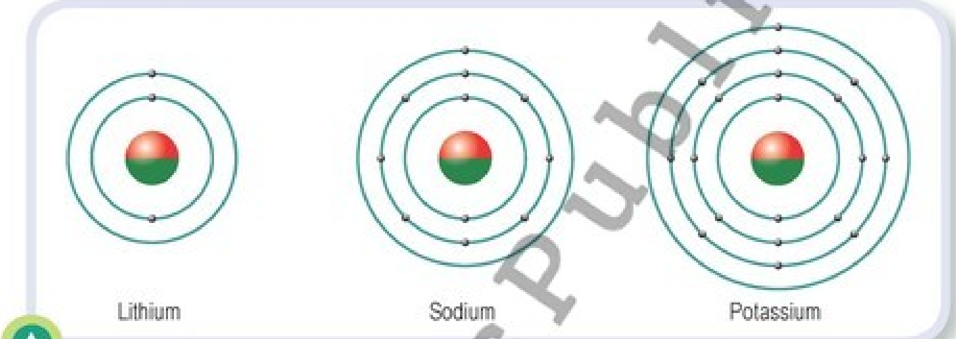


Figure 6.3

The reactivity of the elements increases moving down the group because

- the atoms increase in size
- as the outer electron gets further from the nucleus, the attraction becomes weaker and so the electron is lost more easily.

Alkali metal oxides and hydroxides

Figure 6.4

Group 1 elements react readily with oxygen to produce different ionic oxides.

Complete the equations for each of the reactions described below.

The combustion reaction for lithium forms white lithium oxide.

$4Li + O_2 \rightarrow$ ____________ [lithium oxide]

The combustion reaction of sodium forms a pale yellow mixture of sodium oxide and sodium peroxide.

$4Na + O_2 \rightarrow$ ____________ [sodium oxide]

$2Na + O_2 \rightarrow$ ____________ [sodium peroxide]
Sodium peroxide consists of Na^+ O_2^{2-} ions

The combustion reaction of molten potassium produces the orange coloured solid potassium superoxide:

$K + O_2 \rightarrow$ ____________
This salt consists of K^+ and O^-_2 ions.

The alkali metal peroxides, such as sodium peroxide and superoxides, have many important uses because of their strong oxidant qualities. Sodium peroxide is used to bleach paper and fabrics and like potassium superoxide, is used in respiratory equipment worn by divers and firemen to get oxygen from carbon dioxide.

All forms of oxide from alkali metals readily react with water to form corresponding hydroxides which are white crystalline solids.

6.3 Complete the equations for each of the following reactions:

(a) $2Na + 2H_2O \longrightarrow \underset{\text{sodium hydroxide + hydrogen}}{\underline{\hspace{6cm}}}$

(b) $Li_2O + 2H_2O \longrightarrow \underset{\text{lithium hydroxide}}{\underline{\hspace{4cm}}}$

(c) $Na_2O_2 + 2H_2O \rightarrow \underset{\text{sodium hydroxide + hydrogen peroxide}}{\underline{\hspace{8cm}}}$

(d) $4KO_2 + 2H_2O \longrightarrow \underset{\text{potassium hydroxide + oxygen}}{\underline{\hspace{6cm}}}$

The alkali metal hydroxides are the strongest of all bases. The strength of the hydroxides as highly corrosive bases is reflected in their typical uses. Sodium hydroxide, for example, is used in a wide range of household products, such as oven and drain cleaners.

The alkali hydroxides readily dissolve in water, emitting large quantities of heat in the process. As strong bases, alkali hydroxides react with all acids to form salts.

6.4 Complete the equations for the reactions of alkali hydroxides with acids and name the products produced in each case.

(a) $2NaOH_{(aq)} + CO_{2(g)} \longrightarrow$ ______________

(b) $LiOH_{(aq)} + HCl\,(aq) \longrightarrow$ ______________

(c) $KOH_{(aq)} + HNO_{3(g)} \longrightarrow$ ______________

(d) $2NaOH_{(aq)} + H_2SO_{4(aq)} \longrightarrow$ ______________

6.5 Why is potassium superoxide used in respiratory equipment?

6.6 Name a household use of sodium hydroxide.

6.7 Explain in terms of its atomic structure why potassium is more reactive than lithium.

Alkaline Earth metals

The **Group II** elements are called the alkaline earth elements. They are all reactive elements, with reactivity increasing down the group.

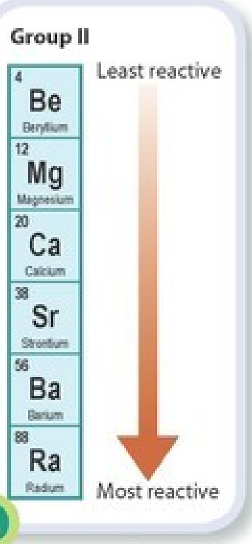


Figure 6.5 *Reactivity trend in Group II*

Physical properties

- They are metals which are harder than the alkali metals.
- They are shiny and silvery-white in appearance.
- They have low densities, low melting and low boiling points.

Chemical properties

- They are less reactive than the corresponding alkali metals. For example, magnesium reacts very slowly with water.
- They are reactive elements and readily lose their two outermost electrons to form cations with a 2^+ charge.

Group II oxides and hydroxides

The oxides of alkaline earth metals can be obtained by heating the metal in oxygen or by heating their carbonates to high temperatures. Their oxides are highly stable white crystalline solids with strong crystal lattice structures.

Calcium oxide occurs naturally in nature in the form of mineral lime, which is made of calcium, magnesium and aluminium carbonates, oxides and hydroxides.

Figure 6.6

The main method of extracting calcium oxide from mineral lime is through heating until the material decomposes into calcium oxide and carbon dioxide.

Calcium oxide has a wide variety of uses which include:

- use in cement and concrete production
- use in production of fertilisers
- use in some bleach products
- an additive in water purification and treatment processes

Hydroxides

All the alkaline earth metals, except beryllium, react with water to form strongly alkaline hydroxides which must be handled carefully.

Q Understanding U2

6.8 Complete the equations for the reactions below.

(a) $CaCO_{3(s)} \xrightarrow{heat}$ calcium oxide + carbon dioxide

(b) $MgO + H_2O_2 \longrightarrow$ magnesium hydroxide

(c) $CaO + H_2O_2 \longrightarrow$ calcium hydroxide

(d) $2Ca + O_2 \longrightarrow$ calcium oxide

Research R2 Research R3 Research R4

Lab Work 8

Portfolio 55

Activity 6.2

Question

How can we compare the reactivity of calcium and magnesium?

Equipment needed

Test-tube rack
4x test-tubes
Splint
2 x beakers
Hydrochloric acid 1 mol dm^{-3}
Small piece of magnesium ribbon
Small piece of calcium
Tongs

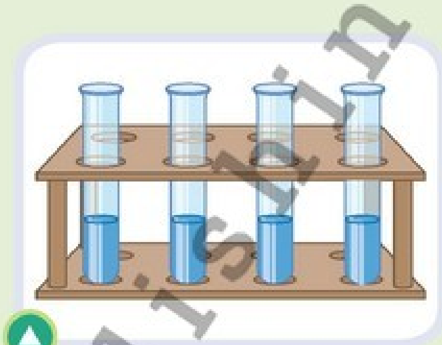


Figure 6.7

Safety

- Be careful when handling all solutions in this activity as they are corrosive.

Conducting the activity

Stage 1

1. Place a small piece of calcium in a half -filled test tube water.
2. Observe the reaction and test for hydrogen using a lighted splint.
3. Repeat steps 1 and 2 with a small piece of magnesium ribbon in water.
4. Record your observations.

Stage 2

1. Half-fill 2 test tubes with hydrochloric acid solution.
2. Place a small piece of calcium in the first, observe the reaction and test for hydrogen with a lighted splint.
3. Repeat steps 1 and 2 replacing a small piece of magnesium in half-filled test tube of hydrochloric acid solution.
4. Record your observations.

Q Understanding U2 Research R5

6.9 What can you conclude from your observations?

Properties of aluminium and its alloys

Q Understanding U_2

6.10 Look at the diagram of the atomic structure of aluminium and answer these questions.

(a) How many valency electrons does aluminium have?

(b) What is the atomic number of aluminium?

(c) What is the electronic configuration for aluminium?

Figure 6.8

There is more aluminium in the Earth's crust than any other metal. It is a silver-grey, soft malleable metal, with low density and a high melting point. It is a highly reactive metal and for this reason is found in nature mainly in compound form.

It has an extremely wide range of uses in manufacturing and engineering because it has high thermal conductivity, can be easily worked and machined and is lightweight and highly resistant to corrosion. These properties make it suitable for use in products such as kitchen foil, cans and window frames. Its alloys with silicon, copper and magnesium are used widely in the manufacture of parts for aircraft and other types of vehicle.

The non-corrosive properties of aluminium are due to an effect known as anodising. Aluminium forms a protective layer of aluminium oxide $Al_2O_3^-$ which prevents it from oxidising further. Anodised coatings with aluminium are used on all sorts surfaces from building exteriors to mobile phones.

Aluminium is a good reducing agent due to its most common +3 oxidation state, reducing $H^+_{(aq)}$ ions to $H_{2(g)}$:

$$\boxed{2Al_{(s)} + 6H_{+(aq)} \rightarrow 2Al_{3+(aq)} + 2H_{2(g)}}$$

Aluminium and its compounds are amphoteric, meaning that they react with both acids and bases. Aluminium hydroxide, for example, reacts

- with OH^- solutions to produce $Al(OH)^{4-}$
- with H_3O^+ solutions to produce $[Al(H_2O)_6]^{3+}$

In **Activity 6.3** we investigate how aluminium has reactions as both a base and an acid.

Lab Work 9

Activity 6.3

Question

How can we investigate whether aluminium reacts as both a base and an acid?

You will observe the reactions of dropping a ball of aluminium foil into a beaker of

(a) a dilute solution of hydrochloric acid
(b) a dilute solution of sodium hydroxide

1. What would be the dangers of carrying out these experiments yourself?
2. Can you predict what the products of these reactions will be?
3. In which of these two experiments does the metal behave more as you would expect it to?

6.11 Complete the equations for the two reactions:

$2Al + 6HCl \longrightarrow$ ______________

$2Al + 2NaOH + 6H_2O \longrightarrow 2Na[Al(OH)_4] +$ ____________

Practical Work 4

Activity 6.4

In this task you are going to devise an experiment to show how you could work out where the metals magnesium, sodium, calcium and aluminium belong, in the incomplete reactivity series in Figure 6.9.

Work in groups and decide on:

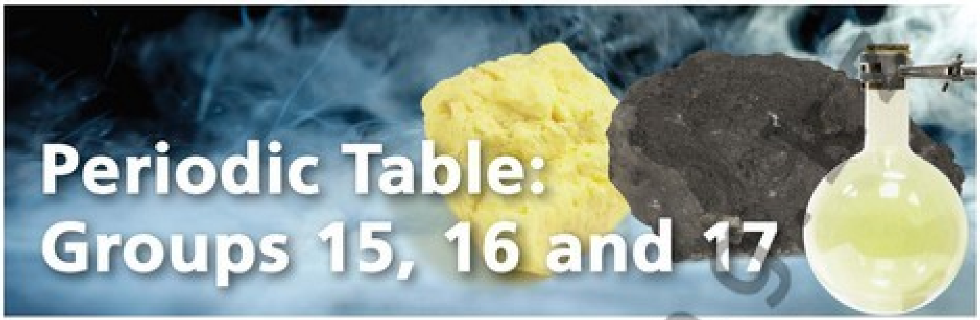
- a hypothesis
- chemicals needed
- equipment needed
- procedure

You will present your ideas to another group.

potassium	most reactive	K
________	↑	___
________		___
________		___
________		___
carbon		C
zinc		Zn
iron		Fe
tin		Sn
lead		Pb
hydrogen		H
copper		Cu
silver		Ag
gold	↓	Au
platinum	least reactive	Pt

Figure 6.9

MODULE 7



Periodic Table: Groups 15, 16 and 17

Learning outcomes

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

- Predict periodic trends for elements in the halogen group 9.2.1.8
- Describe chemical properties of chlorine; interaction with metals, hydrogen and halides 9.2.1.9
- Explore chemical properties of hydrochloric acid solution 9.2.1.10
- Describe common characteristics of elements in Group 16 9.2.1.11
- Compare the physical properties of allotropic forms of sulfur and make equations for reactions involving sulfur 9.2.1.12
- Compare the physical and chemical properties of sulfur oxide (IV) and sulfur oxide (VI) and explain the physiological effect of sulfur dioxide 9.2.1.13
- Explain acid rain as a phenomenon and its environmental impact 9.4.2.1
- Explore the physical and chemical properties of sulfuric acid solution and its salts 9.2.1.14
- Calculate percentage yield from the theoretically possible 9.2.3.3
- Explain the properties of nitrogen and the nitrogen cycle 9.2.1.15
- Explain the molecular, ionic and structural formulae of ammonia 9.1.4.6
- Describe the properties, preparation and uses of ammonia 9.2.1.16
- Produce ammonia by means of the interaction of ammonium salts solution with alkali solution and explore the properties of gaseous ammonia and its solution 9.2.1.17
- Explain the industrial process of ammonia production 9.3.3.5
- Understand the molecular formulae of nitric acid and explain the formation of chemical bonding between atoms 9.1.4.7
- Make equations for the production of nitric acid from nitrogen 9.2.1.18
- Compare the properties of nitric acid with other common acids through investigation 9.2.1.19
- Describe and make equations for the reaction of dilute and concentrated nitric acid with metals 9.2.1.20
- Explain and make an equation for the thermal decomposition of nitrates 9.2.1.21
- Compare the allotrope forms of phosphorous 9.2.1.22
- Identify the names of the deposits of phosphorous compounds found in Kazakhstan 9.4.2.2
- Explain the common chemical properties of phosphorous and its compounds 9.2.1.23
- Classify mineral fertilisers and nutrients according to their chemical composition 9.4.2.3
- Assess the impact of nitric and phosphorous fertilizers on the environment 9.4.2.4

Keywords

✓ diatomic molecules ✓ outer shell ✓ atomic weight ✓ ionic crystals
✓ ionised ✓ allotropes ✓ crystalline ✓ amorphous ✓ contact process
✓ dehydration ✓ inhalation ✓ fractional distillation ✓ nitrogen fixation

The Halogens

The elements in Group 17 of the Periodic Table are called the halogens. Fluorine and chlorine (F_2 and Cl_2) are gases at STP; bromine (Br_2) is a liquid and iodine (I_2) is a solid.

The halogens are reactive non-metals with seven valence electrons.

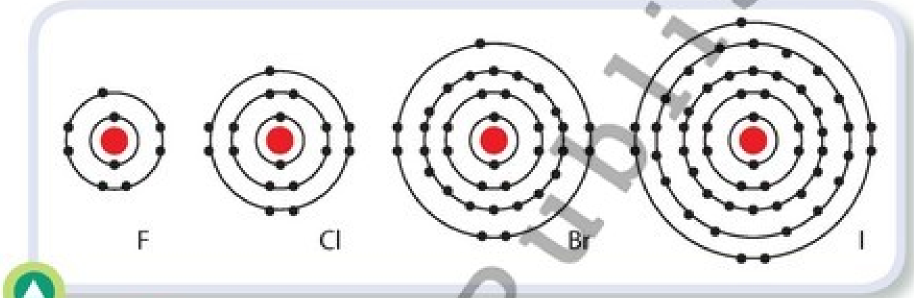


Figure 7.1

As pure elements, they form diatomic molecules with atoms joined by non-polar covalent bonds. By sharing electrons in a covalent bond they achieve full outer electron shells.

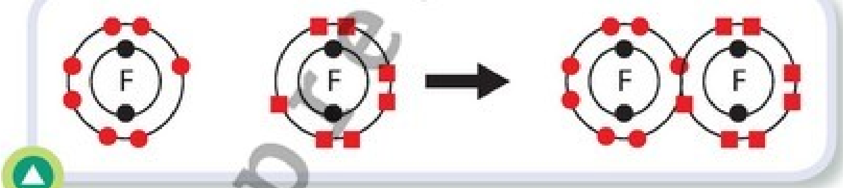
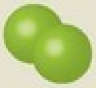
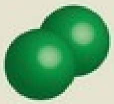
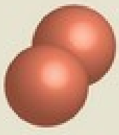
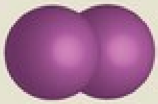


Figure 7.2

Table 7.1

Halogen	Fluorine	Chlorine	Bromine	Iodine
molecule	F_2	Cl_2	Br_2	I_2
model				

The formulas, diatomic structure and relative size of the molecule of the first four members of the halogen group are illustrated in the table. The size of each molecule relates to the number of electron bands in each atom and the atomic weight relating to the number of electrons in each atom. The illustration also shows the colours of the halogens and how they become progressively darker down the group.

7.1 Would you expect the melting and boiling points of the halogens to increase or decrease moving down the group?

7.2 Would you expect the reactivity of the halogens to increase or decrease moving down the group?

7.3 What colour would you expect Astatine the fifth halogen in the group to be?

Physical and chemical properties of halogens

As is typical of non-metals, the melting and boiling points of the halogens are low.

Table 7.2

Halogen	Melting Point (°C)	Boiling Point (°C)
Fluorine	–220	–188
Chlorine	–101	–35
Bromine	–7.2	58.8
Iodine	114	184

Halogens are highly reactive with alkali metals and alkaline earth metals and form stable ionic crystals. The reactivity of the elements decreases moving down the group. We shall take chlorine as an example and look at its interactions with metals, hydrogen and halides.

Chlorine along with the other halogens reacts vigorously with alkali metals to form white salts, for example:

sodium + chlorine → sodium chloride

Chlorine forms strong ionic bonds with metal ions. It forms –1 ions in ionic compounds.

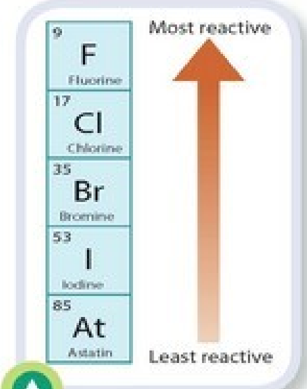


Figure 7.3

7.4 Complete the equations and name the product in each case:

(a) $2Na + Cl_2 \rightarrow$

(b) $Ca + Cl_2 \rightarrow$

(c) $2Al + 3Cl_2 \rightarrow$

(d) $Mg + Cl_2 \rightarrow$

In the halogen group the oxidising power of the elements increases moving down the group, which means that chlorine is capable of releasing bromine and iodine respectively from solutions of their salts:

$Cl_{2(aq)} + 2Br^{-}_{(aq)} \rightarrow 2Cl^{-}_{(aq)} + Br_{2(aq)}$

Reduction: $Cl_{2(aq)} + 2e^{-} \rightarrow 2Cl^{-}_{(aq)}$

Oxidation: $2Br^{-}_{(aq)} - 2e^{-} \rightarrow Br_{2(aq)}$

$Cl_{2(aq)} + 2I^{-}_{(aq)} \rightarrow 2Cl^{-}_{(aq)} + I_{2(aq)}$

Reduction: $Cl_{2(aq)} + 2e^{-} \rightarrow 2Cl^{-}_{(aq)}$

Oxidation: $2I^{-}_{(aq)} - 2e^{-} \rightarrow I_{2(aq)}$

Figure 7.4 *Magnesium chloride*

Hydrochloric acid

There are different ways in which the gas hydrogen chloride can be formed. When hydrogen gas and chlorine gas are exposed to sunlight or temperatures above 250° C, they react violently to produce hydrogen chloride. The gas is also commonly produced through the reaction of sodium chloride and sulfuric acid.

7.5 Complete the two equations:

(a) $H_2 + Cl_2 \rightarrow$

(b) $NaCl + H_2SO_4 \rightarrow$ ___________ $+ HCl$

Hydrogen chloride dissolves in water to produce hydrochloric acid. The acid is fully ionised in solution so it is a strong acid and care must be taken when handling it.

Lab Work 10

Activity 7.1

Question

How can we analyse the properties of hydrochloric acid?

Equipment needed

4 x test tubes
test tube rack
2 x clamps and stands
delivery tube and test tube stopper
hydrochloric acid
universal indicator
magnesium strip
sodium hydroxide solution
splint
matches
marble chips

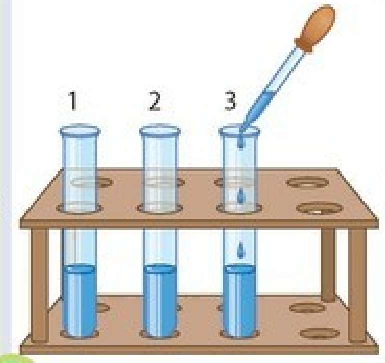


Figure 7.5

Safety

- Take care when handling the hydrochloric acid, and wear protective goggles at all times.

Conducting the activity

1. Pour hydrochloric acid into three test test-tubes so that each test tube is one-third full.

Test 1

2. To test tube 1 add a few drops of universal indicator solution. Record your observations.

Test 2

3. Now add sodium hydroxide solution slowly to test tube 1 using a pipette. Make sure to shake the mixture carefully each time you add the sodium hydroxide solution. Continue adding until the colour of the solution changes to purple.

Test 3

4. Add a strip of magnesium to test tube 2. Record your observations. Collect the gas given off by holding an inverted test tube over the mouth of the test tube and then place a lit splint to the mouth of the inverted tube. Record your observations.

Test 4

5. Place some marble chips into test tube 3. Add a stopper with a delivery tube connected to test tube four containing limewater, as shown in the diagram. Record your observations.

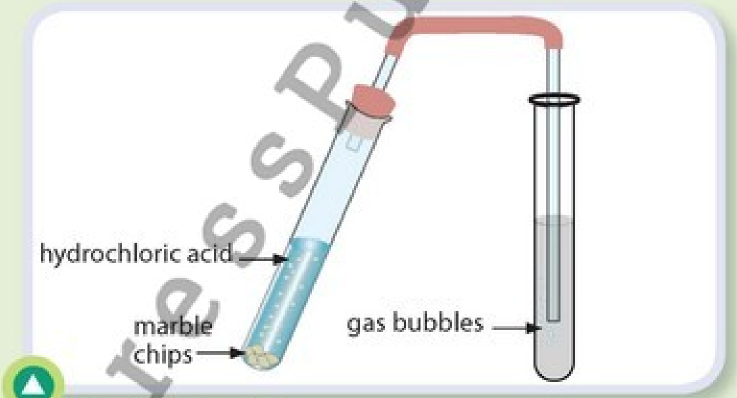


Figure 7.6 *Delivery tube*

	Observations
Test 1	
Test 2	
Test 3	
Test 4	

7.6 What can you conclude about the test with the universal indicator?

7.7 Which gas was produced in test 3?

7.8 Which gas was produced in test 4?

7.9 In test 2, what is the name for the reaction that turned the universal indicator green?

7.10 What can you conclude about the properties of hydrochloric acid overall from the activity.

7.11 Complete these equations of the reactions involved in the different tests:

(a) $NaOH + HCl \rightarrow$

(b) $Mg + 2HCl \rightarrow$

(c) $CaCO_3 + HCl \rightarrow$

(d) $Ca(OH)_{2(aq)} + CO_{2(g)} \rightarrow$

7.12 Name the products of each reaction and in reaction d include the states of the products: $_{(g)\ (s)\ (l)\ (aq)}$

Group 16 Elements

O
S
Se
Te
Pb

Group 16 of the periodic table consists of 5 elements: oxygen, sulfur, selenium, tellurium and polonium.

Going down the group, oxygen and sulfur are non-metals, selenium and tellurium are metalloids and polonium, which is a radioactive element, is a metal. There are allotropes of each element.

The allotropic forms of oxygen, for example, are:

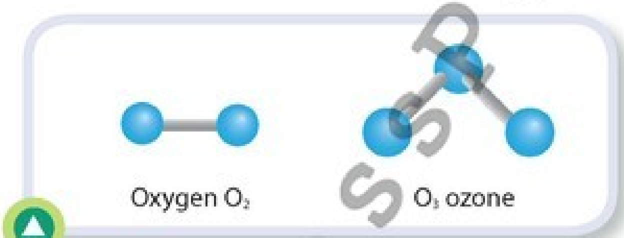


Figure 7.7

Allotropes of sulfur

There are three allotropes of sulfur. The crystalline forms: orthorhombic carbon sulfur and monoclinic sulfur and the amorphous form: amorphous [or plastic] sulfur. The structural arrangement of the atoms is shown in Figure 7.8.

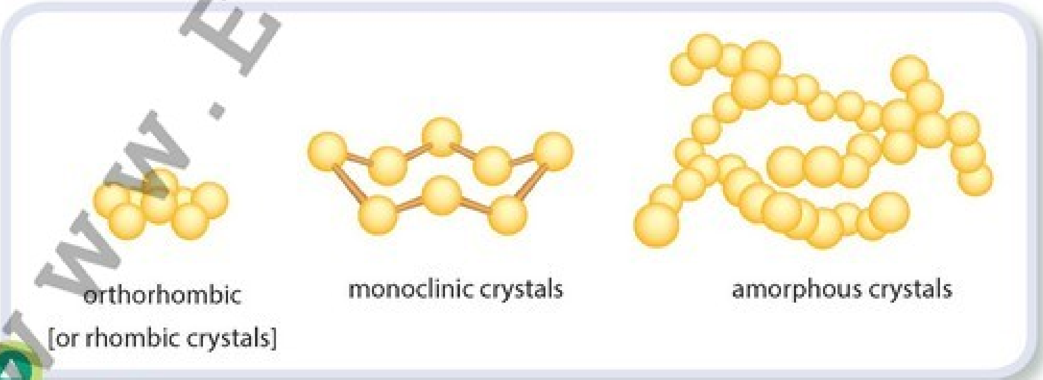


Figure 7.8

The crystalline structures of sulfur are made from octatomic molecules known as the S_8 ring. In the amorphous form the S_8 rings are broken open by heat and form long chains. Table 7.3. highlights the key differences between each allotrope.

Table 7.3

Allotrope	Temperature range	Physical features
Rhombic crystal form	stable below 96°C	• yellow, brittle solid
Monoclinic crystals	stable between 96°C – 119°C below 90°C returns to rhombic form	• found naturally as molten sulfur • yellow solid, needle-like crystals
Plastic sulfur	heated to above 200°C super-cooled by quenching with water	• dark viscous liquid at above 200 °C • soft rubber-like dark brown material when supercooled

Activity 7.2

Your teacher will show a video or demonstrate the changes in colour and form of the allotropes of sulfur.

7.13 Why is the rhombic crystal allotrope of sulfur its most common form?

7.14 How do you explain the difference in the crystalline and amorphous forms of sulfur?

Sulfur (IV) oxide and Sulfur (VI) oxide

Sulfur (IV) oxide is the more common oxide. It is formed when sulfur burns in an excess supply of air or oxygen. It is a toxic gas at room temperature with a strong smell similar to the smell of lighting a match. Sulfur (IV) oxide – more commonly known as sulfur dioxide – is a simple covalent structure made up of 1 sulfur and 2 oxygen atoms.

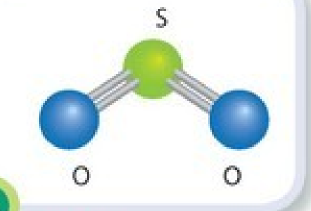


Figure 7.9

7.15 Complete the equation for the reaction between sulfur and oxygen:

$S + O_2 \rightarrow$

It dissolves in water to form a weak acid oxide called sulfurous acid:

$SO_2 + H_2O \rightarrow H_2SO_3$

It reacts with bases to form sulfite salts and water.

7.16 Complete the reaction equation

$SO_2 + 2NaOH \rightarrow$

Did you know?

In sulfites there are three oxygen atoms and one sulfur atom. In sulfates there are 4 oxygen atoms and 1 sulfur atom.

Sulfur (VI) oxide, also known as sulfur trioxide, is a volatile white solid at room temperature. It is a simple covalent structure made up of 1 sulfur atom and three oxygen atoms.

It is produced on an industrial scale using a process known as the contact process, an oxidation reaction between sulfur (IV) oxide and oxygen using a solid catalyst vanadium oxide:

$$\boxed{2SO_2 + O_2 \xrightarrow{V2O5} 2SO_3}$$

Sulfuric acid in made from sulfur (VI) oxide which is a very strong acidic oxide. Sulfuric acid is formed by dissolving sulfur (VI) oxide in water. It dissolves exothermically producing dangerous sulfuric acid mist and water vapour.

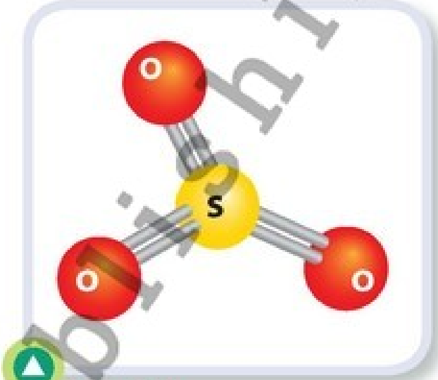


Figure 7.10

Sulfur (VI) oxide reacts with bases to form sulfate salts and water.

Q U₂ Understanding

7.17 Complete the equation for reaction of sulfuric acid and magnesium hydroxide.

$H_2SO_{4(l)}$ +$Mg(OH)_{2(s)}$ →

7.18 What are the names of the products of this reaction?

Acid Rain

Unpolluted rainwater is slightly acidic. It has a pH of approximately 5.7 due to the acidic gas, carbon dioxide, in the atmosphere. Emissions of other acidic gases into the atmosphere lead to a lowering of pH, with the formation of what is referred to as **acid rain.** This acid rain damages soil, causes the poisoning of fish, attacks trees and erodes buildings. The main causes of acid rain are oxides of nitrogen and oxides of sulfur.

Figure 7.11 *Trees damaged by acid rain*

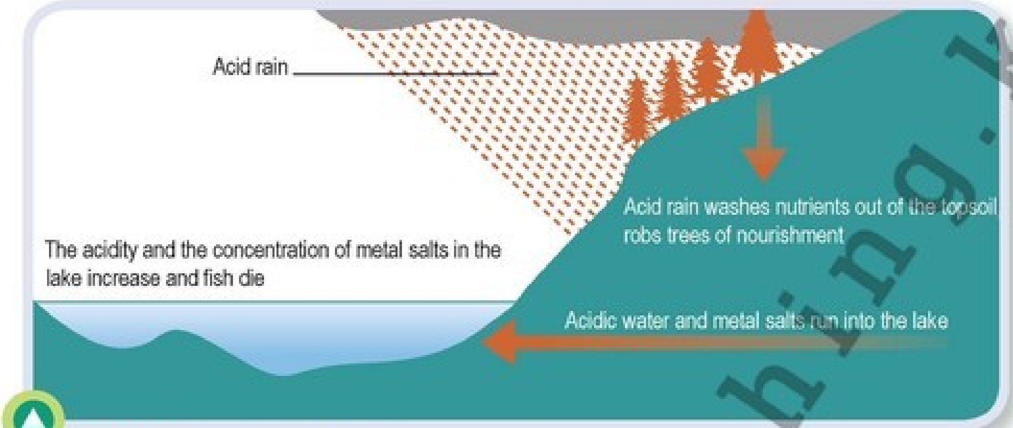


Figure 7.12 *How acid rain kills fish*

Figure 7.13 *Volcanoes release sulfur dioxide into the atmosphere*

Oxides of sulfur are mainly formed by the combustion of fossil fuels, particularly coal. They can also be released by volcanoes and by the decay of organic matter.

Sulfur in fossil fuels forms sulfur dioxide, SO_2, which is a dangerous pollutant. Sulfur dioxide dissolves in water to form sulfurous acid, H_2SO_3

$$SO_{2(g)} + H_2O_{(l)} \rightarrow H_2SO_{3(aq)}$$

In the atmosphere, sulfur dioxide is oxidised to sulfur trioxide, SO_3. Sulfur trioxide dissolves in rainwater to form sulfuric acid, H_2SO_4.

$$SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$$

It is thus the release of SO_2 into the atmosphere that results in rain containing H_2SO_3 and H_2SO_4.

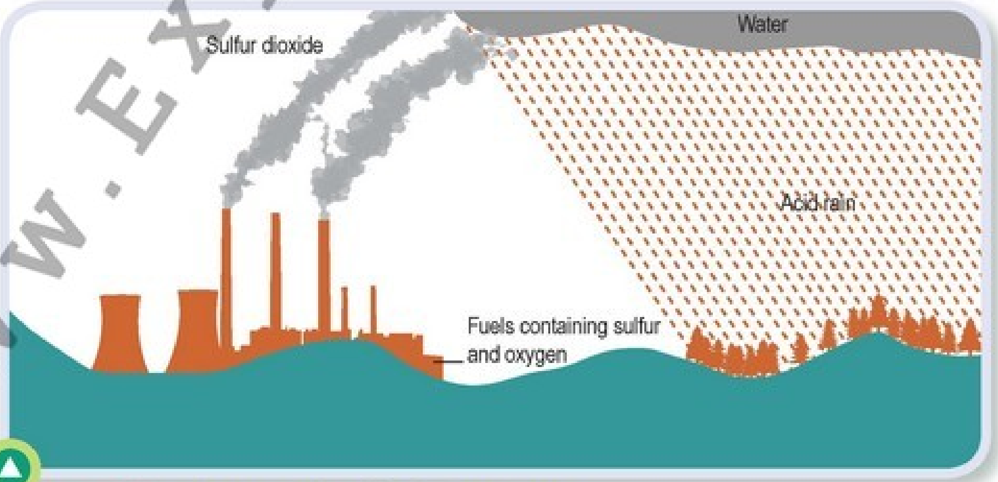


Figure 7.14 *Acid rain caused by sulfur dioxide emissions*

Uses and properties of sulfuric acid

Sulfuric acid is a transparent, oily liquid with no discernible smell. It has a wide range of industrial uses and applications in industry such as the treating of ores and metals and the making of fertilisers such as phosphate fertilisers and ammonium sulfate. Sulfuric acid has a strong affinity for water and is thus used in dehydration processes such as the dehydration of ethanol and to dehydrate hydrated crystals.

Practical Work 5

Activity 7.3

Question

How can we investigate the properties of sulfuric acid?

Equipment needed

beaker
diluted sulfuric acid
pipette
3 sugar cubes
petri dish 2 x test tubes zinc granules splint and matches

Safety

- Be extremely cautious when handling dilute sulfuric acid as it is extremely corrosive and burns skin.

Conducting the activity

Part A

1. This part of the activity should be performed in a fume cupboard.
2. Place three sugar cubes in a pyramid form in a petri dish and using a pipette drop several drops of sulfuric acid onto the cubes.
3. Observe the reaction and record your results.

Part B

4. Pour dilute sulfuric acid into a test tube held in a test tube rack until it is one-third full.
5. Add some zinc granules and observe the reaction.
6. Place the mouth of another test tube over this test tube for five seconds and then test the gas collected with a lighted splint.

7.19 Is the reaction you observed in Part A exothermic or endothermic?

7.20 What were the products produced in the reaction?

7.21 What was the vapour you observed in the reaction?

7.22 What is the word to describe what sulfuric acid does to sugar? Refer to the chemical formula for sucrose: $C_{12}H_{22}O_{11}$ in your explanation.

7.23 What can you conclude from your observations of the reaction in Part B?

7.24 Complete the equation for the reaction in Part B

$Zn + H_2SO_4 \rightarrow$

7.25 What were the products of this reaction?

The reaction observed in Part A actually involved two phases. Firstly the sucrose is dehydrated by the addition of sulfuric acid acting as a catalyst and then because sulfuric acid is a strong oxidising agent, the carbon produced in the dehydration phase is oxidised to CO_2 while the sulfuric acid is reduced to SO_2.

First phase:

$$\rightarrow 12C_{(s)} + 11\ H_2O_{(l)}$$

Second phase:

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$$

Overall equation:

$$C_{12}H_{22}O_{11(s)} + H_2SO_{4(aq)} + 1/2\ O_{2(g)} \rightarrow 11C_{(s)} + CO_{2(g)} + 12H_2O_{(g)} + SO_{2(g)}$$

Group 15 Elements

Nitrogen

Nitrogen is a colourless and odourless diatomic gas with the formula N_2. It is the lightest element in **Group 15** of the period table.

Nitrogen, at 78%, is by far the most abundant substance in the atmosphere. It occurs free in nature as the diatomic gas, N_2, with the two nitrogen atoms joined by a non-polar covalent triple bond.

Lab Work 11

The bond energy of the N ≡ N bond is very large, which means that a lot of energy is needed to break the bond. Consequently, nitrogen is very stable and chemically unreactive. Nitrogen is a co-product of the fractional distillation of air.

Uses of nitrogen

Figure 7.15 *Nitrogen is used to keep foods fresh*

When oil tankers empty their cargoes, nitrogen gas, being inert, is used to flush out the explosive vapours from their tanks. The tanks remain full of nitrogen until they are refilled with their next cargo of oil.

Food is kept fresh by storing it in nitrogen. In the consequent absence of oxygen, micro-organisms are unable to cause food to go off. Thus the food remains fresh for a longer period.

Because of its extremely low boiling point, liquid nitrogen is used as a coolant and refrigerant. For example, in medicine it is used for the long-term storage of viable human cells such as sperm.

Figure 7.16 *Liquid nitrogen being sprayed onto a patient's hand.*

It is also used in the treatment of certain skin conditions where it is sprayed onto damaged skin cells. The extremely low temperature kills the damaged cells, allowing new, healthy cells to grow in their place.

Nitrogen Fixation

Nitrogen fixation is the conversion of atmospheric nitrogen to compounds that can be used by plants.

Nitrogen is essential for the growth of most living things, and is a vital ingredient of proteins.

In order for plants and animals to make use of nitrogen, the unreactive nitrogen gas in the atmosphere must be converted into more reactive nitrogen compounds, which can then find their way into the soil. This process is called **nitrogen fixation.**

One mechanism for this occurs during thunderstorms, when lightning flashes provide the high energies necessary to split the N_2 bond and allow it to react with oxygen to form nitrogen oxides, such as NO and NO_2. These can then dissolve in the water in clouds to form nitric acid (HNO_3) and nitrous acid (HNO_2). These compounds are changed into nitrates in the soil that are then absorbed into the root systems of plants.

The roots of legumes (plants such as clover, peas, and beans) are colonised by nitrogen-fixing bacteria that have the ability to directly convert N_2 into useful nitrogen compounds via the use of a special enzyme called nitrogenase. The virtue of this system is that these plants provide their own fertiliser.

Figure 7.17 *Lightning provides enough energy to allow nitrogen to be changed into nitrogen oxides*

Figure 7.18 *A clover plant can fix nitrogen directly from nitrogen in the air*

The Nitrogen cycle

The nitrogen cycle indicates why the level of nitrogen in the atmosphere does not change. Some nitrogen is removed from the atmosphere by bacterial nitrogen fixation. More is removed by the action of lightning followed by rain. Some atmospheric nitrogen is used in the manufacture of artificial fertilisers. However, certain bacteria convert nitrates in soil into nitrogen gas, which is released back into the atmosphere to begin the nitrogen cycle again.

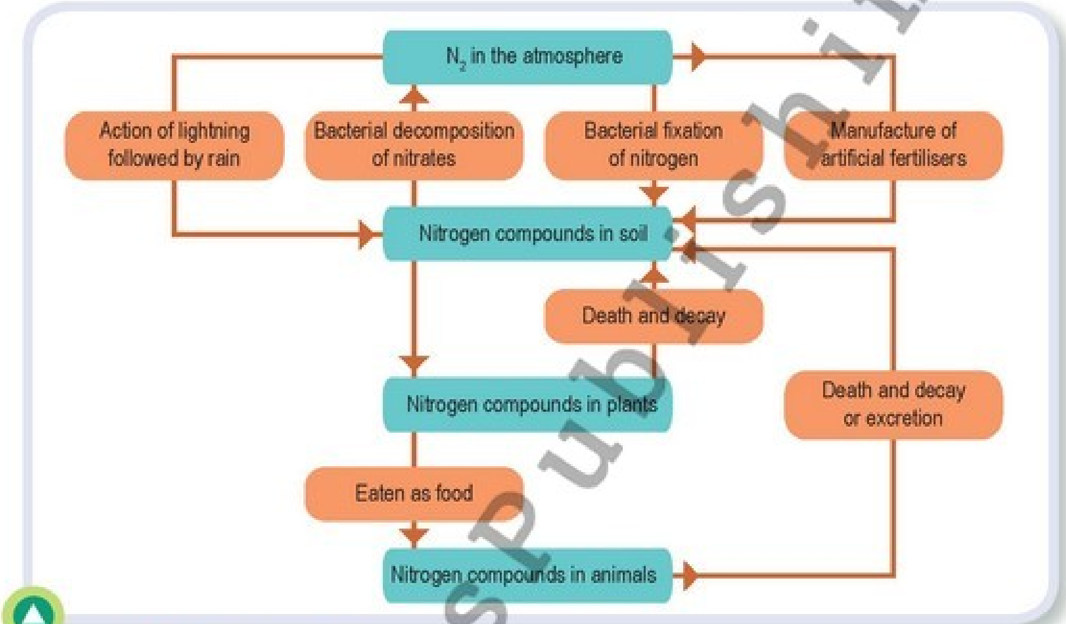


Figure 7.19 *The nitrogen cycle*

7.26 Write the formula of nitrogen gas and explain the structure of the molecule.

7.27 Give two uses each of **(a)** nitrogen gas **(b)** liquid nitrogen.

7.28 Explain the term nitrogen fixation.

7.29 What is the overall effect of the nitrogen cycle?

Ammonia

In the ammonia molecule there are three hydrogen atoms and an unshared pair of electrons attached to the nitrogen atom. Its formula is NH_3.

Ammonia is a colourless gas with a pungent smell. NH_3 fumes are toxic and exposure to and inhalation of ammonia can seriously affect your eyes, nose and throat.

Lab Work 12

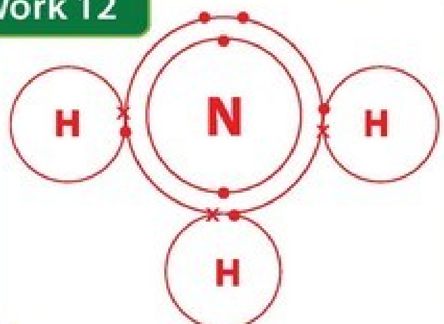


Figure 7.20

Ammonia has a very low boiling point of –33.35 °C, and a freezing point of –77.7 °C. In liquefied form it is stored at a low temperature and high pressure because of its low boiling point.

Ammonia readily dissolves in water in an exothermic reaction. It forms ammonium hydroxide (NH_4OH) which is a weak base solution.

Activity 7.4

Question

How can we prepare ammonia and investigate its properties?

Equipment needed

- 1 x small beaker
- 1 x large beaker
- Spatula
- Retort stand clamp
- Boiling tube with stopper
- Delivery tube
- 2 x test tubes [one with stopper]
- Test tube rack
- Retort stand
- Clamp
- Bunsen burner
- Heat resistant mat
- Ammonium chloride
- Calcium hydroxide
- Calcium oxide
- Concentrated hydrochloric acid
- Universal indicator paper

Safety

- The activity should be carried out in a fume cupboard or well-ventilated laboratory. Students should wear safety goggles at all times.

Conducting the activity

1. Have ready a large beaker two-thirds filled with water. Also have ready a test tube with a few drops of hydrochloric acid with a stopper in it. You will need these later in the experiment.
2. Mix 2 spatulas of ammonium chloride with an equal amount of calcium hydroxide in beaker. You should observe an immediate reaction as they begin to mix.

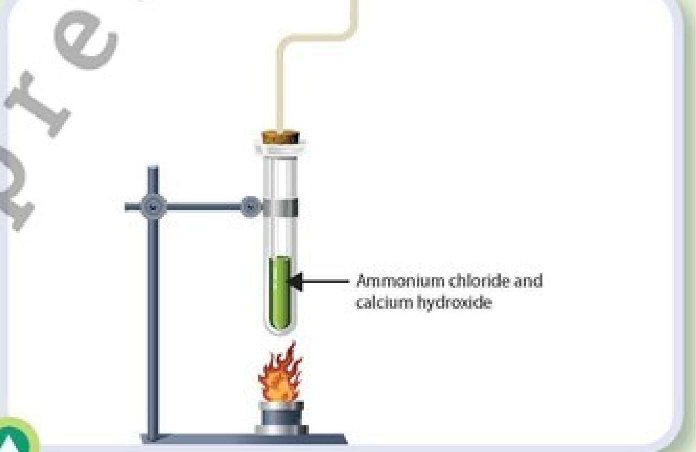


Figure 7.21

3. Place a piece of universal indicator paper over the beaker and record the change in colour.
4. Transfer the contents of the beaker into a boiling tube and add a small amount of calcium oxide [to absorb any water produced in the reaction].
5. Place the boiling tube containing the ammonium chloride and calcium hydroxide mixture in a clamp as shown in **Figure 7.21** and insert a delivery tube and stopper.

6. Heat the mixture gently.
7. Take the test tube containing the hydrochloric acid prepared earlier. Remove the stopper and place near the mouth of the delivery tube. Record your observations and then replace the stopper and return to the rack.
8. Continue heating and then hold another boiling tube inverted over the delivery tube as shown in **Figure 7.22**.
9. Remove the boiling tube holding it upside down and quickly place mouth down in the beaker of water prepared earlier. Record your observations.

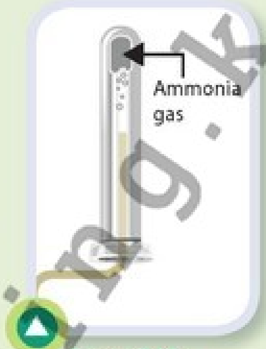


Figure 7.22

Q Understanding U2 Research R4 Research R5

7.30 What did the test with universal indicator paper show?

7.31 Are alkaline gases common?

7.32 Why was it important for the success of the experiment to add calcium oxide?

7.33 What did you observe when placing the hydrochloric acid near the delivery tube?

7.34 Can you complete the equation for the reaction you observed at this stage:
$NH_{3(g)} + HCl_{(g)} \rightarrow$

7.35 What did you observe in the last stage of the experiment when the boiling tube of ammonia was placed upside down in the beaker of water? What can you conclude from this?

7.36 Complete the equation for the reaction used in this experiment to produce ammonia:
$2NH_4Cl_{(s)} + Ca(OH)_{2(s)} \rightarrow$

7.37 What type of reaction is this?

The Haber process

The synthesis of ammonia from its elements:

$$\boxed{N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}}$$

is an important process in the fertiliser industry. The process used to make ammonia is called the Haber process.

In the industrial synthesis of ammonia, the objective is to produce the maximum possible amount of ammonia at the lowest cost and in the shortest possible time. A balance has to be struck between factors that favour a reasonable rate of reaction and those that favour a reasonable yield.

Temperature

$$\boxed{N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)} \qquad \Delta H = -92\ kJ\,mol^{-1}}$$

Since the forward reaction is exothermic, it is favoured by a lowering of the reaction temperature at equilibrium. The extent to which this can be done is limited by the fact that the temperature must be high enough to allow the reaction to proceed at a reasonable rate. If the temperature is too low, not enough nitrogen and hydrogen molecules will have the activation energy for the reaction, and the reaction rate will be too slow. In practice, temperatures of about 673 K are used. This results in a low yield, but unreacted nitrogen and hydrogen can be collected and sent back into the reaction vessel again and again.

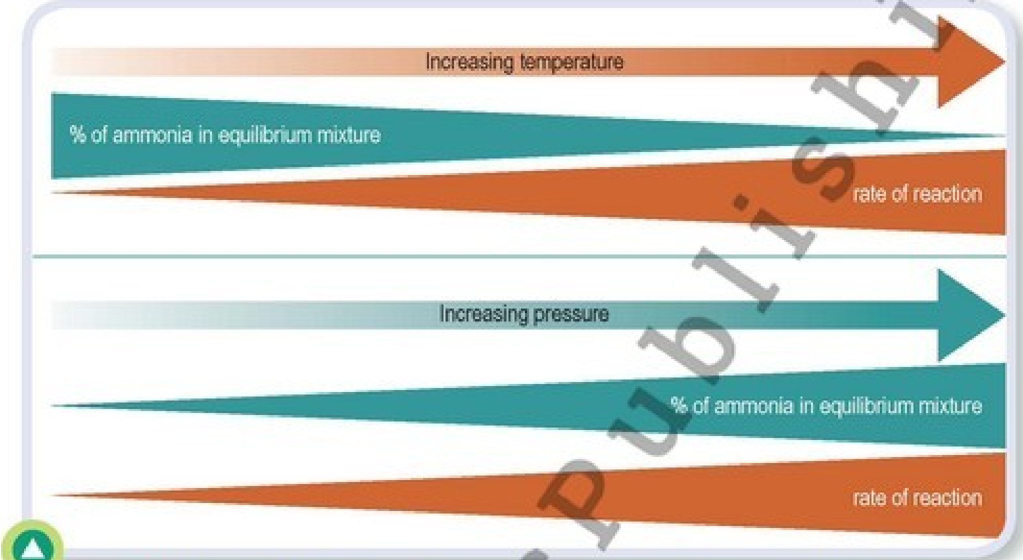


Figure 7.23 *A high yield of ammonia is favoured by (a) lower temperatures and (b) higher pressures*

An increase in pressure at equilibrium favours the formation of ammonia. For this reason the reaction is carried out under pressure. The Haber process uses high pressures (about 200 atmospheres) to increase the yield of ammonia.

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

4 moles 2 moles

Around 80% of the ammonia produced in industrial plants like the one in **Figure 7.24**, is used as agricultural fertilizer. Another important use, because of its extremely low boiling point, is as a refrigerant gas. In manufacturing it is used to make textiles, plastics and explosives. You will also find it in low strength cleaning products at home.

Nitric acid is a clear liquid either colourless or slightly yellow in colour. It is highly corrosive and has a pungent, choking smell.

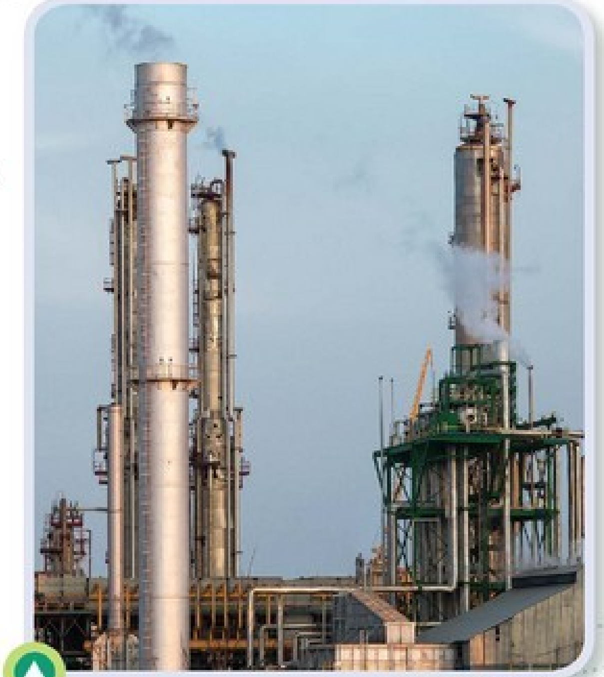


Figure 7.24

The structure of the HNO_3 molecule is shown in **Figure 7.25**. The nitric acid molecule contains a hydrogen atom, a nitrogen atom with a positive charge and three oxygen atoms – one of which carries a negative charge. The nitrogen atom is bonded to all three oxygen atoms.

Pure nitric acid is liquid at room temperature but becomes a white solid below –41°C and it boils at 83°C. Nitric acid is miscible in water forming hydrates at a low temperature.

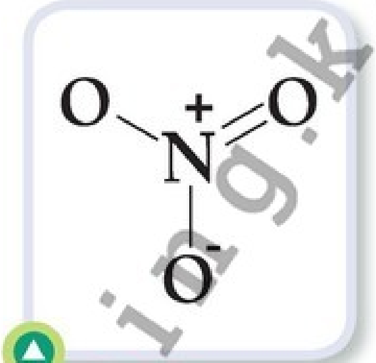


Figure 7.25

There are two common methods for the production of nitric acid.

In the laboratory it can be prepared by the reaction of nitrogen dioxide with water.

$$\boxed{3NO_2 + H_2O \rightarrow 2HNO_3 + NO}$$

One industrial process for the production of nitric acid is a process known as the Oswald process. This involves four stages to produce a concentrated form of nitric acid.

Stage 1:

$$\boxed{4NH_3 + 5O_2 \leftrightarrow 4NO + 6H_2O}$$

In this reaction anhydrous ammonia is oxidised to nitric oxide at high temperatures using a platinum catalyst.

Stage 2:

$$\boxed{2NO + O_2 \leftrightarrow 2NO_2}$$

Nitric oxide from Stage 1 is cooled and oxidises to form nitrogen dioxide.

Stage 3:

$$\boxed{3NO_2 + H_2O \rightarrow 2HNO_3 + NO}$$

The nitrogen dioxide obtained in Stage 2 is reacted in with water in a special chamber. As the nitrogen dioxide is absorbed into the water nitric acid is obtained.

Stage 4:

To increase the concentration of the nitric acid obtained in Stage 3 HNO_3, vapours are passed over concentrated H_2SO_4. You will remember the dehydrating properties of sulfuric acid from earlier in this module.

The main uses of nitric acid are in the production of:

- fertilisers
- chemical dyes
- nitro-organic products such as explosives e.g. nitroglycerin
- plastics

Research R2 Research R3 Research R4

Portfolio 71

Activity 7.5

Question

How do acids differ in their reactivity with metals such as copper?

Equipment needed

1 x 500 ml florence flask
2 x small beakers
ring stand
large ring placed below the flask
small ring over the neck of the flask to hold it in place
one-hole rubber stopper
delivery tube
large container of water
6 pieces of copper wire or six copper coins
dilute nitric acid, hydrochloric acid and sulfuric acid

Figure 7.26

Safety

- Be extremely careful when handling the acid solutions as they are all corrosive and can burn skin.

Conducting the activity

1. Place the 2 pieces of copper wire in one beaker and cover with hydrochloric acid. Observe the reaction and record the results.
2. Place 2 pieces of copper wire in the other beaker and cover with sulfuric acid. Observe the reaction and record results.
3. Set the equipment up as shown below. Place two pieces of copper wire into the Florence flask and add the concentrated nitric acid solution.
4. As the gas given off bubbles into the water container, stir so that the gas dissolves in the water.
5. Record your observations.

Q Understanding U2 Research R5

7.38 What can you conclude about the reaction of copper with hydrochloric acid and sulfuric acid?

7.39 What gas was produced when copper was added to nitric acid?

7.40 What type of reaction was this?

7.41 What did you observe when the gas in the flask began to cool?

7.42 How can you explain this?

The experiment showed that nitric acid is a powerful oxidizing agent. It reacts violently with many non-metallic compounds and it also reacts with metals dissolving them and forming metal oxides.

7.43 Complete the equation for the reaction in the experiment:

$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 +$ ____________

Nitrates and Nitrites

The difference between a nitrate and a nitrite compound is how many oxygen atoms each compound contains. Nitrites have two oxygen atoms and one nitrogen, whereas nitrates have three oxygen atoms.

7.44 Complete these other equations for other common reactions with nitrates and nitrites and name the products

(a) $Cu_{(s)} + AgNO_{3(aq)} \rightarrow$ ____________ + ____________

(b) $3PbO + 2NH_3 \rightarrow$ ____________ + ____________ + water

(c) $3Mg_{(s)} + N_{2(g)} \xrightarrow{\text{Burning}}$ ____________

(d) $MgN_{2(s)} + 6H_2O_{(l)} \rightarrow 3Mg(OH)_{2(aq)} +$ ____________

Phosphorous

There are many different allotropes of phosphorus. Their structural properties give them different degrees of stability and reactivity: for example, their reactions with the air. We shall focus on the two most common ones here: white phosphorus and red phosphorus.

Figure 7.27

White phosphorous which has the appearance of a waxy, yellowy white, solid is the most common form. It is highly reactive and will catch fire spontaneously in the air so it kept under water. Red phosphorus which is much less reactive is best known as one of the components in the red tip of matches.

White phosphorus consists of P_4 molecules, whereas the crystal structure of red phosphorus has a complicated network of bonding. White phosphorus has to be stored in water to prevent natural combustion, but red phosphorus is stable in air.

As **Figure 7.28** shows, white phosphorus atoms bond covalently with the other three atoms in the molecule with single bonds.

Figure 7.28

Red phosphorous has the same tetrahedral structure as white phosphorous but has linked single p-p bonds. The formula for both types of phosphorous is given as P_4 which points to red phosphorus as being more of an intermediary form than a true allotrope.

Red phosphorus can be converted to white phosphorus by careful heating.

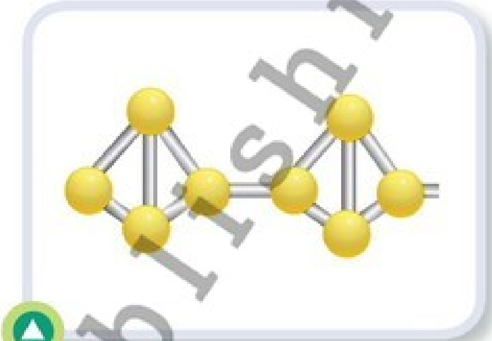


Figure 7.29

The most common phosphates are compounds of phosphate ions in combination with one or more common elements, such as aluminium, sodium, potassium and calcium. Phosphates are used in a wide range of products:

sodium phosphate: mineral supplements and laxatives [medicine]

aluminium phosphate: fertilisers [farming]

calcium phosphate: toothpaste [dental products]

Q U2 R5

Understanding Research

7.45 What is the formula for the phosphorous molecule in white and red phosphorous?

7.46 Phosphorus is in Group 15, the same periodic group as nitrogen. How many valence electrons would you expect it to have in its outer shell?

Potassium deposits in Kazakhstan

Kazakhstan is rich in mineral deposits and is a major producer of phosphorus. Phosphorus is found in several major deposit basins around the country.

Q C1 S2

Communicating Society

7.47 In groups conduct research to find out:

(a) in which regions of Kazakhstan these deposits are found
(b) the chemistry involved in extracting phosphorous from its ores
(c) the chemical form in which phosphorous is mainly found in Kazakhstan
(d) any environmental hazards or health risks linked to the industry.

Figure 7.30

Fertilizers

Figure 7.31

Fertilisers add minerals to the soil which dissolve and can be absorbed by plants allowing them to grow more quickly and produce larger crop yields for farmers.

The main chemicals in fertilizers, as we have seen across this module, are nitrogen, phosphorus and potassium. What is known as the N-P-K ratio is often stated on fertiliser labels indicating the proportion of each of these elements in the fertiliser. For example, if a fertiliser contained 20% nitrogen 5% phosphorous and 5% potassium, the ratio would be given as 20-5-5.

Q U2 Understanding

7.48 Look at the table and match the names of main fertilisers to their chemical formulas.

Fertilizer	Formula
anhydrous ammonia	$(NH_4)_3PO_4$
ammonium phosphate	$(NH_4)2SO_4$
urea	NH_4NO_3
potassium nitrate	NH_3
ammonium nitrate	KNO_3
ammonium sulfate	$(NH_2)_2CO$

One of the causes of eutrophication of waterways is the run off of excess fertiliser from farmland. The fertiliser spread on the land may be artificial, or it may be natural, i.e. animal slurry. Either way, if the quantity used is not carefully controlled, excess nitrates and phosphates may be washed into rivers or lakes by rainwater. Eutrophication results from excessive amounts of these nutrients behaving as fertilisers that increase the growth of plants such as algae in lakes and rivers (see **Figure 7.32** below). Fortunately, however, increased regulation of the use of fertilisers in many countries has brought about a reduction in the amount of pollution of waterways.

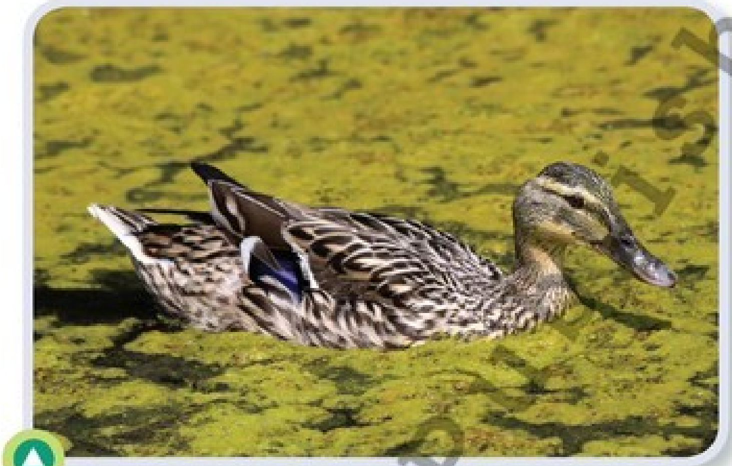


Figure 7.32 *Algal bloom*

Percentage yields

It is not always possible to obtain 100% yield of a product (theoretical yield) in reactions for any of a number of possible reasons:

- a reaction may not go to completion either because it is reversible or some of the product is lost when it is separated from the reaction mixture
- there may be side reactions where reactants react in unexpected ways

The amount of a product actually obtained is known as the yield and this amount when compared to the maximum theoretical amount – calculated as a percentage – is known as the percentage yield.

$$\boxed{\text{percentage yield} = \frac{100 \times \text{actual yield (in grams, kg)}}{\text{theoretical yield}}}$$

The calculation involves working out the predicted maximum theoretical yield from a reacting mass calculation and then comparing this to the actual yield.

Steps in calculating theoretical yield

Step 1: Convert mass of initial value to moles.

Step 2: Use mole ratios (the coefficients) to convert moles of initial substance to moles of final substance.

Step 3: Convert moles of final substance to mass of final substance.

Sample question 1

9.87 g of ammonia (17.031 g/mol) react completely according to the following reaction:

$2NH_{3(g)} + CO_{2(g)} \rightarrow CN_2OH_{4(s)} + H_2O_{(l)}$

If 13.74 g of urea (CN_2OH_4 60.056 g/mol) are produced what is the percent yield for this reaction?

Answer:

Actual yield = 13.74 g

$$\text{Theoretical yield} = 9.87\ \text{g} \times \frac{\text{mol NH}_3}{17.031\ \text{g}} \times \frac{\text{mol CN}_2\text{OH}_4}{2\ \text{mol NH}_3} \times \frac{60.056\ \text{g}}{\text{mol CN}_2\text{OH}_4} = 17.4\ \text{g}$$

$$\text{Percentage yield} = \frac{13.74\ \text{g}}{17.4\ \text{g}} \times 100 = 79.0\%$$

Sample question 2

(a) From the equation $FeBr_2 + 2KCl \rightarrow FeCl_2 + 2KBr$ what is the theoretical yield of iron (II) chloride, if 34 g of iron (II) bromide are used?

(b) What is the percentage yield if only 4 g of iron (II) chloride are actually produced?

Answer:

(a) Theoretical yield =

$$34\ \text{g} \times \frac{\text{mol FeBr}_2}{216\ \text{g}} \times \frac{\text{mol FeCl}_2}{\text{mol FeBr}_2} \times \frac{126\ \text{g}}{\text{mol FeCl}_2} = 0.16 \times 126 = 20.16\ \text{g}$$

(b) $\text{Percentage yield} = \frac{4\ \text{g}}{20.16\ \text{g}} \times 100 = 19.8\%$

7.49 **(a)** Calculate the theoretical yield of aluminum chloride (in grams) that can be produced from 10.00 grams of aluminum metal.

$$\begin{array}{ccccc} 70.90\ \text{g/mol} & & 26.98\ \text{g/mol} & & 133.33\ \text{g/mol} \\ 2Cl_{2(g)} & + & 3Al_{(s)} & \longrightarrow & 2AlCl_{3(s)} \end{array}$$

(b) What is the percentage yield if 27.42 g are actually produced?

7.50 **(a)** In the reaction: $Mg + 2HNO_3 \longrightarrow Mg(NO_3)_2 + H_2$ if 40 g of magnesium is reacted with excess nitric acid, how much hydrogen will be produced?

(b) If only 1.65 g of hydrogen is actually produced, what is the percentage yield?

7.51 **(a)** In the reaction how many grams of iron (III) sulphate are produced if 20g of iron (III) phosphate are reacted with excess sodium sulphate?

$2FePO_4 + 3Na_2SO_4 \longrightarrow Fe_2(SO_4)_3 + 2Na_3PO_4$

(b) If only 24 g are produced, what is the percentage yield?

7.52 **(a)** In the reaction: $LiOH + KCl \longrightarrow LiCl + KOH$ what is the theoretical yield of lithium chloride if 20 g of lithium hydroxide are reacted with excess potassium chloride?

(b) What is the percentage yield if only 12 g are actually produced?

MODULE 8

Silicon and its compounds

Learning outcomes

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

- Explain the various uses and applications of silicon and the meaning of a semiconductor 9.2.1.24
- Describe the type of crystal lattice structure and type of chemical bond in silicon, silicon dioxide and silicon carbide 9.1.4.8
- Describe the main chemical properties of silicon and its compounds and make equations for their reactions 9.2.1.25

Keywords

✓ silica ✓ silicones ✓ macromolecular crystal ✓ crystalline silicon ✓ purity ✓ semi-viscous ✓ sealant ✓ adhesive ✓ lubricant ✓ abrasive ✓ foundry ✓ doping ✓ depletion layer ✓ diode ✓ transistor

Silicon in the form silicon dioxide (silica) and its silicates – minerals consisting of silicon and metal oxides – is estimated to make up around 28% of the Earth's crust, making it the second most abundant element in the Earth's crust after oxygen. Silicon in a pure form is too reactive to be found in nature but silicon dioxide [SiO_2] and silicates formed with the metal oxides from metals such as potassium, magnesium, aluminium and calcium are found in rocks, soils, sand and clay.

The particular physical and chemical properties of silicon and its compounds such as silica and silicon carbide means that it has a wide range of uses in industry, manufacturing and the modern electronics industry as the name Silicon Valley, given to the high-tech area of northern California, attests to.

Figure 8.1 *Silicon Valley home to Apple, Cisco, Google, HP, Intel and Oracle*

8.1 In which group of elements is silicon in the periodic table?

8.2 From the neighbouring elements in the group, what can you predict about silicon?

8.3 What properties of silicon do you think might make it particularly useful to the modern electronics industry?

8.4 Look around the classroom? Can you see anything that you think contains silicon?

Silicon

The electronic configuration of the silicon atom is represented by the two diagrams below.

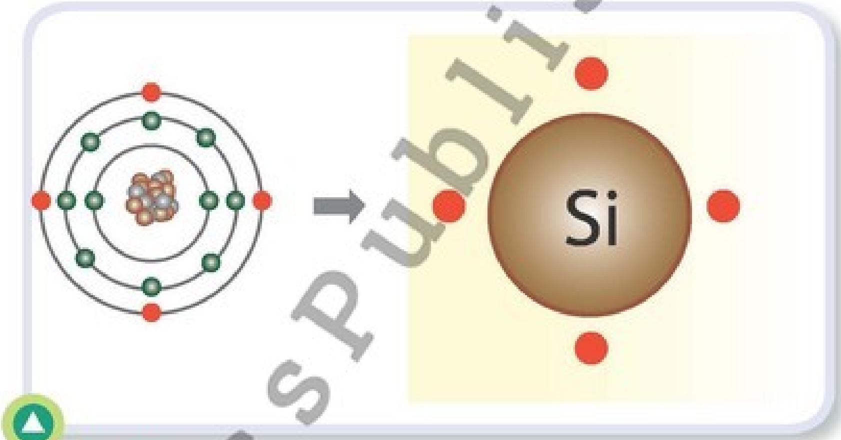


Figure 8.2

8.5 Complete the electron shell configuration table for silicon:

Element	Formula	Electrons in first shell	second shell	third shell
silicon	Si			

8.6 What is the atomic number of silicon?

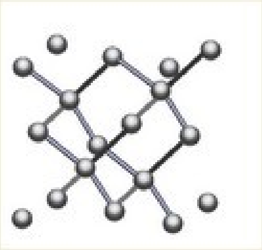
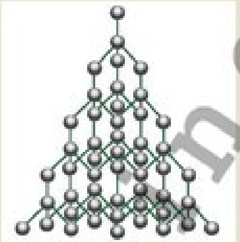
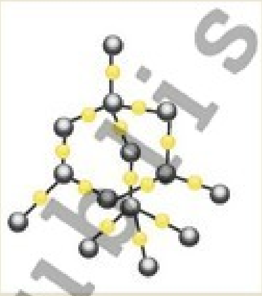
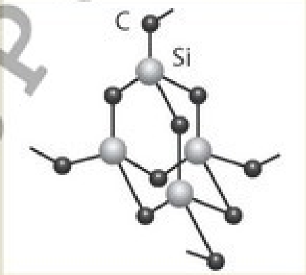
8.7 How many valence electrons does it have?

Covalent macromolecular compounds

Silicon and its compounds silicon dioxide and silicon carbide form **covalent macromolecular crystals**. In such crystals the atoms form lattice points and covalent bonding holds the atoms together. These structures are similar to the structures of diamond and graphite and have a regular three-dimensional structure.

The table shows the different covalent macromolecular structures of silicon, silicon dioxide and silicon carbide.

Table 8.1

Structures	model
Silicon Crystalline silicon shares the same structure as diamond. Each silicon atom is bonded to four other silicon atoms in a tetrahedral arrangement.	 Silicon  Diamond
Silicon dioxide In silicon dioxide each silicon atom is covalently bonded to four oxygen atoms and each oxygen atom is covalently bonded to two silicon atoms in the macromolecular lattice structure.	
Silicon carbide In silicon carbide each silicon atom is covalently bonded to four carbon atoms and each carbon atom is covalently bonded to 4 silicon atoms in the macromolecular lattice structure.	

Q Understanding U1

8.8 Write the formula for **(a)** silicon dioxide and **(b)** silicon carbide.

Physical and chemical properties of silicon and its compounds

The strong bonds in the covalent lattice structure of silicon and its compounds mean that the bonds are extremely difficult to break, as is shown by their high melting and boiling points in Table 8.2.

Table 8.2

Silicon		Silicon dioxide		Silicon carbide	
Melting point	1687 K	Melting point	1986 K	Melting point	3100 K
Boiling point	3538 K	Boiling point	3220 K	Boiling point	sublimes

Silicon

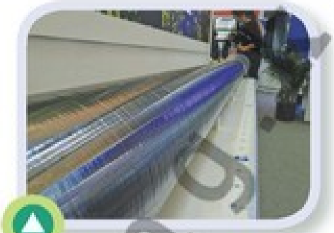


Figure 8.3

In its pure form crystalline silicon is a hard, dark grey solid with a metallic lustre. To produce pure crystalline silicon, silica and coke are heated to extremely high temperatures. High temperatures are required to produce a reaction where the oxygen is removed and the silicon formed displaces the carbon. It is then refined to improve its purity.

Q U2 Understanding

8.9 Complete the reaction:

$SiO_2 + 2C \xrightarrow{\text{Heat}}$ ______________________

Uses of silicon

Silicon is considered as one of the most useful elements on Earth. It is manufactured from silica into a pure form used in modern electronics industries (see below), but most of its different uses relate to its alloy forms with metals such as aluminium and iron and in silicones which are silicon-oxygen polymers with methyl groups attached.


Figure 8.4 *Silicon Carbide*

Alloys such as ferrosilicon and aluminium silicon have many uses in manufacturing and industry. Ferrosilicon is extremely important in the steel-making industry in the smelting of structural, tool and spring steels which exhibit different qualities of strength, hardness and flexibility. Aluminium silicon is an alloy that has highly corrosion-resistant and wear-resistant properties and is used to produce functional parts of engines such as pistons, cylinder liners and coatings for pipelines.

Silicones have an exceptionally wide range of uses. They are produced in many forms, including solids, liquids, oils, semi-viscous pastes and rubber-like forms. They are valued for their heat- and stress-resistant qualities. Silicones are used as sealants for the doors and windows of aircraft and also as sealants in swimming goggles and diving masks. They are used as thermal and electrical insulators and are key compounds in adhesives, lubricants, paints and coatings for bridges, oil rigs and roads. Strong, durable silicones feature in hundreds of products you see around you every day: keyboards, keypads, computers, kitchen equipment and home entertainment devices.


Figure 8.5

Silicon dioxide

Silicon dioxide is commonly found in the form of quartz.

Pure forms of silicon dioxide can also be obtained by either burning silicon in air or through the slow decomposition of silicon in a red hot state when exposed to steam.

Quartz

Pure silicon dioxide

Figure 8.6 *Quartz and pure silicon dioxide*

8.10 Complete the equations:

(a) $Si + 2H_2O \longrightarrow$

(b) $Si + O_2 \longrightarrow$

Uses of silica

Figure 8.7

Items made from silica are all around us. It is an essential element in glass, pottery, the manufacture of bricks, cement and road surfaces. It has an important functional role as a filler in products such as rubber, paints and plastics. In agriculture, it is used to supply silicon to crops which helps reduce transpiration, salinity stress and water stress at times of drought. Silica is used in processes such as sandblasting and hydraulic fracturing in the oil and gas industry and has a wide range of uses in the metallurgical industry. It is also the raw material in the production of pure silicon and ferrosilicon and is used widely in precision casting in foundries.

Silicon carbide

When a mixture of silicon and coke is heated in an electric furnace, silicon carbide is formed. It is an endothermic reaction.

8.11 Complete the equation:

$Si + C \longrightarrow$

Uses of silicon carbide

Silicon carbide is produced at extremely high temperatures using a mixture of pure silica (sand)and finely ground coke (carbon). It is one of the hardest substances known and is used in cutting and grinding tools as an abrasive for smoothing and cutting hard materials. As a wear-resistant material, it is also used in heating elements, car components and engine parts.


Figure 8.8

8.12 What are the products of adding steam to silicon in a red hot state?
8.13 Describe the process by which pure silicon is formed.
8.14 What qualities does adding ferrosilicon in different measures and processes give to steel?
8.15 Why is aluminium silicon used in lining cylinders and coating pipelines?
8.16 Why do silicones have such a wide range of uses?

Use of silicon as a semiconductor

Silicon has a valency of four. This means that it has four electrons in its outermost shell but, like other atoms, would 'prefer' to have eight.

Figure 8.9

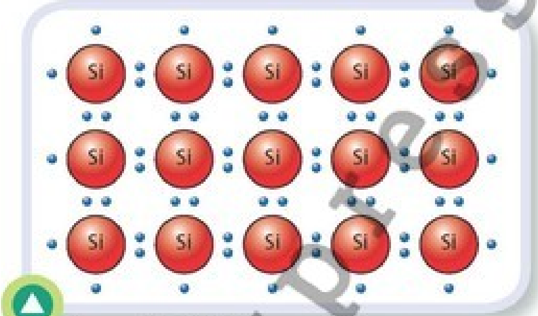


Figure 8.10 *A silicon crystal at low temperature*

In a silicon crystal, in order to achieve a situation in which each atom is surrounded by eight electrons, the atoms arrange themselves in such a way that each atom shares an electron with each of four neighbours.

At a temperature of 0 K, silicon is a perfect insulator. No electrons are free to move and they are therefore not free to carry a current.

Above 0 K, the added heat energy is distributed throughout the crystal as kinetic energy. From time to time, an electron will gain sufficient energy to escape from its place and move through the crystal. It leaves a 'hole' behind it.

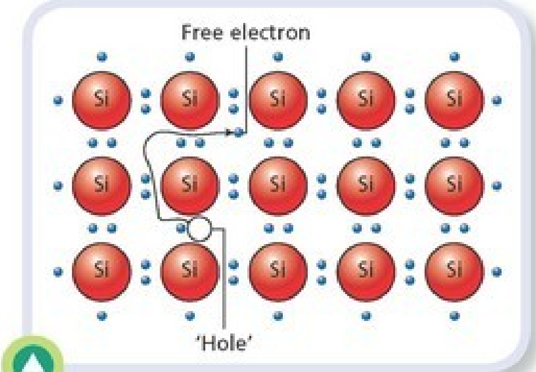


Figure 8.11 *A silicon crystal at high temperature*

Once electrons gain enough energy to escape from their places, they are available to carry a current if the crystal is placed in a circuit. In such a case, the free electrons will move towards a positive connection, and the holes will end up effectively moving towards the negative connection. They have no charge, of course, but as they will always move towards a negative connection, it is useful to think of them as being positive, and they are often referred to as **positive holes**.

The electrical conductivity of a semiconductor is roughly proportional to the number of charge carriers. In the production of silicon semiconductors, different chemicals are added to silicon to adjust the impurities to give a wide range of semiconductors with different conductivities that can be assembled into complex devices.

Conduction can be improved by either adding extra free electrons **(n-type doping)** or extra holes **(p-type doping)**.

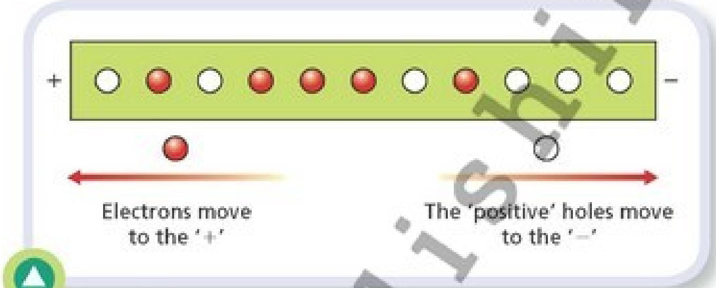


Figure 8.12

n-type doping

When constructing a silicon crystal, we can introduce a small number of phosphorous atoms. These will try to fit into the overall crystal structure, sharing one electron with each of four neighbours. But they have five electrons in their outer shell and will therefore have one 'extra electron'. This is likely to break free from the atom and to become available for conduction. As there are extra free electrons, conductivity is improved.

p-type doping

If we add a small quantity of boron to a growing crystal of silicon, the boron atoms will try to fit into the structure of the crystal. Each one has only three electrons in its outer shell, however, so an extra 'hole' will be created. In p-type material, we even say that holes are the majority charge carriers. In reality, they act as stepping stones for the moving electrons, which improves the conductivity.

Doping is the addition of small quantities of an impurity to a semiconductor in order to improve conductivity.

p–n junction

n-typing and **p-typing** are really only of any use to us when we place the two together at a p–n junction. The simplest example of this is in a **diode**.

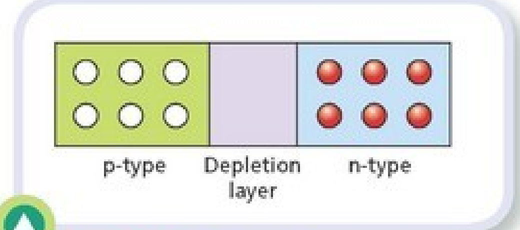


Figure 8.13 *A p–n junction*

At the junction of the two materials, the extra free electrons tend to occupy the extra free holes. This is known as the **depletion layer**. It acts as a block of insulating material in the middle of the diode (see Figure 8.13).

The symbol for a diode is shown in **Figure 8.14**. The significance of the diode is that it allows current to flow in one direction and not the other.

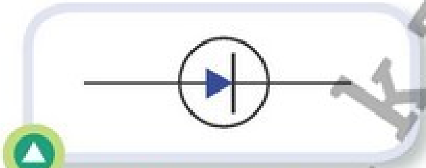


Figure 8.14 *The symbol for a diode*

Forward bias

If the diode is placed in a circuit so that the p-type material is attached to a positive connection, current will flow. This is because the electrons are drawn towards the positive, and the holes to the negative: the depletion layer shrinks and allows current to pass through it (see **Figure 8.15**).

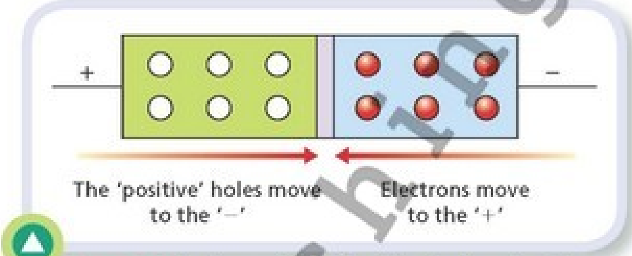


Figure 8.15 *In forward bias, the depletion layer shrinks and a current flows*

Reverse bias

For reverse bias, the p-type material is attached to the negative terminal, and the n-type to the positive. The electrons are drawn backwards from the centre and the depletion layer grows. A current will not be able to move through the device.

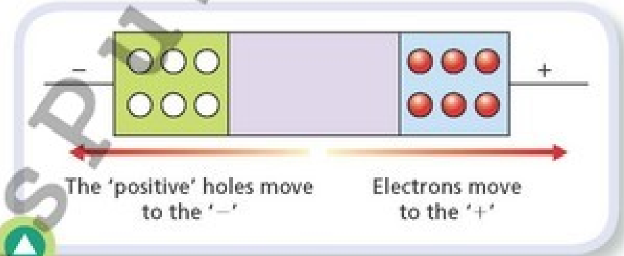


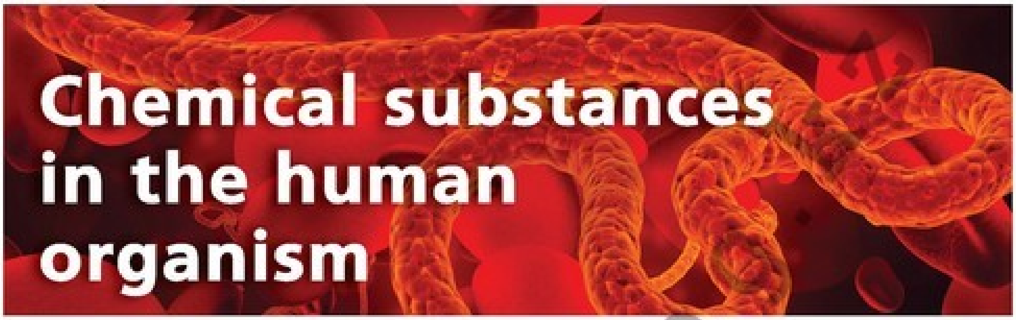
Figure 8.16 *In forward bias, the depletion layer shrinks and a current flows*

Diodes go into transistors and thousands of transistors go into silicon chips. The transistors act as switches, which allow the operation of Boolean gates in microprocessors. Silicon has thus been central to the development of the modern electronics industry and is why the creative hub of the industry is called 'Silicon Valley'.

Q Understanding U1

8.17 What is a semiconductor?
8.18 What element is added to silicon in
(a) n-doping (b) p-doping?
8.19 How do the valencies of these elements compare to silicon?
8.20 What does reverse bias result in?
8.21 Name one biological function of silicon.
8.22 Look at the other elements in Group 14. What is the trend in conductivity?
8.23 Name three uses of silicones.
8.24 What is meant by the term 'doping'?
8.25 Briefly explain the term 'extrinsic conduction'.
8.26 What is a p–n junction?

MODULE
9



Chemical substances in the human organism

Learning outcomes

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

- Identify elements contained in the human organism and recognise their role in biological processes 9.5.1.1
- Evaluate the typical diet of people in Kazakhstan and compare this to the notion of a balanced diet 9.5.1.2
- Explain the role of calcium and iron in the human organism 9.5.1.3
- Identify carbon as a key element in food products 9.5.1.4
- Recognise heavy metals as sources of pollution and explain their impact on the environment 9.5.1.5

Keywords

✓ nutrients ✓ metabolism ✓ metabolic reactions ✓ minerals
✓ reducing sugars ✓ trace elements ✓ dipositive ions ✓ precipitation
✓ filtration ✓ collagen ✓ osteoporosis ✓ anemia

Food and nutrition

Nutrition is the way in which an organism obtains and uses food. Nutrients are the chemical substances present in food that are used by organisms. Nutrients are essential to maintain metabolism and continuity of life for all living organisms. In particular, nutrients are necessary:

- as a source of energy
- to make chemicals needed for metabolic reactions
- as the raw materials for the growth and repair of structures in the organism.

The elements present in food

There are essentially 14 elements that our food can contain. Apart from carbon, hydrogen and oxygen, the rest of these elements are often called minerals.

- The six common elements found in food are: carbon (C), hydrogen (H), oxygen (O), nitrogen (N), phosphorus (P) and sulfur (S). The first four of these elements make up over 96% of the mass and atoms present in your body. See Table 9.1. Most of the chemical compounds found in living things are made from carbon atoms bonded together. Compounds made from carbon are said to be organic compounds.

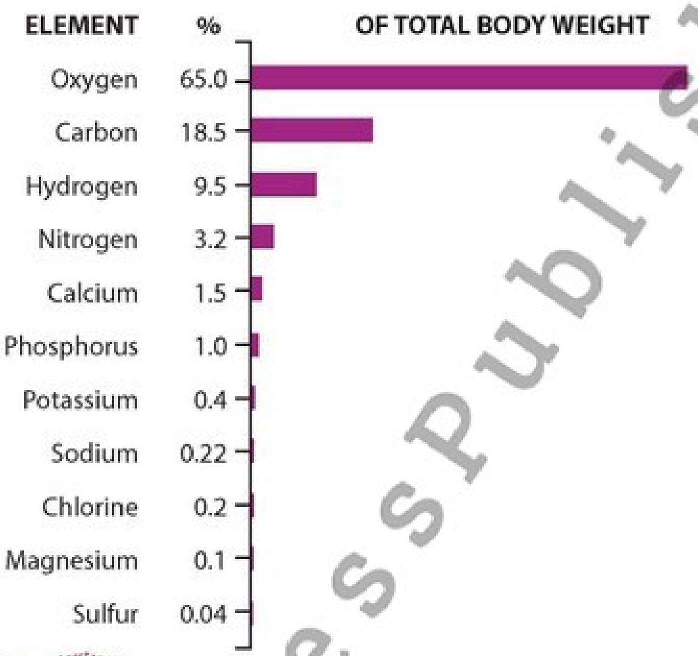
- Five elements that are present as dissolved salts are: sodium (Na), magnesium (Mg), chlorine (Cl), potassium (K) and calcium (Ca).
- The three trace elements are: iron (Fe), copper (Cu) and zinc (Zn).

Trace elements are elements that are only required in tiny amounts in the diet.

Did you know?

Water accounts for between 50% and 65% of the body mass of an average person.

Table 9.1



9.1 What is a trace element?

9.2 What term is given to compounds containing carbon?

9.3 In what form is potassium found in food?

Balanced human diet

There are seven components in a balanced diet: carbohydrate, protein, lipid, vitamins, minerals, fibre and water. These components must be present in our diet in the right amounts and should come from a variety of sources. This ensures that the body gets all the necessary energy and nutrients.

The total amount of food a person requires depends on:

- Age (young people need more food than older individuals)
- Activity levels
- Gender (males need more food than females)
- Health

A balanced diet contains all the necessary food types in the correct proportions.

Food groups

Foods that contain similar nutrients are arranged into four food groups. These groups are:

- cereals, bread and potatoes
- fruit and vegetables
- milk, cheese and yoghurt
- meat, fish and poultry.

We should eat a variety of foods from each of these food groups. Suggested numbers of servings of each food group are given in the food pyramid in **Figure 9.1**.

Figure 9.1 *The food pyramid*

U_5 C_1 C_2 S_2

9.4 Research online the type of nutrients obtained from each of the four food groups and why each is needed by the body.

Food group	Nutrients	Needed for
		
		
		
		

9.5 In your opinion, which elements were most likely to be lacking in a traditional diet in Kazakhstan? What do you think people should eat more/less of in modern day Kazakhstan?

9.6 Research the World Health Organisation profile of Kazakhstan. Are there any particular improvements shown over the last 50 years that you would argue might relate more directly to nutritional health?

Calcium

Calcium is needed to build strong bones and teeth. It is also an important chemical in physiological processes such as blood clotting, muscle contraction, and nerve-cell communication. A dietary intake of calcium which is considerably below recommended levels may impact on the development of bones and the development of diseases such as osteoporosis. Bones increase in size and mass during childhood and adolescence, which means that a healthy intake of calcium and vitamin D are particularly important.

Bone as a natural composite

Bone is comprised of three composite parts:

- water
- an inorganic ionic component : a form of calcium phosphate known as hydroxyapatite
- an organic covalent molecular component: a protein called collagen

In **Activity 9.1** below you will investigate in stages what happens to chicken bones as each of these components is successively removed. You will need to observe and record how the structure and properties of bone change with the removal of each component.

Lab Work 14

Activity 9.1

Question

What do we observe as the composite parts of chicken bones are successively removed?

Equipment needed

beaker
lab balance
measuring cylinder
Bunsen burner
tongs
2 x cleaned chicken bones
1 mol/ dm^3 hydrochloric acid

Safety

- Wear your safety glasses.

Conducting the activity

Stage 1

1. Place Bone 1 in a beaker and cover with hydrochloric acid.
2. Leave for 24-36 hours. Remove the bone with tongs.

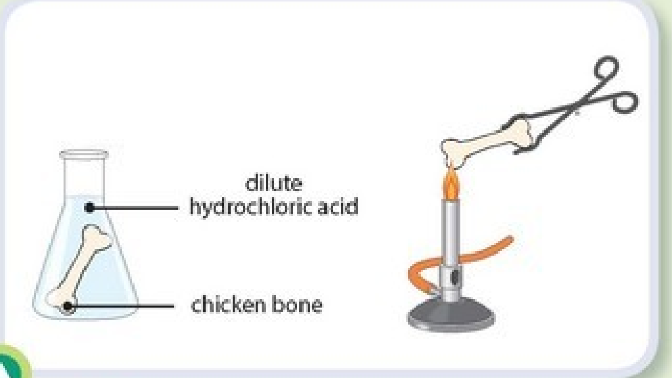


Figure 9.2

9.7 What has been removed from the bone by this process?

9.8 How has this changed the properties of the bone?

Stage 2

1. Measure the mass and volume of Bone 2 and calculate its density. What steps can you take to ensure accurate measurement?
2. Heat the bone at a temperature of 60 °C in order to remove the water.

9.9 What is the mass of the water that has been removed from Bone 2?

9.10 What is the new density of Bone 2?

Stage 3

1. Using a Bunsen burner in a fume cupboard completely roast the bone using the tongs.
2. Weigh the bone and calculate the mass that has been lost in this process.

9.11 What has been lost in this process?

9.12 What is the substance that remains at the end of this process?

9.13 Calculate the mass of collagen that was present in Bone 2.

9.14 How has the removal of collagen affected the properties of the Bone 2?

9.15 From your observations in the three stages of the experiment, what would you say are the main purposes of each component of bone?

Iron

Iron is necessary for the delivery of oxygen to cells and the regulation of cell growth. Iron deficiency is a condition that develops gradually. It is most common in children and women of childbearing age. A lack of iron leads to insufficient supply of oxygen to cells and this can result in a condition known as anemia. The symptoms of anemia include severe fatigue, dizziness, leg cramps and decreased immunity. In children it is linked to slow cognitive and social development.

Carbon in food

It can be said that humans are a "carbon-based" life form. Carbon atoms are a major part of our molecular structure and we need through food intake to replace carbon that we use up in biological processes. Our major food sources such as vegetables, fruit, grains and meat are all rich sources of carbon.

Carbon is the major structural component of all three of our major food groups: carbohydrates, fats and proteins. The elements present in carbohydrates are indicated by the name itself: carbon (C), hydrogen (H) and oxygen (O). These elements are usually present in the ratio $Cx(H_2O)y$, where x and y are the same number (i.e. x = y). This means there is twice as much hydrogen as carbon or oxygen in a carbohydrate. Carbohydrates make up the bulk of organic substances on Earth and perform numerous roles in living things.

Glucose is a simple carbohydrate in which x and y are both equal to 6. The formula for glucose is $C_6H_{12}O_6$.

Glucose, is the carbohydrate which reacts with oxygen during respiration. The energy trapped in a glucose molecule is released through the supply of oxygen in the respiration process:

$$C_6H_{12}O_6 + O_2 \xrightarrow{\text{(respiration)}} CO_2 + H_2O$$

Glucose along with other common carbohydrates such as lactose, fructose and maltose are called reducing sugars because they cause other substances to be reduced e.g. Cu_2+ to Cu while they are oxidised. Carbohydrates that are non-reducing, such as sucrose, are sugars that cannot donate electrons to other molecules and therefore cannot act as reducing agents. We shall explore this distinction in terms of chemical formula further in Module 10.

9.16 What are the elements in all carbohydrates?
9.17 What releases the energy in glucose in respiration?
9.18 What condition is caused by iron deficiency?
9.19 Which sections of the population are most likely to suffer from this condition?

In the qualitative test below we look at how to identify the presence of different forms of carbohydrate in terms of this distinction between reducing and non-reducing sugars.

Lab Work 15

Activity 9.2

Question

How can we test for the presence of reducing sugars?

Equipment needed

4 x test tubes
test tube rack
Dilute fructose, glucose, maltose solution
Benedict's solution
Bunsen burner
Water bath
Tongs

Safety

- Use tongs when placing the test tubes into and removing them from the boiling water bath.

Conducting the activity

1. Label the test tubes with name of the sugar and mix an equal amount of Benedict's solution with each sugar solution.
2. Heat the test tubes in a boiling water bath.
3. Record your observations.

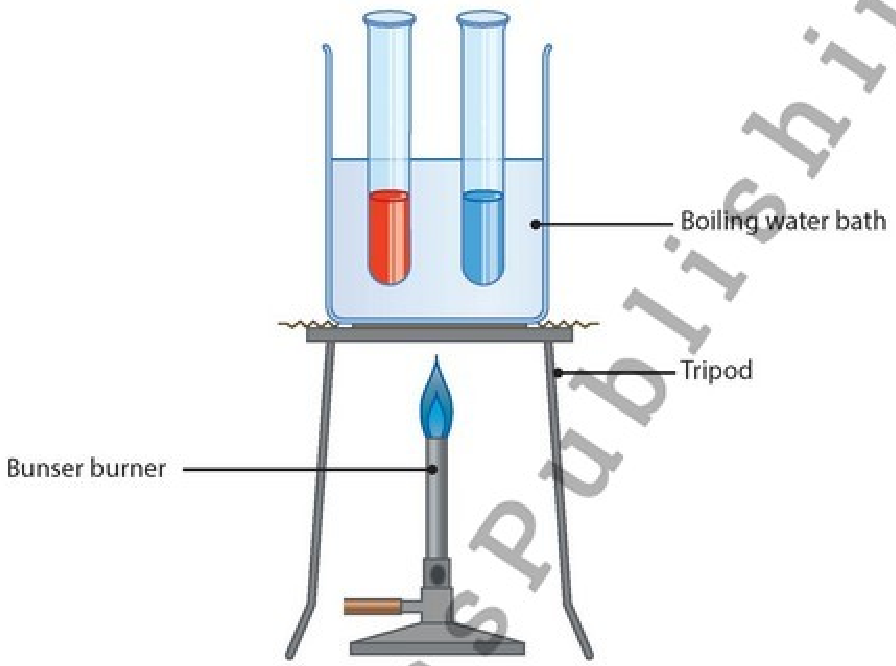


Figure 9.3

Q U5 R5

Understanding Research

9.20 If reducing sugars are oxidized when **heated** with **Benedict's solution** and lose electrons which react with the copper ions in Benedict's solution to form copper oxide - a reddish brown precipitate – what can you conclude from your results?

9.21 Why did the mixtures in **Activity 9.1** need to be heated?

9.22 Can you think of a medical test that involves testing for glucose levels?

Heavy metal pollution in the environment

Metals with high relative atomic masses such as mercury, cadmium and lead are known as heavy metals. Quantities of these elements are sometimes dumped, e.g. car batteries containing lead, or dry batteries containing cadmium. As waste management around the world improves, recycling of elements such as these will lessen the amount of dumping. Dipositive ions of these metals, i.e. Hg^{2+}, Cd^{2+} and Pb^{2+}, sometimes get into waterways from industrial effluent, and consequently into drinking water. These elements are cumulative poisons in that frequent exposure causes a build-up in the body, with consequent serious health damage.

Before the effluent is run into a waterway, **the metal ions can be removed by means of precipitation**. For example, lead(II) hydroxide, $Pb(OH)_2$, is insoluble in water, so that if effluent containing Pb^{2+} ions is treated with calcium hydroxide solution, the lead will precipitate as $Pb(OH)_2$:

$$\boxed{Pb^{2+} + 2OH^-_{(aq)} \rightarrow Pb(OH)_{2(s)}}$$

Hydroxide precipitation is ineffective for mercury removal, but alternatives such as carbonate or sulfide precipitation are available. Flocculating agents are often used to increase the size of the solid particles before their removal by filtration.

Q Understanding U5

9.23 Why is pollution of water by heavy metal ions a problem?

9.24 What steps can be taken to prevent pollution of water by heavy metal ions?

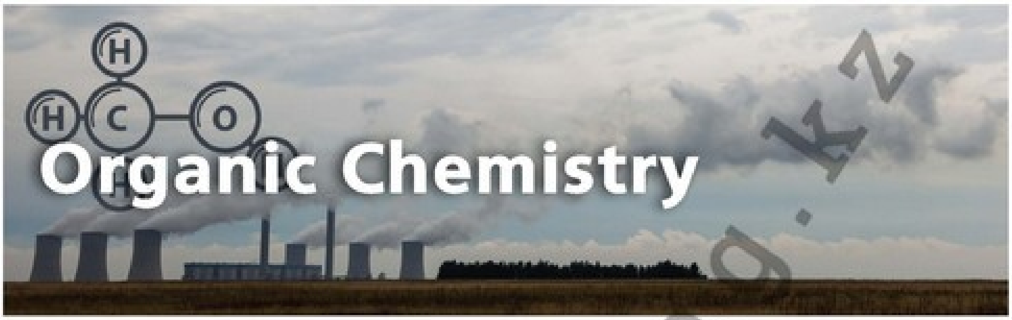
9.25 How can heavy metal ions be removed from water?

Q Understanding U5

9.26 Why is it necessary to set limits for nitrates and phosphates in water?

9.27 Why is it necessary to set limits for heavy metal ions in water?

MODULE 10



Organic Chemistry

Learning outcomes

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

- Explain the specificity of organic compounds in terms of their structure 9.4.3.1
- Recognise the system of classification of hydrocarbons and their derivatives 9.4.3.2
- Explain the concept of functional group as a group determining the typical chemical properties of a class of compounds 9.4.3.4
- Understand and apply the rules of IUPAC nomenclature 9.4.3.5
- Understand the definition of isomerism and make formulas for the structural isomers of hydrocarbons 9.4.3.6
- Calculate the molecular formula of gaseous compounds by relative atomic mass and percentage composition of mass element 9.2.3.4

Keywords

✔ organic compound ✔ saturated ✔ unsaturated ✔ tetrahedral
✔ homologous series ✔ suffix ✔ structural formula ✔ empirical formula
✔ molecular formula ✔ functional group ✔ planar ✔ isomer

Organic chemistry

Organic chemistry is the chemistry of carbon compounds other than carbon monoxide, carbon dioxide, carbonates and hydrogencarbonates. Carbon has the ability to form chains and rings with itself, resulting in the existence of a very large number of compounds. Most organic compounds contain hydrogen as well as carbon, many contain oxygen, and some contain other elements such as nitrogen and the halogens.

Chemical bonds in organic compounds are usually covalent, involving shared pairs of electrons. Organic compounds containing only single bonds are said to be **saturated**. Compounds containing a double or triple bond are said to be **unsaturated**. Such bonds always consist of one sigma bond, with the other bond(s) of the pi type, involving the sideways overlap of p orbitals.

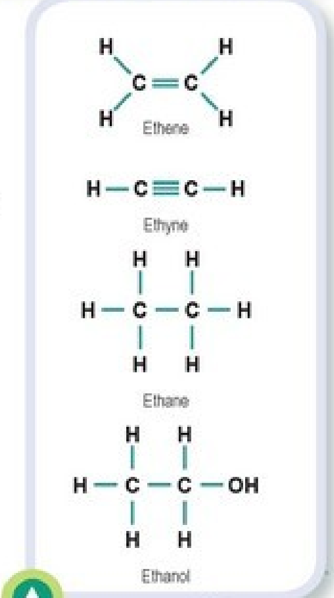


Figure 10.1 *Saturated and unsaturated compounds*

Tetrahedral carbon

A tetrahedral arrangement is particularly important in molecules containing carbon. **In saturated organic compounds, all of the carbon atoms are tetrahedral.** Molecular models may be used in the study of these compounds, so that the tetrahedral arrangement of the carbon atoms can be examined more closely.

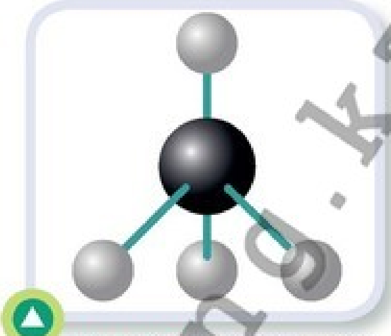


Figure 10.2 *Methane molecules have a tetrahedral shape*

Alkanes, chloroalkanes and alcohols are three groups of saturated organic compounds. All of the carbon atoms in these compounds are tetrahedral.

10.1 Explain the difference between a saturated and unsaturated compound.
10.2 Explain the term organic chemistry.
10.3 Why does carbon form such a large number of compounds?
10.4 Explain what is meant by a tetrahedral shape.
10.5 What is the tetrahedral bond angle?
10.6 Name three groups of compounds where all of the carbons are tetrahedral.

Homologous series

The study of such a large number of compounds would be very difficult if all the compounds had to be studied individually. The study becomes much easier when certain classifications are applied to the compounds. Organic compounds can be grouped into families called **homologous series**. Each family has a general formula that gives the ratio of atoms in any compound in the family. For example, the general formula of a homologous series of hydrocarbons called the alkanes is C_nH_{2n+2}. The formulas of the individual members of the series can be obtained by substituting successive values for n, starting in this case with n=1.

Sample question 1

Use the general formula C_nH_{2n+2} to work out the molecular formulas of the first five members of the alkanes.

Answer:

Substituting n=1 into the formula as follows gives the formula of methane, the first member of the series:

$$C_nH_{2n+2} \longrightarrow C_1H_{2(1)+2} \longrightarrow C_1H_4 \longrightarrow CH_4$$

Substituting n = 2, 3, 4, and 5 in turn produces the formulas C_2H_6 for ethane, C_3H_8 for propane, C_4H_{10} for butane and C_5H_{12} for pentane.

In summary, a homologous series is a family of compounds with:

- the same general formula
- successive members differing by CH_2
- similar chemical properties.

A homologous series is a family of organic compounds with the same general formula, similar chemical properties, and successive members differing by CH_2.

Understanding U4

10.7 Explain the term *homologous* series.

10.8 Write the general formula for the alkane homologous series.

Naming system

Given the very large number of organic compounds, it would be very difficult indeed to remember the names of more than a small number of them without a good naming system. In the IUPAC system, each hydrocarbon name consists of two parts, the root and the suffix. The roots correspond with the number of carbon atoms in the longest chain present. The first ten roots are shown in **Table 10.1**.

Table 10.1

No. of carbon atoms	Root	No. of carbon atoms	Root
1	Meth-	6	Hex-
2	Eth-	7	Hept-
3	Prop-	8	Oct-
4	But-	9	Non-
5	Pent-	10	Dec-

The suffix depends on the homologous series to which the compound belongs. For example, each alkane has the suffix **–ane**, each alkene has the suffix **–ene**, and each alkyne has the suffix **–yne**.

Sample question 2

Write the names of the following:

(a) the alkene with three carbons

(b) the alkane with seven carbons in its longest chain

Answer:

(a) Three carbons means that the root is prop-, and alkenes have the suffix -ene. Thus the name is propene.

(b) Seven carbons means that the root is hept-, and alkanes have the suffix -ane. Thus the name is heptane.

Q Understanding U4

10.9 Write the names of the following:

(a) the alkene with two carbons

(b) the alkyne with two carbons

(c) the alkane with three carbons

(d) the alkane with six carbons

Alkanes

The alkanes form a homologous series of aliphatic hydrocarbons of general formula C_nH_{2n+2}. The names and formulas of the first five alkanes are shown in **Table 10.2**. They are saturated compounds, containing only single bonds. The first four members of the series are gases, those containing from five to fifteen carbon atoms are liquids, and those containing a greater number of carbon atoms are waxy solids.

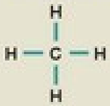
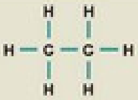
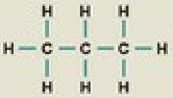
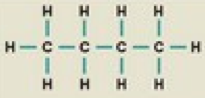
Table 10.2

No. of carbon atoms	Name	Formula	No. of carbon atoms	Name	Formula
1	Methane	CH_4	4	Butane	C_4H_{10}
2	Ethane	C_2H_6	5	Pentane	C_5H_{12}
3	Propane	C_3H_8			

The formula shown in each case is the molecular formula, which gives the number of each type of atom present in a molecule of the compound. For example, a molecule of ethane, C_2H_6, consists of two carbon atoms and six hydrogen atoms. However, the molecular formula gives no information about the way in which the atoms are arranged or bonded together. This information is obtained from the **structural formula**. In the structural formula, single bonds are written as a single line, double bonds are written as a double line, and triple bonds as a triple line.

The structural formulas of the first four alkanes are shown in **Table 10.3**.

Table 10.3

Name	Molecular formula	Structural formula
Methane	CH_4	
Ethane	C_2H_6	
Propane	C_3H_8	
Butane	C_4H_{10}	

10.10 Draw the structural formula of pentane.

Molecular models may be used to examine the structures of molecules of the alkanes. Ball and stick type models may be used, but if these are not available something as simple as plasticine and matches are a satisfactory alternative. In the alkanes, each carbon atom forms four separate single bonds to satisfy the valency of carbon.

Figure 10.3 *The zigzag arrangement of carbon atoms in octane*

The tetrahedral nature of the bonding involving carbon atoms means that the spine of the molecule, the carbon chain, has a zigzag arrangement when there are more than two carbons present. However, all of the alkane molecules studied so far are referred to as straight chain molecules, given the absence of any branches or rings.

Uses of alkanes

Figure 10.4 *Alkanes are the main component of petrol*

Figure 10.5 *Methane is the main component of natural gas*

Alkanes are very suitable as fuels due to their high heats of combustion. Natural gas consists mostly of methane. Liquefied petroleum gas, LPG, is made from propane and butane. Petrol is a mixture of a large number of hydrocarbons, many of which are alkanes containing from five to ten carbon atoms. Fuel oil used in ships and power stations, and diesel, are also mixtures of alkanes.

10.11 State what you know about the physical states of alkanes.
10.12 Give two examples of the uses of alkanes as fuels.

Alcohols

The general formula for alcohols which are another homologous series of organic compounds is $C_nH_{2n+1}OH$. Alcohols are formed when a hydroxyl group replaces a hydrogen atom, OH, in an alkane. The carbon atoms, including that bonded directly to the hydroxyl group, are all tetrahedral.

Alcohols are named by replacing the *-e* at the end of the corresponding alkane name with *-ol*.

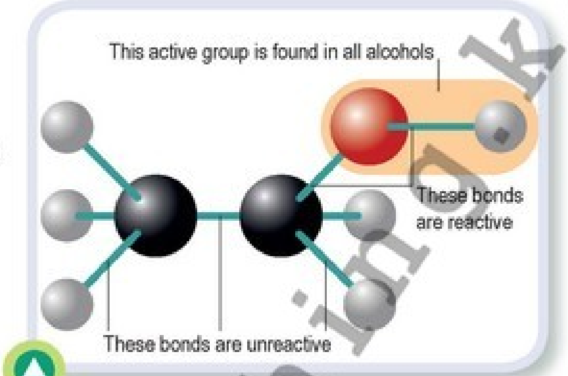


Figure 10.6 *The hydroxyl group is the active part of an alcohol molecule*

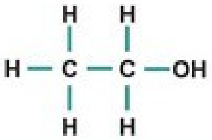
Sample question 3

Write the name, chemical formula and structural formula of the alcohol derived from ethane, C_2H_6.

Answer:

The name of the alcohol is ethanol. The formula is C_2H_5OH.

The structural formula is:

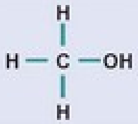
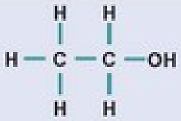
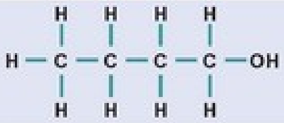


$H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-OH$

The position of the hydroxyl group on the carbon chain must be given when the chain contains three or more carbons. In alcohols with the hydroxyl group at the end of the carbon chain, the position of the hydroxyl group is given as 1. These alcohols are called primary alcohols.

Understanding U4

10.13 Complete the missing parts of the table.

Name of alcohol	Chemical formula	Structural formula
Methanol		 $H-\overset{H}{\underset{H}{C}}-OH$
	CH_3CH_2OH	 $H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-OH$
Propan-1-ol		
Butan-1-ol	$CH_3CH_2CH_2CH_2OH$	 $H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-OH$

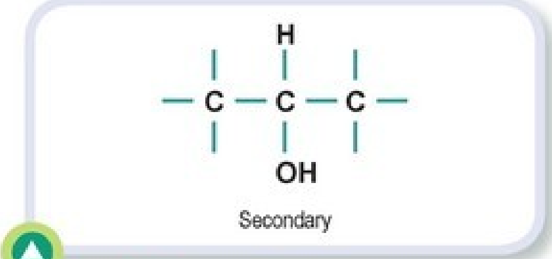


Figure 10.7 *Secondary alcohols*

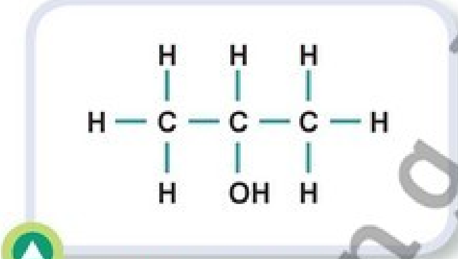


Figure 10.8 *Propan-2-ol*

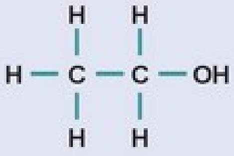
Secondary alcohols have two carbon atoms attached to the carbon with the hydroxyl group. This means that the hydroxyl group cannot be at the end of the carbon chain. In propan-2-ol, $CH_3CH(OH)CH_3$, for example, the hydroxyl group is attached to the middle carbon atom. This carbon is attached in turn to two other carbon atoms.

10.14 What is the general formula of the alcohol homologous series?

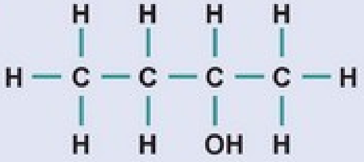
10.15 Explain the difference between primary and secondary alcohols.

10.16 Name the two alcohols shown in the diagrams and write their molecular formula.

(a)



(b)



10.17 Write the name, chemical formula and structural formula of the secondary alcohol containing four carbon atoms.

Comparing the alkanes and alcohols shows how the chemical properties of different homologous groups differ. The alkane propane ($M_r = 44$), and the alcohol ethanol ($M_r = 46$), might, for example, be expected to have similar boiling points. However, propane (B.P. = 231 K) is a gas at room temperature, while ethanol (B.P. = 351 K) is a liquid. **Table 10.4** compares the boiling points of the first four primary alcohols with those of alkanes of similar relative molecular mass.

Table 10.4

Alkane	Formula	B.P.	Phase at 298 K	Alcohol	Formula	B.P.	Phase at 298 K
Ethane	C_2H_6	184 K	Gas	Methanol	CH_3OH	338 K	Liquid
Propane	C_3H_8	231 K	Gas	Ethanol	C_2H_5OH	351 K	Liquid
Butane	C_4H_{10}	272 K	Gas	Propan-1-ol	C_3H_7OH	371 K	Liquid
Pentane	C_5H_{12}	309 K	Liquid	Butan-1-ol	C_4H_9OH	390 K	Liquid

In each case, the alcohol is a liquid at room temperature, and has a higher boiling point than the corresponding alkane.

Planar carbon

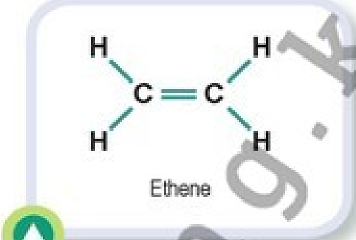


Figure 10.9 *Ethene molecules have a planar shape*

Ethene, C_2H_4, is a planar molecule. It has two carbon atoms joined by a double bond to each other, with two hydrogen atoms bonded to each carbon atom. Molecular models may be used to show that all of the atoms are in the same plane.

Planar carbon atoms feature in a number of homologous series – those whose compounds have a carbon–carbon (C=C) or carbon–oxygen (C=O) double bond. **The atoms in the double bond are planar**, although other carbon atoms in the molecule may be tetrahedral. The presence of a double bond means that these compounds are **unsaturated**.

Alkenes, aldehydes, carboxylic acids, are all groups of unsaturated organic compounds. All compounds of these types possess one or more planar carbon atoms. We shall examine these groups and their properties more closely in Module 11. Here we just show their names and general formula.

Aldehydes

Aldehydes form a homologous series of compounds, of general formula $C_nH_{2n+1}CHO$, where n = 0, 1, 2, 3 etc. All aldehydes contain the **carbonyl group**, which consists of an oxygen atom attached to a carbon atom by a polar double covalent bond, C=O. In aldehydes, there is always a hydrogen atom bonded to the carbonyl carbon. The functional group of the aldehydes is CHO.

Aldehydes are named by replacing the *-e* at the end of the corresponding alkane name with *-al*. The molecular formula is worked out by changing the last carbon in the chain of the alkane from being a part of a methyl group, CH_3, to being part of the aldehyde functional group, CHO.

Understanding U4

10.18 Write the name and chemical formula of the aldehyde containing 2 carbon atoms.

Carboxylic acids

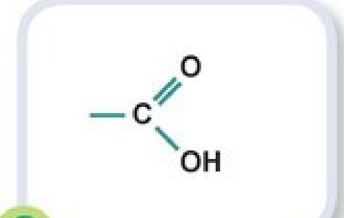


Figure 10.10 *The carboxyl group*

Carboxylic acids form a homologous series of compounds, of general formula $C_nH_{2n+1}COOH$, where n = 0, 1, 2, 3 etc. The carboxyl group, COOH, is the functional group of carboxylic acids. The structural formula of a carboxylic acid shows the carboxyl group to be a combination of the carbonyl group, with its polar double bond, C=O, and the hydroxyl group, OH.

Carboxylic acids are named by replacing the *-e* at the end of the corresponding alkane name with *–oic acid*. The molecular formula is worked out by replacing the last methyl group in the chain of the alkane with a carboxyl functional group.

10.19 Write the name, and chemical formula of the carboxylic acid containing two carbon atoms.

10.20 Give two examples of bond types that involve a planar carbon atom.

10.21 What is the general formula of the aldehydes?

10.22 Write down the structural formula of
  (a) the functional group of the aldehydes
  (b) the carbonyl group.

10.23 Write the name and structural formula of
  (a) ethanal **(b)** propanal

10.24 Name the aldehydes
  (a) HCHO **(b)** $CH_3CH_2CH_2CHO$.

10.25 Draw the structural formula of ethanoic acid.

10.26 Is the carboxyl carbon tetrahedral or planar?

Structural isomers

If a ball and stick model of pentane, C_5H_{12}, is disassembled, it can be reassembled into a different structure while still ensuring that each carbon atom forms four bonds.

The two molecules have the same molecular formula but have different structural formulas. They are **structural isomers** of each other.

> Structural isomers are compounds that have the same molecular formula but different structural formulas.

Naming the new molecule is not as straightforward as naming a straight chain molecule. You need to work out three things:

1. The number of carbons in the longest straight chain in the molecule.
2. The group(s) attached to the longest chain. (A group is formed when an alkane molecule loses a hydrogen atom, e.g. when a methane molecule loses a hydrogen atom the methyl group, CH_3, remains.)
3. The positions on the chain where the group(s) are attached. (The numbering of positions on a carbon chain begins from the end of the chain that gives the lowest numbered labels to the side chains.)

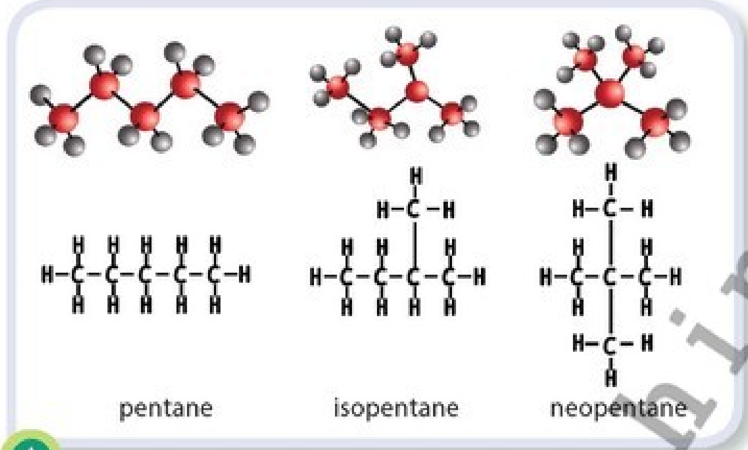


Figure 10.11 *The isomers of pentane, C_5H_{12}*

Sample question 4

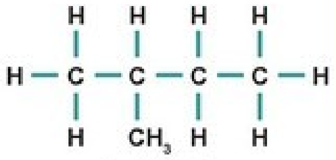
Draw the structural formula of 2-methylbutane.

Answer:

Since there are four carbons in the longest chain, the molecule is regarded as a butane derivative.

The presence of a methyl group attached to the longest straight chain means that the compound is called methylbutane.

Since the methyl group is attached to the second carbon in the chain, the compound is called 2-methylbutane.



2-methylbutane

Sometimes the structural formula is written in abbreviated or condensed form.

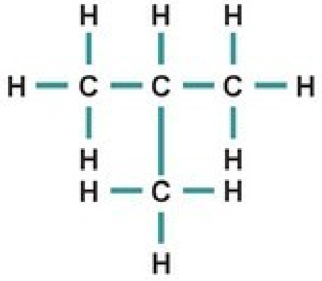
- The structure is not drawn, but can be deduced from the condensed form.
- Atoms joined to carbons on the main chain are written immediately after the carbon to which they are joined.
- In the case of groups such as the methyl group joined to carbons on the main chain, the formula of the group is written in brackets immediately after the carbon to which they are joined.

Sample question 5

Draw the structural formula and name the compound $CH_3CH(CH_3)CH_3$.

Answer:

The methyl group in brackets forms the branch or side chain. The remaining three carbons form the longest straight chain, so the compound is a derivative of propane. The side chain methyl group is written after the second carbon in the chain, which means that it is attached to this carbon. Thus the compound is 2-methylpropane. Its structural formula is as shown.



10.27 Explain the term structural isomers.

10.28 Name the structural isomers of pentane.

10.29 Write the condensed structural formulas and draw the full structural isomers of pentane.

10.30 Name and draw the structural formula of $CH_3C(CH_3)_2CH_3$.

10.31 Name the structural isomers represented by these formulas and draw the structural formulas:

(a) $CH_3CH(CH_3)CH_2CH_3$.   **(b)** $CH_3CH(CH_3)CH_3$.

Molecular formulas

Covalent compounds are made up of molecules, and the simplest ratio of atoms of each element in a molecule – the empirical formula – provides only a limited amount of information about the molecule.

The **molecular formula** of a compound gives extra information about the compound. The molecular formula of a covalent compound shows **the actual number of atoms of each element present in a molecule of the compound.**

Calculating the molecular formula of gaseous substances by mass fractions and relative density of elements

To find the molecular formula of a compound, both its empirical formula and its relative molecular mass must be known. We will perform this type of calculation below relating to some of the hydrocarbon gases we have seen in this Module.

Sample question 6

The relative molecular mass of propene is found to be 42. On analysis, it is found to contain 85.7% carbon and 14.3% hydrogen by mass. Find the molecular formula of propene.

Answer:

Element	Percentage	Percentage / Ar	Simplest Ratio
carbon	85.7	85.7 / 12 = 7.14	1
hydrogen	14.3	14.3 / 1 = 14.3	2

Empirical formula of propene = CH_2

Formula mass of CH_2 = 14

Relative molecular mass of propene = 42

Number of CH_2 units in a propene molecule = 42 / 14 = 3

Molecular formula of propene = C_3H_6

Q Understanding U4

10.32 Butane gas contains only C and H.

(a) If the compound contains 82.8% C, calculate the empirical formula.

(b) If its molecular mass is 58g/mol, calculate its molecular formula.

10.33 A gas is found to consist of 81.8% carbon and 18.2% hydrogen. After using a mass spectrometer, it is shown that it has a molecular mass of 44.

(a) Calculate the empirical and molecular formula.

(b) Identify the hydrocarbon gas.

MODULE 11

Hydrocarbons

Learning outcomes

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

- Describe the chemical properties of alkanes and confirm them through reaction equations 9.4.3.7
- Explain the meaning of the chlorination of alkanes, their use as solvents and dangers associated with these uses 9.4.3.8
- Describe the concept of unsaturation 9.4.3.9
- Investigate the properties of alkenes (combustion, hydrogenation, hydration, halogenation, test reactions) and confirm them by reaction equations 9.4.3.10
- Explain the specific structure of polymers and the mechanism of polymerization reactions in terms of polyethylene 9.4.3.11
- Explain and explore the problem of the destruction of plastic and assess the consequences of the accumulation of plastic materials in the environment 9.4.3.12
- Investigate properties of alkynes and confirm them by reaction equations 9.4.3.13
- Describe the preparation, properties and uses of benzene 9.4.3.14

Keywords

✓ fossilised ✓ deposit ✓ prefix ✓ substitution reaction ✓ stratosphere
✓ hydrogenation ✓ hydration ✓ polymer ✓ polythene
✓ carbon footprint ✓ incineration ✓ landfill ✓ biodegradable

Sources of Hydrocarbons

The fossil fuels, crude oil, natural gas and coal are sources of hydrocarbons. These fuels are formed from dead plants and animals that fossilised over 200 million years ago.

Crude oil and natural gas were formed from the remains of marine animals and plants that sank to the bottom of ancient seas and lakes. The resulting conditions were ideal for the anaerobic decay of this material to form crude oil and the gases associated with it. Natural gas often forms a pocket at the surface of oil reservoirs.

Layers of rock trapped both oil and gas deposits. Coal was produced in similar circumstances, but from dead wood from ancient forests that fell into swamps.

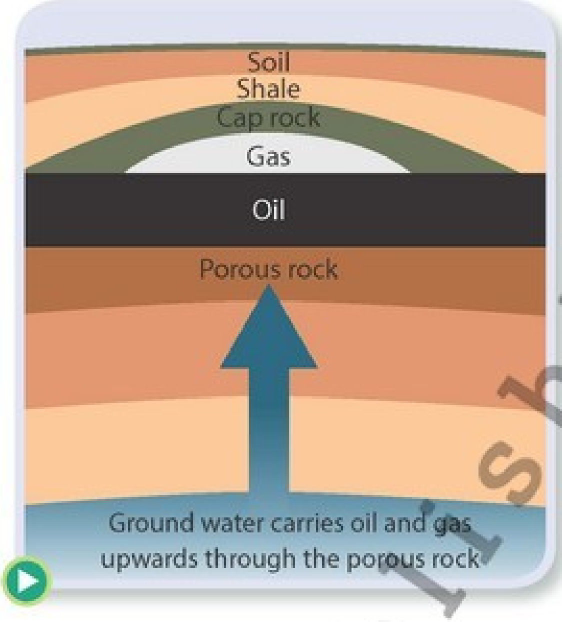


Figure 11.1

Alkanes

Alkanes form a homologous series of aliphatic hydrocarbons, of general formula C_nH_{2n+2}. They are saturated compounds, containing only single bonds.

The alkanes as a group are typically unreactive. Neither the C-HC-H bonds nor the C-CC-C bonds in typical saturated compounds react with strong acids at standard temperatures, nor are they oxidised by oxygen or reduced by hydrogen at such temperatures.

At high temperatures alkanes in the presence of excess oxygen will undergo complete combustion. Here is the equation for the complete combustion of propane:

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

Understanding U4

11.1 Write the balanced chemical equation for the combustion of:
(a) methane **(b)** butane

Chloroalkanes

A chloroalkane is an alkane in which one or more hydrogen atoms are replaced with chlorine atoms. For example, if one hydrogen atom in methane is replaced with a chlorine atom, chloromethane, CH_3Cl, is formed.

Chloroalkanes are named after the alkane from which they are derived. The prefix **chloro-** indicates the presence of a chlorine atom. A further prefix, such as **di-** or **tri**, is needed when more than one chlorine atom is present.

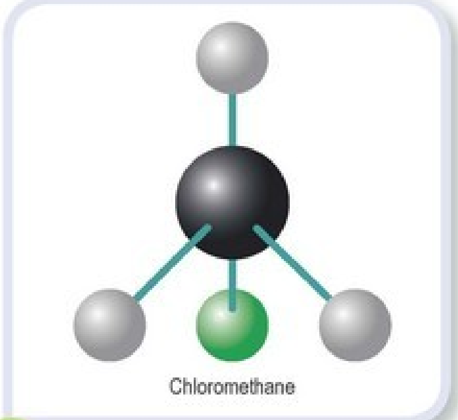
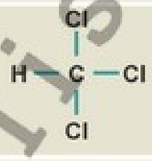


Figure 11.2 *Structure of chloromethane. The central carbon atom is still tetrahedral.*

Q U4 Understanding

11.2 Complete the table below.

Name of compound	Chemical formula	Structural formula
Chloromethane	CH_3Cl	
Dichloromethane		
Trichloromethane		
Tetrachloromethane		

Alkanes undergo a **substitution reaction** with halogens in the presence of light. In the reaction between an alkane such as methane and chlorine, for example, hydrogen atoms are successively replaced by chlorine atoms.

$$\boxed{CH_4 + Cl_2 \rightarrow CH_3Cl + HCl}$$

Q U4 Understanding

11.3 Look at the reactions below and name the products of each one.

(a) $CHCl_3 + Cl_2 \rightarrow CCl_4 + HCl$

(b) $CH_3 + Cl_2 \rightarrow CH_2Cl_2 + HCl$

Chloroalkanes are particularly useful as solvents. Because of their lack of polarity, they readily dissolve grease and oil and other substances that are insoluble in water. For this reason they have been used extensively in the dry-cleaning of clothes.

However, some chlorine compounds, such as chloromethane, are believed to cause damage to the ozone layer on reaching the stratosphere. Alternatives are now used in many cases. For example, for environmental reasons 1,1,1-trichloroethane is no longer used as a solvent for correcting fluid.

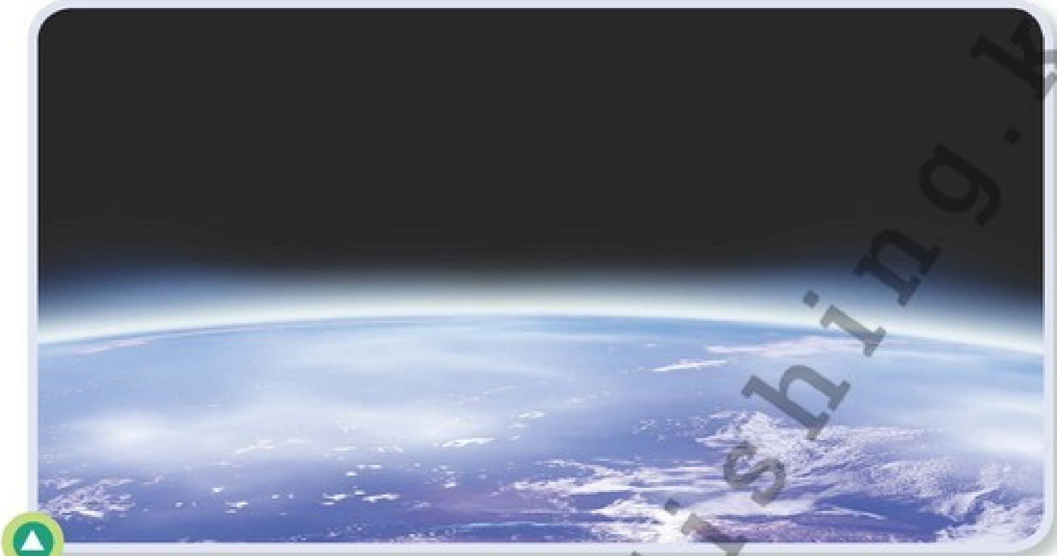


Figure 11.3 *The stratosphere.*

Alkenes

Alkenes form a homologous series of hydrocarbons, of general formula C_nH_{2n}. Each alkene molecule contains a carbon–carbon double bond, C=C, which is the reactive portion or functional group of the molecule. Since a carbon–carbon double bond must be present, it is not possible for an alkene with only one carbon atom to exist. The lowest member of the series has two carbon atoms.

Molecules that contain a double or triple bond are unsaturated.

The first three members of the series are gases at room temperature. Longer chain alkenes are liquids or solids. An increase in boiling point, as the number of carbon atoms increases, is typical of all hydrocarbons.

Reactions of Alkenes

The alkenes, especially ethene, are very important industrial chemicals, being much more reactive than alkanes. The presence of the C=C double bond allows alkenes to react in ways that alkanes do not.

Here we will look at four types of reaction with the alkenes.

Combustion

In a sufficient supply of oxygen, alkenes **combust completely** to produce **carbon dioxide** and **water**. In limited supplies of oxygen, incomplete combustion occurs producing carbon monoxide and carbon (soot).

11.4 Complete the equations for the three reactions:

(a) Complete combustion of ethene $C_2H_4 + 3O_2 \longrightarrow$ ____________

(b) Incomplete combustion ethene $C_2H_4 + 2O_2 \longrightarrow$ ____________

$C_2H_4 + O_2 \longrightarrow$ ____________

Activity 11.1

Question

How does ethene combust and how can we test for the presence of alkenes?

Equipment needed

Bromine water
Dilute acidified potassium permanganate solution
Ethanol
Aluminium oxide
Lime water
Glass wool
Water
Test-tube rack
5 test tubes
5 rubber stoppers for test tube

5 test tubes
Retort stand
Boiling tube
Rubber stopper with hole
Delivery tube
Spatula
Bunsen burner
Glass rod
Splint
Matches

Safety

- Students should wear safety goggles at all times when observing the demonstration.

Conducting the activity

Your teacher will demonstrate the following three procedures:

1. The production and collection of ethene gas.
2. The combustion of ethene.
3. Tests for the presence of alkenes.

Make notes and record your observations.

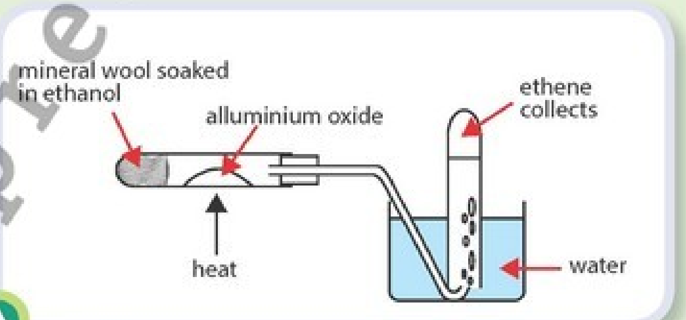


Figure 11.4

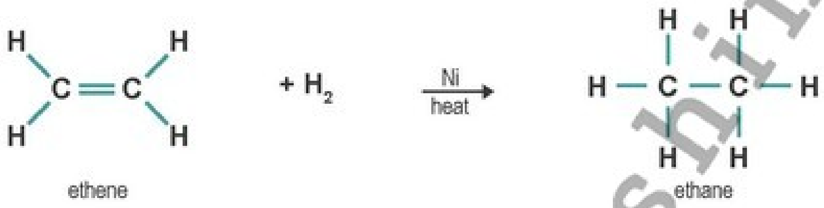
11.5 What can you conclude from your observations?

11.6 Write the equation for the production of ethene in the demonstration.

11.7 What is the difference in alkane and alkene structure that bromine water and dilute acidified potassium permanganate test for?

Hydrogenation

The addition of hydrogen to alkenes, is a very important reaction industrially. Hydrogenation of double bonds is used in industry to convert vegetable oils, which are unsaturated, into solid saturated materials used in margarine and dairy spreads. By controlling the degree of hydrogenation, the spread can be made as soft or hard as is required.

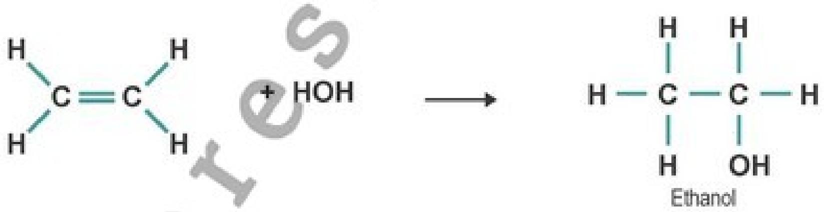


Understanding U4

11.8 Write the chemical formula for this reaction.

Hydration

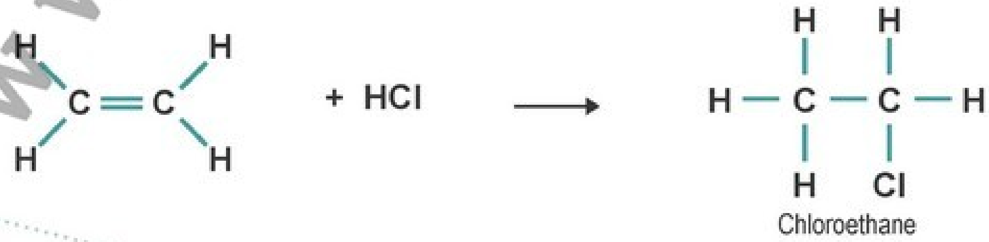
The addition of water is called a **hydration** reaction. It is the most frequently used reaction for manufacturing ethanol, which is very widely used in industry as a solvent.



Understanding U4

11.9 Write the chemical formula for this reaction.

Addition of hydrogen chloride



11.10 Write the chemical formula for this reaction.

The chloroethane made from such reactions is widely used in the manufacture of a substance called ethylcellulose which is used as a thickening and binding agent in paints and cosmetics.

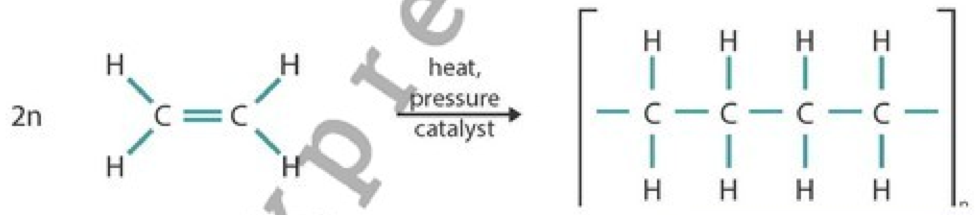
11.11 Look at the changes in structural formula in the above three reactions and complete the sentence.
In addition reactions with alkenes the structure changes from planar to ___________.

This allows us to tell alkenes apart from alkanes using a simple chemical test.
Bromine water is an orange solution of bromine. It becomes colourless when it is shaken with an alkene. Alkenes can decolourise bromine water, but alkanes cannot.

Polymers

A polymer is a large molecule made up of many identical repeating units called monomers.

Alkenes are the raw materials in the industrial manufacture of polymers.
In the polymerisation of ethene, the molecule **adds to itself** to produce a saturated hydrocarbon chain that consists of repeating –CH_2CH_2– units.



The resulting solid substance is called poly(ethene) or polythene and is an example of an addition polymer.

Depending on the conditions used, two forms of polythene can be produced, low-density polythene and high-density polythene.

Low-density polythene

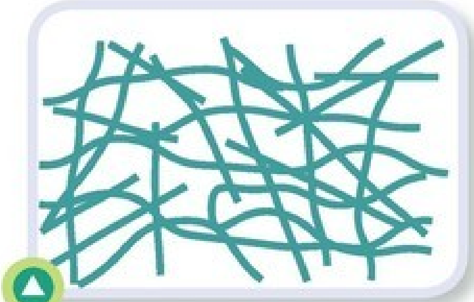


Figure 11.5 *Low-density polythene: more branched chains.*

This type of polythene has a significant degree of branching. This prevents the molecules from packing closely together. Because the molecules take up more space, the density of the polymer is reduced. Low-density polythene is very flexible and melts at about 360 K.

High-density polythene

There are few side branches in this form of polythene and this allows the molecules to pack more tightly. Because the molecules take up less space, the density of the polymer is increased. High-density polythene has a higher melting point, about 400 K. It is more rigid than low-density polythene.

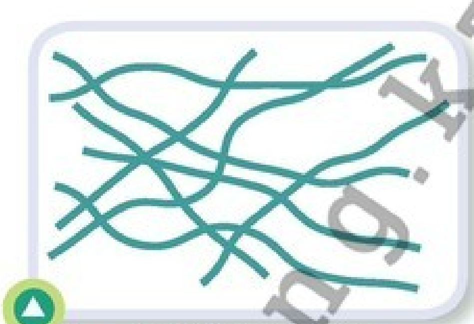


Figure 11.6 *High-density polythene: fewer branched chains*

What is the life cycle of plastic?

Have you ever considered where the plastics we use come from or where they go when you finish with them? Every product we use goes through a life cycle, and each stage of the life cycle has environmental impacts. Today this is particularly apparent as so many of the products we use and consume are packaged in single-use plastics e.g. plastic bags made from LDPE.

One impact is carbon dioxide. Any stage in a product's life that releases carbon dioxide into the atmosphere affects the environment and the total effect is calculated as the **carbon footprint** of the product.

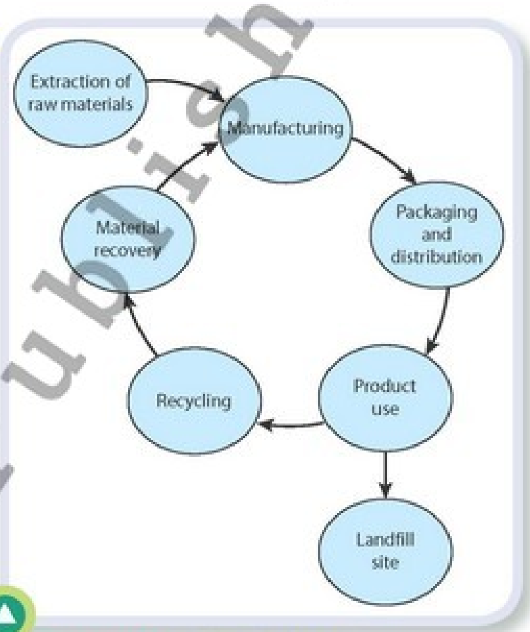


Figure 11.7 *The life cycle of a product.*

However, if we reduce the use of materials at each stage then this will lessen the impact on the environment.

Q C1 Communicating

11.12 Draw a diagram showing the stages in the life cycle of a supermarket plastic bag.

Plastic Pollution

The world's use of plastic is a growing environmental disaster since most plastics are made from petroleum or natural gas, as in the case of polyethylene. Hydrocarbons are non-renewable resources extracted and processed in energy-intensive ways that contribute to environmental pollution.

Figure 11.8 *A landfill waste disposal site.*

The manufacture of plastic, as well as its destruction by incineration, pollutes the air, land and water systems as do plastics that go into landfill waste. Synthetic plastic does not biodegrade. It is not known how long it takes for plastic to break down, but it is believed to take hundreds of years and during photo-decomposition it releases toxins into the environment which pass through food chains through bioaccumulation.

Q Communicating C1 Communicating C2 Society S2

11.13 Explain how sustainable products are good for the environment.

11.14 What is meant by cradle-to-cradle design?

Packaging

In our modern society we have become more dependent on packaging, as households have become increasingly busy and convenience has become an important factor in purchases. While packaging does improve safety and convenience and reduces theft, it also has a number of disadvantages.

It can be expensive and can cause damage to our environment because most of the waste produced by packaging ends up in landfill. For example, some meat products are packed in polyethylene, which is a petroleum-based plastic and therefore cannot be recycled.

Sustainable packaging is very important so that we can reduce the waste in our landfill sites. Some examples are:

- Many food processors have reduced their packaging by offering reusable and refillable containers. For example, branded coffee where the refill packaging is manufactured from cardboard or paper which can be recycled and composted.
- Supermarkets are buying products from suppliers that use recyclable material in their packaging. Polyethylene terephthalate (PET) is 100% recyclable, making it ideal for packaging. It is light and easily crushed and so is easy to transport to the recycling plant.
- The biodegradable plastic polylactide (PLA) is processed from the starch of plants such as corn or sugarcane and is ideal for shopping bags because it is a compostable material.

The key factors when choosing a sustainable material is that it is from easily grown materials, is recyclable and is locally available.

11.15 A company called Ecolean has developed ecologically sustainable food packaging that is not a fossil fuel-based material, but is made with the natural material chalk (limestone).

Research answers to the following questions:

(a) Where is limestone formed in nature?

(b) What is the name of this new packaging material?

(c) Explain why this new material has less impact on our environment.

(d) What products is this new sustainable packaging material used for?

11.16 What three conditions are necessary to produce polyethene from ethene?

11.17 Is polyethene a saturated or unsaturated compound?

11.18 Describe three ways in which the use of plastics harms the environment.

11.19 What is meant by the term bioaccumulation?

Alkynes

The alkynes form a homologous series of aliphatic hydrocarbons of general formula C_nH_{2n-2}. Each alkyne molecule contains a carbon–carbon triple bond, $C \equiv C$, and this is the functional group of the series. Like the alkenes, it is not possible for an alkyne with only one carbon atom to exist. The lowest member of the series has two carbon atoms.

Like most hydrocarbons, ethyne gas burns in air to form carbon dioxide and water vapour.

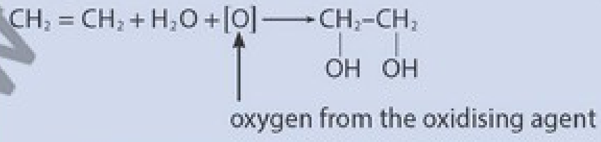
Complete the equation for the complete combustion of ethyne.

$$\boxed{C_2H_{2(g)} + \tfrac{5}{2}O_{2(g)} \rightarrow}$$

As an unsaturated compound, ethyne decolourises both bromine water and acidified potassium manganate(VII) solution.

Did you know?

In organic chemistry we can make equations involving oxidation like this:



$$CH_2 = CH_2 + H_2O + \underset{\substack{\uparrow \\ \text{oxygen from the oxidising agent}}}{[O]} \longrightarrow \underset{OH}{CH_2}-\underset{OH}{CH_2}$$

[O] stands for oxygen from the oxidising agent.

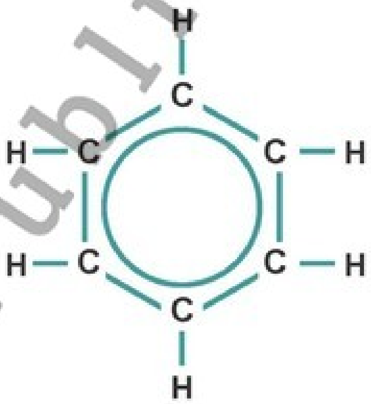
Putting in the whole reaction equation e.g. potassium manganate (VII) solution above would obscure the organic change we wish to focus on. The equation here shows the oxidation of ethene to ethylene glycol.

Ethyne is commonly known as acetylene gas. Due to its high proportion of carbon, it burns in air with a sooty flame. However, when it is burned, while mixed with pure oxygen, it produces a clear flame of about 3000 °C. This makes it suitable for cutting and welding, and it is used for those purposes in an oxyacetylene torch.


Figure 11.9

Benzene

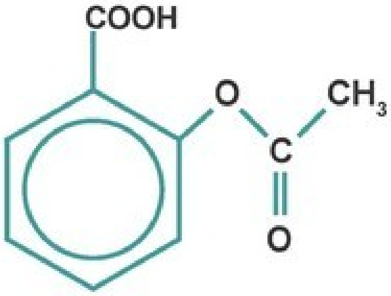
Benzene is part of a group known as the aromatic compounds The benzene molecule consists of six carbon atoms joined to form a hexagonal planar ring. Each carbon atom has four valence electrons. One of these is used to bond with a hydrogen atom, and two others are used to form sigma bonds with the carbon atoms on either side. This leaves one valence electron, which is not involved in sigma bonding, on each carbon atom. Like alkanes, aromatic compounds are obtained from the fractional distillation of crude oil. The difference from alkanes in that aromatic compounds contain a benzene ring.



Aromatic compounds

Hydrocarbons do not dissolve to any extent in water. Compounds such as methylbenzene are, however, themselves good non-polar solvents, dissolving non-polar solutes readily. Methylbenzene, $C_6H_5CH_3$, is a widely used industrial solvent.

A huge range of aromatic compounds have been discovered, or synthesised in the laboratory. Some, such as benzene, are carcinogens, meaning that they are capable of causing cancer in humans. Very strict safety protocols must be followed in handling and storing such products. Many others, however, are non-carcinogenic e.g. the common painkiller, aspirin.



MODULE 12

Groups of organic compounds

Learning outcomes

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

- Recognise the classification of oxygen-containing organic compounds 9.4.3.18
- Understand the classification of alcohols; explain physical and chemical properties and production and uses of ethanol and methanol 9.4.3.19
- Explain the physiological impact of methanol and ethanol on the human organism 9.4.3.20
- Recognise the physical properties and uses of ethylene glycol and glycerol 9.4.3.21
- Investigate the composition of carboxylic acids, their chemical properties and uses of acetic acid 9.4.3.22
- Explain the structure and properties of esters and lipids and describe the role of lipids and their significance 9.4.3.23
- Understand how soaps are made and used 9.4.3.24
- Explain the impact of soaps and detergents on the environment 9.4.3.25
- Explain the classification, biological role and significance of carbohydrates 9.4.3.26
- Describe the formation of peptide bonds between α-amino acids in proteins 9.4.3.27
- Investigate the denaturation of proteins 9.4.3.28
- Explain the biological importance and functions of proteins 9.4.3.29

Keywords

✓ IUPAC system ✓ fermentation ✓ distillation ✓ denaturing agent
✓ esters ✓ non-polar solvents ✓ long-chain hydrocarbons
✓ eutrophication ✓ algae ✓ polysaccharides ✓ peptide bond
✓ substrate ✓ product ✓ globular proteins ✓ fibrous proteins

In Module 10 we considered the molecular and structural formulas of different homologous groups and how these are named in the IUPAC system.

Here we will further explore some of the chemical properties and uses of organic compounds in those groups and others.

Alcohols

The alcohols form a homologous series of organic compounds of general formula $C_nH_{2n+1}OH$. The carbon atoms, including those bonded directly to the hydroxyl group, are all tetrahedral.

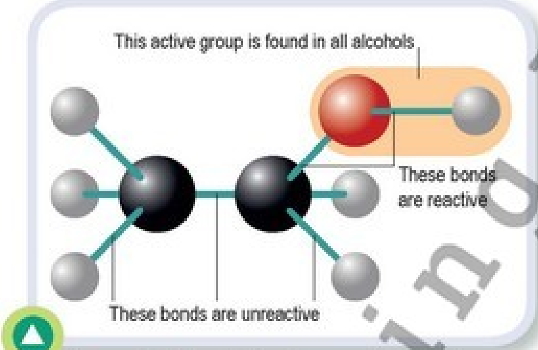
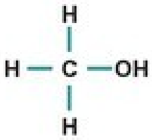


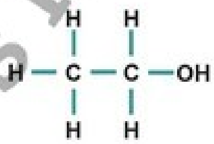
Figure 12.1 *The hydroxyl group is the active part of an alcohol molecule*

In methanol, the single carbon is bonded to three hydrogen atoms, and not to a carbon atom.



$$\begin{array}{ccc} & H & \\ & | & \\ H - & C & - OH \\ & | & \\ & H & \end{array}$$

The structural formula for ethanol, the next alcohol in the homologous group, is:



$$\begin{array}{cccc} & H & H & \\ & | & | & \\ H - & C - & C & - OH \\ & | & | & \\ & H & H & \end{array}$$

12.1 Write the molecular formulas for:

(a) methanol **(b)** ethanol

The presence of the polar OH group in alcohols, results in hydrogen bonding between neighbouring molecules.

Oxygen, being more electronegative, has a partial negative charge, and hydrogen has a partial positive charge.

The oxygen in a hydroxyl group in one molecule attracts the hydrogen in the hydroxyl group on a neighbouring molecule. These strong hydrogen bonds have to be broken before the alcohol boils, so more energy must be supplied and the boiling point is correspondingly raised.

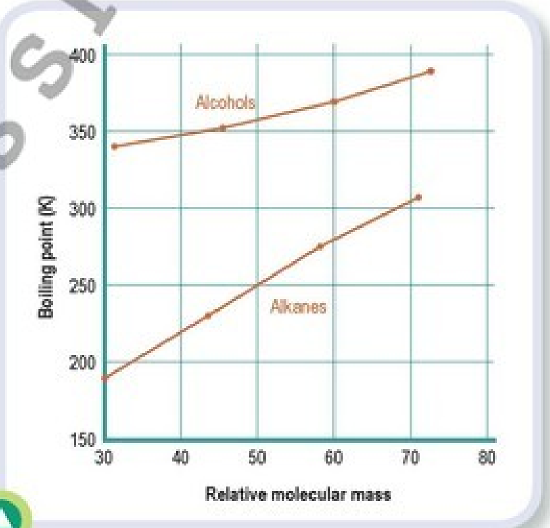


Figure 12.2 *Alcohols have much higher boiling points than alkanes of comparable relative molecular mass*

Wine and beer are mixtures of ethanol, water and some other substances. Ethanol and water in alcoholic drinks mix completely because hydrogen bonds form between the hydroxyl groups in ethanol and the hydroxyl groups in water. Methanol and propan- 1-ol also mix with water in all proportions. However, the solubility decreases as the hydrocarbon chain gets bigger and becomes more significant.

Manufacture and uses of methanol and ethanol

Most of the world's supply of ethanol is made by the hydration of ethene. It is manufactured by reacting ethene (from **cracking** crude oil fractions) with steam. Phosphoric acid is used as a catalyst to speed up the reaction.

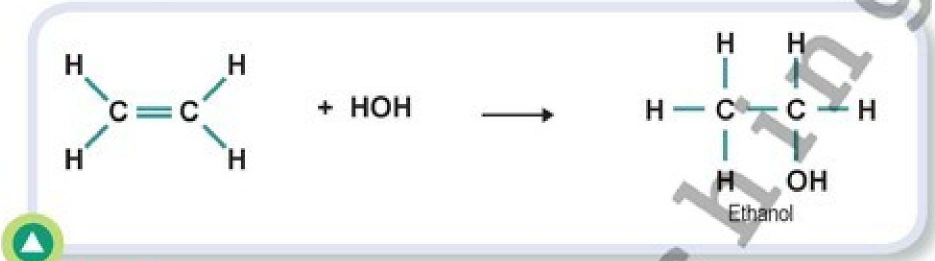


Figure 12.3

$$C_2H_4 + H_2O \xrightarrow[\text{high pressure}]{\text{high temperature}} C_2H_5OH$$

The fact that ethanol is the only product of this reaction, makes this a relatively straightforward and efficient manufacturing process.

The ethanol in alcoholic drinks is made by fermentation of sugars in fruits, such as grapes. Wine is made in this way. Fermentation involves a series of biochemical reactions brought about by enzymes contained in yeast. The enzymes in yeast break down the sugar to give carbon dioxide and ethanol.

Air is excluded during fermentation to prevent oxidation of ethanol to ethanal and ethanoic acid.

$$C_6H_{12}O_6 \xrightarrow{\text{enzymes in yeast}} 2C_2H_5OH + 2CO_2$$

The advantage of producing ethanol in this way, compared to the process of hydrolysing ethene, is that plant material is a renewable resource, whereas crude oil is not and its extraction and transportation are associated with many other environmental problems.

Fermentation is used in the brewing industry to make beer. In this industry, the raw material is malted grain, and the beer produced usually contains about 3% to 5% (v/v) ethanol.

Fermentation is also used in the distilling industry to make spirits. For example, malted barley is the raw material in the manufacture of whisky, and the fermented liquid is distilled to increase the ethanol concentration. Whisky usually contains about 40% (v/v) ethanol.

Figure 12.4 *Beer is made by fermentation*

The drinking of too much alcohol, either on a single occasion or over a longer period of time, can seriously impact a person's health.

Over-consumption of alcohol, particularly over longer periods, can impact various parts of the body and is linked to various conditions detailed in **Table 12.1**.

Table 12.1

Part of the body	Effects
Brain	impacts brain's communication pathways resulting in mood changes, loss of concentration and co-ordination
Heart	irregular heart beat, damage to heart muscle, strokes and high blood pressure
Liver	fatty liver and fibrosis
Pancreas	impacts function of the pancreas causing it to produce toxins
Immune system	impacts immune system increasing risk of infection

Other uses

- In some countries, ethanol made by the fermentation of sugar cane is mixed with petroleum products and used as a motor fuel instead of petrol.
- Ethanol is also a very good solvent and evaporates easily. It is used extensively as a solvent for paints, glues, perfumes and aftershaves.
- Ethanol is the major component of methylated spirits, which is burned in some types of camping stove. However, a small amount of methanol, which is poisonous, and a dye are added to the ethanol to make it undrinkable. Methanol used in this way is said to be a **denaturing agent**.

Figure 12.5 *Ethanol-based products*

Uses of other alcohols

Ethylene glycol is used in many commercial and industrial applications including anti-freeze and coolant. It is also used in the manufacture of polyester fibre for the making of clothes and carpets.

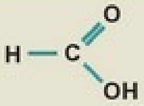
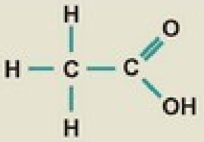
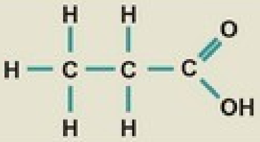
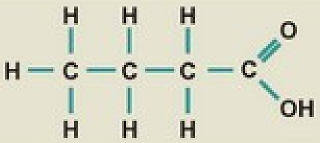
Glycerol is used in a range of personal care products such as toothpaste, shaving cream and hair care products, helping to soften and lubricate the mixtures.

Carboxylic Acids

We have seen that carboxylic acids form a homologous series of compounds, of general formula $C_nH_{2n+1}COOH$, where n = 0, 1, 2, 3 and so on.

Table 12.2 gives the name, formula and structural formula of each of the first four carboxylic acids.

Table 12.2

Name of carboxylic acid	Chemical formula	Structural formula
Methanoic acid	$HCOOH$	
Ethanoic acid	CH_3COOH	
Propanoic acid	CH_3CH_2COOH	
Butanoic acid	$CH_3CH_2CH_2COOH$	

Physical state and properties

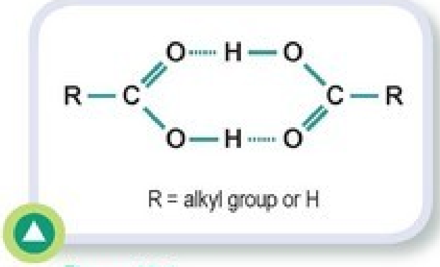


Figure 12.6

The lower carboxylic acids are colourless liquids, usually with sharp or distinctive smells. Boiling points are higher even than those of alcohols of similar relative molecular mass. These relatively high boiling points result from the formation of dimers, where two carboxylic acid molecules are held together by two hydrogen bonds. These arise due to polarity in the C=O bond and in the O–H bond. Each

H atom, which has a partial positive charge, forms a hydrogen bond with the carbonyl oxygen, which has a partial negative charge, on the neighbouring atom. The reverse also occurs, so two hydrogen bonds are formed.

Hydrogen bonding can also occur between water molecules and the lower carboxylic acids. As a result, acids containing up to four carbon atoms are highly soluble in water. However, as the carbon chain increases in length, the molecule becomes less polar and solubility in water decreases.

Figure 12.7 *Ethanoic acid is a constituent of vinegar*

Occurrence and uses

- Methanoic acid, HCOOH, is an irritant fluid emitted by some ants, and also found in nettles. Thus, an antidote to a nettle or ant sting would be to apply a weak base such as sodium hydrogencarbonate (bread soda) to the affected area.
- Ethanoic acid, CH_3COOH, is a very widely used carboxylic acid. It is the principal acid found in vinegar, which is usually made by oxidation by air of ethanol. Wine becomes undrinkable if it is left open to the air for too long, as the ethanol becomes oxidised to ethanoic acid, and a vinegary smell develops. Some fast-food outlets make vinegar simply by diluting ethanoic acid with water.
- Ethanoic acid is also used in the manufacture of cellulose acetate, which is used in lacquers, varnishes, photographic film, non-shatter glass, and as rayon fibre. The acid is one of a number of chemicals reacted with wood pulp to produce cellulose acetate.

Q Understanding U4 Research R1

12.2 What is the general formula of the carboxylic acids?
12.3 Is the carboxyl carbon tetrahedral or planar?
12.4 Explain why carboxylic acids have such high boiling points.
12.5 Give an example of a carboxylic acid found in nature.
12.6 Give an everyday use of ethanoic acid.

Lab Work 16

Activity 12.1

Question

How can we investigate the properties of ethanoic acid?

Equipment needed

Concentrated sulfuric acid
Ethanol
Deionised water
Universal indicator paper
Magnesium ribbon
Ethanoic acid (dilute)
Anhydrous sodium carbonate
Safety glasses
Water bath
Graduated cylinder
Bunsen burner
Tripod
Wire gauze
Tongs

Safety

- Use tongs to hold the test tube when heating over the Bunsen burner.

Conducting the activity

1. Smell ethanoic acid by carefully wafting some of the vapour towards your nose.
2. Compare its smell to that of ethanol and vinegar. Record your observations.
3. Dip some universal indicator paper in ethanoic acid to measure the pH. Record your result.
4. Drop a clean strip of magnesium into some ethanoic acid in a test tube, and swirl. Record your observations.
5. Add 1 g of anhydrous sodium carbonate powder to some ethanoic acid in a test tube, and swirl. Record your observations.
6. Carefully add 2 drops of concentrated sulfuric acid to some ethanoic acid in a test tube. Add 1 cm^3 of ethanol and warm gently. Carefully smell the reaction product. Record your observations.

Results

Test	Observation	Explanation
Odour detected		
Universal indicator paper dipped		
Magnesium strip added		
Anhydrous sodium carbonate added		
Ethanol with concentrated sulfuric acid mixed		

12.7 What did you observe at each stage of the experiment?
Write equations for the reactions in step 4 and step 5 of the activity.
What can you conclude about:

(a) the reaction with magnesium
(b) the reaction with sodium carbonate
(c) the reaction with ethanol

Esters

Esters are compounds derived from carboxylic acids. The formula of an ester can be represented as R_1COOR_2, where R_1 could be a hydrogen atom or an alkyl group, and R_2 is an alkyl group. All esters contain the ester linkage –COO–, which contains a planar carbon. Esters may be regarded as consisting of two parts: one, an alkyl group, R_2, derived from an alcohol, and the other part, R_1COO, derived from a carboxylic acid.

The name of each ester has two parts: firstly the alkyl group derived from the alcohol, and secondly the name of the acid, with the ending **-oic acid** replaced by **-oate**.

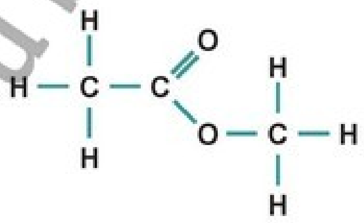
Sample question 1

Write the name, chemical formula and structural formula of the ester derived from ethanoic acid, CH_3COOH, and methanol, CH_3OH.

Answer:

The alcohol, methanol, provides the first part of the name, methyl. The carboxylic acid, ethanoic acid, provides the second part of the name, ethanoate. Thus, the ester is methyl ethanoate.

In writing the formula of the ester, the formula of the acid is written first with the carboxyl hydrogen atom omitted, CH_3COO^-. The alkyl group from the alcohol is then added in place of the omitted hydrogen, so the completed formula is CH_3COOCH_3. The structural formula is built up in the same way.



Methyl ethanoate

Table 12.3 and **Figure 12.8** give the name, formula and structural formula of each of the esters containing up to four carbon atoms.

Table 12.3

Name of ester	Chemical formula
Methyl methanoate	$HCOOCH_3$
Ethyl methanoate	$HCOOC_2H_5$
Propyl methanoate	$HCOOC_3H_7$
Methyl ethanoate	CH_3COOCH_3
Ethyl ethanoate	$CH_3COOC_2H_5$
Methyl propanoate	$C_2H_5COOCH_3$

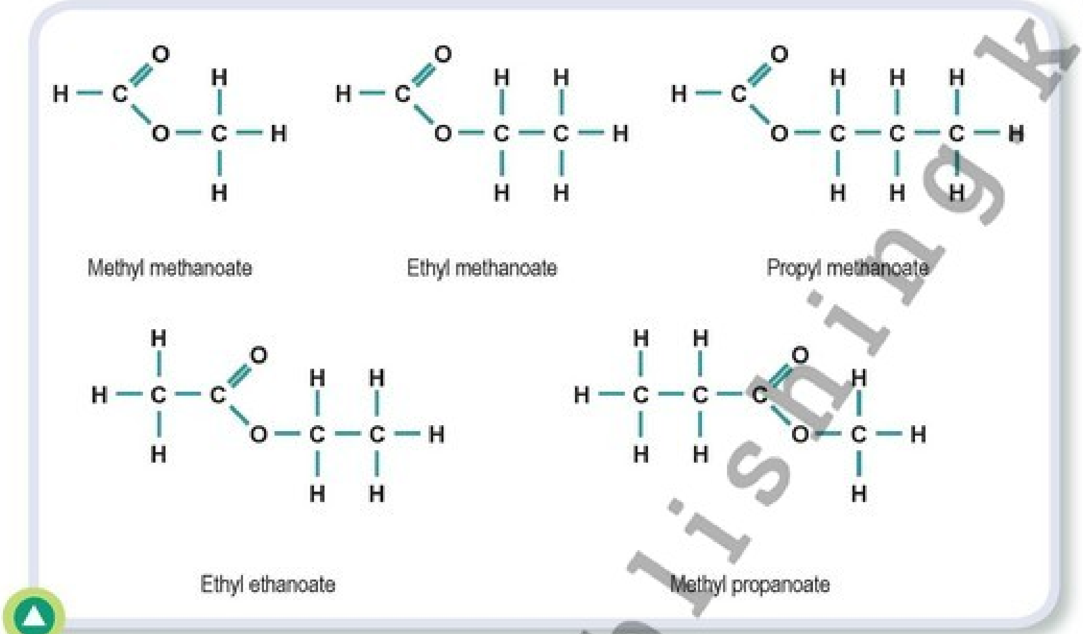


Figure 12.8

Physical state and properties

All of the esters listed in **Table 12.3** are liquids. Thus, the boiling points of esters are similar to those of aldehydes and ketones of similar relative molecular mass.

The polarity of the C=O group allows hydrogen bonding to occur with water molecules, so esters with fewer than five carbon atoms are water-soluble. Esters are soluble in non-polar solvents.

Occurrence and uses

- Members of the ester family have strong and often pleasant, fruity smells. Many of them occur naturally and are responsible for the flavour in fruits and the smells of flowers. An average mango, for example, contains at least eight esters, one of which is ethyl ethanoate.
- Many of the artificial flavours listed as ingredients of processed foods are esters that have been made in the laboratory.

Figure 12.9 *Mango contains a number of different esters*

- Ethyl ethanoate is used as a solvent for printing inks and paints. Its volatility means that its evaporation rate is suitable for this purpose.
- Fats and oils are naturally occurring esters of long-chain carboxylic acids. They are used as energy storage compounds by plants and animals.

Q Understanding U4

12.8 Draw the structure of the ester linkage.

12.9 Fill in the gaps: An ester may be regarded as a derivative of a ______________ and an ______________ .

12.10 Give two examples of esters found in nature.

12.11 Give an everyday use of ethyl ethanoate.

Soap

Soaps are the sodium or potassium salts of long-chain carboxylic acids. Soap is generally made by the base hydrolysis of the esters present in vegetable oils or animal fats. During soap-making, a mixture of vegetable oils and sodium hydroxide or potassium hydroxide is heated with steam.

Figure 12.10

The base hydrolysis of these substances produces glycerol, and the salt of the acid present, i.e. soap. Soap is a long-chain hydrocarbon with an ionic group at the end. The non-polar part dissolves grease, while the ionic end dissolves in water. This combination of properties gives it its cleansing action.

Soaps and detergents which contain phosphates can be harmful for the environment by causing eutrophication.

Eutrophication is the presence of excessive nutrient content in a water body to an extent that causes over-abundant plant growth.

Eutrophication is caused by the over-enrichment of water by nutrients such as phosphates in detergents and nitrates in fertilizers. In effect, the nutrients behave as fertilisers that increase the growth of plants such as algae in lakes and rivers. This problem may be noticed by the appearance of algal bloom on the water.

Figure 12.11 *Algal bloom*

The algae are shortlived. As they decay, micro-organisms use up much of the dissolved oxygen in the water, leading to the death of many forms of animal life. The occurrence of eutrophication has led many countries to limit or ban the use of phosphates in detergents. Highly eutrophic waters, i.e. rich in organic matter, are very poor in oxygen, and not very favourable to biodiversity and promote the proliferation of pathogens.

The bans on phosphate use typically cover household detergents and products for dishwashers.

Q Understanding U4

12.12 What is meant by eutrophication?

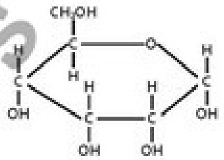
12.13 What is likely to happen to the water in a lake if there are high levels of phosphates from detergents present.

Structure of carbohydrates

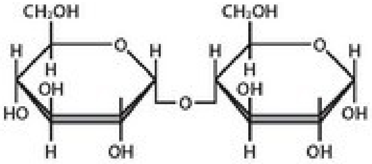
In Module 9 we saw that carbohydrates contain the elements of carbon, hydrogen and oxygen. These elements are present in the ratio $C_x(H_2O)_y$, where x and y are the same number.

There are three categories of carbohydrate:

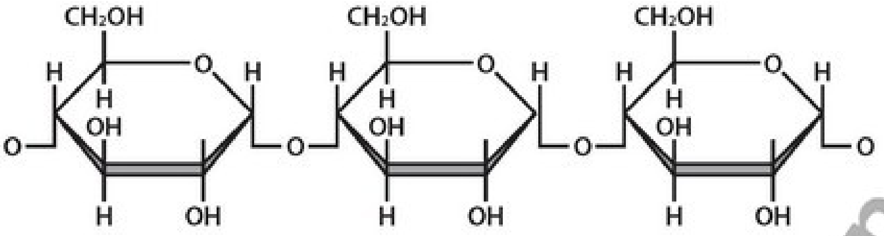
- Monosaccharides which consist of one sugar unit. Examples include fructose, glucose and ribose. They are white crystalline solids and dissolve in water to form sweet tasting solutions.



- Disaccharides which are made of two sugar units joined together. Examples include maltose (glucose + glucose) and sucrose (glucose + fructose). Monosaccharides are formed into disaccharides through a condensation reaction which involves two molecules combining to form a larger molecule, whilst producing a small molecule such as H_2O as a by-product.



- Polysaccharides which are made of many sugar units joined together. Examples include starch (found in flour, bread, potatoes, rice and pasta) and glycogen (stored by animals in the liver and muscles). Starch is insoluble due to its structure and is, therefore, the main storage sugar in plants.



An example of the structural role carbohydrates play in nature is: cellulose forming plant cell walls; their metabolic roles include the breaking down of glucose in animal respiration and photosynthesis in plants.

12.14 What is the ratio of carbon atoms to hydrogen atoms in a carbohydrate?
12.15 What do organisms need that is stored in carbohydrate bonds?
12.16 What is the by product of a condensation reaction between monosaccharides in forming a disaccharide?
12.17 What carbohydrate has a support role in plants?

Structure and shape of proteins

Proteins are composed of long chains of smaller molecules called amino acids. Figure 12.12 shows the relationships between amino acids and a protein.

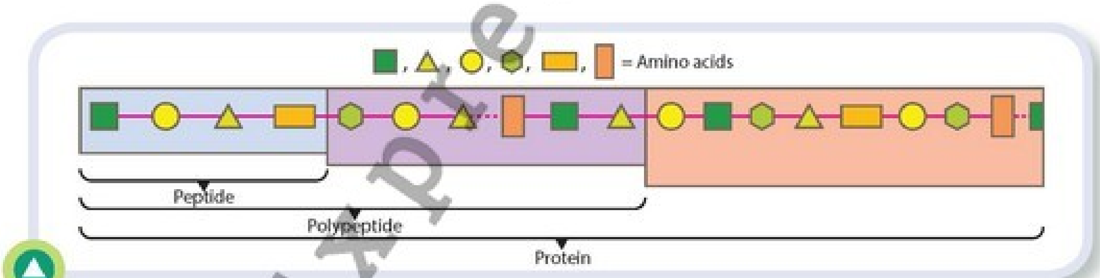
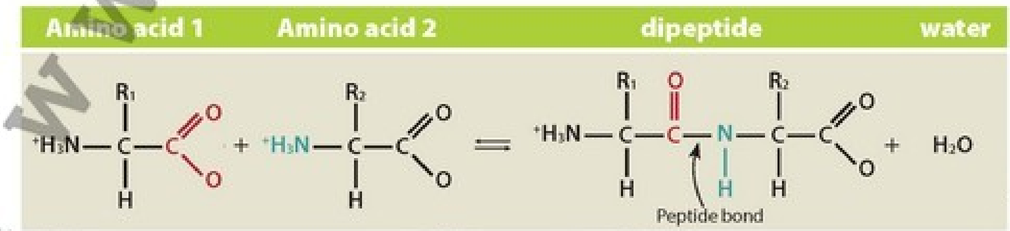


Figure 12.12

The bond between two amino acids is called a peptide bond. When two amino acids bond one loses a hydrogen and oxygen from its carboxyl group (COOH) while the other loses a hydrogen from its amino group (NH_2). The reaction produces a molecule of water (H_2O) and the two amino acids joined by a peptide bond (–CO–NH–).



The amino acids that make up a protein can be thought of as letters in an alphabet. By combining them in different sequences, nature can make them into a huge range of proteins.

It is not just the sequences of amino acids that determines how proteins work. How proteins fold into 3-D shapes is also essential to their effective functioning. Enzymes are proteins that are often referred to as biological (or organic) catalysts. They speed up chemical reactions in the body and allow them to proceed at normal cell temperatures. The specific 3-D shape of an enzyme means that it will fit neatly and react only with a substance – known as the substrate – of a shape that matches the enzyme.

Denaturation of enzymes

When most proteins are heated above 40°C, subjected to unsuitable pH values or treated with certain chemicals or radiation, they gradually lose their 3-D shape. This means that the active site as shown in Figure 12.13 will lose its ability to react with its substrate. Such a change in shape and loss of biological activity is called denaturation.

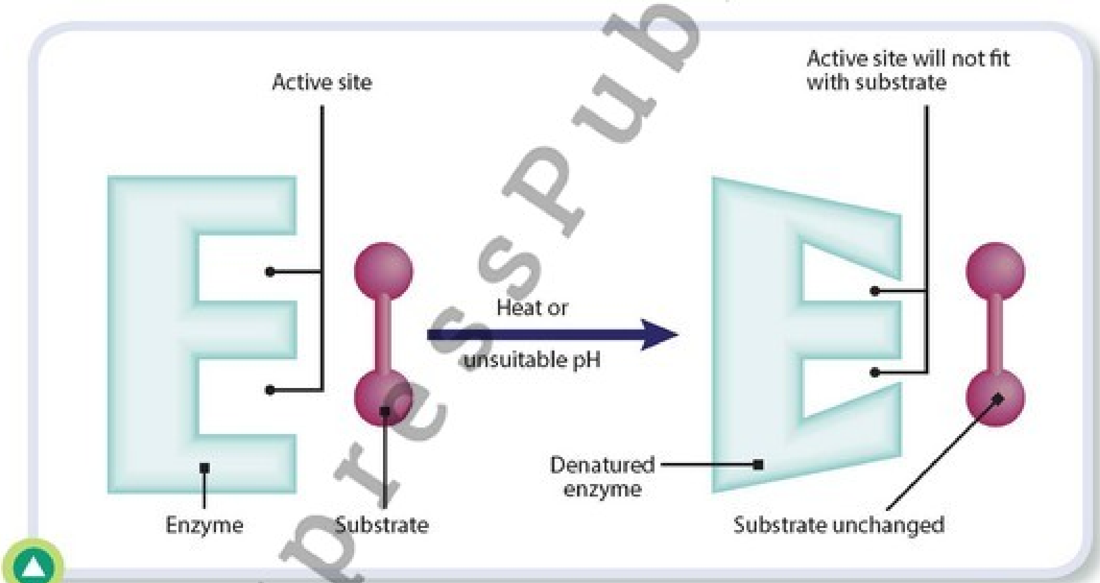


Figure 12.13

In the following experiment the protein peroxidase is extracted from celery. A large volume of celery extract is used as this gives only a small percentage of this enzyme. The **substrate** is hydrogen peroxide, which breaks down to release oxygen. The **product** – oxygen – is trapped by washing-up liquid, forming a foam. The volume of foam produced in a given time is taken as a measure of enzyme activity.

Did you know?

High temperatures or fever in a human can be very serious, as they can affect the shape and the efficiency of enzyme reactions.

Lab Work 17

Activity 12.2

Question

How can we show that enzymes are denatured by an unsuitable pH environment?

Equipment needed

Catalase source (celery/radish/potato/liver) – fresh samples
Blender
Distilled water
Coffee filter paper/coarse filter paper
Hydrogen peroxide (20%)
pH buffers (pH 4, 7, 10, 13)
Washing-up liquid
Dropper
10 cm^3 pipette
100 cm^3 graduated cylinder
Knife
Chopping board
Stopwatch
Water bath
Thermometer

Safety

- Hydrogen peroxide is a corrosive solution; students should use safety glasses and avoid contact with skin.

Conducting the activity

1. Blend three stalks of celery using a hand blender and add 100 cm^3 of distilled water.
2. Filter this solution into a 250 cm^3 beaker using coffee filter paper.
3. Place this solution in a water bath at 25°C.
4. Place 10 cm^3 of this catalase solution into a 100 cm^3 graduated cylinder and place in the water bath.
5. Add 10 cm^3 of pH 4 buffer solution to the graduated cylinder.
6. Using a dropper, add one drop of washing-up liquid.
7. Add 5 cm^3 of hydrogen peroxide into a test tube and place in the water bath.
8. Leave both solutions until they have reached 25°C.
9. Add the hydrogen peroxide to the cylinder.
10. Swirl the cylinder carefully.
11. Start the stopwatch.
12. Note the volume of foam produced after 2 minutes.
13. Repeat using buffers 7, 10 and 13.
14. Draw a graph of the results showing volume of foam against pH.

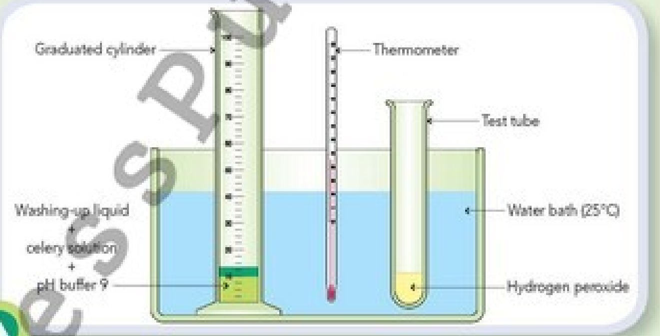


Figure 12.14

12.18 What can you conclude from your results? At what pH is catalase active?

Importance of proteins

Proteins have a wide range of biological functions. They can be categorised into two main types:

- **Globular proteins** which are water soluble have a wide range of metabolic roles. These include: **specific enzymes in all organisms, plasma proteins and the formation of antibodies in mammals**.
- **Fibrous** proteins which are insoluble in water consist of long fibres and mainly repeated sequences of amino acids which are insoluble in water. These proteins tend to have structural roles rather than metabolic roles. These include: the formation of **collagen in bone and cartilage and keratin in the fingernails and hair**.

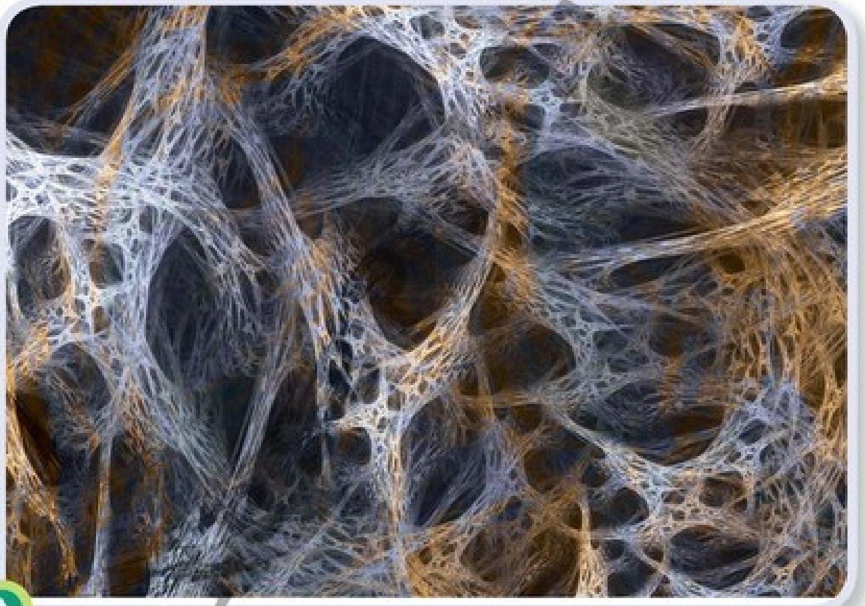


Figure 12.15 *Fibrous proteins in a plant wall cell*

Q U4 Understanding

12.19 Explain the term peptide bond.
12.20 Name two factors that can cause an enzyme to be denatured.
12.21 What happens when a denatured enzyme loses it shape?
12.22 Distinguish between the structural and metabolic roles of proteins.
12.23 Name two differences between fibrous and globular proteins.

Glossary

A

abrasive A very hard, brittle, heat-resistant substance that is used to grind the edges or rough surfaces of an object.

acid An acid is a proton donor. It is a substance which turns blue litmus red and has a pH less than 7.

acid-base indicator A weak acid which is of a different colour from its conjugate base. In solution, it varies in colour as the concentration of hydrogen ions changes.

acid-base reaction This is a reaction involving proton transfer.

acid–base pair A pair of two species that transform into each other by gain or loss of a proton.

acid rain Rainwater with a pH of less than 5.5.

activation energy The minimum energy required to start a reaction by the breaking of bonds.

actual yield The amount of product obtained from a reaction.

adhesive Any substance that is capable of holding materials together in a functional manner by surface attachment that resists separation.

adsorption The process that occurs when a gas or liquid or solute is held to the surface of a solid.

algae Photosynthetic organisms of aquatic or moist habitats, ranging in size from single-celled diatoms to large seaweeds such as kelp, and characterized by a lack of complex organs and tissues.

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.

alkanes The hydrocarbon homologous series with single carbon-to-carbon bonds and the general formula: C_nH_{2n+2}.

alkenes The hydrocarbon homologous series with one double carbon-to-carbon bond and the general formula: C_nH_{2n}.

alkyl group A side chain formed by removing a hydrogen atom removed from an alkane chain, an alkyl group can be represented as R.

alkynes The hydrocarbon homologous series with one triple carbon-to-carbon bond and the general formula: C_nH_{2n-2}.

alloy A metal made by combining two or more metallic elements.

amount of substance The quantity whose unit of the mole is used as a means of counting any species such as atoms, ions and molecules.

amphoteric A substance capable of acting as either an acid or a base.

anhydrous Containing no water molecules.

anion A negatively charged ion with more electrons than protons.

anodising (anodise vb) An electrolytic process for producing thick oxide coatings, usually on aluminium and its alloys.

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

atomic number (Z) This is equal to the number of protons in the nucleus of the atom. It is also equal to the number of electrons in the neutral atom.

atomic orbital A region around the nucleus that can hold up to two electrons, with opposite spins.

Avogadro constant This is the number of atoms in exactly 12g of carbon-12
(6.02×10^{23} mol^{-1})

B

base A base is a proton acceptor.

binary compound A compound containing two elements only.

biodegradable A substance which can be chemically decomposed (broken down to simpler components) by natural biological processes such as soil bacteria, weather, plants and animals.

boiling point The temperature at which a liquid boils.

bond angle The angle between two bonds of an atom.

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

bonded pair A pair of electrons shared between two atoms to make a covalent bond.

Brønsted-Lowry acid A species that is a proton, H + , donor.

Brønsted-Lowry base A species that is a proton, H^+ , acceptor.

buffer solution A solution which is resistant to changes in pH on addition of small quantities of acid or alkali, or upon dilution with water.

C

carbon footprint The total amount of carbon dioxide and other greenhouse gases produced by an individual, organisation, product or event; usually expressed in equivalent tons of carbon dioxide (CO_2).

carbon reservoir A place on Earth where there is a build-up of carbon dioxide.

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

catalyst A substance that increases the rate of a chemical reaction without being used up in the process.

cation A positively charged ion with fewer electrons than protons.

chain reaction A reaction in which the propagation steps release new radicals that continue the reaction.

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.

chlorination The addition of chlorine to water to kill bacteria.

closed system A system isolated from its surroundings.

collide In a chemical reaction, the reactant particles can only react with each other when they bump into one another. When molecules collide, bonds between their atoms can break and new bonds can form to create new molecules.

composites Materials which are made up of at least two different materials.

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.

conjugate acid A species that releases a proton to form a conjugate base.

conjugate base A species that accepts a proton to form a conjugate acid.

conservation of mass When matter is changed from one form into another but there is no change in the overall mass.

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.

cross-contamination The process by which bacteria or other microorganisms are unintentionally transferred from one substance or object to another, with harmful effect.

crystallisation The formation of crystals by cooling a saturated solution or by evaporating off solvent.

D

decompose A decomposition reaction is a type of chemical reaction in which a single compound breaks down into two or more elements or new compounds. These reactions often involve an energy source such as heat, light, or electricity that breaks apart the bonds of compounds.

dehydration An elimination reaction in which water is removed from a saturated molecule to make an unsaturated molecule.

denaturing agent Denaturation – the alteration of a protein shape – can occur when proteins and nucleic acids are subjected to elevated temperature or to extremes of pH, or to non-physiological concentrations of salt, organic solvents, urea, or other chemical agents.

deposit A sediment precipitated out of solution by chemical action.

diatomic molecules Composed of only two atoms, of the same or different chemical elements. Common examples are: hydrogen, oxygen and nitrogen.

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.

diode A passive electronic component that allows current to flow in only one direction.

displacement reaction A reaction in which a more reactive element displaces a less reactive element from an aqueous solution of its ions.

dissociation A general process in which molecules (or ionic compounds such as salts, or complexes) separate or split into smaller particles such as atoms, ions or radicals, usually in a reversible manner.

distillation A process used to separate two miscible liquids with different boiling points, e.g. alcohol and water.

doping The addition of impurities to a semiconductor to control the electrical resistivity.

dynamic equilibrium The equilibrium that exists in a closed system when the rate of the forward reaction is equal to the rate of the reverse reaction and concentrations do not change.

E

electrolysis An electric current is passed through a molten sample of a metal compound causing it to split up and release the metal.

electrolyte A chemical compound that conducts electricity by changing into ions when melted or dissolved into a solution.

electrolytic The addition of a solvent or of energy in the form of heat causes molecules or crystals of the substance to break up into ions (electrically charged particles).

electron A negatively charged particle in an atom.

electron configuration A shorthand representation that shows how electrons occupy sub-shells in an atom.

electronegativity The relative power of an atom in a molecule to attract electrons to itself.

empirical formula The formula that gives the simplest whole-number ratio of atoms of each element present in a compound.

endothermic A reaction where energy is absorbed by the chemicals from its surroundings in order for the reaction to take place.

energy profile diagram Shows the energy transfer in a chemical reaction.

equilibrium constant Kc is a measure of the position of equilibrium. It indicates whether there are more reactants or more products in an equilibrium system.

eutrophication The undesirable overgrowth of vegetation in an ecosystem caused by chemical nutrients, often compounds containing nitrogen, phosphorus, or both. Eutrophication is almost always caused by the discharge of nitrate or phosphate-containing detergents, fertilizers, or sewage into an aquatic system.

exothermic reaction A reaction where energy is transferred from the chemicals to the surroundings.

F

fermentation The process in which a substance breaks down into a simpler substance. Microorganisms like yeast and bacteria usually play a role in the fermentation process, creating beer, wine, bread, kimchi, yogurt and other foods.

fibrous proteins These provide structural support for cells and tissues. Fibrous protein molecules are long, spiral chains that are folded to form strong structures such as hair, fingernails and muscle tissue.

filtration A method used to separate insoluble solids from liquids.

fossil The remains of an animal or plant that have been turned into rock and preserved. They are only found in sedimentary rocks (since the changes that produce metamorphic and igneous rocks would destroy any fossils)

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.

foundry A factory where castings are produced by melting metal, pouring liquid metal into a mould, then allowing it to solidify.

fractional distillation The separation of a mixture into its component parts, or fractions. Chemical compounds are separated by heating them to a temperature at which one or more fractions of the mixture will vaporize.

freezing Where a liquid is cooled and changes into a solid.

fuel A substance that burns in oxygen and produces heat.

functional group An atom or group of atoms joined in a specific manner, which gives the chemical properties of the organic compound and are the centres for chemical reactivity. In organic chemistry it is very common to see molecules comprised mainly of a carbon backbone with functional groups attached to the chain.

G

general formula The simplest algebraic formula of a member of a homologous series, eg. the general formula of the alkanes: C_nH2_{n+2}.

global warming An increase in the temperature in the Earth's atmosphere due to the greenhouse effect.

globular proteins Unlike fibrous proteins, globular proteins have the spiral chains folded into spherical shapes and are responsible for substances in the human body such as haemoglobin, enzymes and certain hormones like insulin.

greenhouse effect When the heat energy of the Sun is trapped within the Earth's atmosphere.

group A vertical column in the periodic table. Elements in a group have similar chemical properties and their atoms have the same number of outer shell electrons.

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.

heterogeneous catalysis A reaction where the catalyst has a different physical state from the reactants, e.g. reactants are gases, whereas the catalyst is a solid.

homogeneous catalysis A reaction in which the catalyst and reactants are in the same physical state, usually gaseous or aqueous state.

homogeneous equilibrium An equilibrium in which all the species making up the reactants and products have the same physical state.

homologous series A group of organic chemical compounds, usually listed in order of increasing size, that have a similar structure (and hence also similar properties) and whose structures differ only by the number of CH_2 units in the main carbon chain.

hydrated A crystalline compound containing water molecules.

hydration A chemical reaction in which a substance combines with water. One of the commonest hydration reactions is the hydration of alkenes to produce alcohols. For example, industrial ethanol is manufactured by the hydration of ethene.

hydrocarbon A compound of hydrogen and carbon only.

hydrogen bond The electrostatic force of attraction between a proton that has been denuded of electrons, by direct attachment to a highly electronegative atom such as N, O or F, and the lone pair on another highly electronegative atom.

hydrogenation The process of hydrogenation involves the use of hydrogen molecules to saturate organic compounds, in the presence of a catalyst.

hydrolysis A reaction with water that breaks a chemical compound into two compounds, the H and OH in a water molecule becomes incorporated into the two compounds.

I

incineration Involves the combustion of organic substances contained in waste materials. Waste destruction takes place in a furnace by controlled burning at high temperatures.

indicators Chemicals that show by changing colour whether a substance is acidic, alkaline or neutral.

initial rate of reaction The change in concentration of a reactant or product per unit time at the start of the reaction: $t = 0$.

insoluble Solids that do not dissolve in a liquid.

intermediate A species formed during a reaction that reacts further and is not present in the final products.

intermolecular force The forces which mediate interaction between molecules, including forces of attraction or repulsion.

intermolecular hydrogen bonds Hydrogen bonds formed between neighbouring molecules.

ion A positively or negatively charged atom or a (covalently bonded) group of atoms.

ionic bond An ionic bond is the electrostatic force of attraction between oppositely charged ions, formed as a result of complete electron transfer.

isomers These are molecules that have the same molecular formula but different structural or spatial arrangements of the atoms within the molecule.

isotopes Atoms that have the same number of protons but different numbers of neutrons.

L

Le Chatelier's principle If the equilibrium of a system is disturbed by a change in one or more of the determining factors (as temperature, pressure, or concentration) the system tends to adjust itself to a new equilibrium by counteracting as far as possible the effect of the change.

limewater A chemical that turns from clear to cloudy (or milky) if carbon dioxide is present.

limiting reagent The reactant that is not in excess, which will be used up first and stop the reaction.

lone pair An outer shell pair of electrons that is not involved in chemical bonding.

lubricant A substance, usually organic, introduced to reduce friction between surfaces in mutual contact, which ultimately reduces the heat generated when the surfaces move.

M

macromolecular structure Covalent structures that have many atoms joined together by covalent bonds. Diamond and graphite are examples of macromolecular structures.

malleable Malleability is one of the unique properties of metals which allow them to change shape by hammering or pressing to form sheets.

mass The amount of matter in an object. This amount never changes.

mass number Equal to the total number of protons plus neutrons in the nucleus of the atom.

matter Anything that occupies space and has mass.

melting point The temperature at which solid materials turn to a liquid.

metabolic reactions The chemical processes that occur in all living organisms, which help them complete their life cycles, such as digestion and respiration.

metabolism The whole range of biochemical processes that occur within a living organism consisting of anabolism (the build-up of substances) and catabolism (the breakdown of substances).

metallic bond The electrostatic attraction between positive metal ions and delocalised electrons.

minerals The chemical substances from which rocks are made.

miscible liquids Liquids that mix, e.g. alcohol and water.

mixture Contains two or more different substances mixed together but not chemically combined.

molar gas volume The volume per mole of gas molecules at a stated temperature and pressure.

molar mass The mass per mole of a substance, in units of $g\ mol^{-1}$.

mole 1 mole is the amount of substance which contains the same number of particles as there are atoms in exactly 12g of carbon- 12.

mole fraction The number of moles of a particular component of a mixture divided by the total number of moles present.

molecular formula This shows the number of each kind of atom in a molecule. For example, CH_4 has one carbon atom and four hydrogen atoms.

molecule The smallest part of a covalent compound that can exist while retaining its chemical identity, consisting of two or more atoms covalently bonded together.

molten An object that is reduced to liquid form by heating.

monomer A small molecule that combines with many other monomers to form a polymer.

N

neutralisation A chemical reaction in which an acid and a base react together to produce a salt.

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.

nomenclature A system of naming compounds.

non-polar With no charge separation across a bond or in a molecule.

non-renewable Sources of energy that cannot be replaced once they are used.

nuclear energy Energy from making and breaking nuclear bonds.

nucleus The central part of an atom that is made up of protons and neutrons.

nutrients Chemical compounds in food that are used by the body to function properly and maintain health, including proteins, fats, carbohydrates, vitamins, and minerals.

O

oxidation Loss of electrons or an increase in oxidation number.

oxidation number A measure of the number of electrons that an atom uses to bond with atoms of another element. Oxidation numbers are derived form a set of rules.

oxidation state The oxidation number.

oxidising agent An oxidising agent (oxidant) is an electron acceptor.

P

π-bond A bond formed by the sideways overlap of two p-orbitals, with the electron density above the plane of the bonding atoms.

partial dissociation The splitting of some of a species in solution into ions.

percentage yield Actual amount, in mol, of product theoretical amount, in mol, of product × 100.

period A horizontal row of elements in the periodic table. Elements show trend in properties across a period.

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.

Periodic trend A repeating trend in properties of the elements across each period of the periodic table.

pH scale The pH of a solution is the negative logarithm to base ten of the molar hydrogen ion concentration.

physical change A change in which no new substance is formed.

planar A planar molecule is one in which the atoms are all on the same plane: the molecule is flat.

plastics Man-made materials made from crude oil.

pollution The addition of harmful or unwanted materials to an environment.

polyatomic ion An ion containing more than one atom.

polymer A large molecule formed from many thousands of repeat units of smaller molecules known as monomers.

polysaccharides A long-chain carbohydrate made up of smaller carbohydrates called monosaccharides, typically used by our bodies for energy or to help with cellular structure.

polythene A light, versatile synthetic resin made from the polymerization of ethylene. Polythene is resistant to chemicals and moisture, has good insulating properties, and can be used especially in packaging and insulation.

position of equilibrium The relative quantities of reactants and products, indicating the extent of a reversible reaction at equilibrium.

precipitates Often formed when two aqueous solutions are mixed together.

precipitation The emergence of the insoluble solid from solution is called precipitation.

precipitation reaction The formation of a solid from a solution during a chemical reaction.

primary alcohol An alcohol in which the OH group is attached to a carbon atom that is attached to two or three hydrogen atoms.

proton A positively charged subatomic particle found in the nucleus of an atom.

proton number The number of protons in the nucleus of an atom; also known as atomic number.

R

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

rate of reaction The change in concentration of a reactant or a product in a given time.

reaction mechanism The sequence of bond-breaking and bond-forming steps that shows the path taken by electrons during a reaction.

redox reaction A reaction involving electron transfer.

reducing agent A reducing agent -reductant- is an electron donor.

reduction Gain of electrons or a decrease in oxidation number.

relative atomic mass The average mass per atom of the natural isotopic composition of the element, relative to the mass of an atom of carbon-12, which is assigned a value of exactly 12.

relative molecular mass The average mass per molecule in the naturally occurring substance, relative to the mass of an atom of carbon-12 which is assigned a value of exactly 12 units. Its value is equal to the sum of all the relative atomic masses which make up the molecular formula.

reversible reaction A reaction that takes place in both forward and reverse.

rigidity The property of a structure that it does not bend or flex under an applied force. The opposite of rigidity is flexibility.

S

σ-bond A bond formed by the overlap of one orbital from each bonding atom, consisting of two electrons and with the electron density centred around a line directly between the nuclei of the two atoms.

salt The product of a reaction in which the H–ions from the acid are replaced by metal or ammonium ions.

saturated hydrocarbon A hydrocarbon with single bonds only.

saturated solution A solution that contains as much dissolved solute as possible at that temperature.

sealant A substance used to block the passage of fluids through the surface or joints or openings in materials.

secondary alcohol An alcohol in which the –OH group is attached to a carbon atom that is attached to two carbon chains and one hydrogen atom.

semi-viscous Property of a liquid with moderate resistance to flow.

shell A group of atomic orbitals with the same principal quantum number, n.

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.

species If one atom is identical to another, we can say they are the same chemical species.

spectator ions Ions that are present but take no part in a chemical reaction.

standard conditions Pressure =100 kPa temperature =298 K concentration (reactions with aqueous solutions) = 1 mol dm^{-3}

strong acid An acid that dissociates completely in solution.

structural formula A formula showing the minimal detail for the arrangement of atoms in a molecule.

structural isomers Molecules with the same molecular formula but with different structural formulae.

substitution reaction A reaction in which an atom or group of atoms is replaced with a different atom or group of atoms.

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.

tetrahedral A term used to describe molecules and polyatomic ions that have one atom in the centre and four atoms at the corners of a tetrahedron.

theoretical yield The yield resulting from complete conversion of reactants into products.

thermal decomposition The breaking up of a chemical substance with heat into at least two chemical substances.

Time It is measured by a basic unit called the second (symbol s).

trace elements A chemical element whose concentration (or other measure of amount) is very low (a "trace amount").

transistor A device that regulates current or voltage flow and acts as a switch or gate for electronic signals.

transition metals A metal which forms one or more stable ions which have incompletely filled d orbitals.

U

unsaturated When referring to chemical solutions, an unsaturated solution is able to dissolve more solute; when referring to organic compounds, unsaturated means a molecule contains double or triple carbon-carbon bonds.

V

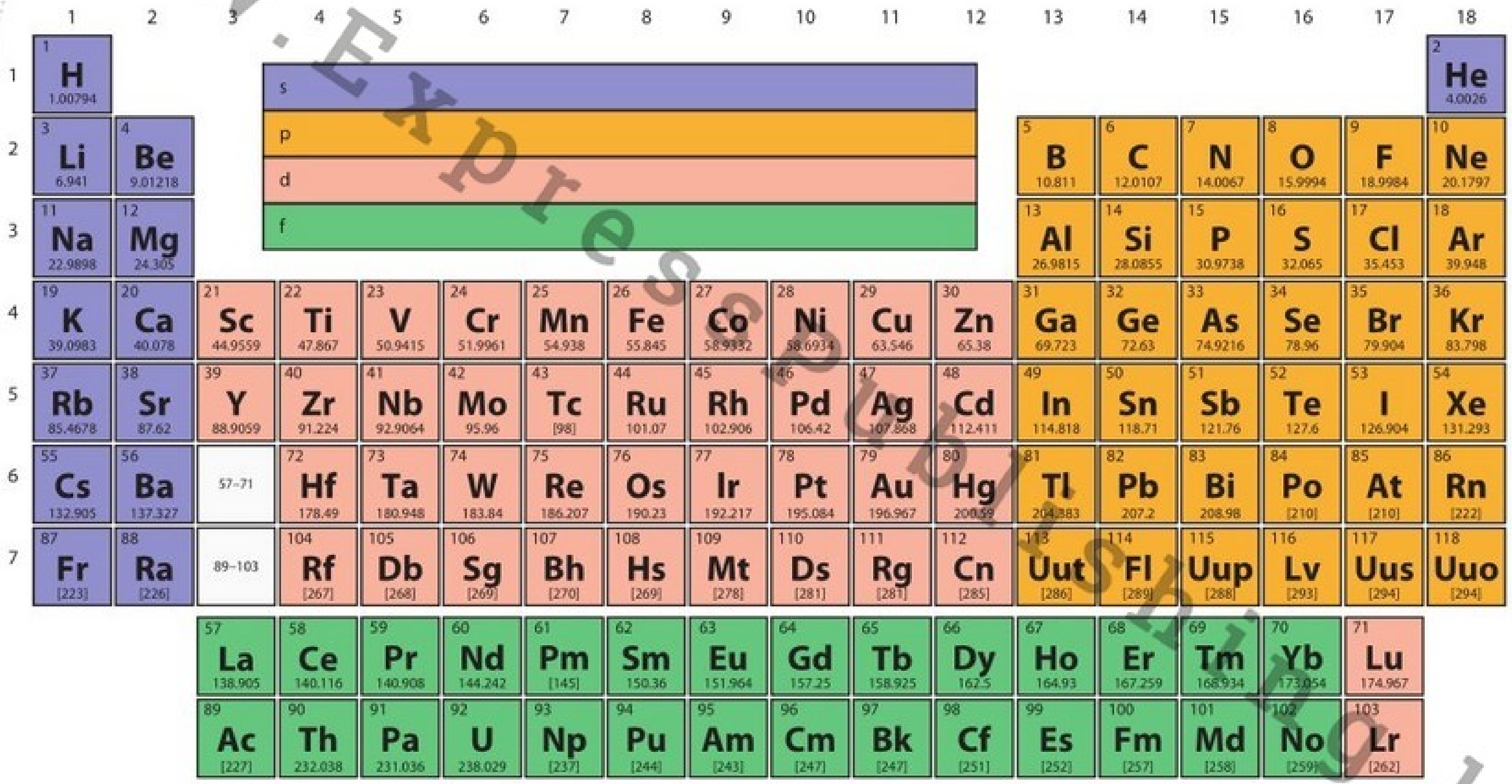
valency of an element The number of electrons an atom of that element wants to gain, lose or share, in order to be chemically stable and have a full outer shell.

volatility The ease with which a liquid turns into a gas. Volatility increases as boiling point decreases.

W

water of crystallisation Water molecules that are bonded into a crystalline structure of a compound.

weak acid An acid that dissociates only partially in solution.



- 🟪 s
- 🟧 p
- 🟥 d
- 🟩 f

	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]	

GRADE

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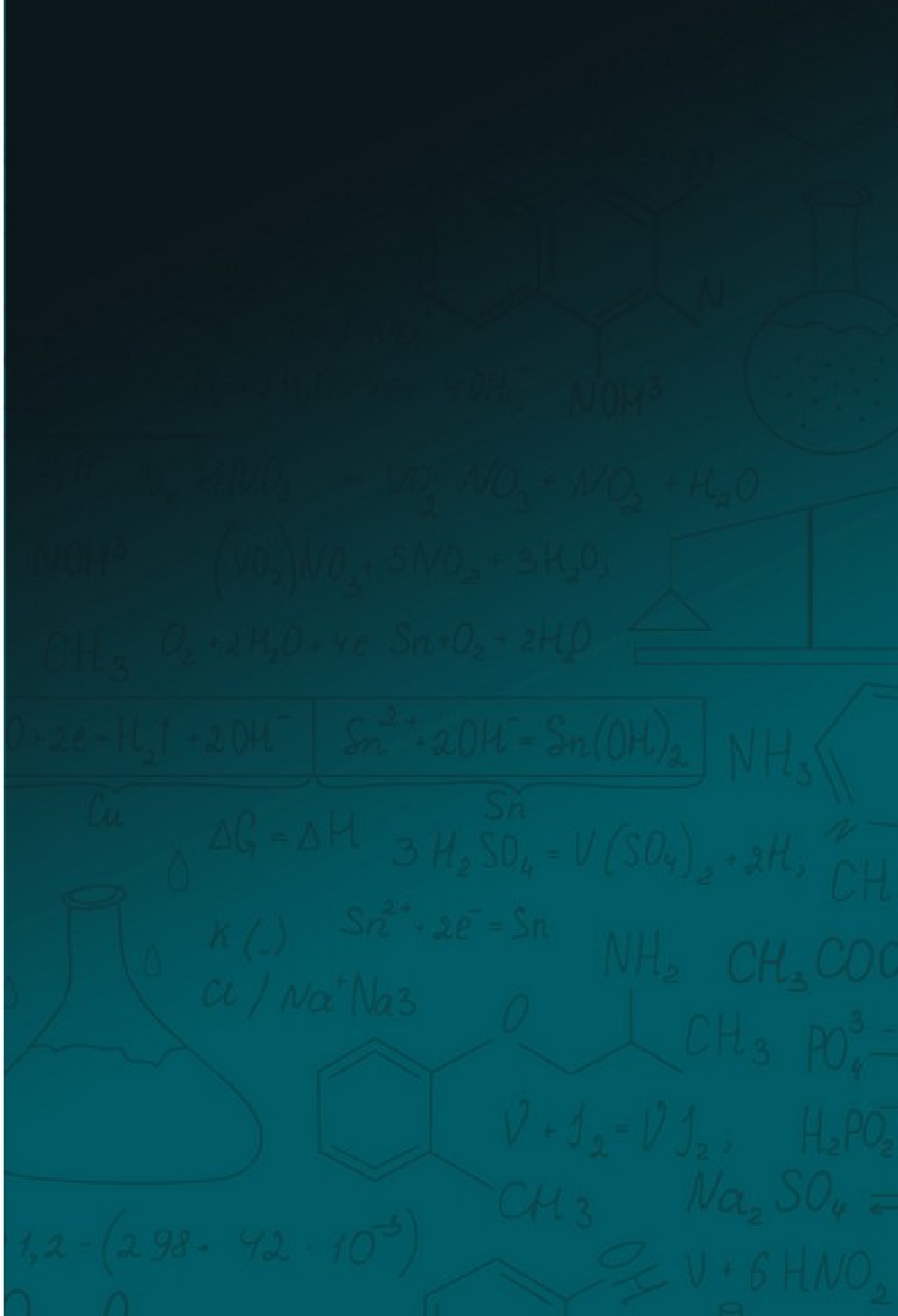
Chemistry



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